The State of Eritrea
Ministry of Education

Chemistry

for

Grade 10

Textbook

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Introduction

Dear learners,

Welcome to the Grade 10 Chemistry Textbook!

This textbook contains five units: Periodic Table of Elements, Forces of Attraction, Solutions and their Properties, Acids, Bases and Salts, and Energy and Chemical Reaction. Each unit is further divided into sub-units. You are already familiar with some of the topics but they are presented here in more detail. Some of the units contain new topics. Therefore you are continuing your journey of discovery in chemistry by enriching the knowledge you already have and by exposing yourself to more advanced topics in chemistry.

This book has been designed and prepared to make chemistry fascinating by making the teaching-learning process activity-oriented, practical, learner-centered and interactive. It contains numerous activities that will be conducted inside and outside your class room, and in your laboratories. These activities will help you to develop the skills and creative abilities necessary for the study of chemistry. While dealing with the activities, you are expected to understand the objective, conduct experiments, observe any changes that may take place, and draw conclusions from the activity-related questions. As so much is expected of you, preparation in advance is required as much as is relating the result to the concept of the subject-matter.

In addition to the activities, the book includes features such as let us discuss, examples, exercises, Did you know?, summary, review questions and key terms. The Let us discuss feature is meant to be used to promote brainstorming what has already been studied with your seat mates, and also to be used to stimulate your recollections of background or previously learned materials in preparation for new topic about to be introduced. The Let us discuss feature may also be a physical example that you encounter in everyday activity and is introduced in the text to guide you towards a better understanding of the particular chemistry concept. Therefore, you are expected to deal with Let us discuss feature before you proceed to the explanation of the subject.

Examples are introduced as supplementary to the theory in order to make a topic clearer. Exercises are questions that you need to do as you come across them. Review Questions are intended to help you revision the subject-matter in each unit. Both Exercises and Review Questions will help you to understand what is being taught and to test to what extent you have understood the topic. The frequent summaries will help you to remember easily what you have learned before proceeding to the next sub-unit or unit. Key Terms will help you to develop your scientific vocabulary and to make quick revision.

Assessment is part and parcel of the teaching-learning process. It is an essential tool for checking your progress in learner-centered interactive pedagogy. Therefore, as the learning proceeds, you will be continuously assessed using various techniques such as homework, class-work, quizzes, tests, examinations, and project work. After each assessment your teacher will give you feedback on your learning progress. This will help you and your teacher to measure whether you are achieving the stated learning outcomes for your grade.

Parents or guardians and your community need to understand to play their role in the teaching-learning process. Their involvement and follow-up motivate learners to develop a positive attitude towards schooling and as a result enhance their achievements.

Your valuable comments and suggestions that could contribute to the improvement of this textbook will be most welcome.

Contents

	Units	Page
1	**Periodic Table of Elements.....**	**1**
	1.1 Historical Development.....	1
	1.1.1 Importance of Classification of the Elements in Chemistry.....	1
	1.1.2 Historical Development of the Periodic Table.....	2
	1.2 The Modern Periodic Table…….....	10
	1.2.1 Modern Periodic law…….....	10
	1.2.2 General Structure of the Modern Periodic Table.....	11
	1.3 Electron Configuration and the Periodic Table...…….....	17
	1.3.1 Electron Configuration …….....	17
	1.3.2 Abbreviation of Electron configuration...	21
	1.3.3 Classification of Elements in the Modern Periodic Table Based on Electron Configuration……	21
	1.4 Periodic Regularities and Characteristics of Elements…….....	24
	1.4.1 Trends in Atomic and Ionic Sizes (Atomic and Ionic Radii).....	24
	1.4.2 Trends in Ionization Energy…….....	28
	1.4.3 Trends in Electronegativity…….....	30
	1.4.4 Trends in Metallic and Nonmetallic Characteristics ….....	32
	1.4.5 Patterns of Chemical Formulas of Compounds……	34
	1.4.6 Summary of Trends in Periodic Properties.....	34

2	**Chemical Bonding and Forces of Attraction ……**	**41**
	2.1. Forces of Attractions……..	41
	2.2 Chemical Bonds.....	42
	2.2.1 Ionic Bonding	42
	2.2.2 Covalent Bonding …….....	43
	2.2.3 Polar and Non polar Covalent Bonds.....	45
	2.2.4 Coordinate Covalent Bonding……	48
	2.2.5 Metallic Bonding……	50
	2.3 Intermolecular Forces of Attraction…...	52
	2.3.1. Ion-Dipole Forces……..	52
	2.3.2. Van der Waals Forces.....…….	53
	2.3.3. Hydrogen Bonding…….……	55
3	**Solutions and Their Properties……**	**63**
	3.1 Solutions…….....	63
	3.1.1 Describing Solutions……	63
	3.1.2 Unsaturated, Saturated and Supersaturated Solutions…….	66
	3.1.3 Solubility and Solubility Curve……..	69
	3.2 Calculations Involving Solutions	73
	3.2.1 Expressing Concentration….....	73
	3.2.2 Diluting of solution	92
	3.3 Properties of Solution…….....	96
	3.3.1 Dissolution of Solutes in Liquids.	96
	3.3.2 Effect of Temperature and Pressure on Solubility…….....	98
	3.3.3 Colligative Properties of Solutions…….	101

4	**Acids, Bases and Salts……**	**113**
	4.1 Acids……	113
	4.1.1 Describing Acids ….....	113
	4.1.2 Classification of Acids …….....	115
	4.1.3 General Methods of Preparation of Acids	119
	4.1.4 General Properties of Acids…….....	121
	4.1.5 Uses of Common Acids……	123
	4.2 Bases …….....	125
	4.2.1 Describing Bases ….....	125
	4.2.2 Classification of Bases …….....	126
	4.3.3 General Methods of Preparation of Bases	127
	4.2.4 General properties of Bases……	127
	4.2.5 Uses of Common Bases……	128
	4.3 Describing Neutralization Reaction…	130
	4.3.1 Titration……	133
	4.3.2 Determining Concentration Using Titration…….	134
	4.3.3 Using Indicators in a Titration…….	138
	4.4 Salts……	139
	4.4.1 Describing Salts ….....	139
	4.4.2 Classification of Salts …….....	140
	4.4.3 General properties of Salts……	141
	4.4.4 General Methods for Preparation of Salts	141
	4.4.5 Uses of Common Salts……..	143
	4.5 The pH Scale. Measurement of Acidity and Alkalinity……	144
	4.5.1 Measuring pH……	147

5	**Energy and Chemical Reactions ….**	**153**
	5.1 Energy Transformation and Chemical Reactions ……...	153
	5.1.1 Law of Conservation of Energy……...	153
	5.1.2 Units of Heat Changes……	154
	5.1.3 Bond Breaking and Bond Formation…..	156
	5.1.4 Enthalpy of a Chemical Reaction…….	157
	5.1.5 Exothermic and Endothermic Reactions	158
	5.1.6 Energy-Level Diagrams……	161
	5.1.7 Common Examples of Exothermic Reactions…….	164
	5.1.8 Common Examples of Endothermic Reactions……	167
	5.1.9 Thermochemical Equations……	170
	5.2 Types of Heat of Reaction and Heat Changes …….....	175
	5.2.1 Heat of Combustion……..…	175
	5.2.2 Heat of Solution……..…	175
	5.2.3 Heat of Neutralization……..…	175
	5.2.4 Heat of Formation…….……	176
	5.2.5 Heat of Fusion…….....	176
	5.2.6 Heat of Vaporization……...…	176

UNIT 1
Periodic Table of Elements

Expected learning outcomes

By the end of this unit, you are expected to:

- describe the historical development of the periodic table and the arrangement of the elements in the modern periodic table
- write the electron configuration of some common elements and classify them into groups according to their electron configuration
- explain the physical and chemical characteristics of elements based on periodic regularities.

1.1 Historical Development

1.1.1 Importance of Classification of the Elements in Chemistry

Let us discuss

Consider the names of the following people together with their respective genders and ages.
Solomon (male, forty years old), Aisha (female, twenty-one years old), Zemichael (male, fifty-two years old), Almaz (female, twenty-one years old), Halima (female, forty years old) and Abdu (male, fifteen years old).

1. Classify the people listed above into groups.
2. What did you base your classifications on?

In grade nine, you learnt that ”matter” is present in the form of elements, compounds and mixtures and that elements contain atoms of only one type. Elements are the fundamental forms of matter.

At present, there are about one hundred and fifteen known elements. 1800s, only thirty elements were known. All of these had seemingly different properties. As new elements were discovered, scientists gathered more and more information about the properties of these elements. They found it difficult to organize the information that was known about the elements. They started looking for a pattern in the

properties of the elements, a base from which they could study such a large number of elements with ease. Scientists made several attempts to classify the elements according to their properties and obtain an orderly arrangement of elements. This orderly arrangement of elements helps us to study, remember and understand the relationships between the elements. It provides a basis for understanding chemistry.

The earliest attempt to classify the elements resulted in grouping the known elements into the metals and the nonmetals. Later, further classifications were attempted as the knowledge of the elements and their properties increased.

Activity 1.1: Classification of Items

Consider the following substances:
Lux, biscuits, rulers, green peppers, Lifebuoy, tomatoes, oranges, pens, onions, bananas, sharpeners, potatoes, pencils, carrots, erasers, sweets, apples, chocolates, and grapes.

1. Classify the above substances into groups?
2. What is the basis for your classifications?
3. What are the advantages of classifying things into groups?

1.1.2 Historical Development of the Periodic Table

Let us discuss

1. List a few means of transport that you are familiar with.
2. Classify each as either a primitive or a modern mode of transportation.
3. Over time, do you think the mode of transportation has improved?

It was difficult to arrange the elements in a meaningful way. It took a long time and many different systems of classification were explored. Some of the important systems of classification are briefly presented according to their historical development as follows.

Dobereiner's Triads

In the year 1829, Dobereiner, a German chemist, attempted to arrange the elements with similar properties into groups. He noticed

that elements with similar properties could be placed into groups of three. He called these groups **'triads'**. Dobereiner noticed that when the three elements in a triad were written in order of increasing atomic mass; the atomic mass of the middle element was roughly the average of the atomic masses of the other two elements. For example, consider the following elements: lithium (Li), sodium (Na) and potassium (K) with their respective atomic mass of 6.9, 23 and 39.The average of the atomic masses of Li and K is 23 which is the atomic mass of Na. Therefore, Li, Na and K form a Dobereiner triad. Similarly, sulphur (S), selenium (Se) and tellurium (Te) with their respective atomic mass of 32, 79 and 127.6 form a Dobereiner triad. This is summarized in Table 1.1.

Triad elements	**Atomic mass**	**Average mass**
Li Na K	6.9 23 39	23
S Se Te	32 79 127.6	80

Table 1.1 Examples of Dobereiner's Triads

Activity 1.2: Dobereiner Triads

In the table below are groups of three elements. These elements are arranged downwards in order of increasing atomic mass.

Group A elements	**Atomic mass**	**GroupB elements**	**Atomic mass**	**Group C elements**	**Atomic mass**
N P As	14 31 74.9	Ca Sr Ba	40.1 87.6 137.3	Cl Br I	35.5 79.9 126.9

1. Can you determine which of these groups form Dobereiner triads? Explain your answer.
2. Do Dobereiner triads apply to all elements? Explain your answer.

The Dobereiner triad system of classification cannot be applied to all elements. A large number of similar elements could not be grouped into triads. For example, nitrogen, phosphorus and arsenic are similar elements but they could not be placed into the triads. This system of classification into triads was not found to be useful for all elements.

Newlands' Law of Octaves

Activity 1.3: Sequential Grouping

Consider the following learners (1, 2, 3 and so on) and their favourite fruits.

Learner	**Favourite fruit**	Learner	**Favourite fruit**
1	Orange	8	Orange
2	Apple	9	Apple
3	Guava	10	Guava
4	Banana	11	Banana
5	Mango	13	Mango
6	Strawberry	14	Strawberry
7	Wine grapes	15	Wine grapes

1. How many different fruits are listed in the table above?
2. Which learners have the same favourite fruit?
3. How often are these favourite fruits repeated?

Dobereiner`s attempts encouraged other chemists to correlate the properties of the elements with their relative atomic mass. In 1864, John Newlands, an English chemist, arranged all elements known at the time into a table in order of relative atomic mass. When these elements were arranged in order of increasing,relative atomic mass, he found that every eighth element had properties similar to those of the first element. He compared this to the octaves found in music. Therefore, he called it the **"Law of Octaves"**. In Newlands' octaves, the properties of lithium and sodium were found to be the same. Sodium is the eighth element after lithium.

Similarly, beryllium and magnesium resemble each other. A part of the original form of Newlands' octaves is given in Table 1.2.

do	**re**	**mi**	**fa**	**so**	**la**	**ti**
H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe
Co and Ni	Cu	Zn	Y	In	As	Se
Br	Rb	Sr	Ce and La	Zr	-	-

Table 1.2 Newlands' octaves

In order to arrange the elements into his table, Newlands placed two elements (Co and Ni) in the same row. He also placed a few unlike elements under the same column. Note that cobalt and nickel are in the same row and these are placed in the same column as fluorine, chlorine and bromine which have very different properties to the elements shown in Table 1.2. Iron, which resembles cobalt and nickel in properties, has been placed far away from these elements. Newlands' Law of octaves only worked well with for elements having a small. It was found that the law could only be applied up to calcium. This was one of limitations because after calcium every eighth element did not possess properties similar to those of the first element.

Did you know?



In the western system of music, there are seven musical notes in a scale :– *do, re, mi, fa, so, la, and ti.* The notes in a scale are separated by whole and half-step frequency intervals of tones and semi-tones. A musician uses these notes to for composing the music for a song. Naturally, there must be a repetition of notes. Every eighth note is similar to the first note and it is the first note of the next scale.

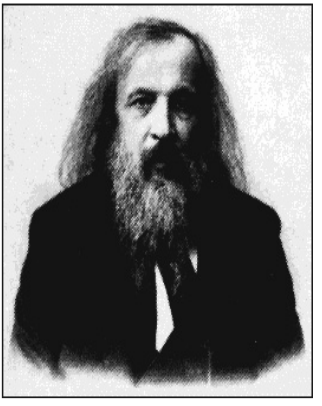
Exercise 1.1

1. What are the limitations of Dobereiner's method of classification?
2. What are the limitations of Newlands' law of octaves?
3. Refer to Table 1.2, and determine whether Dobereiner's triads also exist in the columns of Newlands' octaves?

Mendeleev's Periodic Table

Even after the rejection of Newlands' law of octaves, many scientists continued to search for a pattern that correlated the properties of the elements to their atomic mass.

Did you know?



Dmitri Ivanovich Mendeleev (1834-1907). Mendeleev constructed a periodic table as part of his effort to systematize chemistry and he received many international awards for his work.

The main credit for classifying elements goes to Mendeleev, a Russian chemist. He was the most important contributor to the early development of a periodic table of elements wherein the elements were arranged on the basis of their atomic mass, and also on the similarities of their chemical properties.

When Mendeleev started his work, there were approximately sixty-three known elements. He examined the relationship between the atomic mass of the each element and their physical and chemical properties. Among the chemical properties, Mendeleev concentrated on the compounds formed by the elements oxygen and hydrogen. He selected hydrogen and oxygen as they are very reactive and form compounds with most elements. The formulas of the hydrides and oxides formed by an element were treated as one of the basic

properties of an element for its classification. He then made sixty-three cards and on each card he wrote down the properties of one element. He arranged the cards, first in order of increasing atomic mass, then into groups with similar properties, and attached the cards onto a wall. He observed that most of the elements had a position in a periodic table and were arranged in order of increasing atomic mass. When the elements were arranged according to increasing atomic mass, he observed that a periodic recurrence of elements with similar physical and chemical properties, occurs. This is called Mendeleev's periodic law**.**

Mendeleev's Periodic Law: The properties of elements are the periodic function of their atomic mass

A tabular arrangement of the elements into rows and columns, highlighting the regular repetition of the properties of the elements is called a **periodic table.** Mendeleev's periodic table contains vertical columns called **groups.** The elements in the same group were called **families.** The horizontal rows are called **periods**.

Advantages of Mendeleev's Periodic Table

Mendeleev's periodic table has several advantages compared with previous attempts made to classify the elements.

1. Mendeleev classified elements into groups and periods.
2. Mendeleev left gaps in his periodic table. Instead of regarding these gaps as defects, Mendeleev boldly predicted the existence of elements not yet discovered at this time, for example, gallium was discovered later. When new elements were discovered, they could be placed into a group without disturbing the existing arrangement.
3. He also predicted the properties of these undiscovered elements which were likely to fit into the gaps. Mendeleev named them by prefixing a sanskrit numeral, *eka* (one) to the name of the preceding element in the same group. For instance, scandium, gallium and germanium, which were discovered later, have properties similar to *eka*–boron, *eka*–aluminium and *eka*–silicon, respectively.

For example, the properties of *eka*–Aluminium predicted by Mendeleev and those of the element, gallium which was discovered later and replaced *eka*-aluminium, are shown in Table 1.3.

Property	***Eka*–Aluminium**	**Gallium**
Atomic Mass	68	69.7
Density	5.9 g/cm^3	5.91 g/cm^3
Melting point	Low	30.1 ^{0}C
Boiling point	High	1983 ^{0}C
Formula of an Oxide	E_2O_3	Ga_2O_3
Formula of Chloride	ECl_3	$GaCl_3$

Table 1.3 Predicted and actual properties of gallium

Mendeleev's periodic table is the foundation for the development of the modern periodic table. However, as with previous contributors to the periodic table, it had limitations.

Limitations of Mendeleev's Classifications

1. Position of Hydrogen

Mendeleev did not clearly indicate the position of hydrogen in his periodic table. This was the first limitation of Mendeleev's periodic table.

2. Disorder in Atomic Mass for Some Pairs of Elements

Another problem was that the atomic mass of the elements did not increase in a regular manner when moving from one element to the next. Mendeleev arranged iodine (126.9 atomic mass) after tellurium (127.6 atomic mass), violating his own period law. Table 1.4 is a small section of Mendeleev's periodic table that illustrates the arrangement of tellurium and iodine. Similarly, he arranged potassium (39.1 atomic mass) after argon (39.9 atomic mass).

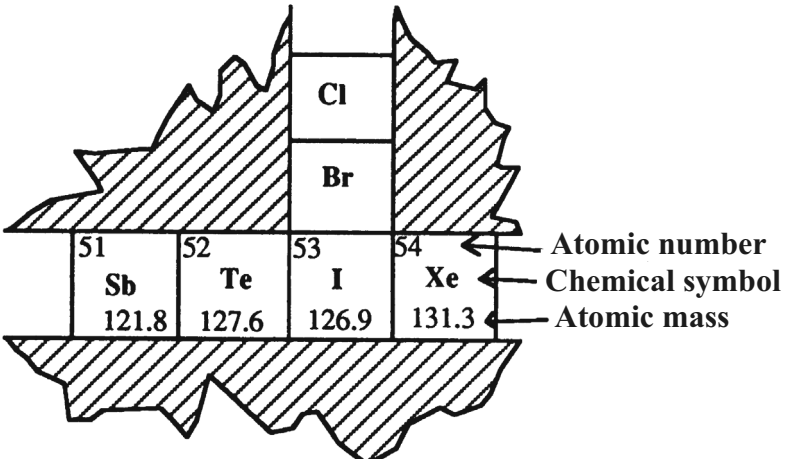


Table 1.4 Irregularities in Mendeleev's periodic table

3. Position of Isotopes

Isotopes were discovered long after Mendeléev had proposed his periodic classification of elements. Isotopes are atoms whose nuclei have the same atomic number but different mass number. According to Mendeleev's periodic law, the isotopes of the same element, for example, Cl-35 and Cl-37, would be placed in different families.

Exercise 1.2

1. What were the criteria used by Mendeleev to create his periodic table?
2. Besides gallium, which other elements have since been discovered that Mendeleev left gaps for in his periodic table? (Name any two elements),
3. What are the advantages and disadvantages of Mendeleev's periodic table?

Summary

There are about 115 elements known today. In order to study, remember and understand the relationships among the elements, scientists have devised the periodic table. Periodic table is a tabular arrangement of elements in rows and columns, highlighting the regular repetition of properties of the elements. In periodic table elements are arranged in an orderly and systematic manner.

Many scientists made several attempts to classify the elements according to their properties and obtain an orderly arrangement of elements.

Dobereiner is one scientist who arranged the elements into groups of three, called the Dobereiner triads, in which the atomic mass of the middle element is roughly the average mass of the other two. Dobereiner's, triad' method of classification did not work for all elements.

Newlands attempted to arrange the elements in order of increasing atomic mass and he found that every eighth element had properties similar to that of the first. Therefore, he called it the "Law of Octaves". Newlands' system of classification failed because it did not work for elements beyond calcium.

Mendeleev was the most important contributor to the early development of a periodic table of elements wherein the elements were arranged on the basis of their fundamental properties and in order of increasing atomic mass. Mendeleev's periodic table contains vertical columns called groups**.** The elements in the same group were called families**.** The horizontal rows were called periods. However, Mendeleev's table had limitations regarding the position of hydrogen, isotopes and a disorder in atomic mass for certain pairs of elements.

1.2 The Modern Periodic Table

1.2.1 Modern Periodic Law

Let us discuss

1. How did Mendeleev arrange elements in his periodic table?
2. What were Mendeleev's limitations?

The need for the modern periodic table arises from the shortcomings of Mendeleev's periodic table. In 1913, Henry Moseley, an English scientist showed that the atomic number of an element is a more fundamental property than its atomic mass because the atomic number of elements is unique. Accordingly, Mendeleev's periodic

law was modified and atomic number was adopted as the basis of modern periodic table. In modern periodic table elements are arranged according to their increasing atomic number and thus the modern periodic law is stated as follows.

Modern Periodic law: The properties of elements are periodic function of their atomic numbers

Let us recall that the atomic number gives us the number of protons in the nucleus of an atom and this number increases by one in going from one element to the next. This leads us to the classification known as the modern periodic table. Refer to Table 1.5. Prediction of properties of elements could be made with more precision when elements were arranged on the basis of increasing atomic number. The position of cobalt, nickel and isotopes (Cl-35 and Cl-37) which were the limitations of Mendeleev's are resolved in the modern periodic table.

In modern periodic table, the position of hydrogen still did not get its right place. In some modern periodic table, it is placed separately.

1.2.2 General Structure of the Modern Periodic Table

In the modern periodic table, elements are arranged horizontally in increasing order of their atomic number. In general, the modern periodic table shows the symbol for each element, its atomic number, atomic mass, period and group in which it is located.

THE PERIODIC TABLE

1 IA												13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
H 1 1.008 Hydrogen	2 IIA																He 2 4.00 Helium
Li 3 6.94 Lithium	Be 4 9.01 Beryllium											B 5 10.81 Boron	C 6 12.01 Carbon	N 7 14.01 Nitrogen	O 8 16.00 Oxygen	F 9 19.00 Fluorine	Ne 10 20.18 Neon
Na 11 22.99 Sodium	Mg 12 24.31 Magnesium	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9	10	11 IB	12 IIB	Al 13 26.98 Aluminum	Si 14 28.09 Silicon	P 15 30.97 Phosphorus	S 16 32.07 Sulfur	Cl 17 35.45 Chlorine	Ar 18 39.95 Argon
K 19 39.10 Potassium	Ca 20 40.08 Calcium	Sc 21 44.96 Scandium	Ti 22 47.88 Titanium	V 23 50.94 Vanadium	Cr 24 52.00 Chromium	Mn 25 54.94 Manganese	Fe 26 55.85 Iron	Co 27 58.93 Cobalt	Ni 28 58.69 Nickel	Cu 29 63.55 Copper	Zn 30 65.39 Zinc	Ga 31 69.72 Gallium	Ge 32 72.61 Germanium	As 33 74.92 Arsenic	Se 34 78.96 Selenium	Br 35 79.90 Bromine	Kr 36 83.80 Krypton
Rb 37 85.47 Rubidium	Sr 38 87.62 Strontium	Y 39 88.91 Yttrium	Zr 40 91.22 Zirconium	Nb 41 92.91 Niobium	Mo 42 95.94 Molybdenum	Tc 43 (97.9) Technetium	Ru 44 101.07 Ruthenium	Rh 45 102.91 Rhodium	Pd 46 106.42 Palladium	Ag 47 107.87 Silver	Cd 48 112.41 Cadmium	In 49 114.82 Indium	Sn 50 118.71 Tin	Sb 51 121.76 Antimony	Te 52 127.60 Tellurium	I 53 126.90 Iodine	Xe 54 131.29 Xenon
Cs 55 132.91 Cesium	Ba 56 137.33 Barium	La 57 138.91 Lanthanum	Hf 72 178.49 Hafnium	Ta 73 180.95 Tantalum	W 74 183.85 Tungsten	Re 75 186.21 Rhenium	Os 76 190.2 Osmium	Ir 77 192.22 Iridium	Pt 78 195.08 Platinum	Au 79 196.97 Gold	Hg 80 200.59 Mercury	Tl 81 204.38 Thallium	Pb 82 207.2 Lead	Bi 83 208.98 Bismuth	Po 84 (209) Polonium	At 85 (210) Astatine	Rn 86 (222) Radon
Fr 87 223.02 Francium	Ra 88 226.03 Radium	Ac 89 227.03 Actinium	Rf 104 (261) Rutherfordium	Db 105 (262) Dubnium	Sg 106 (263) Seaborgium	Bh 107 (262) Bohrium	Hs 108 (265) Hassium	Mt 109 (266) Meitnerium	Unnamed Discovery 110 Nov. 1994	Unnamed Discovery 111 Nov. 1994	Unnamed Discovery 112 1996		Unnamed Discovery 114 1999		Unnamed Discovery 116 1999		Unnamed Discovery 118 1999
ALKALI METALS	ALKALI EARTH METALS													HALOGENS		NOBLE GASES	

( ) = ESTIMATES

LANTHANIDES	Ce 58 140.12 Cerium	Pr 59 140.91 Praseodymium	Nd 60 144.24 Neodymium	Pm 61 (145) Promethium	Sm 62 150.36 Samarium	Eu 63 152.97 Europium	Gd 64 157.25 Gadolinium	Tb 65 158.93 Terbium	Dy 66 162.50 Dysprosium	Ho 67 164.93 Holmium	Er 68 167.26 Erbium	Tm 69 168.93 Thulium	Yb 70 173.04 Ytterbium	Lu 71 174.97 Lutetium
ACTINIDES	Th 90 232.04 Thorium	Pa 91 231.04 Protactinium	U 92 238.03 Uranium	Np 93 237.05 Neptunium	Pu 94 (240) Plutonium	Am 95 243.06 Americium	Cm 96 (247) Curium	Bk 97 (248) Berkelium	Cf 98 (251) Californium	Es 99 252.08 Einsteinium	Fm 100 257.10 Fermium	Md 101 (257) Mendelevium	No 102 259.10 Nobelium	Lr 103 262.11 Lawrencium

Table 1.5 The modern periodic table of elements.

Activity 1.4: Row Number and Column Number of learners

Consider the following seating arrangement of learners in a class room. The row numbers are represented by R_1, R_2, R_3, R_4, R_5 and R_6, and the column numbers are represented by C_1, C_2, C_3 and C_4, as follows:

Teacher

	C_1	C_2	C_3	C_4
R_1	A	B	C	D
R_2	E	F	G	H
R_3	I	J	K	L
R_4	M	N	O	P
R_5	R	S	T	U
R_6	W	X	Y	Z

1 What is the row number and column number for learner N?
Which learners have the same row number with as N?
Which learners have the same column number with as N?
2. Which learner is in C_1R_1? Which learner is in C_4R_4?
List the learner in C_3 and R_5?

A. Period

Each horizontal row in the periodic table is called a **period.** There are seven periods and the period number usually represents number of electron shells which a given element possesses. This means that elements that are found in the same period have the same number of shells, but they have a different number of valence electrons. Table 1.6 shows the number of periods available, the number of elements accommodated in each period along with the range of atomic numbers and type of shells.

Period number	Number of elements	Range of atomic numbers	Main shells
1	2	1 - 2	K
2	8	3 - 10	K, L
3	8	11 - 18	K, L, M
4	18	19 - 36	K, L, M, N
5	18	37 - 54	K, L, M, N, O
6	32	55 - 86	K, L, M, N, O, P
7	Incomplete	87 - continues	K, L, M, N, O, P, Q

Table 1.6 Number of periods and number of elements and shells they accommodate.

Example

The atomic number (Z) of calcium (Ca) is 20. It falls into (period 4 of) the periodic table and has K, L, M, N energy shells.

B. Group

The vertical column in the periodic table is called a **group**. Groups are classified into two. These are A and B block elements. There are eight A block elements numbered from I to VIII, which form the main block of the periodic table. There are also eight B block elements. These are known as transition elements. Elements in a group have similar chemical properties and their behaviour depends mainly on the number of electrons in their outer shell, their valence electrons. They contain the same number of valence electrons. Due to similarities, a groups of elements is sometimes referred to as a **family**.

Group number	I	II	III	IV	V	VI	VII	VIII
Number of valence electrons	1	2	3	4	5	6	7	8

Table 1.7 Relationship between group number and number of valence electrons.

Some of the groups have acquired common names. For example, except for H, the Group IA elements are called the **alkali metals** or **simply the alkalis**. Group IIA elements are known as the **alkaline earth metals**. The group 0 (VIIIA) are the **noble gases** (formerly called **inert gases**). These noble gases are inert and do not react under normal conditions. Finally, the elements in Group VII are called the **halogens** derived from the Greek word "halous" meaning "sea or salt producer".

Some groups are named by their common names (alkali metals, alkaline earth metals and halogens), while others are named after the first element of the group, for example, boron is the first element in group IIIA and the group is named as boron family. The group numbers and their names are given in Table 1.8.

Group number	Name of the group
IA	Alkali metals
IIA	Alkaline earth metals
IIIA	Boron family
IVA	Carbon family
VA	Nitrogen family
VIA	Oxygen family
VIIA	Halogen
VIIIA	Noble gases

Table 1.8 The group numbers and names of the groups

Example

Referring to Table 1.5 and Table 1.7, classify the following elements into groups of alkali metals, alkaline earth metals, halogens and noble gases:
Ne , Mg , Cl , Na , Be, Ar , F , Li

Answer

Na and Li are alkali metals because each has one valence electron. Mg and Be are alkaline earth metals because each has two valence electrons. F is a halogen because it has seven valence electrons. Ne is a noble gas because it has eight valence electrons.

Activity 1.5: Comparing the Properties of Elements

Cut a chart into card-size pieces, on one side of the card write the name of an element and on the reverse side of the card write the electron configuration of that element.

1. Write down the number of valence electrons of the first twenty elements.
2. Do you see any similarity in the number of valence electrons? Which elements have the same number of valence electrons?
3. How does the number of valence electrons vary in a period from left to right? What happens to number of valence electrons moving down a group?

Exercise 1.3

Referring to Table 1.5 and Table 1.7, classify the following elements as an alkali metal, as alkaline earth metal, a halogen or a noble gas: Ar, Ca, Cl, K,

Summary

Henry Moseley showed that the atomic number of an element is a more fundamental property than its atomic mass. In the modern periodic table elements are arranged according to their increasing atomic number.

In the modern periodic table elements are arranged in periods and groups. Each horizontal row in the periodic table is called a period. There are seven periods and the period number usually represents number of electron shells that which a given element possesses. The vertical column in the periodic table is called a group. Groups are classified into two. These are A and B block elements. There are eight A block elements numbered from I to VIII, which form the main block of the periodic table. There are also eight B block elements. Elements in a group have similar chemical properties because they contain the same number of valence electrons.

Group IA elements are called the alkali metals or simply alkalis. Group IIA elements are known as the alkaline earth metals. The Group 0 (VIIIA) are the noble gases and the elements of group VIIA are called the halogens.

1.3 Electron Configuration and the Periodic Table

Let us discuss

1. How are the electron configuration and the position of elements related?
2. A property of an element is greatly influenced by its valence electrons. Explain why?

1.3.1 Electron Configuration

In grade nine, you learnt that electrons are arranged around the nucleus in electron shells (also called energy levels). The shells are denoted as K (1^{st} shell), L (2^{nd} shell), M (3^{rd} shell l), N (4^{th} shell), etc. The maximum number of electrons that each shell can hold is given by the formula $2n^2$, where n is the shell number, as illustrated in the following example.

Example

Calculate the maximum number of electrons that the N energy level can hold.

Answer

The n value for the N energy level is 4; therefore, the maximum number of electrons in the N energy level is, $2n^2 = 2 \times 4^2 = 32$.

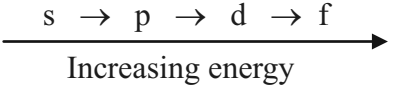
The electron shells consist of one or more subshells that are denoted by the letters s, p, d, and f. The number of subshells contained in each shell is equal to its n value. For example, there is one subshell in n = 1, the s subshell, and because it makes up the first shell, it is denoted as 1s.There are two subshells in the n = 2 shell, the s and p. Since they are part of the n = 2 shell, they are designated 2s and 2p.The s, p, d and f subshells contain a maximum number of 2, 6, 10 and 14 number of electrons respectively. The electron shells, subshells, letter designation and maximum number of electrons are given in Table 1.9.

Electron shell (n)	sub-shells	Letter designation	Maximum number of electrons
1	s	K (1s)	2 (2)
2	s p	L (2s) (2p)	2 (8) 6
3	s p d	M (3s) (3p) (3d)	2 6 (18) 10
4	s p d f	N (4s) (4p) (4d) (4f)	2 6 (32) 10 14

Table 1.9 The relationships beteen the electron shells, subshells, letter designation and the maximum number of electrons.

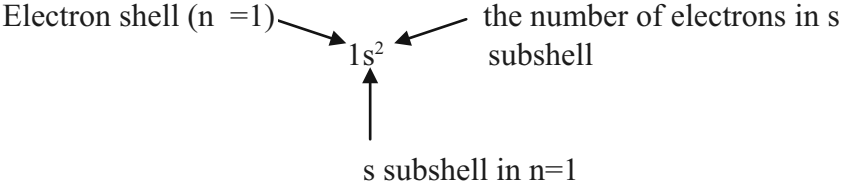
Arranging Electrons in Subshells

In order to arrange the electrons in the atoms or write their electron configuration, one must know the order in which the different energy levels are filled. It is important to remember that the electrons will fill the lowest energy levels first before they start to fill the next, higher energy level. The order of energy of the subshell within a specific electron shell is given as follows:



This means that the s subshell of any of the shells must be filled before adding electrons to the p sub shell, and the p sub shell must be filled before electrons are added to the d sub-shell and so on.

The number and location of the electrons are specified by the following symbols.



The arrangement of electrons in the sub-shells starts by filling the lowest sub shell (1s) then filling the next, higher subshell (2s) and so on. The sequence for filling the subshells starting at 1s is given in Fig 1.1.

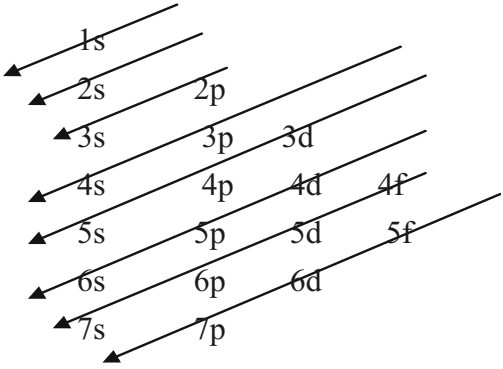


Figure 1.1 Order of filling atomic sub-shells

The following examples illustrate the filling of electrons the subshells.

Example

How do you fill the electron/electrons in the subshells for the following elements? H (Z=1), Li (Z= 3), C (Z= 6), and Fe (Z=26)

Answer

H has 1 electron. The lowest subshell is 1s (n = 1).The electron configuration for H is written as: H: $1s^1$

Similarly, Li has 3 electrons. The lowest energy level 1s is filled. Once the 1s subshell is filled, you move up to the next, higher subshell 2s (n = 2).The electron configuration for Li is written as: Li: $1s^2 2s^1$

C has 6 electrons. The lowest subshell 1s (n =1) is filled. The next higher subshell 2s (n =2) is also filled so you move up to the subshell 2p (n = 2). The electron configuration for C is written as: C: $1s^2 2s^2 2p^2$

Fe has 26 electrons. The lowest subshell 1s (n =1) is filled. The subshells 2s and 2p (n =2) are filled. Once these subshells are

filled you move to subshells 3s, 3p, 3d, (n =3).Only 3s and 3p subshells are filled but the 3d subshell (n = 3) is not filled. Since the 4s subshell comes before the 3d (n = 3), the 4s sub-shell is also filled. The electron configuration of Fe is written as: Fe: $1s^2\ 2s^2\ 2p^6 3s^2 3p^6 4s^2 3d^6$

The electron configuration of the first twenty elements is given in Table 1.10 below.

Element	Symbol	Atomic number(Z)	Electron configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2\ 2s^1$
Beryllium	Be	4	$1s^2\ 2s^2$
Boron	B	5	$1s^2\ 2s^2 2p^1$
Carbon	C	6	$1s^2\ 2s^2\ 2p^2$
Nitrogen	N	7	$1s^2\ 2s^2 2p^3$
Oxygen	O	8	$1s^2\ 2s^2\ 2p^4$
Fluorine	F	9	$1s^2\ 2s^2 2p^5$
Neon	Ne	10	$1s^2\ 2s^2 2p^6$
Sodium	Na	11	$1s^2\ 2s^2\ 2p^6 3s^1$
Magnesium	Mg	12	$1s^2\ 2s^2\ 2p^6 3s^2$
Aluminum	Al	13	$1s^2\ 2s^2\ 2p^6 3s^2 3p^1$
Silicon	Si	14	$1s^2\ 2s^2\ 2p^6 3s^2 3p^2$
Phosphorus	P	15	$1s^2\ 2s^2\ 2p^6 3s^2 3p^3$
Sulfur	S	16	$1s^2\ 2s^2\ 2p^6 3s^2 3p^4$
Chlorine	Cl	17	$1s^2\ 2s^2\ 2p^6 3s^2 3p^5$
Argon	Ar	18	$1s^2\ 2s^2\ 2p^6 3s^2 3p^6$
Potassium	K	19	$1S^2\ 2S^2\ 2P^6 3S^2 3P^6 4S^1$
Calcium	Ca	20	$1S^2\ 2S^2\ 2P^6 3S^2 3P^6 4S^2$

Table 1.10 Electron configurations of the first twenty elements

1.3.2 Abbreviating of Electron Configurations

The electron configuration of elements can be abbreviated by using the symbols of the noble gases that have the nearest lower atomic number. For example, the noble gas that has the nearest lower atomic number to Li is He which has two electrons ($1s^2$). The electron structure of Li has three electrons. It consists of two electrons as in He ($1s^2$), plus one 2s electron ($2s^1$).Therefore, the electron configuration of Li can be abbreviated as follows: [He]$2s^1$

Similarly, the noble gas that has the nearest lower atomic number to Na is Ne which has ten electrons ($1s^2\ 2s^2 2p^6$). The electron structure of Na consists of eleven electrons as in Ne ($1s^2\ 2s^2 2p^6$), plus a $3s^1$ electron. Therefore, the electron configuration of Na can be abbreviated as follows: [Ne]$3s^1$

Exercise 1.4

Using the symbols of noble gases, write the electron configuration of the following elements: N (Z=7), Al (Z=13) and K (Z=19)

1.3.3 Classification of Elements in the Modern Periodic Table Based on Electron Configuration

In modern periodic table, all elements are arranged in increasing order of their atomic number that is on the bases of their electron configurations. As a result, the elements can be categorized into three types as shown in Table 1.11.

A. Representative Elements

These elements belong to Groups I to VIII in the modern periodic table and are denoted by the letter A. Generally, they have valence electrons ranging from one to eight. Representative elements are called s or p block elements based on their electron configuration. For example, the electron configuration of Li is $1s^2 2s^1$. Li is an s-block element because its outermost subshell is an s subshell. Similarly, the electron configuration of C is$1s^2 2s^2 2p^2$, C is a p block element because its outermost subshell is a p subshell.

B. Transition Elements

Transition elements are found in the center of the modern periodic table between IIA and IIIA and are denoted by the letter B. Transition elements are also called d block elements based on their electron configuration. For example, the electron configuration of Fe is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. Fe is a d block element because its outermost subshell is a d subshell.

C. Inner Transition Elements

The elements in the two long rows below the main body of the table are called the **inner transition elements**, and each row is named after the elements that it follows in the main body of the table. For example, elements with the atomic number 58-71 are called the Lanthanide series elements because they follow Lanthanum (La, Z = 57) and elements with the atomic number 90- 109 are called the actinide series of elements because they follow actinium (Ac, Z= 89). Inner transition elements consist of f-block elements.

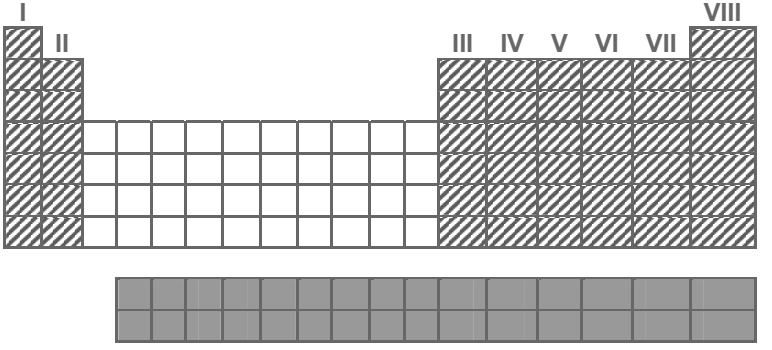


Table 1.11 Classification of elements in the modern periodic table

Key

- Representative elements
- Transition elements
- Inner transition elements

Exercise 1.5

1. Refer to Table 1.5 and classify the following elements as representative elements, transition elements or as inner transition elements.
   Al, Ni, Be, Kr, Ac, Cl, Cu, K, Ne, Sr, Fr, Pa, Fe, Xe, Zn, U.
2. Based on electron configuration, classify the following elements as representative and transition elements.
   Li (Z=3), F (Z=9), Sc (Z=21), P (Z=15), Ar (Z=18), V (Z=23)

Summary

Electron configuration is the arrangement of electrons in energy levels or shells. The main shells or energy levels are designated as K (1^{st} shell), L (2^{nd} shell), M (3^{rd} shell), N (4^{th} shell),etc. The main shells consist of one or more subshells(s, p, d and f).

The order of energy of the subshells within a specific electron shell is given as:

$$s \rightarrow p \rightarrow d \rightarrow f$$

Increasing energy

The arrangement of the electrons in subshells starts by filling the lowest subshell (1s) then filling the next, higher sub shell (2s) and so on.

The electron configuration of the elements can be abbreviated by using the symbols of the noble gases:

In the modern periodic table, the elements can be categorized into three types: representative (s and p block elements), transition (d block elements) and inner transition elements (f block elements). The elements in Groups IA to VIIIA are called representative elements. Transitions elements are found in the center of the periodic table and are denoted by a letter B. The two long rows which are located below the main body of the table are called inner transition elements.

1.4 Periodic Regularities and Characteristics of Elements

The modern periodic table is widely used to predict the physical and the chemical properties and the periodic regularities of the elements. This can be achieved by studying the trends of the elements moving down a group and moving across a period. In this sub-unit, we discuss some of the periodic regularities.

1.4.1 Trends in Atomic and Ionic Size (Atomic and Ionic Radii)

Let us discuss

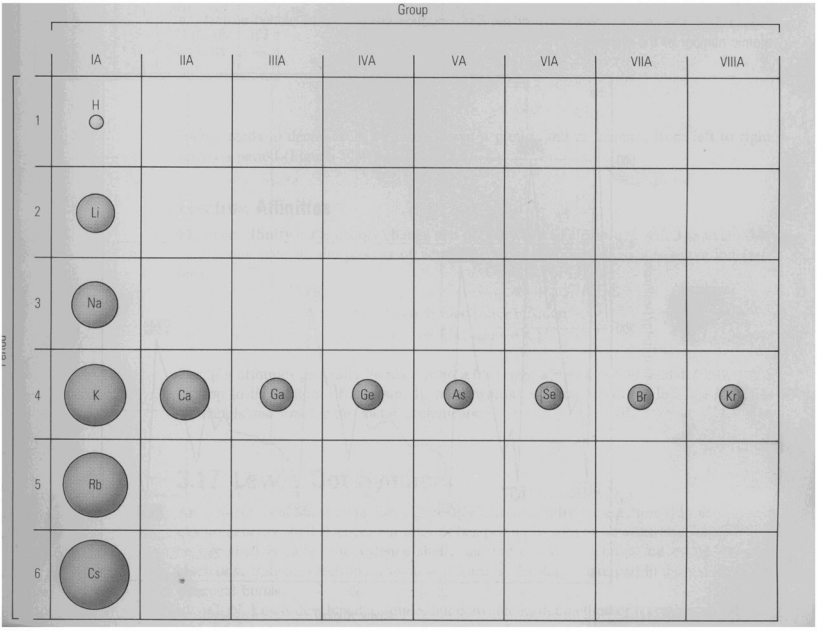
1. What do you observe when you blow up a balloon?
2. What do you observe when the gas is released from a balloon?
3. Suppose you are standing at a fixed point in front of the learners in a class. What can you say about the distance between you and the learners sitting in the first row compared to the middle and last rows?

An atom has a central nucleus with electrons zooming around the outside the nucleus. It is reasonable to assume that the size of an atom depends on how far its outermost valence electrons are from the nucleus. If they are very close to the nucleus, the atom will be very small. If they are far away, the atom will be quite a bit larger. So the atomic size is determined by how much space the electrons take up.

The atomic radius of an atom is the distance from the centre of the nucleus to the outermost shell. The atomic radius can be calculated by measuring the distances between the nuclei of atoms in the chemical bonds of molecular substances. The **ionic radius** is defined as the distance from the nucleus of an ion to its outermost shell.

Activity 1.6: Atomic Size of Elements

Consider the following picture.



What happens to the size of the atoms as we move down a group and move across a period?

Trends in Atomic Radius

A. Down a Group

Generally, in the Periodic Table, the atomic radius of elements increases moving down a group. As you move down a group, the number of shells and the nuclear charge increases. However, the increase in the number of shells overcomes the increase in the nuclear charge. Consequently, the electrons in the valence shells are further and further away from the nucleus because they are in higher shells as we move down a group.

B. Across a Period

An effective nuclear charge is the positive charge that an electron experiences from the nucleus of an atom and it is equal to the

nuclear charge but reduced by any shielding or screening from any intervening electron distribution.

Across a period, the atomic radius tends to decrease with increasing atomic number (nuclear charge). Across a period, the number of shells and the number of core electrons remain constant. Only the number of valence electrons increases. Because there is no effective shielding between valence electrons, there is an increase in the effective nuclear charge as we move across a period, which results in reducing the atomic size by pulling the outer electrons inward.

Example

Arrange the following elements in order of increasing atomic radius.

	C
Al	Si

Answer

Note that C is above Si in Group IVA. Therefore, the radius of C is smaller than that of Si (the atomic radius increases moving down a group of elements). Note that Al and Si are in the same period. Si comes after Al, so the radius of Si is smaller than that of Al (atomic radius decreases within a period). Therefore, the radius of these in elements in order of increasing atomic radius is: C, Si, Al.

Activity 1.7: Atomic Sizes of Period 2 Elements

The atomic radii in a picometer (pm, scale: 1mm = 10^9pm) of the elements of period 2 are given below:

Period 2 elements: B Be O N Li C
Atomic radius (pm): 88 111 66 74 152 77

On a chart draw a circle for each of these elements with their respective atomic radius. Arrange them in increasing order of their atomic radii.

1. Which element has the largest atomic size? Which element has the smallest atomic size?
2. How does the atomic radius change as you move from left to right in a period?

Exercise 1.6

Arrange the following elements in order of their increasing atomic radius.

	Be
Na	Mg

Ionic Radius

An ion is a charged atom. Ions are of two types: cations and anions. Cations are positively charged ions and are formed by losing one or more electrons from the outermost shell of a neutral atom. In cations, the number of protons (nuclear charge) is greater than the number of electrons. Therefore, the electrons are pulled towards the nucleus. Cations are smaller than the parent atom.

Conversely, an anion is a negatively charged ion formed by gaining electrons. This means that there are more electrons than protons. The size of anion increases because the effective nuclear charge decreases. The size of a selection of atoms and ions is given in Table 1.12.

Atomic radius and ionic radius (in Angstroms)			
Atomic radius	**Ionic radius**	**Atomic radius**	**Ionic radius**
Na 1.86	Na^+ 0.95	O 0.74	O^{2-} 1.40
K 2.27	K^+ 1.33	F 0.71	F^- 1.36
Mg 1.60	Mg^{2+} 0.65	S 1.03	S^{2-} 1.84

Table 1.12 Comparison of the atomic and ionic radii of certain elements

For ions derived from elements in different groups, the comparison in size is meaningful only if the ions are **isoelectronic** (ions with the same number of electrons). Generally, for isoelectronic ions, we find that the cations are smaller than the anions. For example, Mg^{2+} is smaller than F^-. Both ions have the same number of electrons,but Mg (Z = 12) has more protons than F (Z = 9).The large nuclear

charge of Mg^{2+} results in a smaller radius. The anion with a greater negative charge is larger than that with a smaller negative charge in the isoelectronic species.

Example

In each of the following pairs indicate which one of the two species is larger:

a) N^{3-} or F^{-} b) Mg^{2+} or Ca^{2+} c) Fe^{2+} or Fe^{3+}

Answer

a) N^{3-} and F^{-} are isoelectronic anions. Since N^{3-} has a lower nuclear charge than F^{-}, N^{3-} is larger.
b) Both Mg and Ca belong to Group IIA (the alkaline metals). The Ca^{2+} ion is larger than the Mg^{2+} ion because Ca lies below Mg.
c) Both ions have the same nuclear charge, but Fe^{2+} has more number of electrons than Fe^{3+}. Therefore, the radius of Fe^{2+} is larger.

Exercise 1.7

For each of the following pairs indicate which one of the two species is larger:

a) Cl or Cl^{-} b) Mg^{2+} or Al^{3+} c) O^{2-} or S^{2-},

1.4.2 Trends in Ionization Energy

Ionization energy is the minimum energy needed to remove the outermost electron from a neutral atom in the gaseous state. This can be represented as follows:

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

The ionization energy of an atom is strongly influenced by the atomic radius and the nuclear charge. When atomic size increases, the distance of the loosely held electrons from the nucleus increases. Therefore, since the force of attraction from the nucleus is less, the ionization energy decreases. When the nuclear charge increases, the valence electrons are attracted more strongly towards the nucleus. This means that a great amount of energy is required to remove the valence electrons, resulting in an increase in the ionization energy,

A. Down a Group

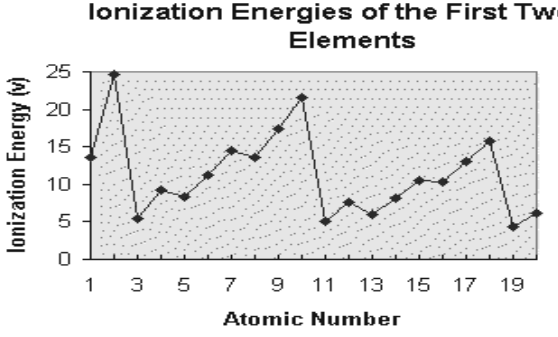
Moving down a group, the effect of an increased nuclear charge is balanced by the effect of increased shielding, and the number of shells becomes the predominant factor. With more shells, the outermost electrons (the valence electrons) are further from the nucleus and are not as strongly attracted to the nucleus. The result is that the ionization energy of elements decreases as we move down a group in the periodic table. This is because the atomic size increases and it is easier to remove the electrons.

B. Across a Period

Moving across the periodic table from left to right, the electrons are held more strongly to the nucleus, causing the atoms to be smaller and the ionization energy to be higher. Thus, within any period, the ionization energy tends to increase with atomic number because nuclear charge increases. The lowest values in a period are found in the alkali metals.

Activity 1.8: Ionization Energy Values of the First Twenty Elements

Consider the following figure that shows the ionization energy of the first 20 elements. Refer to the modern periodic table.



1. Which atomic numbers of the elements represent the highest peaks? Name these elements?
2. Which atomic numbers of the elements represent lowest peaks? Name these elements?
3. What do you understand regarding this pattern?

Example

Arrange the following elements in order of increasing ionization energy.

S		Ar
Se		

Answer

S and Ar are in the same period and Ar comes after S. Ar has a higher ionization energy than S because ionization energy increases moving across a period from left to right. But Se and S are in the same group. Se is found below S. Se has a lower ionization energy than S because ionization energy decreases moving down a group. Therefore, the order of these elements according to their increasing ionization energy is Se, S, Ar.

Exercise 1.8

1. Arrange the following elements in order of increasing ionization energy.

N			Ne
P			

1.4.3 Trends in Electronegativity

Let us discuss

Consider the following: two teachers (A and B) teach different subjects to the same class. Teacher A is enthusiastic, energetic, participative, and incorporates a variety of methods when he teaches. Teacher B is unenthusiastic, not energetic, does not participate, and follows a single method of teaching.

To whom will the learners be more attracted ? Explain your answer.

Electronegativity is a measure of the tendency of an atom to attract electrons in a covalent bond. It indicates the ability of an atom in a molecule to attract the shared electrons towards it. Generally, electronegativity depends on the number of valence electrons (i.e in the outer shell), the size of the atom and the nuclear charge. Atoms with nearly filled outer-most shells (valence shells) have higher elecetronegativities than those atoms with one or two electrons

in their outer-most shells. Electronegativity values of common elements on Pauling's scale are given in Table 1.13.

IA							VIIIA
H							He
2.1	IIA	IIIA	IVA	VA	VIA	VIIA	---
Li	Be	B	C	N	O	F	Ne
1.0	1.5	2.0	2.5	3.0	3.5	4.0	---
Na	Mg	Al	Si	P	S	Cl	Ar
0.9	1.2	1.5	1.8	2.1	2.5	3.0	---
K	Ca	Ga	Ge	As	Se	Br	Kr
0.8	1.0	1.6	1.8	2.0	2.4	2.8	---
Rb	Sr	In	Sn	Sb	Te	I	Xe
0.8	1.0	1.7	1.8	1.9	2.1	2.5	---
Cs	Ba	Ti	Pb	Bi	Po	At	Rn
0.7	0.9	1.8	1.8	1.9	2.1	2.2	---
Fr	Ra						
0.7	0.9						

Table 1.13 Electronegativity values of representative elements

Fluorine is the most electronegative element (4.0), the second is oxygen (3.5), and third is nitrogen (3.0) Cesium and Francium the least electronegative elements (0.7).

A. Down a Group

Generally, electronegativity decreases from top to bottom moving down a group. This is because an atom that has small ionization energy will give away an electron more easily than an atom with greater ionization energy. An atom with small electronegativity will lose its share of an electron pair more easily than an atom with greater electronegativity.

B. Across a Period

Electronegativity increases from left to right moving across a period. This is because the nuclear charge increases (decreasing atomic size) across a period.

Elements in the same region on the periodic table (for example, the nonmetals) have similar electronegativities. But nonmetals have greater electronegativity than metals.

Example

Arrange the following elements in order of increasing electronegativity.

			F
Na			Cl

Answer

Na has lower electronegativity than Cl. This is because electronegativity increases moving across a period from left to right. F has greater electronegativity than Cl because electronegativity decreases moving down a group.
Therefore, the electronegativity in increasing order is: Na, Cl, F.

Exercise 1.9

Arrange the following elements in order of increasing electronegativity.

K			Br
Cs			

1.4.4 Trends in Metallic and Nonmetallic Characteristics

The periodic table organizes all sorts of chemical and physical information about the elements and their compounds.

There are three main classifications of the elements- the metals, the nonmetals and the metalloids.

From a chemical point of view, the metals are defined as those elements that possess 1, 2 or 3 valence electrons. Metals generally have a tendency to lose electrons and form cations. Most elements in the periodic table are metals. Metals are found on the left-hand side of the periodic table.

The nonmetals occupy the upper right-hand side of the periodic Table. From a chemical point of view, nonmetals are defined as those elements that possess 5, 6 or 7 valence electrons. Nonmetals generally have a tendency to gain electrons and form anions.

The metalloids are elements that have properties of both metals and nonmetals and they are located between the metals and the nonmetals in the periodic table.

A. Down a Group

Moving down a group, the effective nuclear charge experienced by the valence electrons decreases because the outermost electrons are further away from the nucleus. Therefore, they can be easily lost.

The metallic characteristics of an element increase moving down a group. However, nonmetallic character decreases down a group. For example, Group IVA shows the most distinct change in metallic characteristics. It begins with the nonmetal carbon (C), followed by the metalloids silicon (Si) and germanium (Ge) and then the metals tin (Sn) and lead (Pb).

B. Across a Period

As the effective nuclear charge acting on the valence electrons increases as we move across a period, the tendency to lose electrons will decrease. The metallic characteristics of an element decrease moving across a period and the nonmetallic characteristics increase.

For example, Period 3 shows the change from metallic characteristics to nonmetallic characteristics. It begins with the metals sodium (Na), magnesium (Mg) and aluminum (Al) followed by the metalloid silicon (Si), then moves to the nonmetals phosphorus (P), sulphur (S) and chlorine (Cl).

Exercise 1.10

1. On which side of the periodic Table are the metals found?
2. On which side of the periodic Table are the nonmetal and the metalloids found?

3. Classify the following elements as metals, nonmetals or metalloids?

   Li, O, B, I , Ca and Ge

1.4.5 Patterns of Chemical Formulas of Compounds

The bonding capacity of an atom is called its **valency**. Groups I, II and III have valencies of 1, 2 and 3, respectively. They tend to lose electrons. Group IV has a valency of 4 and its elements have the ability to share electrons. Groups V, VI and VII have valencies of 3, 2, and 1,respectively. They have a tendency to gain or share electrons.

The valency of an element helps with writing the formulas of their compounds. For example:

Na^+ has a valency of 1, with Cl^- it forms NaCl which contains one Cl^-. Similarly, Al^{3+} has a valency of 3, with Cl^- it forms $AlCl_3$ which contains three Cl^-. Refer to the Table 1.14.

Periods	Groups						
	I	II	III	IV	V	VI	VII
2	LiCl	$BeCl_2$	$BeCl_3$	CCl_4	NCl_3	Cl_2O	ClF
3	NaCl	$MgCl_2$	$AlCl_3$	$SiCl_4$	PCl_3	SCl_2	ClCl
4	KCl	$CaCl_2$	$GaCl_3$	$GeCl_4$	$AsCl_3$	$SeCl_2$	BrCl

Table 1.14 Chlorides of the elements in periods 2, 3 and 4

Exercise 1.11

Which of the following elements have a tendency to lose, share or gain electrons? K, Al, C, S, Br, Si

1.4.6 Summary of Trends in Periodic Properties

Though there are some exceptions: a summary of the trends in periodic properties is shown in Table 1.15.

S.No.	Property	Down a group (top to bottom)	Across a period (left to right)
1	Atomic radius	Increases	Decreases
2	Ionization energy	Decreases	Increases
3	Electron negativity	Decreases	Increases
4	Metallic character	Increases	Decreases
5	Nonmetallic character	Decreases	Increases

Table 1.15 Summary of the trends in periodic regularities

Summary

Elements in the modern Periodic Table show different trends moving down a group or moving across a period. These trends are the atomic radius, the ionic radius, ionization energy, electronegativity and the metallic characteristics. The indicate the chemical and physical properties of an element

The atomic radius is defined as one-half of the distance between the nuclei of two covalently bonded atoms of the same element in a molecule.

The ionic radius is defined as the distance from the nucleus of an ion to the outermost shell.

An effective nuclear charge is the positive charge that an electron experiences from the nucleus of an atom, equal to the nuclear charge but reduced by any shielding or screening from any intervening electron distribution.

Ionization energy is the energy needed to remove an electron from a neutral atom in the gaseous state.

Electronegativity is a measure of the tendency of an atom to attract electrons in a covalent bond.

Most elements in the periodic table are metals. They are found on the left-hand side of the table. Metals tend to lose electrons while forming bonds, that is, they are electropositive in nature. The non_metals are found on the right-hand side of the periodic table towards the top. Nonmetals, are electronegative. They tend to form bonds by gaining electrons.

The metalloids are elements that have properties of both the metals metals and the nonmetals and they are located between the metals and the nonmetals in the periodic table.

Moving down a group from top to bottom, the atomic radius, the ionic radius and the metallic characteristics increase but the ionization energy, electronegativity and nonmetallic characteristics decrease.

Moving, across a period from left to right, the ionization energy, electronegativity and nonmetallic characteristics increase but the atomic radius and the metallic characteristics decrease.

Key Words

Alkali metals
Alkaline earth metals
Atomic radius
Effective nuclear charge
Electron configuration
Electronegativity
Electron shell
Families of elements
Groups
Halogens
Inner transition elements
Ionic radius
Ionization energy
Isoelectronic
Law of octaves
Metalloids
Noble gases
Octaves
Periods
Periodic law
Periodic table
Representative elements
Subshells
Triads
Transition elements
Valency

Review Questions

Part I. Choose the Correct Answer

1. Which of the following statements **<u>is not</u>** correct in terms of the trends when moving from left to right across the periods in the a periodic table?
   a. The elements become less metallic in nature.
   b. The number of valence electrons increases.
   c. The atoms lose their electrons more easily.
   d. The number of valence electrons decreases.

2. Element X forms a chloride with the formula XCl_2, which is a solid with a high melting point. X would most likely be in the same group of the periodic table as:
   a. Na b. Mg c. Al d. Si
3. All of the following are properties of metals **except:**
   a. malleable b. ductile c. brittle d. lustrous
4. What is a common property of the alkali metals?
   a. they have equal numbers of shells.
   b. they have one "s" electron in their valence shells.
   c. they have two "s" electrons in their valence shells.
   d. they are all located in Period 3.
5. Electronegativity is a measure of the tendency of an atom to:
   a. repel electrons
   b. lose electrons to another atom
   c. attract shared electrons
   d. combine with protons
6. Which of the following elements has the highest electronagitivity?
   a. Mg b. K c. B d. F
7. Which of the following properties of the atom of element X and its ion X^{+3} must be the same?
   a. size b. nuclear charge
   c. ionization energy. d. number of electrons
8. Which of the following statements **is not** true?
   a. metallic characteristics increase moving down a group
   b. ionization energy increases from left to right within a period
   c. Electronegativity increases from left to right within a period and decreases from top to bottom moving down a group.
   d. atomic size increases from left to right within a period.
9. Which of the following elements (Al, Si, P, N) has the least metallic character?
   a. Al b. Si c. P d. N
10. Which of the following elements (F, K, S, Ca) has the highest ionization energy?
    a. F b. Ca c. S d. K
11. The correct order for the following ionic radii is:
    a. $N^{3-} > F^-$ b. $Na^+ < Mg^{2+}$ c. $Fe^{2+} < Fe^{3+}$ d. $Mg^{2+} < Al^{3+}$

12. The electron configuration of the element found in period 4 and group 2A is:
    a. $[Ar]4s^1$ b. $[Ar]3d^2$ c. $[Ne]3s^1$ d. $[Ar]4s^2$
13. Which of the following is **not true** regarding the noble gases?
    a. they have complete "s or p" sub shells.
    b. they are inert gases
    c. they do not react under normal conditions
    d. they have incomplete "d" sub shell
14. One of the following pairs is different. Which pair is it?
    a. B and Si b. Na and Li c.O and Cl d. K and F

Part II. Fill in the blank spaces.

1. _____ tried to arrange elements into groups of three.
2._____ states that the properties of elements are the periodic function of their atomic mass.
3. Each horizontal row in the periodic table is called a _____.
4. _____ tend to form bonds by gaining electrons.
5. _____ are the elements that have properties both the of metals and the nonmetals.
6. The electrons in the outer most shell of an atom are called____ .
7._____ is the radius of a cation or anion.
8.Electronnegativity ____ moving across a period and_____ moving down a group.
9. Elements that are denoted by A in the periodic table are called_____.

Part III. Short Answer/ Work out Questions.

1. How did the modern Periodic Table overcome the various anomalies in Mendeléev's periodic table?
2. Name two elements that exhibit similar show chemical reactions as magnesium. What is the basis for your choice?
3. Name:
   a) three elements that have a single electron in their outermost shell.
   b) two elements that have two electrons in their outermost shell.
   c) three elements with a filled outermost shell.

4. a) Lithium, sodium, and potassium are all metals that react with water to liberate hydrogen gas. What are the similarities between the atoms of these elements?
   b) Helium is an unreactive gas and neon is a gas with extremely low reactivity. What, if anything, do these atoms have in common?
5. In the modern Periodic Table, name the metals among the first ten elements,
6. According to their position in the Periodic Table, which one of the following element would have the most metallic characteristic?
   Ga Ge As Se Be
7. Which elements have the following properties?
   a) two shells, both of which are completely filled with electrons?
   b) the electronic configuration 2, 8, 2?
   c) a total of three shells, with four electrons in its valence shell?
   d) a total of two shells, with three electrons in its valence shell?
   e) twice as many electrons in its second shell as in its first shell?
8. a) What property do elements in the same column as born have in common?
   b) What property do all elements in the same column as periodic table as fluorine have in common?
9. An atom has electronic configuration 2, 8, 7.
   a) What is the atomic number of this element?
   b) To which of the following elements would it be chemically similar? (Atomic numbers are given in parentheses.)
   N(7), F(9), P(15), Ar(18)
10. The positions in the Periodic Table of three elements A, B, and C, are shown below:

Group VI	Group VII
-	-
-	A
-	-
B	C

   (a) State whether A is a metal or a nonmetal.
   (b) State whether C is more reactive or less reactive than A.
   (c) Will C be larger or smaller in size than B?
   (d) Which type of ion, cation or anion, will be formed by element A?

11. Nitrogen (atomic number 7) and phosphorus (atomic number 15) belong to Group VA. Write the electronic configuration of these two elements. Which of these will be more electronegative? Explain your answer.
12. How does the electronic configuration of an atom relate to its position in the modern periodic table?
13. In the modern periodic table, calcium (atomic number 20) is surrounded by elements with atomic numbers 12, 19, 21 and 38. Which of these have physical and chemical properties resembling calcium?
14. Compare and contrast the arrangement of elements in Mendeléev's Periodic Table to or with its modern version.
15. The following Table shows the ionization energies for the elements of periods 1 and 2.

Period	Element	Ionization energy ($kJ.mol^{-1}$)
1	H	1312
	He	2372
2	Li	520
	Be	899
	B	801
	C	1086
	N	1402
	O	1314
	F	1681
	Ne	2081

a) What is the meaning of the term "ionization energy"?
b) Identify the pattern of ionization energies in a period.
c) Which two elements exert the strongest forces of attraction on their electrons? Use the data in the table to give a reason for your answer.
d) It is safer to use helium gas than hydrogen gas in balloons. Which property of helium makes it a safer option?
e) Group IA elements readily form positive ions. Is this statement correct? Explain your answer by referring to the table.

UNIT 2
Chemical Bonds and Forces of Attraction

Expected Learning Outcomes

By the end of this unit, you are expected to:

- explain the different types of intramolecular forces
- describe the nature of intermolecular forces of attractions and their influence on the physical properties of substances
- differentiate between the different types of intramolecular and intermolecular forces of attraction.

2.1. Forces of Attraction

Let us discuss

In one high school a football competition was held among the school's Grade 10 classes. The winning class is expected to play with the winning classes of other high schools. This means there will be a competition between the winning classes of different high schools.

1. How many kinds of competition will the winning class be taking part in?
2. What is the difference in the nature of the competitions?

Just at there intraschool competitions and interschools competition that learners may participate in, so there are two types of attractive forces: intramolecular and intermolecular. Chemical bonds are the forces of attraction that hold atoms together: can be described as attractive forces. **Intramolecular** bonds refer to the forces of attraction that hold atoms together within a molecule. There is also another type of attractive force between molecules called an **intermolecular force** of attraction. Generally, intramolecular forces are considerably stronger than intermolecular forces.

While the intramolecular forces keep the atoms in a molecule together and are the basis for the chemical properties, the intermolecular forces are those that keep the molecules together and are responsible for all the physical properties of a material. In

Figure 2.1, there are intramolecular forces of attractions between hydrogen atoms and oxygen atom in one water molecule of which is a typical chemical bond. And there is intermolecular force of attraction between two water molecules.

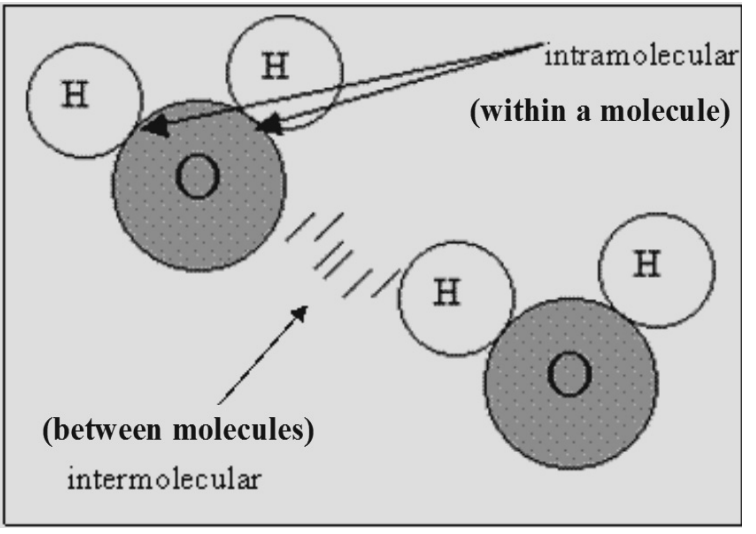


Fig 2.1 Intramolecular and intermolecular forces of attractions in water.

(The prefixes '*intra*' and '*inter'* come from the Latin stem meaning "within or inside." and "between." respectively.)

In grade 9 chemistry you have learned an introductory part of chemical bonds, mainly ionic and covalent bonds. In this unit you will learn the different types of intermolecular and intramolecular forces of attractions.

2.2 Chemical Bonds

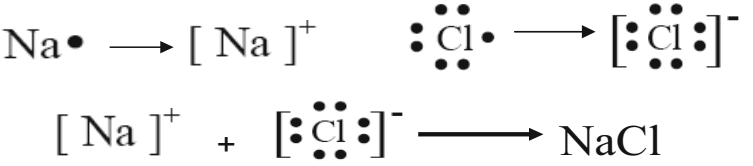
Chemical bonds such as ionic bonding, covalent bonding and metallic bonding are categorized as intramolecular forces.

2.2.1. Ionic Bonding

An ionic bond is an electrostatic force of attraction between a positive ion of a metal and negative ion of a nonmetal. Metals lose electrons to become positive ions; nonmetals gain electrons to become negative ions. This is because both metals and nonmetals want to attain stable noble gas configuration. In the following

example, Na loses an electron to become Na^+ and Cl gains an electron to become Cl^-. This follows the **octet rule** (attaining 8 valence electrons).

$$Na\ (g) \rightarrow Na^+(g) + 1e^- \qquad Cl\ (g) + 1e^- \rightarrow Cl^-(g)$$
$$2,8,1 \qquad 2,8 \qquad\qquad 2,8,7 \qquad 2,8,8$$

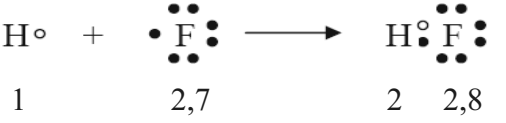


The equation given above shows the formation of ions and it leads to the attraction of the oppositely charged species in ionic bonds. Chlorine has higher electronegativity (2.9) compared to sodium (0.9); chlorine can easily take sodium's one lonely valence electron. This makes chlorine more stable, as it now has eight valence electrons. This also makes sodium more stable, as it also now has eight valence electrons.

2.2.2 Covalent Bonding

A covalent bond is formed by two nonmetallic atoms when they share one or more pairs of valence electrons between them. In covalent bond, the more electronegative atom attracts the shared electron pairs more than the one with less electronegativity value.

Example: Let us look at the bond that is formed between H and F.



Both hydrogen and fluorine want to attain a stable noble gas configuration. But both elements have a tendency to gain an electron and thus neither wants to lose from its valence electron. So the only way of fulfilling their valence electrons is by sharing; both contribute a single electron for sharing.

By sharing a single pair, two pairs and three pairs of electrons; single, double and triple bonds can be formed respectively.

Note that in the H_2 molecule each hydrogen atom has only one valence electron, and thus each hydrogen atom shares a pair of electrons, forming **single bond.** In O_2 molecule each oxygen atom has six valence electrons and needs two more electrons to fulfill the octet configuration. Thus, two pairs of electrons are shared between the two oxygen atoms to form a **double bond**. In the N_2 molecule each nitrogen atom has five valence electrons to fill its outer shell. Thus, by sharing three pairs, both nitrogen atoms form **triple bonds.**

Activity 2.1: Comparison of ionic bonds and covalent bonds

Copy and complete the following comparison of ionic and covalent bonds; find the appropriate words or phrases for the letters of the alphabet, A, B, C, D, and E.

Description	**Ionic Bond**	**Covalent Bond**
Consist of	Ions	A
Formed by	B	sharing electrons
Made from	metals and nonmetals	C
Strength of bond	Strong bond	D
Examples of	E	H_2O, CO_2, NH_3,

Ionic and covalent compounds differ markedly in their general properties because of the difference in the nature of their bonds. Ionic compounds conduct electricity when dissolved or in the molten state, whereas most covalent compounds do not conduct electricity when they dissolve in water. This makes them poor conductors of electricity in aqueous solution. Covalent solids are usually much softer than ionic materials. Ionic crystals are harder but often quite brittle. The melting and boiling points of ionic compounds are usually very high. Covalent compounds melt and boil at much lower temperatures than ionic compounds do.

2.2.3. Polar and Non- polar Covalent Bonds

Activity 2.2: Pulling strength and direction of movement



Two men (A and B) are pulling the woodblock in opposite directions.

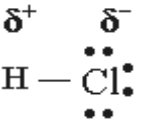
1. If Mr A and Mr B have equal attractive power, what happens to the woodblock?
2. If Mr B is stronger than Mr A, in which direction does the woodblock move? Why?

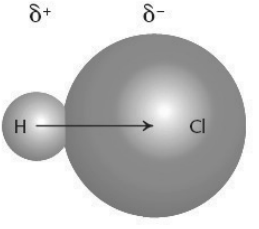
There are two types of covalent bonds: polar and nonpolar.

A Polar covalent bond is a covalent bond in which the electrons are not equally shared. This unequal sharing is due to the electronegativity difference of the two bonding atoms. The more electronegative atom attracts the shared electron pairs more than the one with less electronegative value. This produces a charge difference called **polarity**.

The one end of the polar bond with the larger electronegativity is the partially charged negative end (**δ-)**. The end of the bond with the smaller electronegativity is the partially charged positive end (**δ+)**. Thus a **dipole**, two different poles**,** is formed.

Example: HCl

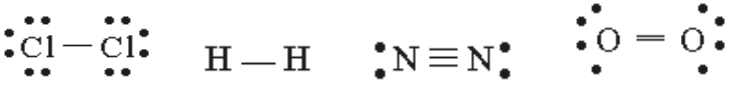




Note: **δ** (read as "delta") means partially, so **δ+** means partially positive and **δ-** means partially negative.

In this H-Cl bond, for example, the bond is covalent, but the electrons are shared unequally, so the bond is polar. The Cl atom has the higher electronegativity so the electrons are drawn more towards the Cl atom and away from the H atom. Since electrons carry a negative charge, this gives the Cl atom a partial negative charge. The H atom, consequently, has a partial positive charge. Other examples of compounds with polar bonds are: H_2O, HF, ClF.

A nonpolar covalent bond is one in which electors are shared equally. Example:-The diatomic elements (F_2, Cl_2, Br_2, I_2, H_2, N_2, O_2). This is because the atoms are sharing the electron pairs equally. Look at the following Lewis structures:



All the halogens and hydrogen are sharing a pair of electrons equally. Oxygen and nitrogen are sharing two pairs and three pairs of electrons respectively. Thus all are forming a stable noble gas configuration.

Bond Polarity and Electronegativity Differences

Generally, compounds can be classified as ionic, polar and nonpolar depending on the electronegativity difference of the bonding atoms. Electronegativity is the attractive power of an atom for shared electron pairs. The electronegativity data of some common elements are given in Unit.

The type of bond formed is largely determined by the difference in electronegativity between the atoms involved. Atoms with similar electronegativities will share an electron pair with each other and form a covalent bond, However, if the difference is too great, the electron will be permanently transferred to one atom and an ionic bond will form. Furthermore, in a covalent bond if one atom pulls slightly harder than the other, a polar covalent bond will form.

Generally, bonds can be classified according to the difference in electronegativities of the atoms (ΔEN). Bonds are:

- **ionic** if $\Delta EN > 1.8$
- **polar covalent** if $0.4 \leq \Delta EN \leq 1.8$
- **nonpolar covalent** if $\Delta EN < 0.4$

The larger the electronegativities difference (ΔEN) of the atoms in a bond, the stronger the bond will be. As the bond becomes stronger, the melting and boiling points generally increase too.

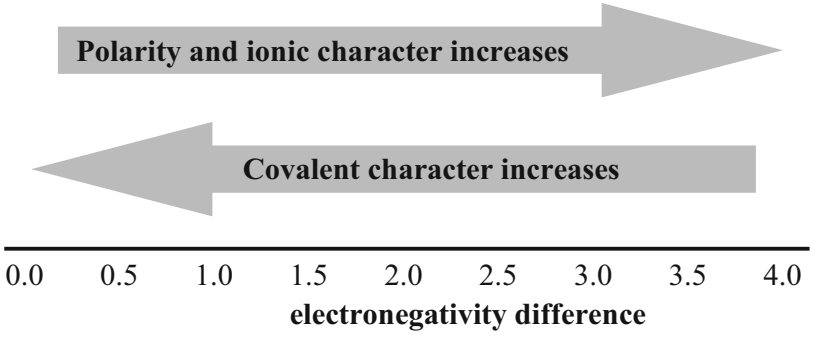


Fig 2.2 Electronegativity versus polarity

Example 1: NaCl and LiF are good examples of ionic compounds. In each case, the electronegativity of the nonmetal is at least two units larger than that of the metal. These compounds are highly polar and are ionic species.

NaCl		**LiF**	
Cl	EN = 3.0	F	EN = 4.0
Na	EN = 0.9	Li	EN = 1.0
ΔEN	**2.1**		**3.0**

Example 2: Examples of covalent compounds include nitrogen dioxide (NO_2) and sulphur dioxide (SO_2). When the electronegativity difference of the elements in a compound is small, the atoms share electrons: and the substance is covalent.

NO_2		SO_2	
O	EN = 3.5	O	EN = 3.5
N	EN = 3.0	S	EN = 2.5
ΔEN	**0.5**		**1.0**

Activity 2.3: Family relationships versus bonding

Mr. Aman and his wife Rahwa have one child called Simon. Both Aman and Rahwa work all day. They leave their child, Simon, with either Aman's or Rahwa's parents. Rahwa's parents have a car and regularly take Simon to the zoo, which Simon really enjoys. Over the period of a month, there are three different ways in which Simon spends time with both sets of grandparents:

1. Simon remains with his mother's parents and never sees his father's parents.
2. Simon spends one week with his mother's parents and the next week with his father's parents.
3. Simon spends most of the time with his mother's parents, but every Wednesday and Saturday with his father's parents.

Which of the above relationships shows a type of ionic, polar or non-polar covalent bonding? Explain your answer.

2.2.4 Coordinate Covalent Bonding

Let us discuss

Mr Hamid usually shares dinner with his best friend, Mr Paulos, at a restaurant. One day, Mr Hamid tried to pay the bill for the dinner. However, Mr Paulos convinced him and thus both paid the bill equally. On another occasion, Mr Hamid received a good sum of money from his job and thus he invited his friend and both shared a dinner together. Mr. Hamid paid the entire bill for that night.
Which of these occasions represents a covalent type of relationship? Why? What type of relationship occurs in the other situation?

Generally, a covalent bond between two atoms is formed by sharing an electron pair; each atom contributes one electron to the pair. If both electrons of the pair are contributed by the same atom ,the bond formed is called a **coordinate covalent bond**. A coordinate

covalent bond is also called **a donor-acceptor bond** or a **dative bond.**

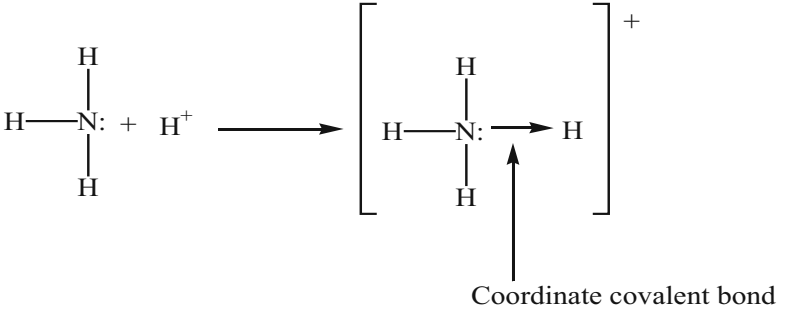
Example

The reaction between ammonia and a hydrogen ion to form an ammonium ion is an example of a coordinate covalent bond:

$$\underset{\text{Ammonia}}{NH_3} + \underset{\text{Hydrogen ion}}{H^+} \rightarrow \underset{\text{Ammonium ion}}{NH_4^+}$$

In ammonia, nitrogen forms three covalent bonds with hydrogen and a lone pair (two electrons) is free for the formation of a dative bond. The hydrogen ion (H^+) does not have any electron and thus H^+ can share the two electrons of nitrogen in ammonia. Thus when the ammonium ion, ($NH_4^{+)}$, is formed, the fourth hydrogen is attached by a dative covalent bond with nitrogen.

In simple diagrams, a co-ordinate bond is shown by **an arrow**. The arrow points from the atom donating the lone pair to the atom accepting it.



All four N-H bonds in the ammonium ion are found by experiment to be equivalent. Once it is formed, a coordinate covalent bond is identical to a normal covalent bond.

Exercise 2.2

1. Classify the following molecules as either ionic, or polar covalent or nonpolar covalent: HF, H_2, LiF, H_2O, O_2.
2. In the following table the melting points and electronegativity differences of NaCl, LiF and KBr are given. Correlate the

melting points of NaCl, LiF and KBr with the electronegativity differences of the compounds.

	NaCl	***LiF***	***KBr***
MP	801 °C	846 °C	657 °C
ΔEN	2.2	3.0	2.1

3. Arrange the following molecules in order of increasing polarity:

HCl, NaCl, H_2O, CsF, Cl_2

4. Show the coordinate covalent bonds in the following compounds. a) In H_3O^+ ($H_2O + H^+$)
b) Between NH_3 and BF_3

2.2.5 Metallic Bonding

Let us discuss

1. Mention some examples of metals and their symbols
2. What are the common properties of metals?

Metals are held together in solid form by metallic bonding. **Metallic bonding** refers to the interaction between the delocalized electrons (freely moving electrons) and the positive metal ions. Valence electrons in metals are mobile (run freely in the entire solid); they are not restricted to certain atoms or bonds.

In Figure 2.3, the positive metallic centre is shown with the circular positive charges and the freely moving electrons using small dots. Each positively charged metal ion is attracted to the negatively charged delocalized electrons, and vice versa.

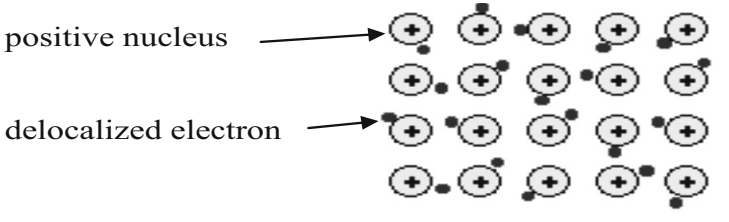


Fig 2.3 A model of metallic bonding

The fact that the electrons flow easily helps to explain some of the properties of metals:

1. Metals are good conductors of heat and electricity. This is due to the mobility of their electrons.
2. If a metal has more delocalized electrons, it results in a stronger metallic bond and a higher melting point.
3. Metals are lustrous (shining). This is due to the uniform way that the valence electrons of the metal absorb and re-emit light energy.
4. Metals are malleable (can be flattened) and ductile (can be drawn into wires) because of the way the metal cations and electrons can move around each other, without breaking the crystal structure

The following Table 2.2 summarizes the model, the basis of attraction and examples of the three types of chemical bonds:

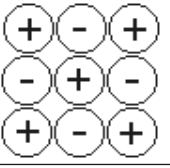
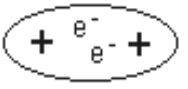
Force	Model	Basis of attraction	Example
Ionic		Opposite charges	NaCl
Covalent		Nuclei and shared electron pair/s	H – H
Metallic		Metal cations and delocalized electrons	Au

Table 2.2 Summary of chemical bonds

Summary

A bond is formed when atoms interact with each other; they interact with each other in order to attain a stable noble gas configuration. There are different types of bond, namely, ionic, covalent and metallic.

Ionic bonds are formed by the transfer of electrons between metals and nonmetals. An example of an ionic bond is NaCl.

Covalent bonds are formed by the sharing of electrons, which commonly occurs when two nonmetals bond together. The H-O bond in H_2O is an example of a covalent bond.

There are two types of covalent bonds; polar and non -polar. A polar bond is formed when electrons are unequally shared between two atoms. An example is HCl. A nonpolar bond is formed when the bonding electrons are equally shared. An example is H_2.

A coordinate covalent bond, also called a dative bond, is a force of attraction resulting from the sharing of a lone pair of electrons from one atom with another atom. Metallic bonding refers to the interaction between the delocalized electrons and the positive metal ions. The physical properties of metals, electrical conductivity, high melting and boiling point, lustre, malleability and ductility are mainly due to the delocalized electrons.

2.3 Intermolecular Forces of Attraction

We will study three types of intermolecular force that can operate between molecules.They are ion-dipole interactions, van derWaals forces, and hydrogen bonds.

2.3.1. Ion-Dipole Forces

These attractive forces occur when an ionic substance dissolves in a polar liquid, e.g. NaCl in H_2O. The positive and negative poles of the polar liquid attract the negative and positive ions of the ionic substance, respectively. See Fig 2.6 below.

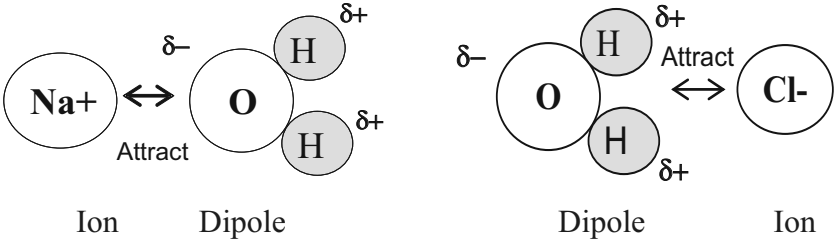


Fig 2.6 Ion -dipole interactions

The strength of these interactions depends on the charge and the size of the ion and on the magnitude of polarity and size of the dipole.

For example, if you compare the ion dipole interaction between Na^+ and Mg^{+2} with the water molecule, Mg^{+2} has a higher charge and a smaller ionic radius than those of Na^+, so Mg^{t2} interacts more strongly with water molecules.This means that the ion-dipole interaction of Mg^{+2} ---H_2O is stronger than that Na^+---H_2O. Similar differences exist for anions of different charge and sizes.

2.3.2. van der Waals Forces

van der Waals forces are weak attractive forces that hold molecules together. There are three types: dipole-dipole, dipole induced dipole and dispersion (London) forces.

1. Dipole-Dipole forces

These forces occur when polar molecules arrange themselves such that their oppositely charged ends can attract each other.

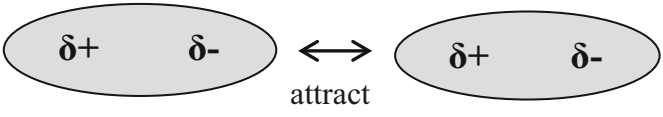
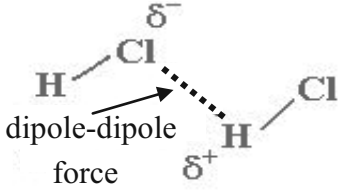


Fig 2.7 Dipole-dipole interactions

For example, HCl molecules are attracted to each other by dipole – dipole attractions. The hydrogen atom has a slightly positive charge and the chlorine atom has a slightly negative charge. Because of the force of attraction between oppositely charged particles, there is a small dipole-dipole force of attraction between adjacent HCl molecules.



2. Dipole-induced dipole forces

This is an interaction between polar and nonpolar molecules.

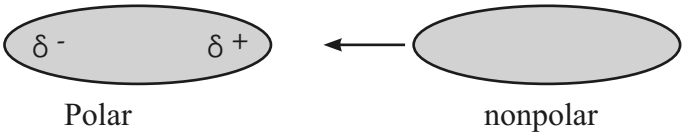


Fig 2.8 Approach of nonpolar to polar molecule

In fig 2.8, as the right-hand molecule approaches, its electrons will tend to be attracted by the slightly positive end of the left-hand one. This sets up an ***induced dipole*** in the approaching molecule, which is oriented in such a way that the partial positive (**δ+)** end of one is attracted to the partial negative (**δ-)** end of the other.

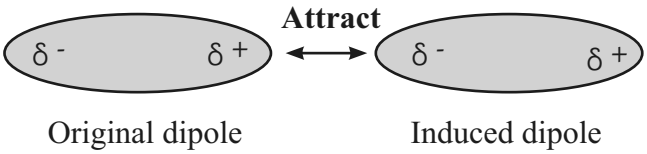


Fig 2.9 Induction of dipole in the nonpolar substance

For example, dipole-induced dipole interaction is observed between HCl and H_2.

3. Dispersion forces

Dispersion forces are also known as "London forces" and they occur between nonpolar molecules. Dispersion forces exist between nonpolar molecules and as when the molecules approach each other, the protons and electrons from different particles attract one another slightly. The electron clouds of these molecules then become distorted and this results in a weak dipole. Both molecules bring a bout temporary polarity. The attraction is thus also called induced dipole- induced dipole interaction".

For example, induced dipole- induced dipole interaction is observed between I_2 and I_2 molecules.

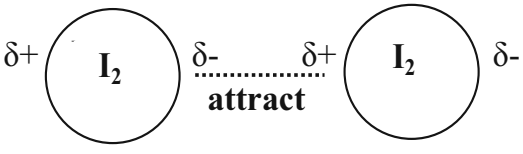


Fig 2.10 Dispersion forces between two iodine molecule

The presence of van der Waals forces affects the properties of certain molecules; gases such as hydrogen, oxygen, and argon liquefy and solidify, whereas molecules, such as naphthalene, crystallize because of the forces of attraction between the particles.

Attractions resulting from van der Waals forces are considered to be very weak. The energy required to break van der Waals is of the order of 5 kJ/mol. That is very weak in comparison to the 436 kJ/mol needed to break a covalent bond in hydrogen molecules, H_2 and 775 k J/mol for the ionic lattice in sodium chloride, NaCl.

2.3.3. Hydrogen Bonding

Activity 2.4: Comparison of water and alcohol

Materials required: Water, alcohol, 2 petri dishes, measuring cylinder, burner, match.

Pour 2 mL of water into a petri dish.
Try to burn it using the burner.
Pour 2 mL of alcohol into another petri dish.
Try to burn it using the burner.

1. Which of the liquids, water or alcohol, evaporates first? Why?
2. What is the common feature in both compounds?

Water
H_2O



Alcohol
CH_3CH_2-OH



A **hydrogen bond** is a particularly strong type of dipole-dipole interaction. It is an attractive force between polar molecules where a hydrogen atom is bonded to a highly electronegative element such as fluorine, oxygen or nitrogen.

In hydrogen bonding, the interaction is between the hydrogen atom in a polar N-H, O-H, or F-H bond and an electronegative atom (F, O, or N). The interaction is written as:

A-H---B or A-H---- A (A and B represent O, N, or F)

The simplest molecules capable of forming hydrogen bond are ammonia, water (H_2O) and hydrogen fluoride, (HF) Hydrogen bonding is strongest between HF molecules where the electronegativity difference is large and it is weakest between NH_3 molecules where the electronegativity difference is small. Examples are given below:

Example: Show the hydrogen bond formation in the following molecules:

A. HF B. NH_3

Answer

A. In the HF molecule the partially positive hydrogen of one HF molecule is attracted to the partially negatively charged fluorine atom of the nearby fluorine atom.

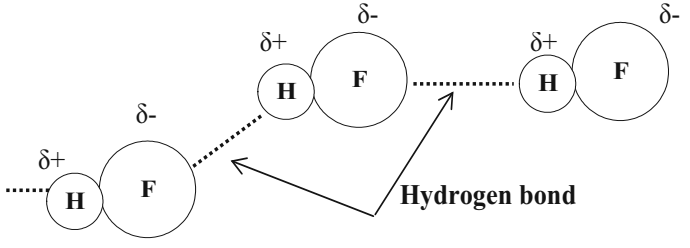


Fig 2.4 Hydrogen bond in hydrogen fluoride molecules

B. The NH_3 molecule is polar,with the nitrogen at its negative end and the hydrogen at its positive end. The hydrogen atom in the N-H bond will be attracted to a nitrogen atom from a nearby NH_3 molecule and should form a strong hydrogen bond.

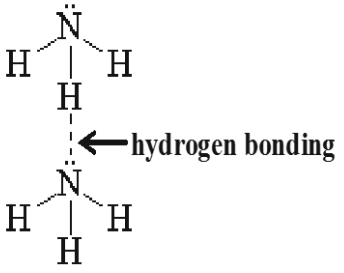


Fig 2.5 Hydrogen bonding in ammonia molecules

Exercise 2.3

1. What are the criteria for the formation of a hydrogen bond?
2. In which of the following can we have hydrogen bonding?
   A. HCl B. CO_2 C. H_2O
3. Describe the hydrogen bond formed in a water molecule?

Hydrogen bonding affects the following properties.

1. Melting point and boiling point

Comparing two similar compounds, water, (H_2O) and hydrogen sulphide, (H_2S); it was expected that H_2S will have a higher boiling point because as you go down a group the boiling point of similar compounds increases. However, the boiling point of water is much higher (100 ^oC) than hydrogen sulphide (-60 ^oC). This is because water molecules are bound by hydrogen bonding whereas hydrogen sulphide molecules are not.

2. Structure in solid state

The presence of hydrogen bonds links the molecules in a regular pattern. This results in the regular structure of these substances when they are in solid state. For example, in ice -the solid form of water, the water molecules are arranged in order due to the presence of hydrogen bonding.

Generally, the relative strengths of intramolecular and intermolecular forces are as follows:

Intermolecular forces (hydrogen bonds, ion-dipole and van der Waals forces) are much weaker than intramolecular forces (ionic bonds, covalent bonds, and metallic bonds). Dispersion forces are the

weakest intermolecular force and hydrogen bonds are the strongest intermolecular force. The order of increasing intermolecular forces is:

dispersion forces < dipole-dipole interactions < hydrogen bonds

Table 2.3 below is a tabular summary of the model, basis attraction, and examples of intermolecular forces of attractions.

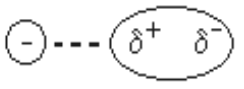
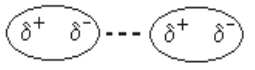
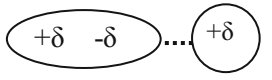
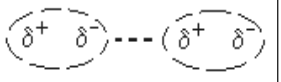
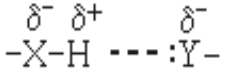
Force	Model	Basis Of attraction	Example
Ion-dipole		Ion and polar molecule	Na^+ and H_2O
Dipole-dipole		Partial charges of polar molecules	HCl and HCl
Dipole-induced dipole		Partial charge of polar molecule and nonpolar species	HCl and Ar
London dispersion		Induced dipoles of polarizable molecules	Xe and Xe
Hydrogen bond		H bonded to N, O, or F, and another N, O, or F	H_2O and NH_3

Table 2.3 Summary of intermolecular forces of attractions

Activity 2.6 : (Word Search)

N	E	G	O	R	D	Y	H	B	I	F	Y	I	M	K	C
O	D	G	E	T	Y	M	**F**	**M**	F	G	B	O	P	L	C
A	C	I	P	Z	D	L	**E**	J	G	E	F	Z	Z	N	O
C	P	V	L	N	D	**T**	C	O	V	A	L	E	O	T	O
I	K	A	L	M	**A**	B	L	D	C	V	B	N	Y	D	R
N	P	I	L	**L**	O	R	Y	O	I	G	P	I	C	A	D
O	U	N	**L**	W	S	P	J	U	D	O	N	T	N	N	I
I	K	**I**	A	O	C	X	H	B	L	F	L	N	E	E	N
B	**C**	A	O	V	D	E	C	A	H	L	C	E	D	S	A
F	F	X	P	X	I	J	R	E	D	W	O	L	X	W	T
Y	Y	Y	R	H	P	A	A	O	O	W	O	A	E	W	E
D	N	O	D	N	O	L	R	R	M	E	C	V	Z	Y	O
P	C	Y	J	K	L	G	K	P	S	T	U	O	S	J	N
D	O	O	T	W	E	Y	X	M	A	S	Y	C	E	X	N

Words about the concepts in this table can be obtained from the definitions given below. You may find the word in the table above either horizontally or diagonally; down words or up words. For example:- The type of bond that holds metals together "**METALLIC"** is shaded in the table. In a similar way, find the correct words or phrases that represent the following and shade it using a pencil on your note- book.

1. In _____ bonding, simultaneous attraction by two atomic nuclei for one or more electrons occurs.

2. A molecule that contains partially charged positive and negative terminals.

3. A relatively strong type of bond between hydrogen and a highly electronegative atom such as F, O or N

4. The weakest type of intermolecular attraction.

5. A type of covalent bond in which two pairs of electrons are shared equally

6. A bond formed between an active metal and a nonmetal.

7. Type of covalent bond where electrons are donated by one atom.

Summary

Hydrogen bonds occur between hydrogen in a polar molecule, being covalently bonded to the most electronegative atoms such as fluorine, oxygen, or nitrogen. For example, hydrogen bonds operate between water (H_2O) molecules, ammonia (NH_3) molecules, and hydrogen fluoride (HF) molecules,

Ion-dipole forces appear between an ion (cation or anion) and a polar molecule. For example, sodium ions and chloride ions are attracted by the polar water molecule.

Dipole-dipole interactions are stronger intermolecular forces than dispersion forces and occur between polar molecules, For example, dipole-dipole interactions occur between HCl molecules and KBr molecules. Dispersion forces usually exist between nonpolar species, such as He atoms and I_2 molecules.

Key terms

Coordinate covalent bond
Covalent bonding
Dipole
Dipole-dipole forces
Dispersion (London) forces
Hydrogen bonding
Polar covalent bond
Polarity
van der Waals forces
Intermolecular
Intramolecular
Ion-dipole forces
Ionic bonding
Metallic bonding
Nonpolar covalent bond

Review Exercise

Part I. Multiple choice questions

1. Which is a nonpolar covalent substance?
   a. Cl_2 b. NH_3 c. H_2O d. KCl.

2. Which of the following pairs shows dipole-dipole forces?
   a. HCl and HCl b. HF and H_2
   c. H_2 and H_2 d .HF and F_2

3. Which of the following molecules shows the greatest effects of hydrogen bonding?
   a. H_2S b. CH_4 c. NH_3 d. H_2O

4. The correct order of increasing attractive strength for weak intermolecular forces is
   a. dipole-dipole, hydrogen bonding, van der Waals forces
   b.van der Waals forces, dipole-dipole, hydrogen bonding
   c. hydrogen bonding, dipole-dipole, van der Waals forces
   d. hydrogen bonding, van der Waals forces, dipole-dipole

5. Which of the following pairs do not form a dative covalent bond to each other?
   a. NH_3 and H^+ b. H_2O and H^+
   c. NH_3 and BF_3 d. NH_3 and NH_3

6. Which one of the following is not true of metallic bonding?
   a. It gives rise to excellent electrical conductivity
   b. Electrons are free to move throughout the structure
   c. The strength of metallic bonds increases down a group
   d. The strength of metallic bonding affects the boiling point of metals:

7. The bond between two identical nonmetal atoms has a pair of electrons_____:
   a. unequally shared between the two.
   b. equally shared between them
   c. transferred fully from one atom to another
   d. none of the above

8. The hydrogen bond is strongest in:
   a. O-H…S b. S-H…O c. F-H…S d. F-H...O

9. The atoms in a molecule of hydrogen chloride (HCl) are held together by
   a. ionic bonds b. van der Waals forces
   c. dipole-dipole attraction d. polar covalent bonds

10. The attraction between nonpolar molecules is called
   a. van der Waals forces: b. hydrogen bonds
   c. electrovalent forces. d. covalent bonds.

11. Which type of bonding involves positive ions being immersed in a sea of mobile electrons?
   a. ionic b. nonpolar c. metallic d. polar

12. The abnormally high boiling point of HF as compared to HCl is primarily due to intermolecular forces of attraction called
    a. network bonds b. electrovalent forces
    c. van der Waals forces d. hydrogen bonds.

Part II. Fill in the blank spaces

1. _____ forces exist within each molecule of a substance (the chemical bonds),
2. _____ forces exist between molecules of a substance.
3. The force of attraction between an ion and a polar molecule is_____.
4. The force of attraction between one polar molecule and another polar molecule is _____.
5. Hydrogen bonding occurs whenever there is _____ attached to a _____ _____ element.
6. _____ forces are due to forces of attraction between instantaneous dipoles.

Part III. Essay type Questions

1. Use electronegativities to decide whether the following compounds are best described as ionic, polar or nonpolar bonds.
    a) Sodium chloride (NaCl)
    b) Carbon monoxide (CO)
2. Why do molecules attract one another?
3. How are London forces formed?
4. What is a hydrogen bond?
5. Identify the strongest intermolecular (force present in pure) samples of the following substances:
    a) H_2O b) Cl_2 c) HI
6. Describe whether the following compounds are likely to be ionic or not ionic based on the properties given. Explain your reasoning.
    a) Compound 1 has a melting point of 545 ^oC and dissolves well in water.
    b) Compound 2 is a brittle material that is used to melt road ice during storms.
    c) Compound 3 melts at 85 ^oC and catches fire when heated

UNIT 3

Solutions and their Proporties

Expected Learning Outcomes

By the end of this unit you are expected to:

- describe types of solutions
- calculate the concentration of solutions and prepare solutions of different concentrations
- explain the properties of solutions
- describe colligative properties and their application.

3.1 Solutions

3.1.1 Describing solutions

Let us discuss

Classify the following mixtures as either homogeneous or heterogeneous.

1. A mixture of table salt and water.
2. A mixture of water and ethanol.
3. A mixture of oil and water.
4. A mixture of sand and water.
5. What is the basis for the classification in numbers 1-4?

When table salt is added to water the salt disappears. This is because the solid crystals of the salt break apart to form individual ions and spread uniformly throughout the water. When ethanol is added to water, it also spreads uniformly throughout the water. However, when sand or oil is added to water it is not spread uniformly throughout the water and one observes that the constituents of the mixture are physically separated.

Mixtures are classified into two based on whether the constituents are uniformly spread uniformly throughout mixture or remain physically separate. Mixtures in which the constituents are physically separated are called **heterogeneous** and those in

which the components are spread uniformly throughout are called **homogenous mixtures** or **solutions**.

A solution is composed of a solute and a solvent. A **solvent** is the substance that dissolves the solute. The material dissolved is called a **solute**.

solute + solvent = solution

Solids, liquids and gases can dissolve in a solvent. Solutions in which the solvents are liquids are called **liquid solutions**. When the solvent is water the liquid solution is called an **aqueous solution**.

It is very important to note that the physical state of a solution is not limited to the liquid state. A gas can dissolve in a gas to produce a **gaseous solution** or a solid can dissolve in a solid to give a **solid solution**. The air we breathe in is a mixture of various gases and it is an example of a gaseous solution. Coins that we use in our everyday transactions, and jewellery such as rings and necklaces are examples of mixtures of metals. Mixtures of metals are called **alloys** and they are solid solutions. Table 3.1 shows examples of common types of solutions.

Type of solution	Solvent	Solute	Example
Liquid	water	table salt	sea water
Liquid	water	ethanol	alcoholic drinks
Liquid	water	carbon dioxide	carbonated beverages such as soft drinks and carbonated waters
Gaseous	nitrogen	oxygen, water vapour, carbon dioxide, etc.	air
Solid	iron	carbon	steel
Solid	mercury	silver	amalgam (used as dental filler)

Table 3.1 Common types of solutions

In Table 3.1 a mixture of water and ethanol is given as an example of a liquid-liquid solution. Two liquids such as water and ethanol, which are able to mix together, produce a solution that is said to be miscible. Liquids that do not mix are said to be immiscible. A mixture of water and oil is an example of an immiscible solution. Solids that dissolve in a liquid are said to be soluble in that solvent. For example, sodium chloride is soluble in water. Solids that do not dissolve in liquids are said to be insoluble

It is difficult to classify the constituents of a solution as a solute or a solvent when both are in the same physical state. In such solutions, the one constituent present in a relatively greater quantity is classified as the solvent and the one constituent present in a smaller quantity is classified as the solute. For example, if 5 mL of alcohol is added to 20 mL of water, the alcohol is the solute and the water is the solvent.

Some characteristics of a liquid solution are:

1. The solution is clear and transparent; or it may be the colour of either of the constituents.
2. It has a variable composition. In a solution, the ratio of solvent to solute is not fixed, and it can vary over a wide range.
3. A solute breaks up into tiny particles. The dissolved solute is molecular or ionic in size.
4. Sedimentation of the particles does not occur in a solution. The particles do not sink to the bottom of a solution when it is left to settle.
5. A solution will pass through filter paper because the dissolved particles are extremely small in size. Therefore filtration cannot be used to separate the substances that make up the solution.
6. Physical properties such as the boiling point and the freezing point of a solution are different from those of a pure solvent.

3.1.2 Unsaturated, Saturated and Supersaturated Solutions

Activity 3.1: Demonstrate the variable composition of a solution
Materials required: Water, sodium acetate ($Na_2C_2H_3O_2$), a scale (50 mL), beaker, a stirrer, a burner.

Pour 10 mL of water into the beaker.
Add 3 g of sodium acetate to the water in the beaker and stir the solution.

1. What happens to the sodium acetate when it is added to the water in the beaker?

Add another 2g of sodium acetate to the same beaker and continue stirring.

2. What happens to the sodium acetate when it is added to the beaker?

Add a further 3g of sodium acetate to the same beaker and stir.

3. What happens to the sodium acetate when it is added to the beaker?

If there is any undissolved sodium acetate remaining in the beaker, heat the solution to 40 ^oC. Continue stirring.

4. What happens to the sodium acetate when it was heated to 40 ^oC?

Allow the heated solution to cool to room temperature without disturbing it.

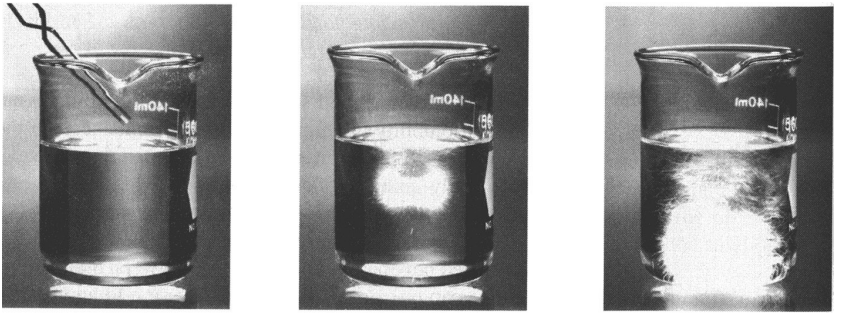
5. What do you observe when the heated mixture cooled to room temperature?

Add a tiny crystal of sodium acetate to the cooled solution and wait for 2 minutes?

6. What do you observe?

A solute will dissolve in a solvent if the amount of solute is added in small quantities. Further, a solute will still dissolve in a solvent if you continue adding solute in small amounts at a particular temperature. However, a point will be reached when no more solute can dissolve in the solution. This means there is a limit to the amount of solute that can dissolve in a given amount of solvent.

A solution which can dissolve more solute is called an **unsaturated solution**. A solution that cannot dissolve any more solute is called a **saturated solution**. If more solute is added to the saturated solution, it will remain at the bottom of the container. In general, the excess solute dissolves when the saturated solution is heated. Sometimes, when such a solution is cooled, without being disturbed, the solute will remain in the solution. A cooled solution that contains more solute than a saturated solution is called a **supersaturated.** It is important to note that only some solutes form a supersaturated solution. Supersaturated solutions are extremely unstable. If disturbed or if a seed of crystal is added they readily revert to the point of saturation by forming crystals as shown in Figure 3.1.



A B C

Fig 3.1 (A) A seed crystal is being added to a saturated solution: (B) crystallization begins and (C) crystallization continues as the solution converts to saturation.

Activity 3.2: Compare the capacity of a bus with the capacity of a solvent to dissolve a solute.

A minibus has the capacity to transport 24 passengers. Twelve passengers are sitting "homogeneously" at bus stop 1. Six passengers each get onto the bus at bus stops 2 and 3 respectively and sit homogeneously.

Four men are sitting in the hot sun at bus stop 4. The driver decides to pick them up and asks them to sit homogeneously amongst the seated passengers. The men who boarded at bus stop 4 got off at bus stop 5 because this bus stop provides shade.

1. How many passengers are there in the bus between stop 3 and stop 4? What is the relationship between the number of passengers and the capacity of the bus?
2. Suggest a possible homogeneous seating arrangement at bus stop 3. Assume there are four seats in a row.
3. Match the analogue features to the target features (characteristics of solution).

Analogue features	Target features
a. Bus capacity and passengers	i. Saturated solution
b. Bus stops 1 and 2 before reaching full capacity	ii. Supersaturated solution
c. Bus stop 3 reaching full capacity	iii. Solvent and solute
d. Men sitting in the hot sun being taken to the next.	iv. Dissolution
e. Sitting homogeneously on the seats of the bus	v. Unsaturated solution

3.1.3 Solubility and the Solubility Curve

Activity 3.3: Comparison of solubility of salts.

Materials required: Potassium iodide (KI), lead (II) chloride ($PbCl_2$), sodium chloride (NaCl), water, stirrers, four beakers (50 mL), a scale.

Label the four beakers 1, 2, 3 and 4 respectively and pour 5 mL water into each of them.
Add 1 g of $PbCl_2$ to beaker 1 and 1g of NaCl to beaker 2 and stir both mixtures.
Add 3 g of NaCl to beaker 4 and 4g of KI to beaker 4 and stir both mixtures.

1. What do you observe when the salts were added to beakers 1 and 2?
2. What do you observe when the salts were added to beakers 3 and 4?
3. Which of the three salts is the most soluble in water?
4. Which of the three salts is the least soluble in water?

The amount of a solute that can dissolve in a given volume of solvent is different for different substances. Some are highly soluble while others are less soluble. Solubility is defined as the maximum amount of solute that can dissolve in a specified amount of solvent at a given stated temperature.

Solubilities are most commonly expressed as the number of grams of solute that can dissolve in 100 g or 100 mL of solvent at a given temperature. The solubility of NaCl in water is 35.9 g/100 mL of water at 25 °C. This means the maximum amount of NaCl that you can dissolve in 100 mL of water is 35.9 g.

Solute	Formula	Grams of solute/100 mL of water at 25 °C
Potassium nitrate	KNO_3	40
Ammonium bromide	NH_4Br	97
Carbon dioxide	CO_2	1.45
Sucrose	$C_{11}H_{22}O_{11}$	211.5
ethanol	C_2H_6O	infinite

Table 3.2 Solubility of some common substances in water at 25 ° C

Activity 3.4: Determining solubility of a salt at different temperatures

Materials required: Potassium nitrate, water, a scale, test tubes (25 mL), beakers (400 mL), test tube clamps,a burner, a ring stand, wire gauze, a thermometer, a 10 mL graduated cylinder, a test tube rack and stirring rods.

Prepare a hot water bath by filling the beaker with water to approximately two-thirds of its capacity. Place the beaker on a wire gauze above a burner and begin heating the water to just below boiling. Label the test tubes 1 through 4. Accurately measure out the following masses of potassium nitrate, placing each quantity of the salt into different test tubes.

test tube 1: 2.5 g KNO_3
test tube 2: 4.0 g KNO_3
test tube 3: 6.0 g KNO_3
test tube 4: 8.0 g KNO_3

Add exactly 5.0 mL of water into each of the test tubes. Place each of the test tubes into the water bath in order to dissolve the solid KNO_3. You may find it necessary to use a stirring rod to speed up the dissolving process, particularly for test tubes 3 and 4.

Remove test tube 1 from the hot water bath once the KNO_3 has fully dissolved and place a thermometer in the tube. Watch the solution carefully. Record the temperature as soon as you see crystals forming within the test tube (you will need to wait awhile for crystals to form in this first test tube).

Remove test tubes 2, 3 and 4 one by one from the water bath and repeat the above process as for test tube 1.

1. Copy the table below in your notebook and complete it by filling in the blanks.

Test Tube	Mass of KNO3 (g) used	Volume of Water (mL	Convert to g/100 mL	Saturation Temp (°C)
1				
2				
3				
4				

2. Using graph paper, construct a graph by plotting solubility against temperature. Connect your data points by drawing a smooth curve. You may add the additional below on data to your curve:

   Solubility at 0 °C: 13 g/100 mL
   Solubility at 100 °C: 247 g/100 mL

3. Based on your graph, predict the solubility of KNO_3 at the following temperatures: a. 50 °C b. 70 °C

A key factor affecting the solubility of a solute is temperature. For most solid solutes increasing the temperature will increase solubility; that is, more solute will be able to dissolve in the same volume of solvent. A good number of solid solutes change only slightly with an increase in temperature and a few show decreased solubility. One simple way of illustrating solubility at different temperatures is by plotting solubility versus temperature on a graph. A graph of solubility against temperature is called a solubility curve. Figure 3.2 shows the solubility curve for certain common substances in water. From the graph, one can easily observe that the solubility of KNO_3 increases as temperature increases and that of $Ce_2(SO_4)_3$ decreases as temperature increases from 10 °C to 40 °C.

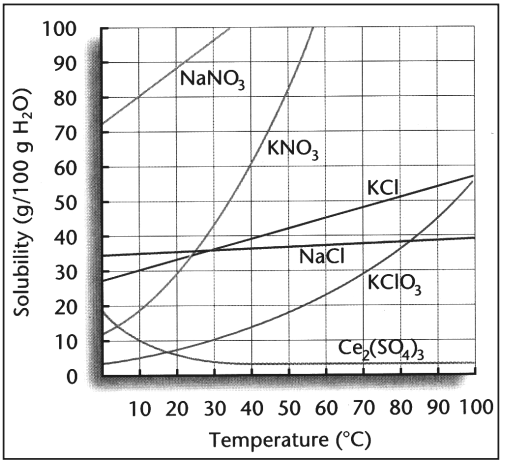


Fig 3.2 Solubility curve for certain substances in water.

A solubility curve has useful applications. Some of these are:

1. It illustrates the ability of a solute to dissolve in water at different temperatures. Each curve in the figure represents the concentration of a saturated solution at various temperatures.
2. The solubility of a solute at any temperature between the given lowest and highest temperatures can be read from the curve.
3. The mass of the crystals deposited when a saturated solution is cooled can be calculated from the curves.
4. The solubility of different substances at different temperatures can be compared by looking at their solubility curves.

Exercise 3.1

Refer to Fig, 3.2 to answer the following questions.

1. Copy the table below into your notebook and complete it by writing the mass of solute that will dissolve in 100 g of water at the temperatures indicated.

Type of solute	Temperature, ^oC	Mass of solute
KNO_3	70	
NaCl	100	
KCl	90	

2. Determine which of the above three substances is most soluble in water at 10 °C?
3. How many grams of KNO_3 will be deposited at 25 ^oC from 70 g dissolved in 100 g of water at 50 ^oC?
4. Compare the solubilities of NaCl and $KClO_3$ below and above 85 ^oC.

Summary

A solution is a homogeneous mixture of two or more substances. The substance that dissolves is called a solute and the substance the solute dissolves in is called a solvent. When both the solute and the solvent have the same physical state, the one present in a smaller quantity is classified as a solute and the one present in a larger quantity is classified as a solvent.

A solution can be solid, liquid or gas. However, most solutions exist in a liquid state. A solution can have variable amounts of solute; but there is a limit to the amount that can dissolve in a given volume of solvent. A solution in which more solute can still be dissolved is called unsaturated. A solution that cannot dissolve anymore solute is called saturated. In some cases, a solution cooled from a higher temperature can contain more solute than a saturated solution,and such a solution is called supersaturated.

The amount of solute that can dissolve in a given volume of solvent is temperature dependent. The solubility of most solids increases with an increase in temperature. The amount of solute that can dissolve in a given mass or volume of solvent at a specific temperature is the called solubility of that particular solute. A curve that shows how the solubility of a solute in a given solvent varies over a temperature range is called a solubility curve.

3.2 Calculations Involving Solutions

3.2.1 Expressing Concentration

Let us discuss

Eden, Halima, Yisak and Roble are neighbours and they usually sit together while one of them boils coffee. It is customary to boil the coffee in three rounds, and drink one cup from each round.

1. Which of the two cups of coffee, first or second round, will be stronger? Why?
2. Which of the two cups of coffee, second or third round, will be stronger? Why?

In chemistry, **concentration** refers to the amount of solute contained in a certain amount of solution. Because varying amounts of solute can dissolve in a solution, concentration is a variable property. Therefore, it is becomes important to be able to express the concentration of solutions in a certain way.

Two methods are used to express the concentration of a solution: qualitative and quantitative. The qualitative method of expressing concentration compares the relative quantities of the solute and the solvent in the solution, whereas the quantitative way gives the exact quantity of solute in a known volume or mass of the solution, or a known mass of the solvent.

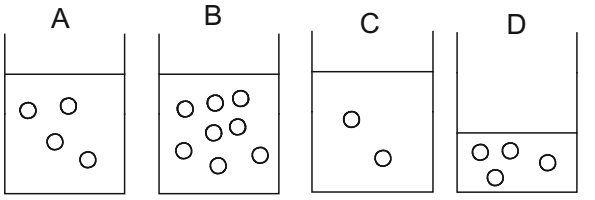
A. Qualitative Expressions

Solutions are described qualitatively as dilute or concentrated. **A dilute solution** is a solution in which the quantity of the solute is relatively small and a **concentrated solution** is a solution in which the quantity of the solute is relatively large. It is important to realize that dilute and concentrated are relative terms. A litre of solution that contains 5 g salt is concentrated when compared to a solution that contain 1g salt; but it in turn is dilute relative to a litre of solution that contains 15g of salt.

Exercise 3.2

Use the illustration below to answer the following questions. Assume that the open circles represent molecules of solute, and solutions A, B, and C have the same volume. Solution D has half the volume of A, B, and C.

1. Which of the solutions is relatively concentrated?
2. Which of the solutions is relatively dilute?
3. Which of the solutions have the same concentration?
4. How does the concentration of A compare to the concentrations of D, B, and C?



B. Quantitative Expressions

In scientific or technical applications, a qualitative account of concentration is not sufficient; therefore quantitative measures are needed to describe concentration.

When solutions are prepared for specific purposes, they usually, have a known concentration. For example, a patient who uses a cough syrup (a solution) will take in a known volume of the syrup of specific concentration for it to work effectively. You may have come across people who say that the problem with traditional medicines is that the healers do not know the correct quantity to give to the patient. These examples show the importance of describing concentrations quantitatively.

Over the years, various methods have been developed to calculate and quantitatively express the concentration of solutions. Some of the common and important methods used to express concentration are mass percent, volume percent, **molarity, molality** and **normality.**

1. Mass Percent of Solute

Let us discuss

Assume you dissolved:

a. 10 g of sugar in 90 g of water
b. 10 g of sugar in 40 g of water

1. What is the amount of sugar dissolved in 100 g of solution in a?
2. What is the amount of sugar dissolved in 100 g of solution in b?
3. Which of the two solutions is concentrated?

Mass percent (% m/m) is the mass of a solute present in 100 g of solution or the fraction of a solute in a solution multiplied by 100. Mathematically mass percent of a solute is expressed as:

$$mass\ percent\ of\ solute = \frac{mass\ of\ solute}{mass\ of\ solution} \times 100$$

Usually, mass percent is applied to describe a solution in which a solid is dissolved in a liquid solvent. The mass of the solution equals the mass of the solvent plus the mass of the solute. Therefore, the equation for mass percent can be expressed as:

Example 1

A solution is produced by dissolving 10 g of table salt (NaCl) in 50 g aqueous solution. Calculate the mass percent of NaCl?

Solution

Step1. Analyse the problem

Given: mass of the solute (NaCl) =10 g
mass of the solution = 50 g

Required: mass percent of NaCl

Step 2. In order to calculate mass percent of NaCl substitute the known values into the equation.

$$mass\ percent\ of\ solute = \frac{mass\ of\ solute}{mass\ of\ solution} \times 100$$

$$= \frac{10g \times 100}{50g} = 20\%Nacl$$

This means dissolving 10 g of NaCl in 50 g of a solution gives a 20% solution of NaCl.

Example 2

What mass of water is required to prepare a solution using 10 g of glucose that will produce a 1%?

Solution

Step1. Analyse the problem

Given: mass of the solute (glucose) = 10 g
mass percent of solute = 1%

Required: mass of the water

Step 2. In order to calculate the mass of the solvent, it is necessary to first calculate mass of the solution from the relationship

$$mass\ percent\ of\ solute = \frac{mass\ of\ solute}{mass\ of\ solution} \times 100$$

$$mass\ of\ solution = \frac{mass\ of\ solute}{mass\ percent\ of\ solute} \times 100$$

$$= \frac{10g}{1} \times 100 = 1000g$$

Step 3. The mass of the solution is the mass of the solute plus the mass of the solvent. Therefore,the mass of the solvent can be calculated as follow:

$$mass\ of\ solution = mass\ of\ solvent + mass\ of\ \ solute$$

$$mass\ of\ solvent = mass\ of\ solution - mass\ of\ solute$$

$$= 1000g - 10g = 990g$$

10 g of solute and 990 g of solvent produce 1000 g of solution. Therefore, dissolving 10 g of glucose in 990 g of water gives 1% (m/m) of glucose.

2. Volume Percent of Solute

Let us discuss

Look at the label on Asmara beer and discuss its alcohol composition.

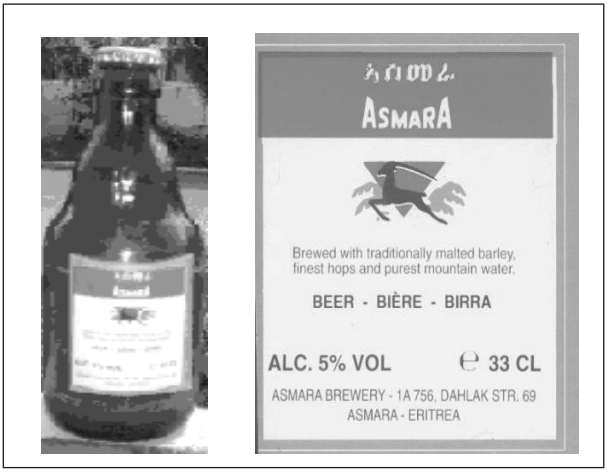


Fig 3.3

The percentage composition of solutions containing liquid solutes dissolved in solvents can be expressed as volume percent (% v/v). Such concentrations state the number of millilitres (mL) of solute dissolved in 100 mL of a solution. Mathematically the volume percent of a solute is expressed as:

$$volume\ percent\ of\ solute = \frac{volume\ of\ solute}{volume\ of\ solution} \times 100$$

Example

Rubbing alcohol is 70% isopropyl alcohol by volume. This solution is used to disinfect medical instruments and human skin or to prepare the skin for injection. What volume of the alcohol is present in 1 litre solution?

Solution

Step 1. Analyse the problem

Given: volume percent of the alcohol = 70 %
volume of the solution = 1 L = 1000 mL

Required: Volume of the alcohol

Step 2. The alcohol is the solute, and to calculate its volume the equation

$$volume\ percent\ of\ alcohol = \frac{volume\ of\ alcohol}{volume\ of\ solution} \times 100$$

$$70 = \frac{volume\ of\ alcohol \times 100}{1000mL}$$

$$volume\ of\ alcohol = \frac{70 \times 1000\ mL}{100} = 700\ mL$$

This means that in a 1 L solution of 70% isopropyl alcohol by volume contains 700 mL of alcohol and 300 mL of water.

Exercise 3.3

A 600 mL solution, A, is 5 % alcohol by volume. Another 100 mL solution, B, is 40 % alcohol by volume.

a. What is the volume of pure alcohol in each solution?
b. Which solution has greater amount of alcohol?

3. Molarity

Activity 3.5: Comparing the concentration of two solutions.

Materials required: Sugar, water, a 10 mL volumetric flask, a 5 mL volumetric flask, a scale, a wash bottle, a stirrer.

Preparation of Solution 1
Add 3.42 g of sugar into the 10 mL volumetric flask. Pour water in small amounts into the flask while shaking it until the sugar dissolves. Then pour more water into the flask up to the 10 mL mark. You now have 10 mL a solution.

Preparation of solution 2
Prepare a 5 mL solution of sugar in the 5 mL volumetric flask this time by dissolving 1.71 g of sugar.

Molar mass sugar (sucrose) = 342 g

1. Which of the two solutions is sweeter? Never taste the solution. Predict the sweetness of the solution from the quantity of the solute used and the volume of the solvent.
2. Convert the mass 3.42 g and 1.71 g into moles.
3. Calculate the moles of sugar per litre of solution.
4. Which of the two solutions is more concentrated in terms of its sugar content?

Molarity is the expression of concentration most frequently used in the chemical laboratory. It is based on the unit of chemical measurement, the mole. The **molarity** of a solution, symbolized by M, is defined as the number of moles of a solute dissolved in exactly one litre of a solution. Mathematically molarity is expressed as:

$$Molarity = \frac{no\ of\ moles\ of\ solute}{volume\ of\ solution\ in\ litres}$$

A solution that contains one mole of solute in one litre of solution has a molarity of one, or we say its concentration is one molar (1 M). The letter M stands for moles/L and represents the word molar.

Note that solutions that contain, for example, 2 moles of sugar in 2 L solution and 0.5 moles of sugar in 0.5 L of solution are both one molar solutions. You can now imagine that these two solutions have the same concentration, of sugar, and therefore they should be equally sweet.

The amount of solute dissolved in a solvent is generally given in units of gram rather than in moles. In this case, it is necessary to first change the mass in grams into the number of moles in order to calculate molarity. The number of moles is given as:

$$no\ of\ moles\ of\ solute = \frac{mass\ of\ solute}{molar\ mass\ of\ solute}$$

Consequently, the equation that expresses molarity will be rearranged as:

$$M = \frac{mass\ of\ solute}{molar\ mass} \times \frac{1}{volume\ of\ solution\ in\ litres}$$

Example 1

What is the molarity of a solution prepared by dissolving 1.5 mole of sodium chloride in enough water to prepare 500 mL of solution?

Solution

Step 1. Analyze the problem

Given: amount of the solute = 1.5 mole
volume of the solution = 500 mL

Required: molarity of the solution

Step 2. Convert the volume into litre

$$500mL \times \frac{lL}{1000mL} = 0.5L$$

Step 3. Calculate the molarity by substituting the given values into the equation

$$M = \frac{no\ of\ moles\ of\ NaCl}{volume\ of\ solution\ in\ litres} = \frac{1.5\ moles}{0.5L} = 3M$$

This means, dissolving 1.5 mole of solute into 0.5 L produces a 3 M solution.

Example 2

How many moles of HCl are neede to prepare 150 mL of a 1.2 M solution?

Solution

Step1. Analyze the problem

Given: molarity of the solution = 1.2 M
volume of the solution =150 mL = 0.15 L

Required: no of moles

Step 2. Calculate the number of moles by substituting the given values into the formula

$$M = \frac{no\ of\ moles}{volume\ of\ solution}$$

$$no\ of\ moles = M \times volume\ of\ solution$$

$$= 1.2M \times 0.15L$$

$$= 0.18\ moles$$

This means 0.18 moles of HCl are required to prepare a 150 mL solution with a concentration of 1.2 M.

Example 3

What is the molarity of a solution prepared by dissolving 11 g of $CaCl_2$ to make 100 mL of solution? Molar mass $CaCl_2$ = 111 g

Solution:

Step 1. Analyze the problem

Given: mass of the solute = 11g
volume of the solution = 100 mL

Required: molarity of the solution

Step 2. In calculating molarity, the volume of the solution must be in litres; therefore it is necessary to change the volume from mL to L.

$$100mL \times \frac{1L}{100mL} = 0.1L$$

Step 3. The mass of the solute in grams must be changed into units of a mole when calculating molarity. To find the number of moles divide the given mass by the molar mass.

$$No\ of\ moles = \frac{mass\ of\ solute}{molar\ mass} = \frac{11\,g}{111\,g} = 0.1\,mole$$

Step 4. Divide no of moles by volume to calculate the molarity.

OR

Molarity is given as:

$$M = \frac{mass\ of\ solute}{molar\ mass} \times \frac{1}{volume\ of\ solution\ litres}$$

Substitute the given values as follows:

$$M = \frac{11g}{111g} \times \frac{1}{0.1L} = 1M$$

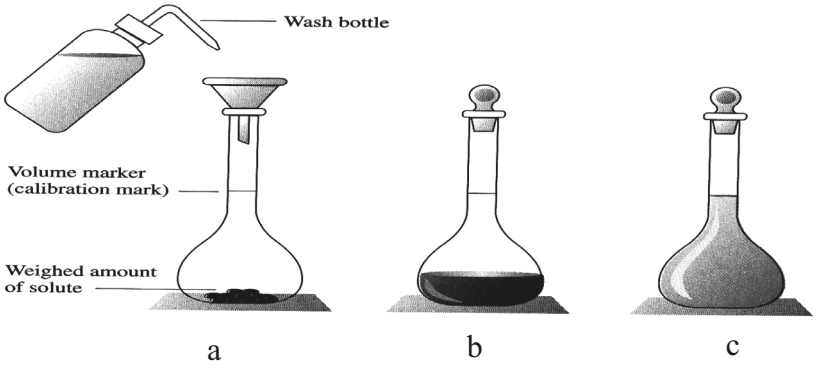
This means that dissolving 11g of $CaCl_2$ in 100 mL of solution gives a 1 M solution.

Activity 3.6: Preparing a solution with a known molarity

Materials required: Potassium hydroxide (KOH), water, a scale, a 100 mL volumetric flask.

Measure out 5.6 g of KOH and transfer it to the 100 mL volumetric flask.
Pour approximately 25 mL of water into the volumetric flask in order to dissolve the salt.
Pour in more water up to the 100 mL mark.

1. What is the number of moles in the solute?
2. What is the total volume of the solution?
3. Calculate the molarity of the solution?
4. Label the solution (write the formula of the solute and the concentration of the solution). Use a sticker or a piece of paper for labelling.



Preparing a liquid solution with a particular molarity. (a) add weighed amount of solute to a volumetric flask, (b) a small amount of water is added to dissolve the solid, (c) More water added until the volume of the solution is up to the mark.

Exercise 3.4

1. What total volume of a 0.5 M solution of sodium chloride in water can be prepared from 14.63 g of the salt?
   Molar mass of NaCl = 58.5 g

2. What is the mass of sugar that is dissolved in 100 mL of a 2 M solution? Molar mass of sugar (sucrose) = 342 g

3. How would you prepare 1.0 L of a 2.0 M solution of sodium chloride?

4. Molality

Up to this point, the concentrations of solutions have been expressed in terms of the quantity of the solute in a known volume or mass of solution. Molality is different in that it gives the quantity of the solute in a known mass of solvent. The molality of a solution is defined as the number of moles of solute dissolved in exactly one kilogram of solvent.

Mathematically, molality is expressed as:

$$Molality = \frac{no\ of\ moles\ of\ solute}{mass\ of\ solvent\ in\ kg}$$

Molality is represented using a small letter "m" which stands for the word molal.

Example

What is the molality of a solution of 10 g NaOH in 500 g water?
Molar mass NaOH = 40 g

Solution

Step 1. Analyze the problem

Given: amount of NaOH =10g
mass of water = 500g

Required: molality of the solution

Step 2. Calculate the number of moles of NaOH from the mass and molar mass.

$$\frac{mole\ of\ NaoH}{40g} \times 10g = 0.25mole$$

Step 3. In calculating molality, the mass of the solvent is given in kg unit. Therefore convert the mass into Kilograms.

$$500g \times \frac{1kg}{1000g} = 0.50kg$$

Step 4. Divide the mass by the number of moles to calculate the molality.

$$m = \frac{number\ of\ moles}{mass\ of\ solvent} = \frac{0.25\ mole}{0.5kg} = 0.5m$$

This means, dissolving 10 g (0.25 mole) of NaOH in 0. 50 kg of water gives 0.5 m solution.

Exercise 3.5

1. Sea water contains 2.7 g of NaCl per 100 g of water. What is the molality of NaCl in the ocean?
2. Calculate the molality of a solution consisting of 25 g of KCl in 250.0 mL of pure water at 20 ^{o}C?

Density of the water =1 g/mL

5. Normality

Let us discuss

1. What is the mass of:
   a) 1 mole of H_2SO_4? b) 0.5 mole of H_2SO_4?
2. What is the relationship between molar mass and a mole?

Normality is defined as the number of equivalents of a solute dissolved in a litre of solution. Normality is represented by the letter N, which stands for the word "normal". Mathematically, normality is expressed as:

$$Normality = \frac{no\ of\ equivalents}{volume\ of\ solution(L)}$$

It is necessary to define **equivalent mass** or **gram equivalent mass** and **equivalent** or **gram equivalent** to understand and make use of the concentration unit normality. Though the concept of gram equivalent is broad, in this unit, we will limit our discussion on equivalents to acids and bases.

The equivalent mass of an acid is defined as the molar mass of the acid divided by the number of replaceable hydrogen atoms in the formula,

$$Equivalent\ mass\ acid = \frac{molar\ mass\ of\ acid}{no\ of\ replaceable\ H^+}$$

Consider HCl and H_2SO_4 as an example. HCl has one replaceable hydrogen atom, whereas H_2SO_4 has two replaceable hydrogen atoms, and the molar mass of the two acids are 36.5 g and 98 g respectively. Therefore, the equivalent mass of HCl is 36.5 g and of H_2SO_4 is 49g.

$$Equivalent\ mass\ of\ HCl = \frac{36.5\ g}{1} = 36.5\ g$$

$$Equivalent\ mass\ of\ H_2SO_4 = \frac{98\ g}{2} = 49\ g$$

The number of equivalents of an acid is defined as the mass of the acid divided by its equivalent mass.

$$no\ of\ equivalents\ of\ acid = \frac{mass\ of\ acid}{equivalent\ mass\ of\ acid}$$

We know that the equivalent mass of HCl is 36.5 g. Therefore 36.5 g and 73 g of HCl correspond to 1 and 2 equivalents respectively.

$$\frac{36\ .5\ \text{g HCl}}{36.5\ \text{g/equivalent}} = 1\ \text{equivalent}$$

$$\frac{73\ \text{g HCl}}{36.5\ \text{g/equivalent}} = 2\ \text{equivalents}$$

Like wise, 98 g of H_2SO_4 corresponds to 2 equivalents.

$$\frac{98\ \text{g } H_2SO_4}{49\text{g/equivalent}} = 2\ \text{equivalents}$$

Table 3.3 shows the relationship between equivalent mass, molar mass, the number of replaceable hydrogen ions and the number of equivalents per mole of hydrochloric acid, sulphuric acid and phosphoric acid.

Acid	Molar mass, g	Number of replaceable H^+	Equivalent mass	Number of equivalents/ mole
HCl	36.5	1	36.5	1
H_2SO_4	98	2	98/2 = 49	2
H_3PO_4	98	3	98/3 = 32.3	3

Table 3.3: Relationship between molar mass, equivalent mass, number of replaceable hydrogen ions and the number of equivalents per mole of hydrochloricacid, sulphuric acid and phosphoric acid.

The equivalent mass of a base is defined as the molar mass of the base divided by the number of hydroxide ions in the formula.

$$Equivalent\ mass\ of\ base = \frac{molar\ mass\ of\ base}{no\ of\ OH^-\ in\ formula}$$

The molar mass of NaOH and of $Ba(OH)_2$ are 40g and 171 g. Therefore, their equivalent masses respectively are:

$$Equivalent\ mass\ of\ NaOH = \frac{40\ g}{1} = 40\ g$$

$$Equivalent\ mass\ of\ Ba(OH)_2 = \frac{171\ g}{2} = 85.5\ g$$

The number of equivalents of a base is defined in a similar way to that of acids. It is defined as the mass of a base divided by the equivalent mass of that base. For example, 171g of $Ba(OH)_2$ corresponds to 2 equivalents.

$$\frac{171\ g\ Ba(OH)_2}{85.5\ g/equivalent} = 2\ equivalents$$

Table 3.4 shows the relationship between molar mass, equivalent mass, the number of replaceable hydroxide ions and the equivalents per mole of potassium hydroxide, calcium hydroxide and aluminium hydroxide.

Base	**Molar mass,g**	**Number of replaceable OH^-**	**Equivalent mass**	**Number of equivalents/ mole**
KOH	56	1	56/1 = 56	1
$Ca(OH)_2$	74	2	74/2 = 37	2
$Al(OH)_3$	78	3	78/3 = 26	3

Table 3.4 Relationship between molar mass, equivalent mass, the number of replaceable hydroxide ions and number of equivalents per mole of potassium hydroxide, calcium hydroxide and aluminium hydroxide.

Example 1

Calculate:

a) The equivalent mass of H_3PO_4.
b) The number of equivalents in 50 grams of H_3PO_4.
c) The normality of a solution prepared by dissolving 50 g of H_3PO_4 in two litres of a solution.

Atomic mass of H = 1.0; P = 31; O = 16.0

Solution

Step1. Analyze the problem

Given: molar mass = 98 g mass of H_3PO_4 = 50 g
volume of solution = 2 L

Required: Equivalent mass; number of equivalents and normality.

Step 2. Calculate the required values by substituting the know into the equation.

$$a.\ equivalent\ mass = \frac{\text{molar mass}}{H^+ \text{ per mole}}$$

$$= \frac{98\ g}{3} = 32.7\ g$$

$$b.\ no\ of\ equivalents = \frac{mass\ of\ H_3PO_4}{equivalent\ mass}$$

$$= \frac{50\ g}{32.7\ g / equivalent} = 1.53$$

$$c.\ Normality = \frac{no\ of\ equivalents}{Volume\ of\ solution\ (L)}$$

$$= \frac{1.53}{2\ L} = 0.76\ N$$

50 g of H_3PO_4 is 1.53 equivalents. Dissolving 1.53 equivalents of an acid in 2 litres of a solution produces 0.76 normal solution.

Example 2

Calculate the normality of a solution prepared by dissolving 30 g of H_2SO_4 in 0.5 L of solution. Molar mass H_2SO_4 = 98 g

Solution

Step 1. Analyse the problem

Given: mass of H_2SO_4 = 30 g
volume of the solution = 0.5 L

Required: normality of the solution

Step 2. Calculate the number of equivalents of H_2SO_4 by dividing mass by equivalent mass

$$30g \times \frac{equivalent}{49g} = 0.61\ equivalent$$

Step 3. Calculate normality by substituting the number of equivalents and the volume of the solution into the equation.

$$Normality = \frac{0.61\ equivalent}{0.5\ L} = 1.2\ N$$

This means, dissolving 30 g of H_2SO_4 in 0.5L of a solution gives 1.2 N.

Relationship between normality and molarity

The number of equivalents of an acid or a base can be expressed in terms of the number of moles and the number of replaceable hydrogen atoms or hydroxide ions, if we equate the two formulas that describe equivalent mass.

$$\frac{malar\ mass}{number\ of\ replaceable\ H^+\ or\ OH^-} = \frac{mass}{number\ of\ equivalents} = equivalent\ mass$$

$$number\ of\ equivalents = \frac{mass}{molar\ mass} \times number\ of\ H^+\ or\ OH^-$$

$$number\ of\ equivalents = number\ of\ moles \times no\ of\ replaceble\ H^+\ or\ OH^-$$

Divide both sides of the equation by volume

$$\frac{number\ of\ equivalents}{volume(L)} = \frac{number\ of\ moles}{volume(L)} \times no\ of\ replaceable\ H^+ or\ OH^-$$

The above equation can be written as:

$$Normality = Molarity \times number\ of\ replaceable\ H^+ or\ OH^-$$

Example

What is the normal concentration of 2 M H_3PO_4?

Step 1. Analyze the problem

Given: molarity of H_3PO_4 = 2 M

Required: normality of H_3PO_4

Step 2. Calculate the normality using the relation

$$Normality = Molarity \times number\ of\ replaceable\ H^+ or\ OH^-$$

Normality = $2 \times 3 = 6$ N

Exercise 3.6

1. What is the molar concentration of 36 N H_2SO_4?
2. a) What is the equivalent mass of $Al(OH)_3$?
   b) What is the number of equivalents in one mole $Al(OH)_3$?

Molar mass $Al(OH)_3$ = 78 g

Did you know?

The carat is a unit of concentration that describes the amount of gold in an alloy. Pure gold is extremely soft; about 30 gram of gold can be stretched into a wire several kilometers long or it can be flattened into gold leaf with an area of 30 square metres. Because of this quality, jewellers combine gold with other metals to form alloys of greater durability. The gold content of an alloy is expressed in parts of 24. Pure gold is 24 carats, whereas 18 carat gold is 18/24 or 75% gold.

3.2.2 Diluting a solution

Activity 3.7: Diluting a solution



Solution 1: 10 mL of alcohol in a 50 mL solution

50 mL of water

Solution 2: 100 mL of a solution

1. What is the difference between the amounts of alcohol in solution 1 and solution 2? (What is the volume of alcohol present in solution 2?)
2. What is the volume percent concentration of solution 1?
3. What is the volume percent concentration of solution 2?
4. Which of the two solutions is relatively dilute?
5. Suppose you take a 5 mL sample from solution. What is its concentration relative to the total 50 mL of the solution?

Most laboratories and industries keep **stock solutions**. Stock solutions are solutions of high concentration frequently used after dilution. Very often, you will need to make a specific volume of known concentration from a small volume of stock solutions. This process is called **dilution**. Dilution is simply used to make a solution less concentrated. Dilution is not limited to use in laboratories and industries; you make use of dilution at home in several different ways. If your tea is too strong or your juice is too thick, you dilute

it by adding some water**.** The following points are important in understanding dilution.

1. A small volume of a solution taken from a stock solution has the same concentration as the original bulk solution.
2. The amount or the number of moles of solute in a concentrated solution does not change when the solution is diluted.

In order to prepare a dilute solution whose volume is $\mathbf{V_f}$ and concentration is $\mathbf{C_f}$ from a stock solution whose concentration is $\mathbf{C_i}$**,** we have to know the volume $\mathbf{V_i}$ that we have to take from the stock solution. In order to calculate $\mathbf{V_i}$**,** we use the following relationship.

$\mathbf{C_iV_i}$ = the amount of solute in the concentrated (stock) solution

$\mathbf{C_fV_f}$ =the amount of solute in the dilute solution

$\mathbf{C_i}$ and $\mathbf{V_i}$ refer to the initial concentration and the initial volume before dilution, and $\mathbf{C_f}$ and $\mathbf{V_f}$ refer to the final concentration and the final volume after dilution respectively.

Because the amount of solute does not change during dilution, the amount of solute in the concentrated solution is equal to the amount of solute in the dilute solution.

$$\mathbf{C_iV_i = C_fV_f}$$

All the three variables are known except $\mathbf{V_i}$. Therefore one can easily calculate $\mathbf{V_i}$ by substituting the known values. However, it is important to note that the units for concentration and volume should be the same on both sides of the equation and $\mathbf{C_i}$ and $\mathbf{C_f}$ represent molarity, normality and percent by volume.

Example

Suppose you have 3 mL of a stock solution of 100 mg/mL ampicillin and you want to make 0.2 mL of solution having 25 mg/ mL.

a). What volume of the stock solution is used to prepare the dilute ampicillin solution?

b). What volume of the solvent is needed to prepare the dilute solution?

Solution

Step 1. Analyze the problem

Given: C_i = 100 mg/ mL in the stock solution
V_f = total volume of the dilute solution = 0.2 mL
C_f = concentration of the dilute solution = 25 mg/ mL

Required:

V_i, volume from stock used to prepare dilute solution;
Volume of solvent

Step 2. Answer "a" by calculating V_i using the equation

$$C_iV_i = C_fV_f$$

$$V_i = \frac{V_f \ x \ C_f}{C_i} = \frac{0.2\,mL \ x \ 25\ mg/mL}{100\ mg/mL} = 0.05\ mL$$

V_i is the volume of solution taken from the stock in order to prepare the dilute solution,

Step 3. Answer "b" as follows: the total volume of the solution is equal to V_i plus the volume of the solvent used to prepare the dilute solution. Mathematically this is given as:

b. $V_f = V_i + V_{\text{added solvent}}$

$$V_{\text{added solvent}} = V_f - V_i$$
$$= 0.2\ mL - 0.05\ mL = 1.95\ mL$$

Therefore, a 0.2 mL solution of 25 mg/mL can be prepared by adding 1.95 mL of water to 0.05 mL of a 100 mg/mL solution.

Exercise 3.7

a) Calculate the volume of 1.0 M NaOH aqueous solution needed to make 100 mL of 0.5 M NaOH aqueous solution.
b) Calculate the volume of the solvent (mL) that must be added to make a total volume of 100 mL.

Summary

The concentration of a solution can be described either qualitatively or quantitatively. The qualitative method uses the terms "dilute" "and" concentrated" to describe the quantity of solute in a solution. A dilute solution is one that has a relatively low concentration of solute whereas a concentrated solution is one that has a relatively high concentration of solute.

The quantitative method gives the exact quantity of the solute in a known mass or volume of solution or solvent. The common ways of describing concentration are given in tabular form below:

Concentration unit	Formula
Mass percent (% m/m)	$\%m/m = \frac{Mass\ of\ solute}{Mass\ of\ solution} \times 100$
Volume percent (% v/v)	$\%v/v = \frac{volume\ of\ solute}{volume\ of\ solution} \times 100$
Molarity (M)	$M = \frac{no\ of\ moles\ of\ solute}{volume\ of\ solution(L)}$
Molality (m)	$m = \frac{no\ of\ moles\ of\ solute}{Mass\ of\ solvent\ (kg)}$
Normality (N)	$N = \frac{no\ of\ equivalents\ of\ solute}{volume\ of\ solution(L)}$

3.3 Properties of Solution

3.3.1 Dissolution of Solutes in Liquids

Activity 3.8 Investigate the dissolution of solutes in liquids

Materials required: Table salt, acetic acid, paraffin wax, cyclohexane, water, beaker (100 mL), a graduated cylinder, a stirrer.

Add 5 g salt to 50 mL of water and stir the mixture.
Add 5 mL acetic acid to 50 mL cyclohexane and stir the mixture.
Add 5 mL paraffin wax to 50 mL water and stir the mixture.
Add 5 mL paraffin wax to 50 mL cyclohexane and stir the mixture.

1. Copy the following table into your notebook and complete the blank spaces.

Solute	Polar, non-polar or ionic solute	Solvent	Polar, non-polar or ionic solvent	Does solute dissolve?
Salt		Water		
Acetic acid		Cyclohexane		
Paraffin wax		Water		
Paraffin wax		Cyclohexane		

2. What do you conclude about the solubility of the solutes in polar and nonpolar solvents?

Dissolution is the term used to describe the dissolving of a solute in a solvent. If a solvent dissolves a solute, the particles of the solvent separate the particles of the solute and occupy the spaces in-between them. This happens if a force of attraction that overcomes the attraction between solute-solute and solvent-solvent molecules subsequently develops between solute and solvent molecules in the solution.

Polar solvent molecules can effectively separate the molecules of other polar solutes because a reasonable force of attraction exists between the solute and the solvent molecules. Water and ethanol are examples of polar substances that readily dissolve in each other. There is H-bonding between the ethanol molecules, and of course between water molecules when they are in a pure form. When the two mix, H-bonding develops between the water and the ethanol molecules. Because the strength of the intermolecular forces in the solution is similar to that of the pure liquids, the solute dissolves in the solvent.

In addition to dissolving polar solutes, polar solvents can generally dissolve ionic compounds. Also in here, an ion-dipole force of attraction that develops between the ions of the solute and the molecules of the polar solvent is responsible for the dissolution.

Paraffin is soluble in cyclohexane. Paraffin molecules are nonpolar and the forces between the molecules are weak van der Waals forces. There are also weak van der Waals forces between the cyclohexane molecules. Because the intermolecular forces in both the solute and the solvent are similar, it is easy for the intermolecular forces in the solute to be broken, allowing the solute to move into the spaces between the molecules of the solvent. The solute dissolves in the solvent. This process is illustrated shown in Fig 3.3.

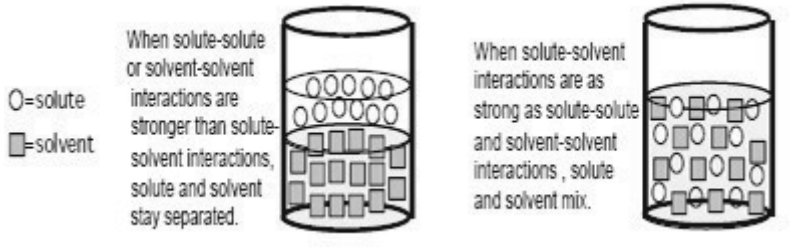


Fig 3.3: Separation of solute and solvent molecules in solutions.

The effect of the polarity of the molecules is also noticeable when gases dissolve in water. H_2 and O_2 are nonpolar, whereas NH_3 and HCl are polar. Owing to the facts discussed above, HCl and NH_3 are highly soluble in water whereas H_2 and O_2 are only slightly soluble in water. Table 3.5 shows the solubility of gases in 1.0 litre of water at 0 ^oC and 1 atmospheric pressure.

Gas	Polarity	No of moles of gas that dissolve in 1 L of H_2O	Grams of gas that dissolve in 1 L of H_2O
NH_3	polar	52.5	900
HCl	polar	22.6	824.9
H_2	nonpolar	9.6×10^{-4}	1.92×10^{-3}
N_2	nonpolar	1.05×10^{-3}	2.94×10^{-2}
O_2	nonpolar	2.18×10^{-3}	6.98×10^{-2}

Table 3.5: Solubility of certain common gases in water

Based on above facts, the popular statement that scientists use to predict solubility is **"like dissolves like"**. This statement indicates that a solute will dissolve best in a solvent that has similar properties. Therefore, polar solutes dissolve readily in polar solvents and nonpolar solutes dissolve readily in nonpolar solvents.

Exercise 3.8

Predict which of the following substances will be soluble/miscible in water. Give explanations for your answers.

Benzene, potassium chloride, sugar, ethanol, iron, sodium hydroxide, gold, oil, bromine.

3.3.2 Effect of Temperature and Pressure on Solubility

Effect of Temperature on Solubility

When a solid dissolves in a liquid, heat is required to break the forces holding the solid particles (molecules or ions) together. At the same time, heat is given off during the formation of the new solute-solvent forces of attraction. Therefore, the overall dissolution process can release or absorb heat depending on whether the heat required or the heat released is greater.

If the heat given off in the dissolution reaction is less than the heat required to break apart the solid, the net reaction is endothermic (heat required). In an endothermic dissolution process, an increase in temperature increases the solubility of a solute. Because

dissolving a solid in a liquid is usually an endothermic process, cooling a saturated solution usually causes a solid to crystallize out of a solution.

One example of such a situation is ammonium nitrate. When ammonium nitrate dissolves in water, the solution becomes very cold because the dissolving process is endothermic. Because cold things are nowadays used to relieve pain, it is used as first-aid instant cold packs in sport fields.

If the heat given off in the dissolution process is greater than the heat required to break the solid apart, the net process is exothermic (heat given off). In an exothermic reaction, an increase in temperature will decrease the solubility; this means that heating will drive a solute out of a saturated solution. Only few solids such as Na_2SO_4 show a decrease in solubility with an increase in temperature

Dissolving a gas in water is always an exothermic process. Therefore, gases become less soluble as temperature rises. This is shown by the downward trend in Figure 3.4. More gas is present in a solution with a lower temperature compared to a solution with a higher temperature.

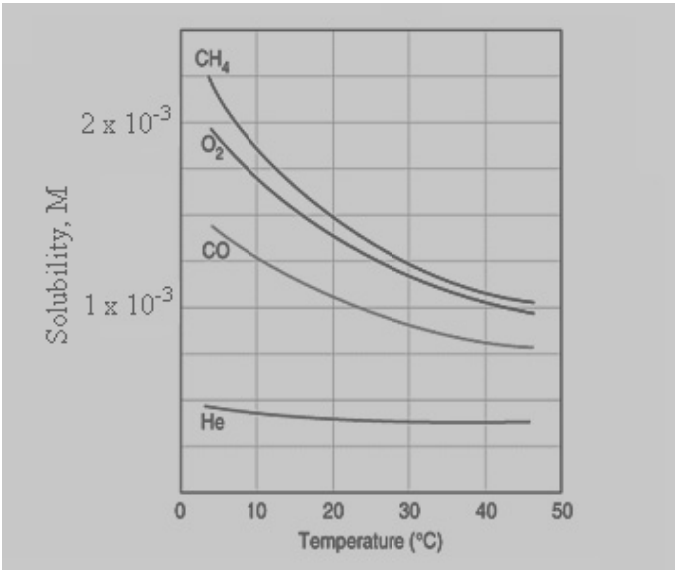


Fig 3.4: Solubility of gases vs temperature

The decrease in the solubility of gases with an increase in temperature is explained in terms of the increase in the kinetic energy of the molecules. The higher kinetic energy results in greater movement between the molecules, which causes the intermolecular bonds to break and escape from the solution

A inverse relationship between the solubility of a gas and temperature exists in our day-to-day activities. The gas in warm carbonated water is fizzier than in cold carbonated water. Boiled water tastes flat because the dissolved oxygen and carbon dioxide gas molecules have been removed when the water is heated.

Exercise 3.9

Refer to Fig 3.4 to answer the following questions.

1. What is the approximate solubility of O_2 at 40 °C?
2. What is the approximate solubility of CH_4 at 20 °C?
3. Which of the four gases shows the greatest decrease in solubility between 4 °C and 10 °C?

Effect of Pressure on Solubility

Liquids and solids exhibit practically no change in solubility with changes in pressure while gases do show an increase in solubility with an increase in pressure. As shown in Fig 3.5, an increase in pressure increases the contact between a gas and a solvent. This results in a greater number of gas molecules entering the liquid and therefore greater solubility.

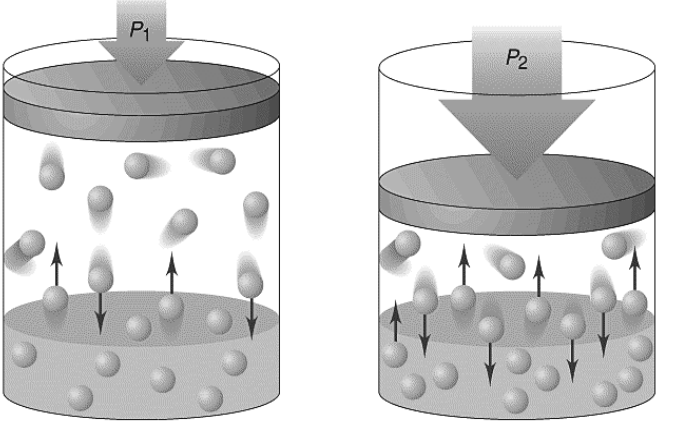


Fig. 3.5: Solubility of a gas: (A) at low pressure and (B) higher pressure

The effect of pressure on gas solubility is used in the preparation of carbonated beverages. All carbonated beverages such as soft drinks are bottled under pressure to increase the quantity of carbon dioxide dissolved in the solution. When the bottle is opened, the pressure above the solution decreases and as a result, the gas begins to bubble out of the solution. This is more evident when the beverage is poured into a glass, as shown in Fig 3.6.



Fig 3.6: The effervescence of carbon dioxide from a soft drink

3.3.3 Colligative Properties of Solutions

Let us discuss

You know that a sugar solution and a salt solution have different tastes, and that both solutions are colourless. You may also know that the colour of a copper sulphate solution is blue.

Which of these properties (taste and/or colour) depend on the identity but not quantity of the solutes?

Solutions have different properties than their solvents. The properties of solutions can be divided into two main groups: **colligative** and **non-colligative** properties. Non-colligative are those that depend on the identity of the dissolved solute and the solvent. Properties such as colour, taste and solubility are examples of non colligative properties.

All physical properties of a solution that depend on the number of solute particles dissolved in the solution are called colligative properties. These include the freezing point, vapour pressure, the boiling point, and osmotic pressure.

1. Vapour Pressure Lowering

Let us discuss

1. Why do we dry our wet washing in the sun?
2. Over a period of time, will there be more vapourization at 25 ^{o}C or 50 ^{o}C ?

Vaporization is a process in which a liquid is converted to a vapour. The molecules escape from the liquid state and change to the gaseous state. Consider vaporization (evaporation) of a liquid in an open container. When a liquid evaporates from an open container, the molecules are steadily leave the liquid and escape into the air. The liquid eventually disappears. See Fig 3.7A.

If a liquid is placed in a closed container, the molecules escape into the confined space, called the vapour space, above the liquid. However, there is a limit to the number of solvent molecules in the vapour space, and no extra molecules can be accommodated once the vapour space contains the maximum number of solvent molecules. The pressure that this maximum number of molecules can exert on the liquid is known as the vapour pressure of the liquid. Note that the level of the liquid decreases during the initial stage and then becomes constant once the vapour space is saturated with the vapour molecules. This is illustrated in Fig. 3.7 B.

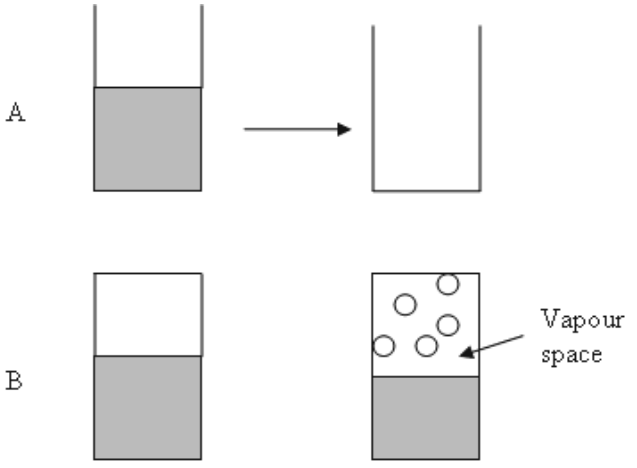


Fig 3.7: Vaporization of solvent molecules from: (A) an open container and (B).a closed container

Now consider a pure solvent, and a solution of the same solvent and, a solid. For the same vapour space, the number of solvent molecules in the vapour state of a pure solvent is more than the number of solvent molecules in the vapour state of a solution, as shown in Fig 3.8. Because vapour pressure depends on the number of molecules, we say that the solution has a lower vapour pressure than the pure solvent.

The number of solvent molecules in the vapour space depends on the number of solute particles dissolved in the solvent; but it is independent of the identity the solute. Two different types of solute that have the same number of particles in a solution lower the vapour pressure by the same amount. Therefore, vapour presure lowering is a colligative property.

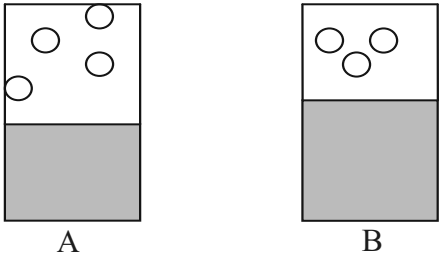


Fig 3.8: Vapourization in closed containers of:- (A) a pure solvent and (B) a solution.

2. Boiling Point Elevation

Let us discuss

You commonly use the term boiling. You say water boils at 100 ^{o}C. But what does boiling mean?

Boiling is a phenomenon of formation of bubbles inside a liquid that occurs when the vapour pressure of the liquid equals the atmospheric pressure. The temperature at which boiling occurs is called boiling point.

Experiments show that a solution of a non-volatile (generally solid) solute boils at a higher temperature than that of a pure solvent. The difference in temperature between the boiling point of the solution and the boiling point of the solvent is called **boiling point elevation**

(ΔT_b). For example, pure water boils at 100 °C whereas a 1 molal aqueous solution of sugar boils at 100.52 °C. Therefore, the boiling point elevation for 1 molal aqueous solution of sugar is:

ΔT_b = boiling point of solution – boiling point of solvent
= 100.52 °C – 100 °C = 0.52 °C

Boiling point elevation, like vapour pressure lowering is a colligative property. It is directly proportional to solute concentration expressed as molality. The mathematical expression that relates boiling point elevation to molality is given as:

$$\Delta T_b = K_b m$$

The proportionality constant K_b is the molal boiling point constant and is a characteristic property of a solvent; that is different solvents have different K_b values. Table 3.6 gives the K_b values for certain common solvents.

Solvent	Molal boiling point constant, K_b (°C/m)
Water	0.52
Ethanol	1.22
Benzene	2.53
Acetic acid	2.93

Table 3.6: Molal boiling point constant (K_b) values of certain solvents

Boiling point elevation is a result of vapour pressure lowering. When the temperature of a solution that contains a non-volatile solute is raised to a temperature of the boiling point of a pure solvent, the vapour pressure of the solution is still less than the atmospheric pressure. Therefore a solution cannot boil at a temperature where pure liquid boils.

Exercise 3.10

1. Arrange the following aqueous solutions in order of increasing boiling point.
   a) 0.2 m glucose, b) 0.5 m sucrose, c) 0.3 m glycerin, d) 0.5 m ethylene glycol and e) 0.6 m glucose solutions.

2. Calculate the boiling point elevation of a 0.4 m aqueous solution of glucose. The K_b for water is 0.52 $^oC/m$.

3. Freezing Point depression

When a solute is dissolved in a solvent, the freezing point temperature of the solution will be lower than the freezing point of the pure liquid. The lowering of the freezing point caused by the dissolved substance is called **freezing point lowering or freezing point depression (ΔT_f)**. For example, pure water freezes at 0 ^oC and a 1 molal aqueous solution of sugar freezes at -1.86 ^oC. Therefore, the freezing point lowering for 1 molal aqueous solution is given as:

ΔT_f = freezing point of the solvent – freezing point of the solution

= 0 ^oC – (-1.86 ^oC)

= 1.86 ^oC

The relationship between the lowering of the freezing point and the molal concentration of a solution is mathematically expressed as:

$$\Delta T_f = K_f m$$

$\mathbf{K_f}$ is the proportionality constant called the molal freezing point constant. $\mathbf{K_f}$ is equal to the depression in the freezing point when 1 mol of a solute is dissolved in 1 kg of solvent. $\mathbf{K_f}$ values for certain solvents are given Table 3.7..

Solvent	K_f ($^oC/m$)
Water	1.86
Ethanol	1.99
Benzene	5.12
Acetic acid	3.98

Table 3.7: K_f values for certain solvents

Example

What is the freezing point of a 0.4 m aqueous solution of glucose?

Solution

Step 1. Analyze the problem

Given: concentration of solution = 0.4 m; $K_f = 1.86$
Required: the freezing point

Step 2. Calculate the freezing point depression

$\Delta T_f = K_f m$

$\Delta T_f = 1.86\ ^oC/m \times 0.4\ m = 0.74\ ^oC$

Step 3. Find the difference in the freezing point between the pure solvent and the solution.

ΔT_f = freezing point of solvent – freezing point of solution

$0\ ^oC - 0.74\ ^oC = -0.74\ ^oC$

Therefore the freezing point of a 0.4 m aqueous solution of glucose is -0.74 ^oC.

Exercise 3.11

Calculate:

a) the boiling point and
b) the freezing point of a solution containing 12 g of naphthalene, $C_{10}H_8$, in 50 g of benzene.

The molar mass naphthalene = 128 g

K_b (benzene) = 2.53 $^oC/m$ K_f (benzene) = 5.12 $^oC/m$

bp of benzene = 80 ^oC, fp of benzne = 5.5 ^oC

4. Osmotic pressure

Osmosis is the net flow of a solvent from a pure solvent or a dilute solution into a more concentrated solution through a semi-permeable membrane ("permeable" means "allowing liquids or gases to pass through"). A semi-permeable membrane is a thin film that allows solvent molecules to pass through it but that restricts the passage of solute molecules or ions. Cellophane and parchment are examples of semi-permeable membranes.

In Fig 3.9, a U-tube is separated by a semipermeable membrane. The left-hand the right-hand sides of the membrane are partially

filled with concentrated and dilute solutions of equal volumes respectively. Over time, the volume of the concentrated solution increases and causes the liquid to rise. This is because the rate of passage of the solvent molecules through the membrane into the concentrated solution is greater than the rate of passage in the opposite direction, because the solvent concentration at the surface of the membrane is greater in the dilute solution.

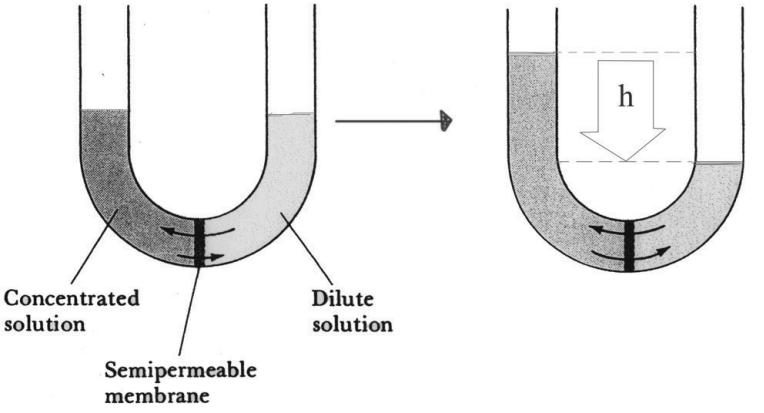


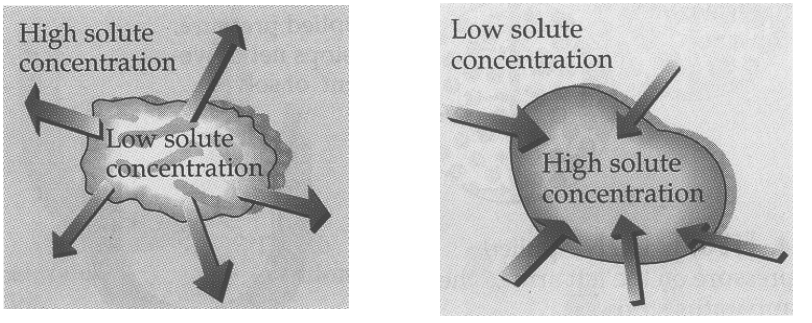
Fig 3.9:- Osmosis and osmotic pressure

The column of the liquid that rises above the surface of the dilute solution exerts a pressure. This pressure tends to increase the rate of passage of the solvent from the concentrated solution into the dilute solution, causing the net rate of osmosis to slow down as the liquid level climbs higher and higher. This continues until eventually the osmosis stops. The pressure on the concentrated solution that prevents the net flow of the solvent to the solution is called **osmotic pressure.** Osmotic pressure is proportional to the height of the liquid, h, shown in Fig 3.9.

The osmotic pressure, like other colligative properties, depends on the number of particles in the solution but it is independent of the identity of the solute.

Osmosis plays an important role in living systems and in our day-to day activities. The principle of osmosis is used in determining the concentration of intravenous fluids administered to patients through their veins, the preservation of food, the growth of plants, and other areas.

People needing nutrients but who cannot be fed orally are given solutions via intravenous, infusion which feeds nutrients directly into the veins. To prevent a net flow of water from or to the cell, the infusion is prepared in such a way that it has the same osmotic pressure as the cell so that no osmosis occurs. A net flow of water through the cell membrane which is semi-permeable will either swell or shrink the cell, depending on the direction of the flow, as shown in Fig 3.10. The net flow affects the proper functioning of the cell.



(a) (b)

Fig 3.10: Osmosis through a semi-permeable membrane of the cell: (a) water flow out of the cell; (b) water flow into the cell.

Did you know?

How antifreeze works in a car?

One of the most common applications of freezing point depression and boiling point elevation is the use of antifreeze solutions in the radiator of a motorcar. The solute is usually ethelyene glycol and the solvent is water. When dissolved in water, etheylene glycol both lowers the freezing point and raises the boiling point of water. In cold weather, it prevents the liquid in the radiator from freezing, and in hot weather it protects the liquid in the radiator from boiling.

Why does salt work as a preservative?

The major effect of salt as a preservative is that it draws water from micro-organisms by a process of osmosis if the external salt concentration is high enough. This causes the micro-organisms to dehydrate, which eventually kills them.

Summary

A general rule in determining the ability of a solvent to dissolve a specific solute is "like dissolves like". That is, polar and ionic compounds dissolve in polar solvents; nonpolar solvents dissolve in nonpolar solvents.

Two other factors that affect the solubility of a solute are temperature and pressure. For most solid solutes, their solubility in liquid solvents increases with an increase in temperature, some show negligible changes and a few show a decrease in solubility as the temperature increases.

The properties of solutions that depend on the number of solute particles, but not the type, in a given amount of solvent are called colligative properties. These properties are freezing point depression, vapour pressure lowering, boiling point elevation and osmotic pressure.

Key Terms

Alloy
Aqueous solution
Boiling point elevation
Colligative property
Concentrated solution
Dilute solution
Freezing point depression
Gaseous solution
Heterogeneous mixture
Homogeneous mixture
Immiscible
Liquid solution
Mass percent
Miscible
Mixture
Molal freezing point constant
Molal boiling point constant
Molality
Molarity
Normality
Osmotic pressure
Saturated solution
Solubility
Solubility curve
Solute
Solution
Solvent
Unsaturated solution
Supersaturated solution
Vapour pressure lowering
Volume percent

Review Questions

Part I. Multiple Choice Questions

Answer questions 1-3 after referring to Fig. 3.2 (solubility curve).

1. Which saturated solution at 10 ^{o}C do you label as dilute?

   a) potassium chlorate b) sodium nitrate
   c) potassium nitrate d) ammonia

2. What is the solubility of sodium nitrate at 80 ^{o}C?

   a) 147 g b) 120 g c) 110 g d) 25 g

3. With an increase in temperature, the solubility of NH_3

   a) increases b) decreases
   c) remains the same d) none of the above

4. When a crystal of solute is added to a solution it dissolves. The solution is:

   a) saturated b) unsaturated
   c) supersaturated d) none of the above

5. When a crystal of solute is added to a solution it increases in size. The solution is:

   a) unsaturated b) saturated
   c) supersaturated d) none of the above

6. The most common type of solution is:
   a) liquid b) solid c) gas d) plasma

7. Kibrom noticed a grease stain on his jacket. Which solvent should he use to remove the stain?

   a) water b) ethanol c) benzene d) HCl

8. Which of the following is an application of a colligative property?

   a) adding table salt to water when a sauce is prepared
   b) adding sugar to a glass of juice

c) spreading calcium chloride on an icy road
d) none of the above

9. Which of the following statements is true?

a) A saturated solution is always a concentrated solution
b) The solubility of solid solutes in water always increases as the temperature increases.
c) The density of a solution is a colligative property
d) Sea water has a higher boiling point than fresh water

10. If 50 mL of a 2 M solution is diluted in 150 mL, the concentration of the dilute solution is

a)half of the concentration of the original solution
b) a third of the concentration of the original solution
c) three times the concentration of the original solution
d)1 M

Part II. Work out and Short Answer Questions

1. What is meant by "like dissolves like"?
2. Why do oil and water not mix?
3. What would you do in order to dissolve a given quantity of large potassium nitrate crystals as quickly as possible?
4. Calculate the percentage concentration, v/v, of a solution that contains 12.0 mL of alcohol dissolved in 50 mL of solution.
5. What mass of KCl is present in 1 kg of a solution that is 5% m/m?
6. Calculate the molarity of each of the following solutions:

a. A solution containing 0.36 mole of Na_2SO_4 dissolved in 1.5 L of solution
b. A solution prepared by dissolving 10 g of sodium hydroxide in enough water to produce 200 mL of solution

7. A solution is prepared by mixing 40 of gram methanol (Molar mass 32) with 60 g of water. What is the molality of the solution?

8. What volume, in litres, of a 2.0 M KCl solution contains 2.5 g of KCl (molar mass = 74.5 g)?

9. How many equivalents are there in 0.2 mole of $Ba(OH)_2$ (molar mass = 171)?

10. What is the normality of 22 g of $Ba(OH)_2$ in 800 mL of a solution?

11. How much 2.0 M of a NaCl solution would you need to make 250 mL of 0.15 M NaCl solution?

12. What would the concentration of a solution made by diluting 45.0 mL of 4.2 M of KOH to 250mL be?

13. Calculate the boiling point of a solution in which 5 g of solute (non-volatile and non-electrolyte) is added to 100 g of water. (assume molar mass of the solute = 101- g)

14. What is the freezing point of a solution that contains 20 g of a solute (nonvolatile and non-electrolyte) dissolved in 250 g of water? (assume molar mass of the solute = 46)

UNIT 4
Acids, Bases and Salts

Expected Learning Outcomes

By the end of this unit, you are expected to:

- classify acids, bases and salts into different classes and explain the application of acids, bases and salts in daily life
- perform an experiment to demonstrate the general methods of preparation and the general properties of acids, bases and salts
- explain a neutralization reaction and use the pH scale to measure the acidity and alkalinity of a solution.

4.1 Acids

4.1.1 Describing Acids

Whenever we drink lemon juice without sugar, it tastes sour. Similarly, vinegar has a sour taste and it is used as a flavouring agent for salads. The sour taste of lemon and vinegar is due to the presence of different types of acids. The same holds true for the sour taste of common fruits and vegetables. Acids are very important in our daily lives.

Activity 4.1: Classification of different household substances using litmus paper

Materials required**:** Lemon juice, bleach, vinegar, milk, sodium chloride, calcium hydroxide (lime water), water, a solution of wood ash, blue and red litmus, test tubes.

Label eight 10 mL test tubes, each with a sample name as listed in the table below. In each of the test tubes prepare a water solution of each of the substances.

Copy the table below in to your notebooks and fill in the blank spaces with your data.

1. Using blue litmus paper, test the solutions and record your results in the table.
2. Again test each of the solutions using red litmus paper and record your results.

Name of the sample	**Colour with blue litmus paper**	**Colour with red litmus paper**
Lemon juice		
Bleach		
Vinegar		
Milk		
Sodium chloride		
Calcium hydroxide		
Water		
Wood ash		

3. Which of the solutions change colour when using the blue litmus paper, and which change colour with the red litmus paper?
4. What do you understand from the experiment?

Caution: Do not taste any chemicals in the laboratory!

The word "acid" is derived from the Latin word **"acidus"** which means **"sour"**. Acids can be defined in many ways. Two common ways of doing so are: the Arrhenius theory and the BrØnsted-Lowry theory.

1. **The Arrhenius theory**: According to Arrhenius, an acid is a compound that produces a hydrogen ion (H^+) on ionization in an aqueous solution. Hydrochloric acid, HCl, is an example of an acid. HCl breaks down into H^+ and Cl^- ions in an aqueous solution. Similarly, nitric acid, HNO_3 and sulphuric acid, H_2SO_4 break down to produce H^+ and NO_3^- and $2H^+$ and SO_4^{-2} ions, respectively.

$HCl\ (aq)$	$\longrightarrow$	$H^+\ (aq)$	+	$Cl^-(aq)$
Hydrochloric acid		Hydrogen ion		Chloride ion
$HNO_3\ (aq)$	$\longrightarrow$	$H^+\ (aq)$	+	$NO_3^-(aq)$
Nitric acid		Hydrogen ion		Nitrate ion
$H_2SO_4(aq)$	$\longrightarrow$	$2H^+(aq)$	+	$SO_4^{2-}(aq)$
Sulphuric acid		Hydrogen ion		Sulphate ion

2. **BrØnsted-Lowry theory:** According to BrØnsted and Lowry, an acid is a substance that donates a hydrogen ion (also called a proton) during a chemical reaction. For example, when hydrogen chloride, gas, HCl, dissolves in water, it donates a hydrogen ion to the water molecules. In this reaction, since hydrochloric acid is a **proton donor**, it is a BrØnsted acid.

$HCl\ (g)$	+	$H_2O\ (l)$	$\rightarrow$	$H_3O^+\ (aq)$	+	$Cl^-(aq)$
acid		**base**		**hydronium ion**		**chloride ion**

4.1.2 Classification of Acids

Some acids ionize more easily than others in aqueous solution. Depending on the degree to which acids are dissociated in water to produce hydrogen ions, they can be divided into two categories: **strong acids** and **weak acids**.

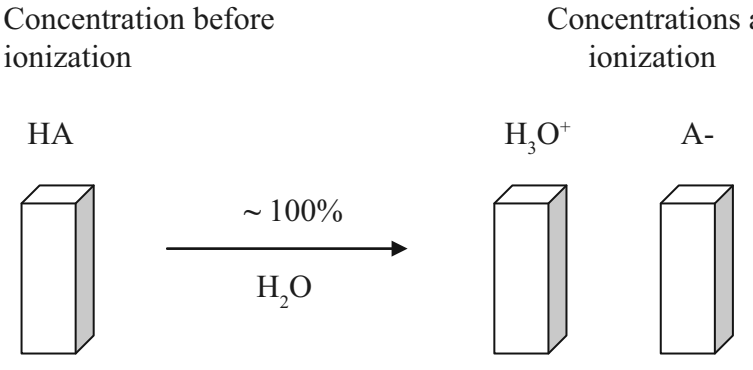
1. **A strong acid** is completely dissociated or ionized in a water solution. For example, hydrochloric acid is a common strong acid. When 1000 hydrogen chloride, HCl, gas molecules are dissolved in water, 1000 hydrogen ions and 1000 chloride ions will form. This shows that hydrochloric acid dissociates completely in water solution.

$$HCl\,(aq) \longrightarrow H^{+}(aq) + Cl^{-}(aq)$$

2. **A weak acid** is slightly dissociated in a water solution. It contains only a small percentage of hydronium ions. Acetic acid is an example of a weak acid. Only 1.3% of the molecules of acetic acid will be ionized in aqueous solution. If 1000 acetic acid molecules are dissolved in water, then, only 13 out of them will ionize into ions. Thus the resulting solution will have 987 molecules of acetic acid, 13 acetate ions and 13 hydronium ions.

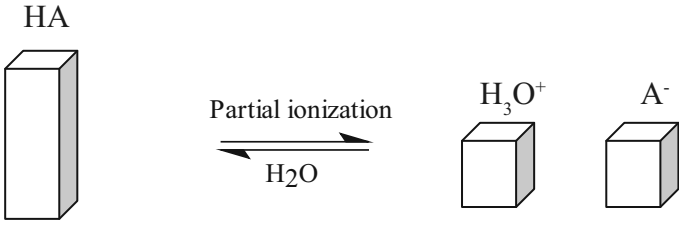
$$\underset{\text{987 molecules of acetic acid}}{H_3CCOOH\,(aq)} + H_2O(aq) \rightleftharpoons \underset{\text{13 hydronium ions}}{H_3O^{+}(aq)} + \underset{\text{13 acetate ions}}{H_3CCOO^{-}(aq)}$$

Fig 4.1 shows the degree of dissociation of a strong acid and a weak acid. A solution of a weak acid contains a higher concentration of un-ionized acid molecules, in addition to a small number of hydronium ions, and the corresponding negative ions. Suppose, HA is a hypothetical representation of an acid.



a) strong acid

Concentration before ionization | Concentrations after ionization



b) weak acid

Fig. 4.1 Ionization of strong and weak acids.

Activity 4.2: To test the reaction of acid solutions of various substances with a universal indicator

Materials required: 2M HCl (hydrochloric acid), $HC_2H_3O_2$ (acetic acid), 2M HNO_3 (nitric acid), lemon juice, a universal indicator, a dropper, 5mL test tubes.

Label the test tubes A1 to A4 such that A1 corresponds to HCl and so on.
Add five drops of each acid solution into the labelled test tubes.
Test each solution with a universal indicator.
1. Copy the following table into your notebooks and record the colour of each acid solution.

Test tube	**Solution**	**Colour of solution after adding an indicator**
A-1	2M HCl	
A-2	$HC_2H_3O_2$	
A-3	HNO_3	
A-4	Lemon juice	

2. Identify which of the acid solutions have almost similar colour after testing with the indicator?

As explained before, when an acid dissolves in water, a hydrogen ion (H^+) is transferred from the acid to the water molecule to form a hydronium ion. Any hydrogen that can be transferred to water is called an acidic hydrogen. Since a hydrogen ion contains only a proton, the prefixes **-protic** are used to indicate the number of acidic hydrogens that are present in an acid. On the basis of the number of acidic hydrogens which are formed from one molecule of the acid, acids may be classified as either **monoprotic (also called monobasic)**, or **polyprotic (polybasic)** acids.

Monoprotic acids: These acids contain only one acidic hydrogen. They produce only one H^+ per molecule of the acid when dissolved in water. Hydrochloric acid is an example of a monoprotic acid.

$$HCl\,(aq) \longrightarrow H^+\,(aq) + Cl^-\,(aq)$$

Polyprotic acids: All acids that contain more than one acidic hydrogen per molecule of the acid are called polyprotic acids. **Diprotic acids** are polyprotic acids that contain two acidic hydrogens, whereas those which contain three acidic hydrogens are called **triprotic** acids. Sulphuric acid, and phosphoric acid are diprotic and triprotic acids respectively.

$$H_2SO_4\,(aq) \longrightarrow 2H^+\,(aq) + SO_4^{2-}\,(aq)$$

$$H_3PO_4\,(aq) \longrightarrow 3H^+\,(aq) + PO_4^{3-}\,(aq)$$

Most acids are oxyacids in which the acidic hydrogen is attached to an oxygen atom. The strong acids mentioned above, except hydrochloric acid, are typical examples.

Exercise 4.1

Identify each of the following acids as monoprotic, diprotic or triprotic.

a) Perchloric acid ($HClO_4$)
b) Phosphoric acid (H_3PO_4)
c) Hydrofluoric acid (HF)
d) Acetic acid ($HC_2H_3O_2$)
e) Carbonic acid (H_2CO_3)

Did you know?

Certain coloured dyes change colour in the presence of an acid or a base. Dyes such as red litmus, and blue litmus are often used to indicate whether a substance is an acid or a base. Likewise, substances such as cabbage juice, and the juice of red rose petals change colour in the presence of an acid or a base. These dyes can also be used to test the acidity or basicity of a substance such as litmus.

4.1.3 General Method of Preparation of Acids

Activity 4.3 To demonstrate the preparation of an acid

Materials required**:** Sodium chloride, (NaCl), concentrated sulphuric acid ($H_2SO_{4,}$) water, a round-bottom flask, a thistle funnel, a Bunsen burner, a delivery tube, a beaker, a funnel.

Set up the apparatus for the experiment as shown in Fig 4.2.
Pour water into the second container.
Add sodium chloride to the round-bottom flask fitted with a thistle funnel and a delivery-tube.
Then add the concentrated sulphuric acid to the flask.

1. What do you observe?
   Test the water in the container with a universal indicator before the reaction starts and record your result.
   Then gently warm the salt and acid mixture.
2. What do you observe?
   At the end of the reaction test the water in the container again with a universal indicator.
3. What do you think has happened in this reaction?

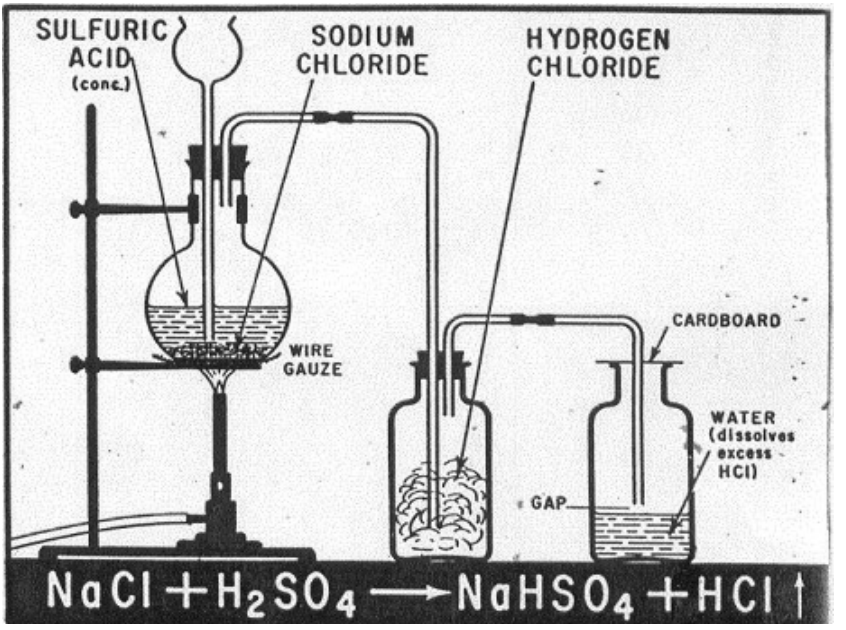


Fig. 4.2. Preparation of an acid

The following are a few general methods used in the preparation of acids.

1. **Reaction of water with nonmetallic oxides**

Nonmetallic oxides are acidic oxides which can react with water to form an acid. A nonmetallic oxide is also called an acid anhydride ("anhydride" means “without water”).

Nonmetallic oxide + water → acid

For example, carbon dioxide and sulphur dioxide react with water to form carbonic acid and sulphurous acid respectively:-

$$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(aq)$$

Carbonic acid

$$SO_2(g) + H_2O(l) \longrightarrow H_2SO_3(aq)$$

Sulphurous acid

2. **The direct combination of elements (synthesis)**

Only binary acids can be prepared using this method. For example, hydrochloric acid can be prepared by combining hydrogen gas (H_2) with chlorine gas (Cl_2) and then dissolving the product in water:

$$H_2\,(g) + Cl_2\,(g) \longrightarrow 2HCl\,(g)$$

Hydrogen chloride gas

$$HCl\,(g) + H_2O\,(l) \longrightarrow HCl\,(aq) + H_2O\,(l)$$

Hydrochloric acid

3. **Reaction of a salt of an acid with concentrated sulphuric acid**

This is the most common method of preparing acids. Sulphuric acid is non-volatile, thus the product can be collected by distillation. In this reaction the sulphuric acid produces the required hydrogen ions. The anion of the salt reacts with the hydrogen ion from the sulphuric acid to form a volatile acid. When metallic nitrates such as $NaNO_3$, KNO_3 and $Mg(NO_3)_2$ react with sulphuric acid, nitric acid is produced. For example:

$$NaNO_3\,(s) + H_2SO_4\,(aq) \longrightarrow NaHSO_4\,(s) + HNO_3\,(g)$$

Sodium nitrate (salt of nitric acid)

nitric acid

Likewise, hydrochloric acid forms when metallic chlorides such as NaCl, and KCl react with sulphuric acid. For example:

$$NaCl\,(s) + H_2SO_4\,(aq) \longrightarrow Na_2SO_4\,(s) + HCl\,(aq)$$

4.1.4 General Properties of Acids

Activity 4.4: To test some properties of hydrochloric acid

Materials required: A dropper, hydrochloric acid, 2M HCl, blue litmus, a piece of zinc, calcium carbonate, copper, and four 10 mL test tubes.

Note: Wear eye goggles.
Using a dropper, add 10 drops of 2M HCl to each of four test tubes labeled A1-A4.
Insert the litmus paper into test tube A1.

Add calcium carbonate to test tube A2. Likewise, add pieces of zinc to test tube A3 and copper wire to test tube A4.

1. Copy the table below and describe your observation for each reaction.
2. Which of the substances reacted with HCl?
3. Can you guess what the products are for the reactions that took place in test tubes A2, A3, A4?

Test tube	Reaction	Observation
A1	Blue litmus + HCl	
A2	Calcium carbonate + HCl	
A3	Zinc + HCl	
A4	Copper + HCl	

Caution – You will need to be very careful when handling acids
At the end of the experiment dispose of all the materials as instructed by the teacher.

The following is a list of some of the general properties of acids.

1. **Taste**

Aqueous solutions of acids have a sour taste. For example, lemon, vinegar, and the strong acids, hydrochloric acid, HCl, nitric acid, HNO_3 and suphuric acid, H_2SO_4 taste sour. **Caution: Do not taste any acid or chemical in the laboratory.** Many acids are poisonous and it is dangerous to taste them.

2. **Action on litmus**

Acids turn blue litmus to red.

3. **Corrosive Property**

Most acids are corrosive; they damage clothes and human flesh. The acids, hydrochloric acid, HCl, nitric acid, HNO_3 and sulphuric acid, H_2SO_4 are corrosive.

4. **Acids neutralize bases**

When acids react with bases, salt and water are formed. This reaction is **neutralization reaction.** During neutralization reaction, the products are neither acidic nor basic.

Acid + Base ⟶ Salt + Water

$$HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$$

5. **Action on metals**

One of the characteristic properties of acids is to react with active metals. For example, dilute hydrochloric acid reacts with sodium, magnesium, iron, zinc or aluminum to give salt and hydrogen gas. For example, the reaction between zinc and hydrochloric acid, produces zinc chloride and hydrogen gas. On the other hand, gold, platinum, silver and copper are inactive metals and do not react with aqueous solutions of acids.

$$Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)$$

6. **Action on metal carbonates and bicarbonates**

Most carbonates and bicarbonates react with acids to produce carbon dioxide gas. For example, when sodium carbonate reacts with hydrochloric acid, carbon dioxide, water and sodium chloride are formed.

Acid + Carbonate ⟶ Gas + Water + Salt

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$$

4.1.5 Uses of Common Acids

1. Source of vitamin C

Acids are found in many common fruits. Citrus fruits such as lemons and oranges contain citric acid. Citrus fruits are good sources of vitamin C.

2. Helps during digestion

After we have eaten, the food needs to be digested. Hydrochloric acid is present in the gastric juice of the stomach. It is important for the digestion of the food we eat.

3. Useful in industries

Acids are also important industrial chemicals. Sulphuric acid is used to manufacture car batteries, fertilizers, explosives, and other chemicals. Hydrochloric acid is used to clean metals, to prepare sugar from starch, and make metal chlorides. Similarly, acetic acid is used in vinegar, and in the dyeing of silk and wool.

Exercise 4.2

1. Give an example of an oxide of sulphur that reacts with water to form an acid.
2. Write an equation for the ionization of an acid, HA (HA is a hypothetical acid), in water.
3. Write an equation for the preparation of hydrochloric acid from barium chloride.
4. Which of the following metals will react in an aqueous solution of hydrochloric acid? Explain your answer.
   Ca , Zn , Au , Cu

Summary

Acids are very important compounds that are present in many house hold products such as vinegar and lemon juice.

According to the Arrhenius theory, an acid is defined as a compound that produces a hydrogen ion (H+) in an aqueous solution. However, according to the Brønsted and Lowry theory, an acid is a substance that donates a proton during a chemical reaction.

Acids are classified into strong and weak acids on the basis of the degree to which they are dissociated in water. Acids are also classified into monoprotic, diprotic, and polyprotic acids on the basis of the number of hydrogen ions, H+, which are formed from one molecule of the acid.

The general methods of preparing acids are: the reaction of water with the oxide of a nonmetal, the direct combination of elements, and the reaction of a salt of an acid with concentrated sulphuric acid. Some of the general properties of acids are: acids have a sour taste, they turn blue litmus paper to red, they are corrosive, they react with metals to produce hydrogen gas, they neutralize bases, and they react with metal carbonates to produce carbon dioxide (CO_2).

Acids have very important applications. Many fruits and vegetables contain acids. Gastric juices, which are present in the stomach are useful in the digestion of food. In industry, acids are used to manufacture fertilizers, explosives, dyes, and other important chemicals).

4.2 Bases

4.2.1 Describing Bases

Bases are chemically opposite to acids. This is because the reaction between an acid and a base produces a neutral substance that has the properties of neither an acid nor a base.

The following two definitions of a base were given by Arrhenius, and Bronsted and Lowry theories.

The Arrhenius Theory- states that a base is a substance that produces a hydroxide ion, OH^- when dissolved in water. For example, NaOH is basic, because, it produces Na^+ and OH^- ions in water.

$$NaOH\,(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

$$KOH\,(s) \xrightarrow{H_2O} K^+(aq) + OH^-(aq)$$

$$Mg(OH)_2(s) \xrightarrow{H_2O} Mg^{2+}(aq) + 2OH^-(aq)$$

The Arrhenius theory has some limitations, because it fails to explain why compounds such as NH_3 are basic in aqueous solution, though it does not have a hydroxide (OH) group.

The Bronsted-Lowry theory - states that a base is defined as a substance that can accept a proton from any other substance:

$$HBr\,(aq) + NaOH(aq) \longrightarrow H_2O(aq) + NaBr\,(aq)$$

HBr donates a proton to NaOH to produce water molecules and a hydronium ion. Therefore, the donor, HBr is considered an acid whereas the NaOH, the proton acceptor, is the base in this reaction. The Brøonsted–Lowry theory explains how non-hydroxide substances, when added to water, produce basic solutions. This has

resolved the shortcomings of the Arrehnius theory in explaining about the basic behaviuor of ammonia in an aqueous solution;

$$\underset{\text{Acceptor}}{NH_3} + \underset{\text{Donor}}{HOH} \rightleftharpoons NH_4^+ + OH^-$$

4.2.2 Classification of Bases

Depending on the degree to which they are either ionized or dissociated in water to produce hydroxide ions, bases are divided into two classes: **strong bases** and **weak bases**.

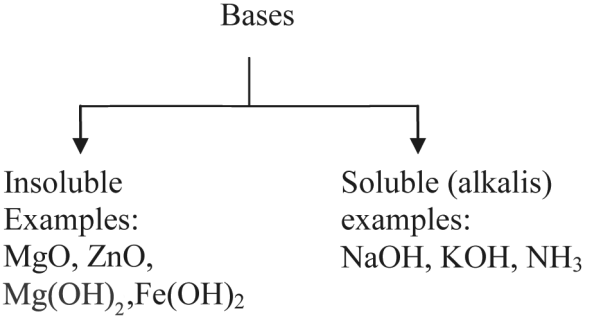
1. A **strong base** - is completely dissociated in a water solution. Sodium hydroxide (NaOH) is an example of a strong base. Sodium hydroxide, NaOH dissociates into separate sodium and hydroxide ions in an aqueous solution:-

$$NaOH\,(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

2. A **weak base** – partially dissociates into ions, when dissolved in water. Weak bases yield low concentrations of hydroxyl ions. For example, ammonia is a weak base. Only, 0.5% of the ammonia reacts with water, in an aqueous solution of 1M ammonia;

$$NH_3 + H_2o\,(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Bases are also classified as **insoluble bases** and **soluble bases**. Soluble bases are also called **alkalis**.



4.2.3 General Methods of Preparation of Bases

1. Reaction of very active metals with water

Active metals, sodium, potassium and calcium react with water to form metal hydroxides. An example of this is the reaction between potassium and water to produce potassium hydroxide and hydrogen:-

$$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$$

2. Reaction of a metal oxide (basic anhydride) with water.

Metallic oxides, referred to as basic oxides, react with water, to form a base. Potassium oxide (K_2O), calcium oxide (CaO) and sodium oxide (Na_2O) are examples of basic oxides. They produce the corresponding hydroxides when dissolved in water.

Metal oxide + water ⟶ metal hydroxide

$$CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$$

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

3. Combination of a salt and a base.

During the reaction of a salt and a base, we can prepare a soluble base such as sodium hydroxide or an insoluble base, such as ferric hydroxide.

$$Na_2CO_3(aq) + Ca(OH)_2(aq) \longrightarrow 2NaOH(aq) + CaCO_3(s)$$

$$FeCl_3(aq) + 3NaOH(aq) \longrightarrow Fe(OH)_3(s) + 3NaCl(aq)$$

4.2.4 General Properties of Bases

The following are some of the properties of bases.

1. Aqueous solutions of bases have a bitter taste and feel slippery. For example, NaOH on skin, has a slippery feel, because it is a base. (**Do not taste any bases**)
2. Bases turn red litmus paper blue.

3. Bases neutralize acids.

Base + Acid ⟶ Salt + water

Activity 4.5: Studying some of the properties of bases

Materials required: Soap, a very dilute sodium hydroxide solution, red and blue litmus paper, a beaker, test tubes,water a stirrer.

Prepare a soap solution by placing a small piece of soap into a beaker containing water and stirring it. Place the soap solution into one test tube and the sodium hydroxide solution into the other.

Feel the soap solution and the sodium hydroxide solutions with your hands.

1. What do you feel?

Test each solution using both red and blue litmus paper.

2. What do you observe after you test the solutions with both the red and the blue litmus papers?

Caution: Do not taste a base

4.2.5 Uses of Common Bases

The following table provides examples of the uses for some common bases.

Base	Formula	Important Uses
Ammonium hydroxide	NH_4OH	House hold cleaners, in the manufacture of fertilizers such as urea.
Aluminum hydroxide	$Al(OH)_3$	To purify water and to fix dyes in clothes.
Calcium hydroxide	$Ca(OH)_2$	Soil treatment, softening of water and for construction.
Magnesium hydroxide	$Mg(OH)_2$	Sold as a tablet or as a suspension “milk of magnesia” as an antacid.
Sodium hydroxide	NaOH	Used in the manufacture of soap. It is the most common laboratory base.

Table 4.1 : Uses of some Common bases

Exercise 4.3

1. What is the common factor for the bases in the following equations?

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

$$KOH(s) \xrightarrow{H_2O} K^+(aq) + OH^-(aq)$$

$$Mg(OH)_2(s) \xrightarrow{H_2O} Mg^{2+}(aq) + 2OH^-(aq)$$

2. Complete the following table by filling in the blank spaces using the information provided.

Name of a substance	Alkaline or insoluble base	Strong or weak base	Colour of red litmus paper	Ions present in solution
Calcium hydroxide				
Ammonia				
Lithium hydroxide				
Potassium hydroxide				

Summary

According to the Arrhenius, theory, a base is a substance that produces hydroxide ions, OH^-, in an aqueous solution. The Bronsted-Lowry definition states that a base is a proton acceptor during a chemical reaction.

Bases are classified into strong and weak bases depending on the extent to which they are ionized in water. Another important division divides bases into soluble bases (alkalis) and insoluble bases. The three general methods of preparing bases are the reaction of a metallic oxide with water, the combination of a salt and a base and the reaction of water with a very active metal.

Four general properties of bases can be mentioned. Bases have a bitter taste and feel slippery, they turn red litmus to blue, they neutralize bases,and their aqueous solutions conduct electricity.

Bases have important applications in the manufacture of soap, and are useful in industry. Certain bases are also useful as medicines such as antacids.

4.3 Describing a Neutralization Reaction

Activity 4.6: To demonstrate acid-base reactions

Materials required**:** 1M NaOH, 1M HCl, blue litmus and red litmus paper, three 100 mL beakers, stirring rod, a measuring cylinder.

Pour 5 mL of HCl into a100 mL beaker and test the solution with both blue and red litmus.
Then pour 5 mL of NaOH into the second beaker and test with both blue and red litmus.
In to the third beaker pour 5 mL of HCl and 5 mL of NaOH and mix the resulting solution with a stirring rod.
Test the mixture in the beaker with a blue and red litmus as you did before.

1. Copy the table below in your notebook and record your results.

Solution	**Test with blue litmus**	**Test with red litmus**
5 mL HCl		
5 mL NaOH		
5 mL HCl + 5 mL NaOH		

2. What colour change did you observe when you tested each of the solutions with red and blue litmus papers?
3. Determine whether, HCl, NaOH and the reaction mixture are acidic, basic or neutral. Explain your answer.

When an acid reacts with a base all the properties of both the acid and the base disappear or are neutralized. The reaction between an acid and a base to form salt and water is called **neutralization.**

Acid + Base ⟶ Salt + Water

For example, HCl reacts with NaOH, to produce NaCl and H_2O :

$$HCl(aq) + NaOH(aq) \rightarrow NaCl\,(aq) + H_2O(l)$$

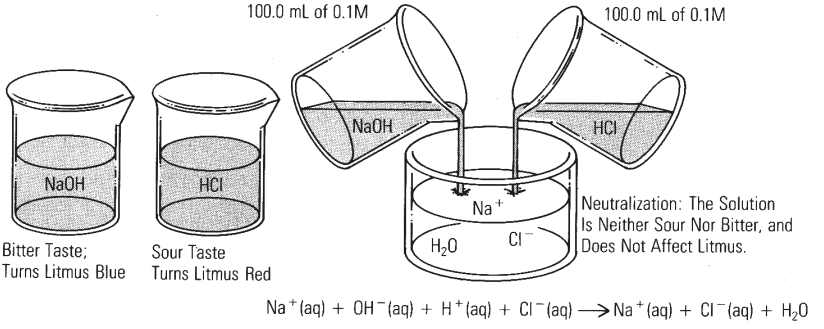


Fig 4.2 Neutralization reaction between sodium hydroxide and hydrochloric acid

A neutralization reaction plays important roles in the lives of human beings, some of which are:

a. To treat excess acid in the stomach

The HCl which is present in the stomach is useful for the digestion of food. But when HCl in the stomach is present in excessive amounts we get heart burn. Antacids are taken in order to treat the stomach. They counteract the acid by neutralizing it. Some of these antacids are shown in Fig. 4.3.



Fig 4. 3 Some common antacids

b. To fight tooth decay

When eating, some of the food remains in between the teeth. Bacteria in the mouth can change the sugars from the food into acids. The acid produced can attack the teeth and can cause tooth decay. Tooth paste is a weak alkali and fights tooth decay by neutralizing the acid that is produced by the bacteria in the mouth. This is why we need to brush our teeth daily.

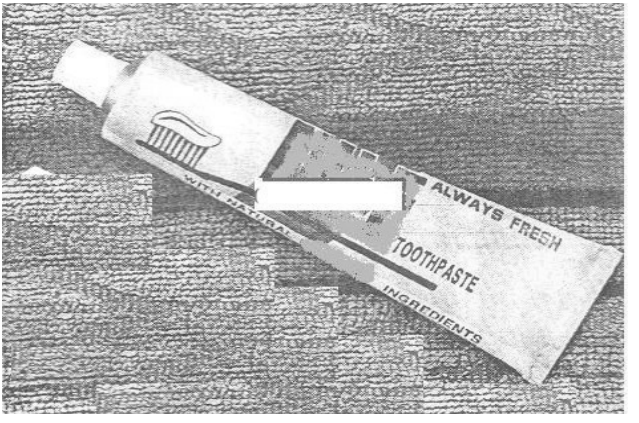


Fig 4.4 Toothpaste,

c. To reduce the acidity or alkalinity of the soil

Certain plants grow well in acidic soil, whereas others grow better in an alkaline or neutral soil. So, whenever, the soil for growing crops becomes acidic, the farmers need to add lime (a base) to the

soil in order to neutralize the acidity of the soil. On the other hand, when the soil is basic, the farmers need to add ammonium sulphate. This neutralization reaction plays an important role in increasing the yield of crops.

d. To treat insect bites

Neutralization reactions can be used to treat insect bites. We feel pain if we are stung by a bee. This is because the bee injects an acidic liquid into our skin. Applying calamine lotion onto the sting will neutralize the acid because the lotion contains a weak base. On the other hand, stings injected by wasps are alkaline, and such stings are neutralized by applying vinegar (a solution of acetic acid) to the site of the sting.

4.3.1 Titration

In a reaction between an acid and a base, the hydrogen ion (H^+) from the acid and the hydroxide ion from the base combine to form water. Since water has no basic or acidic properties, the resulting solution is neutral to litmus. The concentration of an acid or a base can be determined by titration. **Titration** is a process of adding a solution of known concentration (the standard solution or titrant) that is required to react completely with a known volume of a reagent of unknown concentration.

If the concentration of either an acid or a base is unknown, then an acid–base titration (neutralization) is carried out. In the process of titration, a measured volume of the unknown acid or base solution is transferred to a flask and the appropriate indicator is added. If the unknown solution is an acid, then the titrant is the base; and if the unknown solution is a base, the titrant is an acid.

In the titration process, the titrant is added to the unknown solution using a burette, until the indicator shows slight changes in colour. This shows that the neutralization reaction has been complete. This is called the **end point**. The end point of a neutralization reaction is called the **equivalence point**. This is the point at which all the hydroxide ions are neutralized by the hydrogen ions of the titrant, or all the hydrogen ions are neutralized by the hydroxide ions. At this point we stop the titration.

Figure 4.5 summarizes the procedure used in the titration of hydrochloric acid against 0.1 M sodium hydroxide solution.

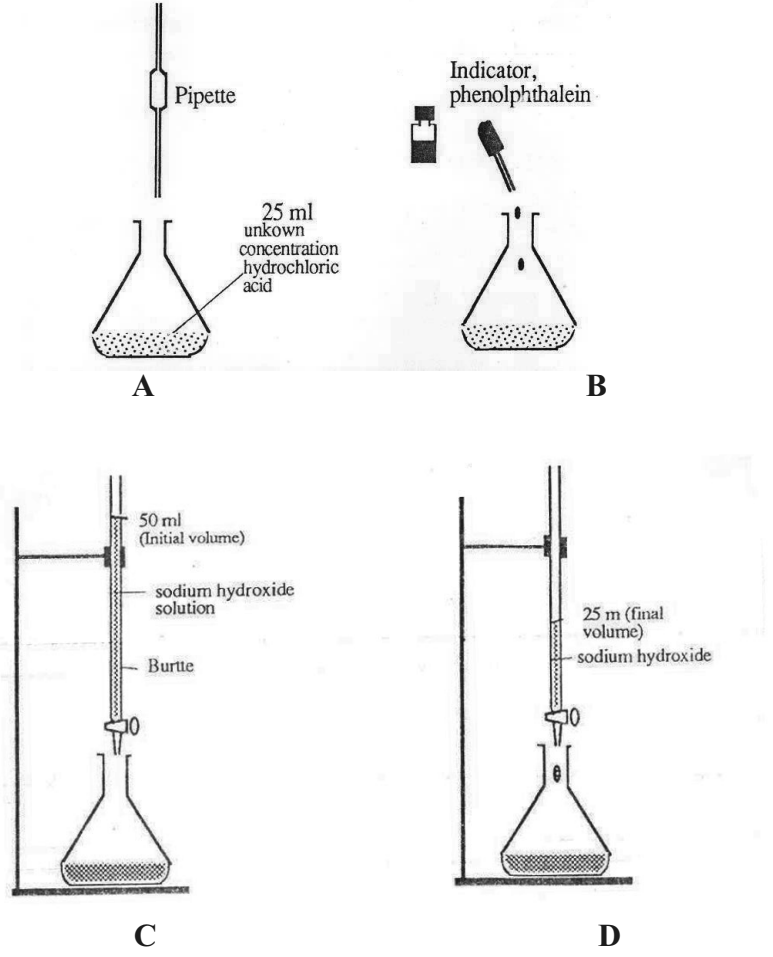


Fig 4. 5 Summary of the procedure followed in titration

4.3.2 Determining Concentration Using Titration

In order to determine the concentration of an unknown solution of hydrochloric acid, HCl, you will need to perform an HCl-NaOH titration and obtain data. If normality (N) is used as a concentration unit in acid-base titrations, the number of equivalents can be used to determine the concentration or volume of an unknown base or acid. At equivalence point (at neutralization point) the number of equivalents of an acid and base are equal.

Number of equivalents of an acid = Number of equivalents of a base

The number of equivalents of a substance is equal to normality (N) multiplied by the volume (V) of a solution in litres; therefore we can express the number of equivalents of an acid or a base in terms of normality multiplied by the volume of a solution in litres.

Number of equivalents of an acid $= N_{acid} \times V_{acid}$

Number of equivalents of a base $= N_{base} \times V_{base}$

At equivalence point,

$$\mathbf{N_{acid} \ x \ V_{acid} = N_{base} \ x \ V_{base}}$$

Example 1

25 mL of 0.1N standard NaOH solution is needed to neutralize a 5 mL solution of H_2SO_4. Calculate the concentration of H_2SO_4 in normality.

Solution

Given: $N_{NaOH} = 0.1N$ $V_{NaOH} = 25$ mL

$V_{H2SO4} = 5$ mL

Required : $N_{H_2SO_4}$

Analyzing:

Step 1: When complete neutralization of the H_2SO_4 by NaOH, takes place, we use the equation

$$N_{H_2SO_4} \text{ x } V_{H_2SO_4} = N_{NaOH} \text{ x } V_{NaOH}$$

Step 2: By rearranging the above equation we can calculate the concetration of H_2SO_4 as follows;

$$N_{H_2SO_4} = \frac{N_{NaOH} \times V_{NaOH}}{V_{H_2SO_4}}$$

$$= \frac{0.1N \times 25 \text{ mL}}{5 \text{ mL}}$$

$$= \frac{2.5 \text{ N}}{5}$$

$$= \mathbf{0.5 \text{ N}}$$

Therefore, the concentration of sulphuric acid is 0.5 N

Example 2

How many milliliters of 0.125 M of H_2SO_4 are required to neutralize 45.0 mL of 0.340 M KOH solution?

Solution:

Given: $M_{KOH} = 0.34$ M $V_{KOH} = 45.0$ mL
$M_{H_2SO_4} = 0.125$ M

Required : $V_{H_2SO_4}$

Analyzing:

Step 1: When complete neutralization of the H_2SO_4 by KOH, takes place, we use the equation

$$N_{H_2SO_4} \times V_{H_2SO_4} = N_{KOH} \times V_{KOH}$$

Step 2 : Rearranging the above equation we can calculate the concentration of H_2SO_4 as follows;

$$V_{H_2SO_4} = \frac{N_{KOH} \times V_{KOH}}{N_{H_2SO_4}}$$

Therefore, in order to solve this problem, the concentration of sulphuric acid must be converted to normality.

N = no of eqt/mole $\times$ M

$N_{H_2SO_4}$ = 2 eqt/mole $\times$ 0.125 M = 0.250 N

N_{KOH} = 1eqt/mole $\times$ 0.34 M = 0.340 N

Substituting the given quantities into the equation, the volume of H_2SO_4 is.

$$= \frac{0.340\ N \times 45\ mL}{0.250\ N}$$

$$= 61.2\ mL\ H_2SO_4$$

Activity 4.7: Titration of an unknown concentration of hydrochloric acid with a standard solution of sodium hydroxide

Materials required: 0.1M of sodium hydroxide, hydrochloric acid, a burette, an Erlenmeyer flask, phenolphthalein indicator, a burette stand and a pipette.

First rinse out the burette with water and then with 0.1 of M NaOH solution.

Clamp the burette vertically on the stand

Fill the burette with 0.1 M of NaOH solution below the zero mark and record the volume to at least 0.05 mL. (Your teacher will show you how to read a burette)

Using a pipette transfer 25 mL of the unknown concentration of HCl into a flask and add two drops of phenolphthalein indicator. Slowly pour the NaOH solution into the acid solution, Gently shake until a pink colour appears.

Then add the NaOH drop by drop until the pink colour persists even after shaking. At this stage the end point has been reached. Read the burette again and record the volume of the sodium hydroxide used by calculating the difference between the intial and final burette readings.

Repeat the titration two more times and record the average of the volume of the NaOH used in the three titrations.

1.Copy the table below into your note book and complete the blank spaces with experimental data.

Titration	Volume of HCl	Burette reading		Volume of 0.1M NaOH used
		First	Second	
First				
Second				
Third				
Average volume of NaOH used				

2. Calculate the concentration of hydrochloric acid from the titration data.

4.3.3 Using Indicators in a Titration

Indicators are acids or bases that change colour by losing or gaining hydrogen ions. An acid-base indicator demonstrates the end point of a titration. For example, in the titration of sodium hydroxide, NaOH, against hydrochloric acid, HCl, the end point of a titration is reached when the resulting solution is neutral to litmus.

The most commonly used indicators in titration involving acid-base reactions are methyl orange, litmus and phenolphthalein. Table 4.2 shows the colour change of the three indicators in an acidic or basic medium.

Type of Indicator	Colour in Acidic solution	Colour in Basic solution
Methyl orange	Pink	Yellow
Litmus	Red	Blue
Phenolphthalein	Colourless	Pink

Table 4. 2 Colour changes of common indicators

Exercise 4.4

1. Fill in the products of the following equations. Write also a word equation for each Reaction.
   a) $KOH(aq) + HNO_3(aq)$
   b) $Ba(OH)_2(aq) + HCl(aq)$
   c) $NaOH(aq) + H_3PO_4(aq)$
   d) $Ca(OH)_2(aq) + H_3PO_4(aq)$
2. Which acid-base reactions produce each of the following products?
   a) NaCl b) $MgCl_2$ c) $CaSO_4$ d) KBr
3. Why is a neutralization reaction important?
4. 50.0 mL of 0.1N of NaOH solution is needed to titrate 20.0 mL of HCl. Determine the normality of the HCl solution.

Summary

The reaction between an acid and a base to form a salt and water is called neutralization. The reaction between HCl and NaOH to produce NaCl and water is an example of a neutralization reaction.

Neutralization reactions also play an important role in our daily lives. Antacids work by neutralizing of the excess acids in the stomach. Tooth pastes are used to neutralize acids in the mouth and this prevents tooth decay.

Titration is a process of adding a standard solution (titrant) of an accurately known concentration to determine the concentration of an unknown solution. The concentration of an acid or a base may be determined by titration. In an acid-base titration (neutralization), a measured volume of a solution of a base (or an acid) of unknown concentration is transferred to a flask and the appropriate indicator is added. The titrant (an acid or a base) is added to the unknown solution until the indicator just changes colour slightly. This is called the end point of the titration. The equivalence point is the exact point at which all the hydrogen ions are neutralized by the hydroxide ions.

4.4 Salts

Let us discuss

Write the chemical name for table salt. What is the source of the table salt that is manufactured in Eritrea? Can you mention a few common uses of table salt?

4.4.1 Describing Salts

We have already seen that a salt is formed in the reaction of a base and an acid during a neutralization reaction. A salt is a compound which consists of positive metallic ions derived from the base, and negative ions which are derived from the acid. For example, in the preparation of NaCl from NaOH and HCl, the Na^+ ion comes from the base (NaOH), whereas the Cl^- ion is derived from the acid (HCl).

Exercise 4.5

The following table shows a list of the acids, bases and their corresponding salts. Copy and complete the table for the compounds **A**,**B**,**C**,and **D** that are missing from the table.

Acid	Base	Salt
HCl	**B**	NaCl
HCl	KOH	**C**
A	NH_3	NH_4NO_3
H_2SO_4	$Mg(OH)_2$	**D**

4.4. 2 Classification of Salts

There are two types of salts: normal salts and acid salts.

1. Normal salts are formed when all the ionizable hydrogen ions of an acid are replaced by other cations (a metal or an ammonium group). Examples are, sodium chloride, NaCl, sodium nitrate, $NaNO_3$, potassium sulphate, K_2SO_4, and magnesium suphate, Mg_2SO_4. Nitric acid has only one replaceable hydrogen ion, therefore, it is monobasic and it can form only a normal salt, that is $NaNO_3$.

$$HNO_3\,(aq) + NaOH\,(aq) \longrightarrow NaNO_3\,(aq) + H_2O\,(l)$$

2. Acid salts (or Acidic salts) are formed when an acid has more than one ionizable hydrogen (polybasic acid). Therefore, some of the salts that are formed may retain one or more hydrogen atoms from the parent acid. In such cases, only part of the acidic hydrogen is replaced by a metal. These salts are called acid salts. Sodium hydrogen sulphate, ($NaHSO_4$,) and potassium hydrogen carbonate, ($KHCO_3$), are examples.

Since sulphuric acid has two replaceable hydrogens (a dibasic or a diprotic acid), it can form two types of salts, namely, an acidic salt and a normal salt.

$$H_2SO_4\,(aq) + NaOH\,(aq) \longrightarrow \underset{\text{Acid salt}}{NaHSO_4\,(aq)} + H_2O\,(l)$$

$$H_2SO_4\,(aq) + 2NaOH\,(aq) \longrightarrow \underset{\text{Normal salt}}{Na_2SO_4\,(aq)} + H_2O\,(l)$$

Table 4.3 illustrates the different types of salts of different acids.

Acid	Acid salt	Normal salt
H_2SO_4	$KHSO_4$	K_2SO_4
HCl	---	KCl
H_2CO_3	$NaHCO_3$	Na_2CO_3
HNO_3	---	$NaNO_3$

Table 4.3 Different types of salts of different acids

4.4.3 General Properties of Salts

1. Salts are the neutralization products of acid-base reactions. Salts consist of the positive ion of the base and the negative ion of the acid.

2. Generally salts are solids at room temperature, have a high melting point, and have good conductivity in an aqueous solution and in their molten state.

3. Salts are differentiated according to their solubility in aqueous solutions. Some salts are soluble in water, whereas others are insoluble in water. Table 4.4 summarizes the solubility of salts in water.

Salt	Solubility in water
Chlorides	All soluble except silver, mercury(I) and lead(II) chlorides
Nitrates	All soluble
Sulphates	All soluble except barium, calcium and lead(II) sulphates
Carbonates	Only sodium, potassium and ammonium carbonates are soluble
Sodium salts	All soluble
Potassium salts	All soluble
Ammonium salts	All soluble

Table 4.4 Solubility of salts in water

4.4.4 General Methods for Preparation of Salts

There are many methods or preparing salts. Some of the more important ways are given below.

1. **Neutralization (action of acids on alkali).**

Acid + Base ⟶ Salt + water

$NaOH\,(aq) + HCl\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$

2. **Action of acids on a metal**

A metal more active than hydrogen reacts with an acid to form salt.

Acid + metal ⟶ Salt + Hydrogen

$Mg(s) + H_2SO_4\,(aq) \longrightarrow MgSO_4\,(aq) + H_2\,(g)$

3. **Action of acids on metallic carbonates and bicarbonates**

When a metal carbonate or a metal bicarbonate reacts with a dilute acid, a salt, carbon dioxide and water form.

Acid + Metal carbonate → Salt + Water + Carbon dioxide

$BaCO_3(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

$HNO_3(aq) + NaHCO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$

In general, soluble salts are obtained using the methods mentioned above. These salts are separated from the solution by evaporation and crystallization. However, insoluble salts are obtained by precipitation or double decomposition. And these salts are separated from a solution by filtration.

Activity 4.8: To demonstrate the reaction between a metallic oxide and sulphuric acid

Materials required: Dilute sulphuric acid (H_2SO_4), copper oxide, CuO, a 100 mL beaker, a stirring rod, Bunsen burner.

1. Observe the copper oxide (CuO) and record its colour.
Add dilute sulphuric acid into the 100 mL beaker and gently warm the acid solution gently without boiling it.
Then add the copper oxide into the warmed acid in small portions. Continue to add the copper oxide until the oxide no longer reacts.

2. What is the colour of the solution after the reaction is complete?
3. Can you predict the chemical reaction that has taken place?

4.4.5 Uses of Common Salt

Salts are very important compounds. They have been used by mankind since ancient times. Some of the uses of some common salts and their formulas are given in the table below.

Salt	Formula	Important uses
Ammonium nitrate	NH_4NO_3	Fertilizer
Copper sulphate	$CuSO_4$	Pesticide and fungicide
Iron sulphate	$FeSO_4.\ 7H_2O$	To treat anaemia
Sodium bicarbonate	$NaHCO_3$	Baking powder
Sodium carbonate	Na_2CO_3	|Manufacture of glass, soap
Sodium chloride	NaCl	Manufacture of sodium hydroxide and flavouring of food

Table 4.5 List of some salts and their important uses

Did you know?

When you immerse a raw egg into salt water, it may either float or sink. This is because the density of the egg is either greater or smaller than that of the salt solution. If the egg sinks, it means it is fresh; but if it floats, it means the egg is not fresh.

Exercise 4.6

1. Explain the difference between normal salts and acid salts.
2. Give two examples each of a soluble salt and an insoluble salt.
3. Name three general properties of salts.

Summary

Salts are compounds produced by acid-base reactions. They consist of metallic ions derived from the base and negative ions derived from the acid. There are two types of salts: normal salts and acid salts.

Three of the general methods of preparation of salts are: neutralization, the action of an acid on a metal oxide, and the action of acids on metallic carbonates or metallic bicarbonates.

Salts play an important role in agriculture as fertilizers and pesticides, in the home as food flavouring and as preservatives, in medicine to repair broken bones, and in industry for the manufacture of soap and glass.

4.5 The pH Scale: Measurement of Acidity and Alkalinity

The $[H^+]$ or the $[OH^-]$ can be used to indicate the acidity or basicity of a solution. The square brackets indicate concentration. There are three possibilities regarding the concentration of hydrogen ions and hydroxide ions in an aqueous solution. When the concentration of the hydrogen ion $[H^+]$ is higher than the concentration of the hydroxide ion $[OH^-]$, the solution is acidic. When the concentration of the hydrogen ion $[H^+]$ is equal to the concentration of the hydroxide ion $[OH^-]$, then the solution is neutral. However, when the concentration of the hydrogen ion $[H^+]$ is less than the concentration of the hydroxide $[OH^-]$, the solution is basic or alkaline.

$[H^+] > [OH^-]$ acidic solution

$[H^+] = [OH^-]$ neutral solution

$[H^+] < [OH^-]$ basic solution

Therefore, the molar concentration of the hydrogen ions can be used to identify whether a solution is acidic, basic or neutral. Since the hydrogen ion concentration in solutions is very small, it is more practical to express the concentration of hydrogen ions (or acidity) in terms of the pH-value.

The pH of a solution is expressed as the negative logarithm of the molar concentration of a hydrogen ion.

$$\mathbf{pH = -\log [H^+]}$$

The term "log" is defined as the logarithm to the base ten or log $[H^+]$.

Pure water ionizes slightly, as shown below:

$$H_2O \longrightarrow H^+ + OH^-$$

One litre of water contains 10^{-7} moles of H^+ and 10^{-7} moles of OH^-. The concentration of either H^+ or OH^- may be used to indicate the acidity or basicity of solution. The pH of pure water is:

$$\begin{aligned} pH &= -\log[H^+] \\ &= -\log [1.0 \times 10^{-7}], \text{ since } [H^+] \text{ is } 1.0 \times 10^{-7} \\ &= -\log 1 + (-\log \times 10^{-7}); \log 1 = 0 \\ &= 0-(-7 \log 10); \log 10 = 1 \\ \mathbf{pH} &= \mathbf{7} \end{aligned}$$

Therefore, the pH of pure water is 7.

Example

A sample of blood has a hydrogen ion concentration of 3.98×10^{-8} M. What is the pH of the solution?

Solution:

Given: $[H^+] = 3.98 \times 10^{-8}$ M

Required: pH

$$\begin{aligned} pH &= -\log [H^+] \\ pH &= -\log 3.98 + (-\log 10^{-8}) \\ &= -0.60 - (-8) \\ &= 7.40 \end{aligned}$$

Therefore, the pH of blood is 7.40

pH values of acidic neutral and basic

Solution type	**pH value**
Acidic	less than 7
Neutral	7
Basic	greater than 7

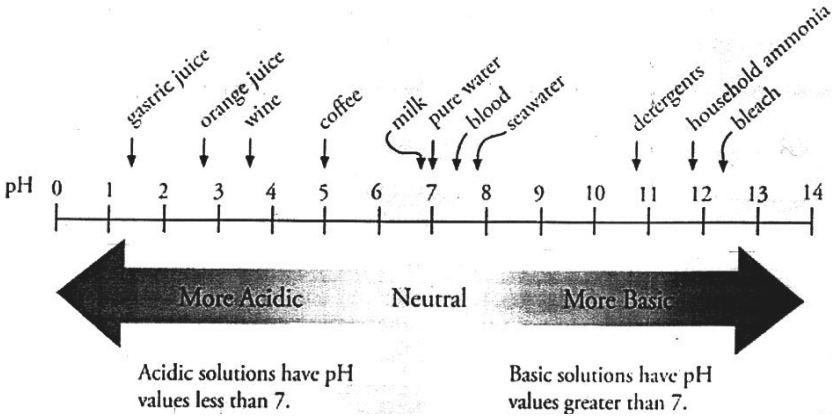


Fig 4.6 The pH scale

Activity 4.9: To measure the pH values of certain substances,

Materials required: Six 10 mL test tubes, water solutions of lemon juice, acetic acid, salt, soap, KOH, water, a universal indicator, a graduated cylinder.

Prepare approximately 5 mL of solution of each substance in each of the 10 mL test tubes.
Then using a universal indicator, test the pH of each solution.

1. Copy and fill the following table.

Solution	**Colour of universal indicator**	**Approximate pH**
Lemon juice		
Acetic acid		
Salt		
Soap		
KOH		
Water		

2. Which of the solutions are acidic, which ones are basic and which ones are neutral?
3. Can you determine the relationship between the pH and acidic, basic and neutral solutions ?

Did you know?

Some gases such as sulphur trioxide, (SO_3) and nitrogen dioxide, (NO_2) are emitted into the atmosphere from industries. Whenever, these gases come in contact with rain water, they form **"acid rain".** During the reaction between sulphur trioxide and rain water, sulphuric acid is formed, and nitrogen dioxide produces nitric acid when it reacts with rain water.

Acid rain damages metals, and buildings and pollutes rivers and oceans. Acid rain can make soil acidic and less fertile and can kill fish and other organisms in the oceans. The pH of acid rain ranges from 2 to 5.

Exercise 4.7

1. The H^+ ion concentration of solutions A and B is 1×10^{-5} M, and 2.5×10^{-9} M respectively. Calculate the pH of each solution. Identify whether each solution is acidic, basic or neutral.
2. a) Is coffee acidic or basic? See Fig. 4.6 for the pH scale.
   b) With reference to the pH scale, determine the approximate pH of orange juice, milk, blood and bleach.

4.5.1 Measuring pH

There are various ways of measuring the pH of solutions. The most common are:

a. The pH meter
b. Indicators

Measuring pH plays an important role in our daily lives and in the laboratory. For example, we may occasionally need to measure the pH of the water in a swimming pool. We can easily measure the pH with an indicator or a pH meter. Measurement of the pH of a pool is important, because if the pH is not within a certain range, it can cause skin problems or may cause bacteria grow. Both will have an adverse effect on our health.

a. The pH- Meter

The pH meter is an instrument that is used to measure the pH of a solution quickly and accurately. It consists of a pair of electrodes connected to a meter which measures small voltages. When the electrodes are immersed in a solution of an unknown pH, a voltage which depends on the pH is generated between the electrodes and is read by the meter, which is calibrated to give pH levels.

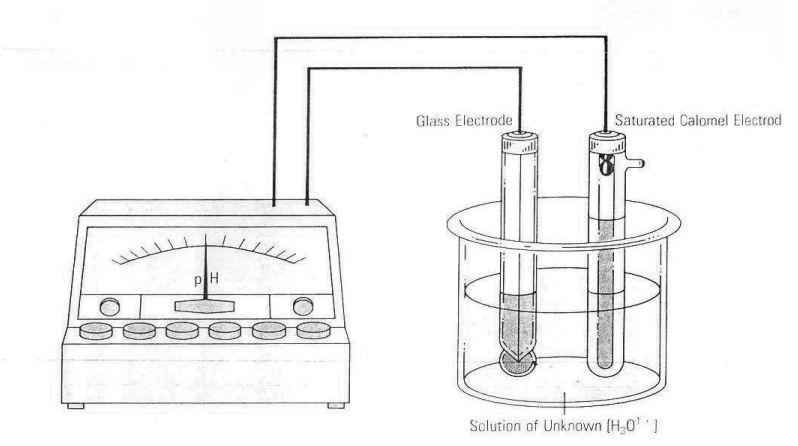


Fig 4.7 A pH-meter

b. Indicators

Indicators are often used to measure the pH of a solution. Acid–base indicators are less precise, because they change colour within a small pH range. These types of indicator are coloured substances that can exist in either an acid or a base form. The indicator changes to one colour in an acid and changes to another colour if placed in a base. Most indicators are extracted from plants and flowers such as red cabbages, roses, beets, grapes, and the petals of sunflowers. They are prepared in solutions or absorbed onto paper

A universal indicator that shows a series of colour changes at different pH values, and indicates the acidity or alkalinity of a solution that is under investigation. Table 4.7 shows the pH values and the corresponding colour on a universal indicator.

pH value	1-4	5	6	7	8	9	10	11-14
Colour	Red	Orange	Yellow	Green	Bluish-green	Blue	Bluish purple	Purple

Table 4.7 pH values and the colour change of universal indicator

Summary

In aqueous solutions, the range of possible hydronium or hydroxide ion concentrations is from 10^{-14} M to 1 M. The pH scale was developed in order to handle this very large range of concentrations. The P^H is a simpler way of expressing the H^+ ion concentration.

The pH is defined as the negative logarithm of the hydrogen ion concentration. pH less than 7 is acidic, pH 7 is neutral, whereas pH greater than 7 is basic.

The acidity or basicity of a substance can be identified by measuring the p^H of the solution. The pH meter or indicators are used to measure the pH of a solution. Measuring the pH of a solution plays an important role in our daily lives and in the laboratory.

Key Terms

Acidic hydrogen
Acid salt
Alkali
Arrhenius acid
Arrhenius base
Bronsted acid
Bronsted base
Basic oxide
Diprotic acid
Equivalence or end point
Hydronium ion
Hydroxide ion
Indicator
Insoluble salt
Monoprotic acid
Neutralization
Normal salt
Polyprotic acid
pH
pH-meter
Strong acid
Strong base
Titration
Triprotic acid
Weak acid
Weak acid
Weak base
Universal indicator

Review Questions

Part I. Multiple Choice Questions

1. Which of the following metals will react readily with an acid?
   a) Hg b) Zn c)Au d) Cu
2. Which of the following is a strong acid?
   a) Acetic acid b). Hydrochloric acid
   c) Formic acid d). Nitrous acid
3. Which of the following is not true about bases?
   a) Bases turn blue litmus to red
   b) Bases have sour taste
   c) Bases have a pH greater than 7
   d) a and b
4. An alkali is
   a) A solid salt b) A weak base
   c) A soluble base d) An insoluble base
5. In the laboratory preparation of $ZnSO_4$, the reacting substances are:
   a) Calcium sulphate and hydrochloric acid
   b) Carbon dioxide and zinc
   c) Zinc and sulphuric acid
   d) Sulphuric acid and calcium oxide
6. One of the following substances can be made from NaOH
   a) NH_4NO_3 b) HNO_3 c) Soap d)Vinegar
7. Which of the following is a normal salt?
   a) KCl b) $CaCO_3$ c) $NaHCO_3$ d) a and b
8. When 0.5 M of NaOH is added to 1M of HCl, then the pH of the solution will:
   a) increase b) decrease
   c) remain the same d) none of the above,
9. Which of the following salts is used as a pesticide?
   a) Na_2CO_3 b) $CaSO_4$ c) $CaCl_2$ d) $CuSO_4$
10. During a neutralization reaction:
    a) P^H of solution is 7
    b) Concentration of H^+ and OH^- is the same
    c) A salt is formed
    d) All of the above occur.

Part II. Fill in the Blank Spaces.

1. An example of a diprotic acid is _____.
2. According to the Bronsted-Lowry theory, an acid is _____.
3. Bases that are soluble in water are called _____.
4. One of the methods of preparing an insoluble salt is _____.
5. Ammonium hydroxide is an aqueous solution of _____.
6. The formula of lead sulphate is_____.
7. Bases are formed when oxides of _____ react with water.
8. If the pH of a solution is zero, then the hydroxide ion concentration is_____.
9. The instrument that is used to measure the pH of a solution is _____.
10. One of the simplest methods of checking whether neutralization has taken place in an acid-base reaction is_____.

<u>Part III</u>. Work out Questions and Problems

1. a) Write an equation for the reaction of water and an oxide to form an acid.
   b) Write an equation for the reaction of water and an oxide to form a base.
2. Give at least two physical properties of bases.
3. a) Write an equation for the reaction between sulphuric acid and barium hydroxide.
   b) The solutions of both sulphuric acid and barium hydroxide are good conductors of electricity. But when the two chemicals react, then the conductivity of the mixture is decreased. Explain.
4. What is the difference between a monoprotic acid and a diprotic acid?
5. Suggest a method of preparing NaCl from two other substances.
6. What is an indicator?
7. Identify each of the following substances as an acid, base or salt.

   a. HF b. CH_3CN c. $NaOCl_4$ d. $Ba(OH)_2$

8. Give an example of a salt that is used to manufacture fertilizers.
9. How would you distinguish salts from acids and bases?
10. Write two methods of preparing acids. Write an equation for each method of preparation.
11. Compare the concentration of hydrogen ions at pH = 2 and pH = 4 Which one is larger and why?
12. All alkalis are not bases but all bases are alkalis. Is this statement true or false? Explain why.
13. Calculate the hydroxide ion concentration of a solution if the pH is
    a) 6 b) 12 c) 2
14. Write the products of the following reactions and balance the equations.
    a) Copper hydroxide + Perchloric acid $\rightarrow$
    b) Calcium hydroxide + Nitric acid $\rightarrow$
    c) Ammonia + Sulphuric acid $\rightarrow$
15. Write three uses of neutralization reaction.
16. Name the following salts and classify them as either normal or acidic salts:-
    a) K_2CO_3 b) $MgBr_2$ c) $NaHCO_3$ d) $LiNO_3$
17. What is meant by a standard solution?
18. A 20.0 mL portion of a solution of 0.0200M HNO_3 is titrated with 0.0100M KOH.
    a) What is the pH at the equivalence point?
    b) How many mL of KOH are required to reach equivalence point?
19. When 1.05g of NaOH is added to 249 mL of distilled H_2O to make a 0.1M NaOH, what is the pH of the solution?
20. As the pH decreases, the concentration of hydronium ion increases, the concentration of hydroxide ions decreases. The hydronuim ion concentration increases by a factor of 10 for every one unit decrease in pH. If the pH of a solution decreases from a pH of 5 to a pH of 4, what would be the H^+ ion concentration?
21. The pH of a cup of coffee (at 25^{o}C) was found to be 5.12. What is the hydrogen ion concentration?

UNIT 5
Energy and Chemical Reactions

Expected Learning Outcomes

By the end of this unit, you are expected to:

- explain the energy transformations and changes involved in chemical reaction
- differentiate between exothermic and endothermic reactions
- write and describe thermochemical equations of different reactions.

5.1 Energy Transformation and Chemical Reactions

5.1.1 Law of Conservation of Energy

Activity 5.1: Energy transformations

Match the following diagrams with the type of energy transformation listed below:



Sun



Battery in torch



Iron



Plant



Candle



Light Bulb

1. Chemical energy to light and heat energies.
2. Chemical energy to electrical energy and then to light energy
3. Solar energy to thermal and light energies
4. Electrical energy to light and heat energies
5. Light energy to chemical energy
6. Electrical energy to heat energy

All matter possesses energy in one form or another. **Energy** is the capacity to do work. Energy can be broadly classified as **kinetic energy** and **potential energy**. Kinetic energy is the energy due to the motion of the particles of a substance. The different forms of kinetic energy include thermal energy and electrical energy. Potential energy is energy due to the position of the particles contained in a substance or it is defined as the stored energy of a substance. The different forms of potential energy include chemical energy and gravitational energy.

Energy can be converted from one form into another. For example, sunlight (or solar energy) falling on a surface is absorbed and converted into heat (or thermal energy) that warms the surface; the kinetic energy of running water and the thermal energy of steam are converted into electrical energy by turbines. Experiments have shown that energy is never lost during these transformations.

The **law of conservation of energy** states that, "Energy is neither created nor destroyed in chemical reactions, but it may be converted from one form to another".

The most common form of energy change in chemical reactions is heat change, and thus the major part of this sub-unit will be concerned with the heat changes that accompany chemical reactions.

5.1.2 Units of Heat changes

Activity 5.2: Direction of heat flow

A. Look at figure A and B below.

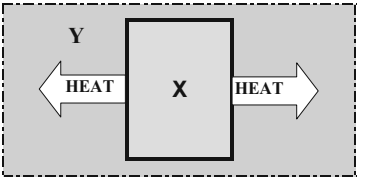


Fig A

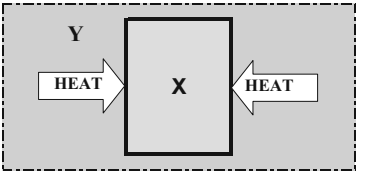


Fig B

"X" represents a vessel for reaction and "Y" represents part of the universe outside the reaction vessel (surroundings).

1. Describe the direction of heat flow in Figs A and B.
2. What could be the possible reason for the direction of flow?

B. Why do you feel warm when you touch a cup of hot coffee and feel cool when you touch a glass of ice?

The most common form of energy transformation in chemical reactions is in the form of heat. Heat is defined as the energy that flows into or out of a system because of a difference in temperature between a system and its surroundings. And temperature is the measure of the hotness or coldness of an object. Temperature determines the direction of heat flow. Heat flows from a region of high temperature to lower temperature. The commonly used temperature scales are Celsius (^oC), Fahrenheit (^oF) and kelvin (K). The International System (SI) unit of temperature is the kelvin (K).

Heat it measured using different units. Here we will mention the common units, namely calorie and joule. For many years the unit of heat change for chemical reactions has been the calorie (cal). A calorie is the heat required to raise the temperature of one gram of water by 1 ^oC.

Joule (J) is now replacing the calorie as the international (**SI)** unit of heat energy. The relationship between calorie and joule is given as follows:

1 calorie (cal) = 4.2 J
1 Kilojoule (kJ) = 1000 J

Exercise 5.1

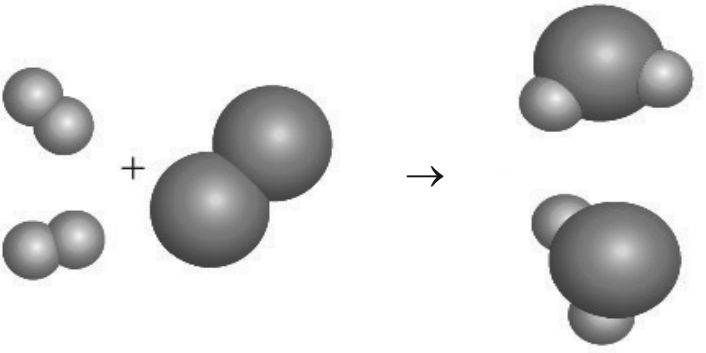
Convert the following units of heat energy:-

A. 100 calories = _____ joules
B. 4830 joules = _____ calories
C. _____cal = 210 kJ

5.1.3 Bond Breaking and Bond Formation

Activity 5.3. Formation of water molecules

Observe at the diagram below; compare the equation given with the corresponding diagram:



$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$

1. Does the diagram represent combination or decomposition?
2. How are the water molecules formed? What is happening to the bonds in H_2 and O_2?

Chemical reactions usually involve the making or breaking of chemical bonds. To break a chemical bond, energy is required to pull atoms apart because atoms in molecules are held together by chemical bonds. When atoms form a chemical bond, they release energy to become more stable in the molecule. The energy of the product molecules may therefore be greater or less than the energy of the reactant molecules. Thus a chemical reaction is accompanied by an overall energy change.

The average energy required to break a chemical bond is called the **bond energy** and is expressed in kJ/mol of bonds. Each bond has a typical value of bond energy. For example breaking 1 mole of C-H bonds requires on average about 413 kJ. Table 5.1 summarizes the bond energy of some covalent bonds.

Covalent bond	Bond energy (kJ/mol)	Covalent bond	Bond energy (kJ/mol)
N≡N	945	C-H	413
O=O	498	C-O	360
H-F	565	C-C	348
H-Cl	432	C=C	614
H-Br	366	C≡C	839

Table 5.1 Bond energy of different bonds

Activity 5.4. Analysis of bond energy

Refer Table 5.1

1. Identify the bonds having the greatest and the lowest bond energies?
2. In the following, compare the bond energies of the given species and justify your answer?
   i. N_2 and O_2,
   ii. C-C, C=C, and C≡C

5.1.4 Enthalpy of a Chemical Reaction

Activity 5.5: Heat involved during chemical reaction

Materials required: Quick lime (CaO), a 25 mL measuring cylinder, a 250 ml beaker, balance, water and stirrer.

Pour about 25 mL of water into the beaker.
Touch the beaker containing the water.

1. What do you feel?

Add about 50 g of the quick lime into the water and stir the mixture. Touch the beaker.

2. What do you feel?
3. Is heat released from the beaker to your hand or from your hand to the beaker?

In describing energy changes during chemical reactions, it is believed that each substance possesses a certain amount of energy.

The energy or heat content of a particular chemical substance is said to be **enthalpy**, and is denoted by H. It is difficult to measure directly the actual enthalpy of a molecule or an atom. However, we can measure the enthalpy change when atoms or molecules undergo reactions. The enthalpy change of a reaction is designated as **ΔH**; that is, the "Δ" represents change and is read as delta. The enthalpy change of a reaction is thus the difference between the heat content of products and the heat content of reactants

Heat change can therefore be expressed as:

Enthalpy change of a reaction	=	**Heat content of products**	−	**Heat content of reactants**
ΔH	=	$H_P - H_R$		

The enthalpy change in a chemical reaction originates from the making and/or breaking of chemical bonds. Thus, there are two stages in any reaction:

1. The breaking of chemical bonds in the reactants. This stage does not occur by it self. Energy, usually in the form of heat, must be supplied to break the bonds.
2. The formation of bonds in the products. This stage occurs entirely of its own accord and energy is released.

The unit of ΔH is the kilojoule per mole (kJ/mol)

5.1.5 Exothermic and Endothermic Reactions

Activity 5.6. Heat release and heat gain in chemical reactions

Materials required: hydrochloric acid (3M), calcium oxide (CaO), ammonium chloride (NH_4Cl), barium hydroxide ($Ba(OH)_2$), spoon, water, stopper, thermometer, 2 test tubes, 25 mL measuring cylinder and stirrer.

Label two test tubes as A and B.
Copy the following table in your notebook and record your results as follows:

	T_i	T_f	ΔT
Experiment A			
Experiment B			

Experiment A

Put about 20 mL of 3.0 M HCl in test tube A.
Measure and record the temperature of the HCl as T_i in the table.
Add a spoonful of CaO and shake to dissolve.
Touch the test tube. Measure and record the temperature of the mixture as T_f in the table.

Experiment B

Put a spoonful of solid NH_4Cl in test tube B.
Measure and record the initial temperature of the sample.
Add two spoonfuls of solid $Ba(OH)_2$.
Stopper and shake well.
Record the final temperatures in the table.

1. In which of the above reactions do we have an increase in temperature?
2. What do you understand from the above results

The majority of chemical reactions are accompanied by an observable heat change. Two types of heat change are distinguished. These are exothermic and endothermic reactions.

An exothermic reaction is a chemical reaction that releases heat energy to the surroundings. ("**exo**" means out and '**thermic**' means heat). The heat content of products is lower than the heat content of reactants ($H_P < H_R$); because energy is given off in the form of heat ΔH is negative. In an exothermic reaction, the energy released is written as part of the product. It can be represented as:

$$\textbf{Reactants} \rightarrow \textbf{Products + Energy}$$

$$\mathbf{H_{products} < H_{reactants}}$$

$$\mathbf{Therefore, \Delta H = H_{products} - H_{reactants} = -ve}$$

An endothermic reaction is a chemical reaction that absorbs heat energy from the surroundings ("endo" means in and indicates an intake of energy). In this type of reaction, the heat content of products is greater than the heat content of the reactants and so ΔH is positive. In endothermic reactions the energy is written as part of the reactants. It can be represented as:

Reactants + Energy → Products

$$H_{products} > H_{reactants}$$

Therefore, $\Delta H = H_{products} - H_{reactants} = +ve$

Table 5.2 below summarizes the difference between exothermic and endothermic reactions.

Exothermic Reaction	Endothermic Reaction
Heat is released.	Heat is absorbed.
Heat is written as part of products. ($R \rightarrow P + heat$)	Heat is written as part of reactants. ($R + heat \rightarrow P$)
Reaction vessel becomes warmer. (Temperature inside reaction vessel increases.)	Reaction vessel becomes cooler. (Temperature inside reaction vessel decreases.)
Heat of the reactants is greater than the heat of the products. $H_{(reactants)} > H_{(products)}$	Heat of the reactants is less than the heat of the products. $H_{(reactants)} < H_{(products)}$

Table 5.2 Comparison of exothermic and endothermic reactions

Activity 5.7. Identification of exothermic and endothermic reactions

Materials required: 10 mL measuring cylinder, hydrochloric acid (HCl) solution, magnesium (Mg) ribbon, large test tube, thermometer.

Copy the following table in your notebook and record your results as follows:

	T_i	T_f	ΔT
Mg + HCl			

Measure out 10 mL of HCl solution into a large,clean test tube.
Place the thermometer into the HCl solution and record the temperature (T_i).
Leave the thermometer in place and add a small piece of magnesium ribbon to the HCl solution,
Gently stir the solution with a stirrer to ensure good mixing.
(**Caution:** Do not breathe in the vapours.)
When the reaction is nearly complete, remove the thermometer and record the new temperature (T_f).

1. Is there an increase or a decrease in the temperature of the reaction?
2. Is the reaction exothermic or endothermic. Why?

5.1.6 Energy- level Diagrams

Energy-level diagrams can be used to illustrate overall exothermic and endothermic changes. They show the difference in total enthalpy of the reactants and products for a reaction.

In the energy diagram of an exothermic reaction, the reactants have more heat content than the products. The extra energy is released to the surroundings and thus the final energy level of the reaction is reduced.

In the energy diagram of endothermic reactions, the reactants have less heat content than do the products. Energy is supplied to the reactants to raise the particles up to the higher energy level. See Figs 5.1 and 5.2 below.

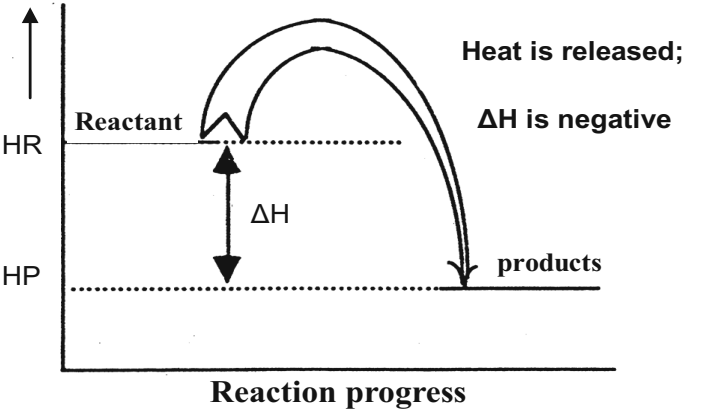


Fig 5.1: Energy diagram for exothermic reaction

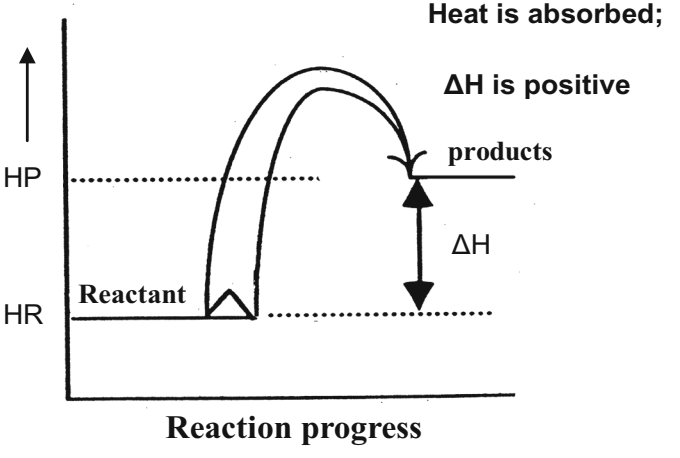
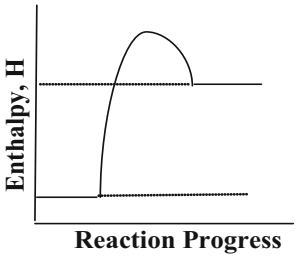


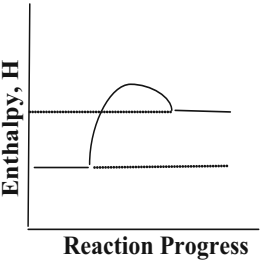
Fig 5.2: Energy diagram for endothermic reaction

Activity 5.8. Comparison of energy diagrams

Look carefully at the energy diagrams A and B below.



A



B

1. Indicate the reactants, products and ΔH in each diagram.
2. Identify the similarity and difference between A and B

Summary

Energy is a property of matter that allows work to be done. The two types of energy are potential (the energy due to position) and kinetic (the energy due to motion). Energy can be converted from one form into another. For example, potential energy can be converted to kinetic energy. The law of conservation of energy states that "Energy is neither created nor destroyed in chemical reactions, but it may be converted from one form to another".

Chemical reactions are accompanied by energy changes, usually in the form of heat. Heat is defined as the energy that flows into or out of a system to the surroundings. Temperature is the measure of the hotness or coldness of an object. The common units of heat are calorie and joule (1 cal = 4.2 J).

Chemical reactions usually involve the breaking and making of bonds. The breaking of bonds requires or absorbs energy;the formation of bonds releases energy. The average energy required to break a chemical bond is called the bond energy.

The heat content of a particular chemical substance is said to be enthalpy (denoted by H) and enthalpy change, ΔH, is the amount of energy absorbed or released by the reaction and can be found from the difference between the heat content of products and the heat content of reactants. that is, $\Delta H = H_P - H_R$

In an exothermic reaction, energy is released by the reaction; the energy of the reactants is greater than the energy of the products. (ΔH is negative). In an endothermic reaction, energy is absorbed by the reaction; the energy of the reactants is less than the energy of the products. (ΔH is positive). On an energy-level diagram, the enthalpy change is measured according to the difference in the enthalpy of the reactants and the products.

5.1.7 Common Examples of Exothermic Reactions

Activity 5.9. Energy-releasing and energy-requiring processes

Look at columns A and B.
1. Match set A with its opposite, set B.
2. Classify each of the following as either energy-requiring or energy-releasing processes:

SET A	SET B
1. making ice cubes	A. producing sugar by photosynthesis
2. formation of snow in clouds	B. conversion of frost to water vapour
3. condensation of rain from water vapour	C. melting ice cubes
4. burning sugar	D. evaporation of water

There are here are many examples of exothermic reactions; below are some of them:

1. Combustion

We need fuels to heat our homes, to cook our food, to run cars, and to generate electricity. When a fuel burns in oxygen, usually heat is released. This oxidation reaction in which energy is released is called a **combustion reaction.** Combustion requires a certain amount of heat to begin, but once started the reaction proceeds, generating the products along with large amounts of energy in the form of heat and light. For example, burning charcoal in air produces heat. That is why it is used as a fuel.

$$C(s) + O_2(g) \rightarrow CO_2(g) + \text{heat}$$

Did you know?

Ethanol (alcohol) is a **biofuel** which is found in sugarcane, corns and other cereals. It has advantages and disadvantages over other common fuels. Ethanol combustion produces a quiet flame that gives off few major pollutants when compared to petroleum-based fuels. Ethanol is also relatively easily made from fermenting plant materials, while other fuels require significantly more difficult processes to be produced. Ethanol combustion, however, produces less heat energy than many other commonly used combustible materials

2. Neutralization Reactions

Activity 5.10: Study of heat changes in acid-base reactions

Materials required: two 25mL measuring cylinders, hydrochloric acid (2M solution), sodium hydroxide (2M solution), beaker thermometer.

Label the two measuring cylinders; one "Acid" and the other "Base".
Copy the following table into your notebook and record your results as follows:

	T_i	T_f	ΔT
Acid + Base reaction			

Measure out 25 mL of 2M hydrochloric acid in the "Acid" measuring cylinder and pour it into a beaker.
Measure the temperature of the acid and record it in the table as T_i.
Measure out 25 mL of 2M sodium hydroxide in the "Base" measuring cylinder.
Add the base quickly to the acid in the beaker.
Stir and record the highest temperature reached (T_f).

1. Is there a change in temperature during the reaction? How?
2. Is the reaction exothermic or endothermic? Why?
3. What do you conclude from this experiment?

Neutralization is a type of double replacement reaction and usually involves release of heat, that is:

Acid + Base → Salt + Water + heat

For example, when dilute hydrochloric acid reacts with sodium hydroxide solution, energy is given out. It can be written as:

$$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l) + heat$$

3. Dilution of Concentrated Acids

Activity 5.11: Dilution of an acid

Materials required: concentrated sulphuric acid, 100 mL beaker, 10 mL measuring cylinder water.

Add 50 mL of water to the beaker. Touch the beaker.
Take 2 mL of concentrated sulphuric acid ,using the measuring cylinder.
Add the 2 mL of concentrated sulphuric acid to the water drop by drop.
Touch the beaker.
(**Caution:** Do not add the water to the acid and great care should be done in handling strong acids such as sulphuric acid.)

1. Has the beaker become warm or not?
2. Does this dilution represent an exothermic reaction? Why?

In the dilution process water is usually added to the solution to be diluted. However, in the dilution of concentrated acids such as sulphuric acid, the acid should be added to the water in small quantities. This is because the dilution of concentrated acids is highly exothermic and as a result much heat is released. Sometimes a characteristic fire or explosion occurs if the water is added to the concentrated acids. This can be represented as

$$H_2SO_4\ (concentrated) + H_2O(l) \rightarrow H_2SO_4(diluted) + heat$$

4. Hydration

If water is slowly added to a white, anhydrous copper sulphate, a characteristic sound and steam result, showing that heat is generated. A blue solid, hydrated copper sulphate is formed. The heat generated is called heat of **hydration**. The complete equation is represented as:

$CuSO_4(s) + H_2O(l) \rightarrow CuSO_4.5H_2O(s) + heat$

Anhydrous copper sulphate — hydrated copper sulphate

5. Respiration

Respiration is the combustion of food in living tissues. The process of respiration involves the breakdown of complex compounds. When the bonds in these compounds are broken, energy is released. The following chemical reaction illustrates the overall reaction that occurs in respiration:

$C_6H_{12}O_6 + 6O_2 \rightarrow CO_2 + 6H_2O + energy$

glucose

Other examples of exothermic reactions include adding an alkali metal to water, many corrosion reactions such as the oxidation of metals and the explosion of bombs.

5.1.8 Common Examples of Endothermic Reaction

Examples of endothermic reactions include:

1. Thermal Decomposition

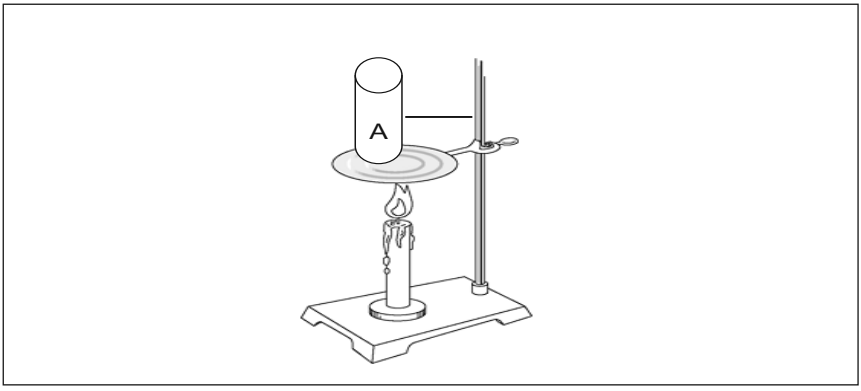
Activity 5.12. Effect of heating on decomposition reaction

Materials required: potassium chlorate, $KClO_3$, manganese dioxide (MnO_2), fire splint (or match stick), two medium sized test tubes, burner.

Label the test tubes as A and B.
Add 1 spoon of $KClO_3$ to each of the two test tubes (A and B).
Add a small amount of MnO_2 to both test tubes (A and B).
Gently heat test tube A using a burner (or spirit lamp) and keep test tube B at room temperature.
After few minutes, light splint or match stick and immediately put it in the mouth of each test tube.

1. What happens to the splint (or match stick) in both test tubes?
2. Which one shows the characteristic reaction? Why?

The set up for the experiment looks like this:



Usually compounds decompose when they are heated. Thermal decomposition reaction is one in which a compound breaks down into simpler substances by heating. Decomposition often requires an input of energy.

For example, when solid mercury(II) oxide is heated, it is decomposed to liquid mercury and oxygen gas, as follows:

$$2HgO(s) \xrightarrow{heat} 2Hg(l) + O_2(g)$$

In the thermal decomposition of potassium chlorate ($KClO_3$), potassium chloride and oxygen gas are produced. This is represented as:

$$2KClO_3(s) \xrightarrow{heat} 2KCl(s) + 3O_2(g)$$

2. Photosynthesis

Photosynthetic organisms such as plants use light energy to produce glucose (food) from carbon dioxide and water. This process is called photosynthesis. **Photosynthesis** is therefore an energy requiring process which occurs only if there is direct sun light. The general equation for photosynthesis is:

$$6CO_2(g) + 6H_2O(l) \xrightarrow{sun\ light} C_6H_{12}O_6(aq) + 6O_2(g)$$

Photosynthesis is almost the exact opposite reaction to respiration. These two reactions work together to maintain a biological energy balance on Earth.

3. Dissolving of Some Salts

Activity 5.13. Study of heat changes in dissolving process

Materials required: solid ammonium nitrate, (NH_4NO_3) measuring balance, water, thermometer, large beaker, measuring cylinder stirrer.

Copy the following table in your note book and record your results as follows:

	T_i	T_f	ΔT
Dissolving ammonium nitrate			

Pour about 100 mL of water into a large beaker.
Record the temperature of the water as T_i.
Weigh exactly 8 g of ammonium nitrate.
Add the 8 g of solid ammonium nitrate to the water and stir to dissolve.
Touch the beaker and record the temperature of the solution as T_f.

1. Is there an increase or a decrease in the temperature of the beaker?
2. Is this process exothermic or endothermic? Why?

Dissolving some salts such as ammonium nitrate, potassium chloride, and ammonium chloride in water is an endothermic process. These salts need the input of energy in order to dissolve. For example, when ammonium nitrate is dissolved in a beaker containing water, the temperature of the reaction vessel decreases. The equation is given as:

$$NH_4NO_3(s) + H_2O \xrightarrow{heat} NH_4NO_3(aq)$$

To identify those salts that dissolve endothermically in water, dissolve some amount of those salts in water and measure the temperature of the water before and after dissolving. A decrease in confirms the endothermic process.

4. Dehydration

Dehydration is the process of removing water of crystallization from a substance. It requires heat to remove the water. It is the opposite of the hydration process. For example, when hydrated copper sulphate is heated, water is removed from the solid crystal.

$$\underset{\text{hydrated copper sulphate}}{CuSO_4.5H_2O(s)} + \text{heat} \rightarrow \underset{\text{anhydrous copper sulphate}}{CuSO_4(s)} + H_2O(l)$$

Other examples of endothermic processes include melting, boiling, and evaporation.

Exercise 5.2

Classify the following chemical reactions as either exothermic or endothermic reactions.

A. Explosive reactions.
B. Hydration of Magnesium sulphate
C. The reaction of antacid in stomach
D. The reverse reaction of respiration.
E. The decomposition of limestone.

5.1.9 Thermochemical Equations

Let us discuss

What is the similarity and the difference between the following hypothetical equations?

1. $A(g) + B(g) \rightarrow AB(g)$
2. $A(g) + B(g) \rightarrow AB(g) + \text{heat}$

The quantity of heat absorbed or released during a chemical reaction can be shown in a chemical equation. An equation that contains the information about the heat changes involved in a reaction is called a **thermochemical equation**.

The following guidelines are helpful in writing thermochemical equations:-

1. The equation should be balanced

2. It should include the enthalpy change (ΔH value).
3. The physical states of reactants and products should be stated. That ,is the phases of each reactant and product must be indicated because the enthalpy of liquid water at 25 °C is different from that of gaseous water at the same temperature.

In thermochemical equations, the changes in energy can be considered as either reactant or product. In exothermic reactions heat is included with the products to show that energy is released. In endothermic reactions the heat is written with the reactants to show that heat is absorbed.

Table 5.3 shows the different ways of representing thermochemical equations for exothermic and endothermic reactions. Of the three kinds of representation, the third type (C) is the most formal. The equation is written and the enthalpy change is given as a separate ΔH on the right side of the equation.

Thermochemical equations of exothermic and endothermic reactions can be represented as follows:

Exothermic Reaction
A. $C(s) + O_2(g) \rightarrow CO_2(g) + 405\ kJ$ B. $C(s) + O_2(g) - 405\ kJ \rightarrow CO_2(g)$ C. $C(s) + O_2(g) \rightarrow CO_2(g);\ \Delta H = -405\ kJ$
Endothermic Reaction
A. $C(s) + 2S(s) + 117\ kJ \rightarrow CS2\ (l)$ B. $C(s) + 2S(s) \rightarrow CS2(l) - 117\ kJ$ C. $C(s) + 2S(s) \rightarrow CS2(l);\ \Delta H = +117\ kJ$

Table 5.3 Thermochemical equations of exothermic and endothermic reactions

Example

Given the following energy diagrams (Figs A and B), describe the diagrams and write their thermochemical equations.

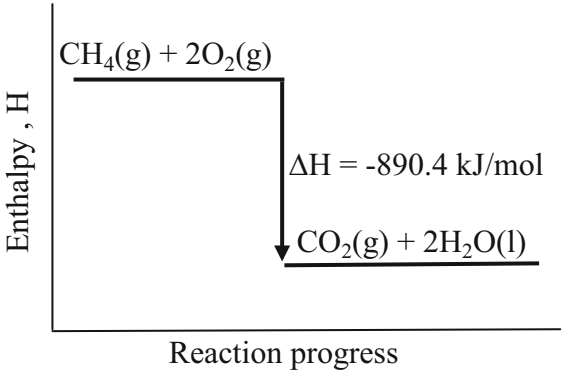


Fig A

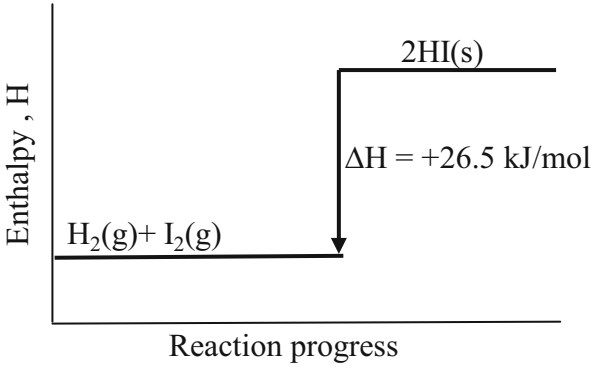


Fig B

Answers

Fig A illustrates an exothermic reaction. The heat of reactants is greater than heat of products and thus 890.4 kJ energy is released when 1 mole of CH_4 is burned in 2 moles of oxygen. The thermochemical equation is given as:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$; $\Delta H = -890.4$ kJ/mol

Fig B illustrates an endothermic reaction. Heat of products is greater than heat of reactants because 26.5 kJ energy is absorbed by the reactants to give the product. The thermochemical equation is given as:

$H_2(g) + I_2(s) \rightarrow 2HI(g)$; $\Delta H = +26.5$ kJ/mol

Interpretation of Thermochemical Equations

The value of ΔH given as kJ/mol refers to kJ per 1 mole of reactant or product as written in the equation:-

i. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$; $\Delta H = -92.4$ kJ/mol

In this equation 92.4 kJ of energy is released when 1 mole of N_2 (g) and 3 moles of H_2 (g) react with each other, and 2 moles of NH_3 (g) is produced. If the amount of reactants increases, the heat change also increases. Thus if the above equation is multiplied by two, the value of ΔH for the new equation obtained is multiplied by the same value (times 2),that is:

2 x $[N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)]$; $\Delta H = 2 \times (-92.4$ kJ/mol)
ii. $2N_2(g) + 6H_2(g) \rightarrow 4NH_3(g)$ $\Delta H = -184.8$ kJ/mol)

If the amount of reactant decreases, the heat changes also decreases. If only 1 mole of ammonia (NH_3) gas is produced, half of the -92.4 kJ/mol (i.e. -46.2 kJ/mol) energy will be released. The thermochemical equation may be written as:

iii. $\frac{1}{2} N_2(g) + 1\frac{1}{2} H_2(g) \rightarrow NH_3(g)$; $\Delta H = -46.2$ kJ/mol

Exercise 5.3

1. Classify the following chemical equations as either exothermic and endothermic reactions:

A. $4Na(s) + O_2(g) \rightarrow 2Na_2O(s) + 831$ kJ
B. $PbCO_3(s) + 2212kJ \rightarrow PbO(s) + CO_2(g)$
C. $3H_2(g) + N_2(g) - 92kJ \rightarrow 2NH_3(g)$
D. $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(aq)$; $\Delta H = -56.2$kJ
E. $CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + O_2(g) - 986$kJ
F. $\frac{1}{2} N_2(g) + \frac{1}{2} O_2(g) \rightarrow NO(g)$; $\Delta H = +903.3$kJ
G. $NO + \frac{1}{2}(O_2\ g) \rightarrow NO_2(g)$ $\Delta H = -56.52$kJ
H. $\frac{1}{2} N_2(g) + O_2(g) \rightarrow NO_2(g) - 33.85$kJ

2. Based on this reaction: $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$; $\Delta H = -110.5$ kJ
   A. How much energy is released if only 3 moles of CO gas are produced?
   B. How much energy is released if 1/4 mole of carbon is used in the reaction?

C. How much energy is released if 10 moles of oxygen is used in the reaction?

3. In the following reaction:

$$N_2(g) + O_2(g) \rightarrow 2NO(g) \; ; \; \Delta H = +180.6 \text{ kJ}.$$

A. What is the energy required when 5 moles of oxygen reacts with nitrogen?
B. How much energy is required (in kJ) to burn 2/3 moles of N_2 gas?

Summary

Common examples of exothermic reactions include combustion reactions, neutralization reactions, and respiration.

Combustion is the burning of a substance in the presence of oxygen and neutralization is the reaction of acids and bases to produce neutral substances, salt and water. The dilution of concentrated acids in water, such as sulphuric acid, is highly exothermic; hydration involves the addition of water molecules to a crystal and involves a release of energy. The burning of food in living tissues is called respiration; energy is released when complex molecules break down.

Common examples of an endothermic reaction include thermal decomposition, photosynthesis, the dissolving of some salts, and dehydration.

The breaking down of compounds into two or more substances in the presence of heat is said to be thermal decomposition and involves an input of energy. In photosynthesis, plants use light energy to produce glucose (food) from carbon dioxide and water and thus is an energy requiring process.

Some salts dissolve endothermically; they require an input of energy for dissolving**.** Dehydration is the process of removing water of crystallization from a substance. It requires heat to remove the water. It is the opposite of hydration.

The quantity of heat absorbed or released during a chemical reaction can be shown in a chemical equation. A complete thermochemical

equation should be balanced, the enthalpy change (ΔH) must be indicated and the physical states of the reactants and products should be stated. In the thermochemical equations of exothermic reactions heat is included with the products to show that energy is released and in endothermic reactions the heat is written with the reactants to show that heat is absorbed.

5.2 Types of Heat of Reaction and Heat Changes

There are different types of heat of reaction. They involve various amounts of heat change, depending on the heat content of the products and the reactants. Some important enthalpies of reactions and heat changes that describe the properties of substances are discussed below:

5.2.1 Heat of Combustion

Heat of combustion is the heat released when 1 mol of a substance is completely burnt in oxygen. For example, when 1 mol of methane gas (CH_4) is burnt completely, 890.4 kJ of energy is released. Thus the heat of combustion of methane is –890.4 kJ/mol"

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l); \quad \Delta H = -890.4 \text{ kJ/ mol}$$

5.2.2 Heat of Solution

Heat of solution is the enthalpy change when 1 mol of substance is completely dissolved in water. When 1 mole of solid sodium hydroxide is dissolved completely in an excessive amount of water, 40 kJ of energy is released. The thermochemical equation of the heat of solution of sodium hydroxide is given below

$$NaOH(s) \xrightarrow{H_2O} NaOH(aq); \quad \Delta H = -40 \text{ kJ/ mol}$$

5.2.3 Heat of Neutralization

Heat of neutralization is the heat released when 1 mol of water is formed during the neutralization of an acid by an alkali. Thus the heat of neutralization is essentially the energy change for the reaction: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$.

Let us look at the reaction between hydrochloric acid and sodium hydroxide:

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Net equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$; $\Delta H = -57.1$ kJ/ mol

Hence the heat of neutralization of HCl with NaOH is -57.1 kJ. The neutralization of any strong acid by a strong base usually produces the same final equation and similar value for the change in enthalpy. For example, the reaction of HNO_3 and NaOH also releases 57.1 kJ of energy.

5.2.4 Heat of Formation

Heat of formation is the enthalpy change when 1 mol of a compound is formed from its constituent elements. For example, when 1 mol of carbon dioxide gas is formed from its elements (carbon and oxygen), 393 kJ of energy is released. Hence the heat of formation of carbon dioxide is - 393 kJ/mol.

$$C(s) + O_2(g) \rightarrow CO_2\ (g)\ ; \qquad \Delta H = -393 \text{ kJ/ mol}$$

On the other hand, when 1 mol of carbon disulphide (CS_2) is formed from its elements (carbon and sulphur), 117 kJ of energy is required. Thus, the heat of formation of carbon disulphide is + 117 kJ/mol.

$$C(s) + 2S(s) \rightarrow CS_2(l)\ ; \qquad \Delta H = +117 \text{ kJ/ mol}$$

5.2.5 Heat of Fusion

Heat of fusion is the amount of heat needed when 1 mol of a solid melts (or fuses) to a liquid at its melting point. When 1 mol of ice is melted, we find from experiment that 6.01 kJ are needed. Thus the molar enthalpy of the fusion of ice is +6.01 kJ/ mol.

$$H_2O\ (s) \rightarrow H_2O\ (l)\ ; \qquad \Delta H = +\ 6.0 \text{ kJ/mol}$$

5.2.6 Heat of Vaporization

Heat of vaporization is the heat absorbed when 1 mole of a liquid is converted to gas at its boiling point. When 1 mole of water is boiled, 40.7 kJ energy is required. Thus the molar enthalpy of vaporization of water is 40.67 kJ/ mol.

$$H_2O\ (l) \rightarrow H_2O\ (g)\ ; \qquad \Delta H = +\ 40.7 \text{ kJ/mol}$$

Selected heats of fusion and vaporization of different substances are given in Table 5.4.

Substance	Formula	H_{fusion} (kJ/mol)	$H_{vaporization}$ (kJ/mol)
Oxygen	O_2	0.44	6.82
Methane	CH_4	0.94	8.18
Mercury	Hg	2.3	58.6
Water	H_2O	6.01	40.67
Sodium	Na	2.6	98

Table 5.4 Molar enthalpies of fusion and vaporization of selected substances**.**

As a substance changes from one phase to another, there will be a change in heat energy but no temperature change. We can summarize the changes of the states of matter from one physical state to another by plotting the temperature versus heat supplied to the system. In Fig 5.3, the heat of fusion of water (6.0 kJ/mol) and the heat of vaporization of water (40.7 kJ/mol) are shown.

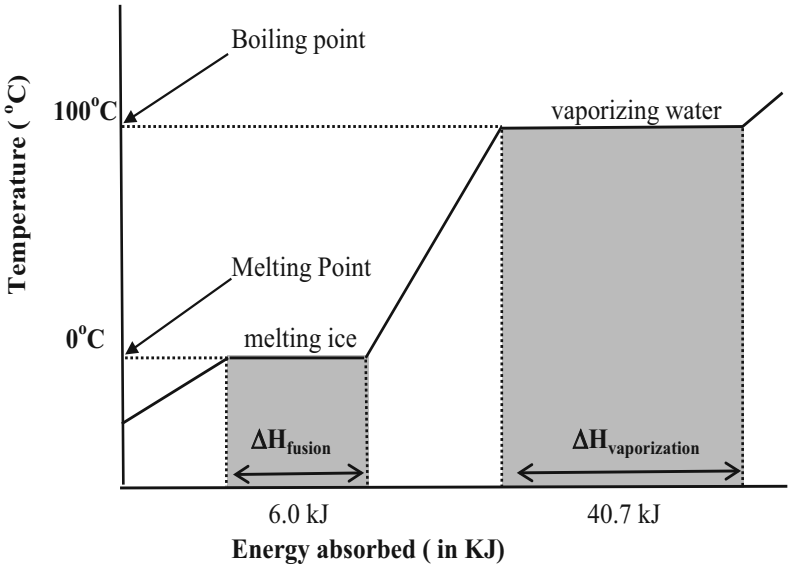


Fig 5.3 Temperature vs. Energy graph of the inter-conversion of the states of water

Summary

There are different types of heat of reaction and heat change, namely heat of combustion, heat of solution, heat of neutralization, heat of formation, heat of fusion and heat of vaporization.

Heat of combustion is the heat released when 1 mol of a substance is completely burnt in oxygen and heat of neutralization is the heat released when 1 mol of water is formed during the neutralization of an acid by an alkali. The heat of neutralization of a strong acid and a strong base is the same - 57.1 kj/mol. Heat of combustion and heat of neutralizaton are exothermic.

Heat of solution is the enthalpy change when 1 mol of substance is completely dissolved in water and heat of formation is the enthalpy change when 1 mol of a compound is formed from its constituent elements. Heat of solution and heat of formation can be either exothermic or endothermic.

Heat of fusion is the amount of heat needed when 1 mol of a solid melts (or fuses) to a liquid at its melting point and heat of vaporization is the heat absorbed when 1 mol of a liquid is converted to gas at its boiling point.

Key terms

Bond breaking
Bond energy
Bond formation
Calorie
Calorimeter
Combustion
Dehydration
Dissolving
Dilution of Acids
Endothermic reaction
Exothermic reaction
Energy diagram
Enthalpy (H)
Heat
Heat of combustion
Heat of formation
Heat of fusion
Heat of neutralization
Heat of reaction
Heat of solution
Heat of vaporization
Hydration
Joule
Law of Conservation of Energy
Neutralization
Photosynthesis
Respiration
Thermal decomposition

Review Questions

Part I. Choose the correct answer

1. The stored chemical energy in a molecule is a kind of ____energy.

a) kinetic b) Potential c) Electrical d) Electrostatic

2. A base ball weighing 143 g has a kinetic energy of 6.84 J, this is equivalent to ____cal.

a) 16.4 b) 1.64 c) 4.184 d) 28.6

3. In an exothermic reaction:
   a) **Δ**H is negative.
   b)The products have less heat energy than the reactants.
   c)Heat is given off to the environment
   d)All of the above

4. A solid (A) at a temperature of 80 ^{o}C is immersed into a liquid (B) with a temperature at 60 ^{o}C. Which statement correctly describes the heat changes between A and B?
   a) Solid A releases heat and B absorbs heat.
   b) Solid A absorbs heat and B releases heat.
   c) Both solid A and B absorb heat.
   d) Both solid A and B release heat.

5. When ammonium chloride crystals are dissolved in water the temperature of the water decreases. What does this temperature change indicate about the dissolving of ammonium chloride in water?
   a) It is an endothermic reaction because it absorbs heat.
   b) It is an endothermic reaction because it releases heat.
   c) It is an exothermic reaction because it absorbs heat.
   d) It is an exothermic reaction because it releases heat.

6. One of the following is not an exothermic process

a) Combustion b) Respiration

c) Photosynthesis d) Neutralization

7. The amount of heat required to change one mole of solid to liquid at its melting point to is _____.

a) Heat of reaction b) Heat of fusion
c) Heat of vaporization d) Heat of neutralization

8. Which of the following is an example of exothermic reaction?

a) removal of water from hydrous sample
b) when nitric acid is reacted with caustic soda.
c) hydrolysis of water by electricity
d none of the above.

9. Choose the true statement among the following.

a) In an exothermic reaction, Hp > HR.
b) In an endothermic reaction, heat is lost in product formation is greater than heat needed to break the bond.
c) $CO + O_2 \longrightarrow CO_2$, is an example of an exothermic reaction.
d) A chemical substance contains kinetic energy in any condition.

Part II. Fill in the blank space

1. Classify the following into exothermic and endothermic reactions.
   a) Dissolving of NH_4NO_3 in water _____
   b) Adding sodium hydroxide solution to sulfuric acid_____
   c) Burning iron in air _____
   d) Dehydration of $CaSO_4.2H_2O$ _____
   e) Dry ice (solid CO_2) changing to carbon dioxide gas_____

2. For each of the following, tell whether the process is exothermic or endothermic:
   a) $H_2O\ (l) \rightarrow H_2O\ (s)$
   b) $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O\ (g)$
   c) $H_2O(l, 25\ ^oC) \rightarrow H_2O\ (l, 15\ ^oC)$
   d) $H_2O\ (g) \rightarrow H_2O\ (l)$

3. _____ is the enthalpy change when 1 mol of a compound is formed from its constituent elements in their standard states.

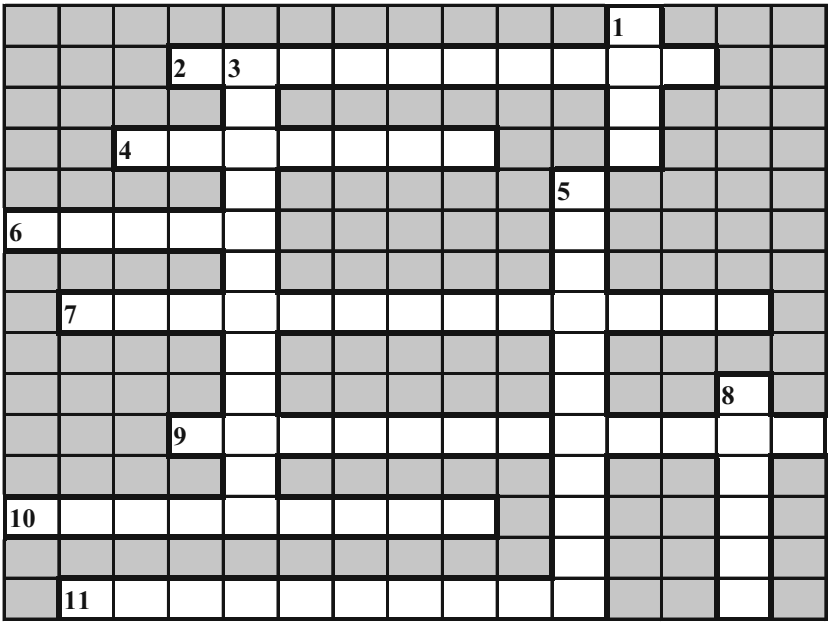
4. Complete the following table.

S.No.	Heat change	Definition	Sign of ΔH
1	enthalpy of combustion		
2	enthalpy of solution		
3	enthalpy of vaporization		
4	enthalpy of neutralization		

5. The three basic requirements in writing a complete thermochemical equations are_____, _____and _____.

Part III Essay type questions.

1. State the Law of conservation of energy and mention two examples of energy transformations.
2. In an old book you read that the oxidation of 1.00g of hydrogen to form liquid water produces 3800 cal. What is this energy in joules?
3. The label of a cereal box indicates that one serving provides 250 cal. What is this energy in joules?
4. What determines the quantity of heat change in a chemical reaction?
5. Describe some units of heat energy and temperature,
6. List the main differences between exothermic and endothermic reactions
7. Why is the enthalpy change of exothermic reactions negative while that of endothermic reactions is positive?
8. Explain why the heat of neutralization of sodium hydroxide by hydrochloric acid is the same as the heat of neutralization of potassium hydroxide by nitric acid?
9. Complete the word puzzle below. Use the descriptions given following the puzzle.



1. Energy is defined as the ability to do _____.
2. _____ energy is the energy required to begin a chemical reaction.
3. The Law of _____ of energy states that energy is neither created nor destroyed in a normal chemical reaction
4. _____ energy is the energy of matter in motion.
5. This type of chemical reaction absorbs energy.
6. _____is the unit for energy in the International System of Measurements.
7. _____energy is the energy that an object has due to its weight and its position with respect to some reference point.
8. _____ is defined as the ability to do work.
9. The amount of work done while moving an object is equal to the product of the force and the _____.
10. _____energy is sometimes called “stored” energy.
11. This type of reaction releases energy.