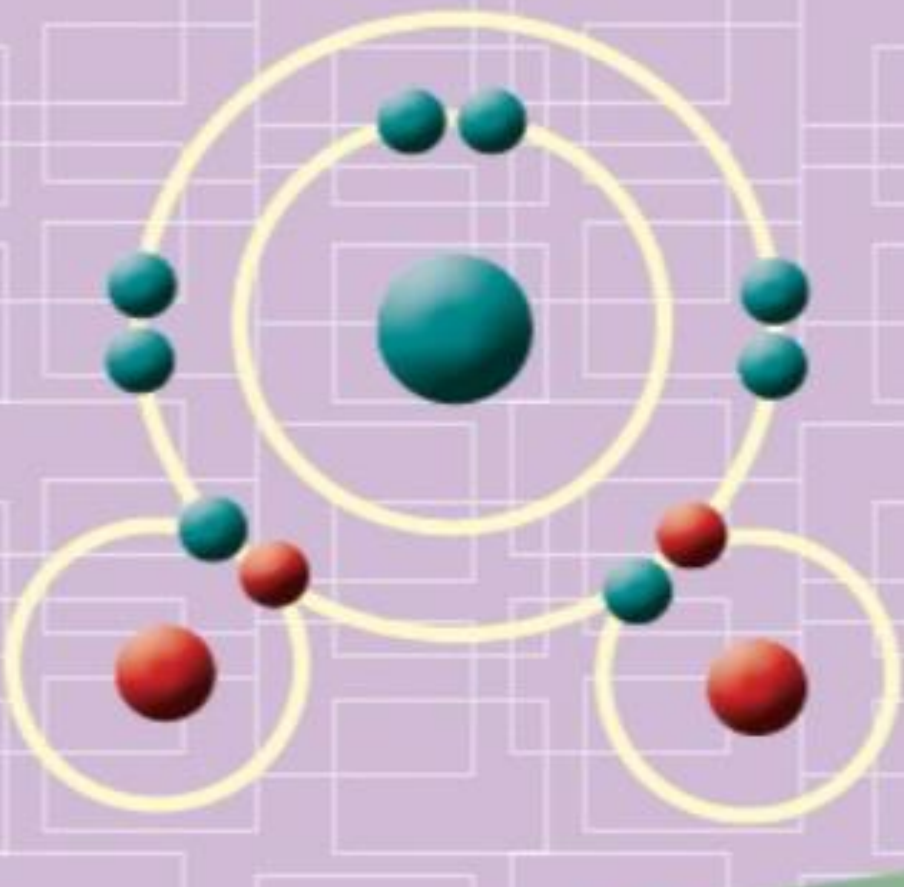
Chemistry

for Secondary Schools

Student's Book
Form Two



Tanzania Institute of Education

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Student's Book

Form Two


THE UNITED REPUBLIC OF TANZANIA
MINISTRY OF EDUCATION,
SCIENCE AND TECHNOLOGY

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Dr. Lyabwene M. Mtahabwa
Commissioner for Education

Tanzania Institute of Education

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Akomba.
Dr Aneth A. Komba
Director General
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Preface

This textbook, *Chemistry for Secondary Schools Student's Book Form Two*, has been written specifically for Form Two Students in the United Republic of Tanzania. The book is prepared following the 2023 Chemistry Syllabus for Ordinary Secondary Education, Form I–IV, issued by the Ministry of Education, Science and Technology (MoEST). It is a revised edition of Chemistry for Secondary Schools Student's Book for Form Two that was published in 2021 in accordance with the 2009 Chemistry Syllabus issued by the then Ministry of Education and Vocational Training (MoEVT).

The book consists of five chapters: Atomic structure, Periodic classification, Chemical bonding, formula and nomenclature, Chemical reactions, and Acids, bases, and salts. In addition to the contents, each chapter contains activities, tasks, illustrations and exercises. Moreover, a project is included in Chapter Five. You are encouraged to do all the activities and tasks found in this book and other assignments provided by your teacher and attempt all the questions in the exercises. You are also required to prepare a portfolio for keeping records of activities performed in different lessons. Doing so will enable you to develop the intended competencies.

Additional learning resources are available in the TIE e-Library at *https://ol.tie.go.tz* or *ol.tie.go.tz*

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Chapter One
Atomic structure

Introduction

Substances are made up of tiny particles called atoms. Atoms are the smallest particles of matter that carry an element's properties. In this chapter, you will learn about the concept of atomic structure and the determination of the atomic number and mass number of an element. The competencies developed will enable you to analyse the composition, behaviour, and properties of different chemical substances, enhancing your understanding of their interactions and applications in real-life situations.

Think

Atomic structure in the foundation of modern life

Concept of atomic structure

Task 1.1

Use an interactive simulation, video, or any other reliable resource to visualise the structure of an atom and examine its sub-atomic particles.

Atoms are fundamental units that make up all matter, forming everything in the surroundings, including the air, materials, and objects used daily. Understanding atomic structure helps scientists predict the behaviour of substances and design new materials and cleaner energy. For example, understanding the behaviour of atoms helps in food preservation, water purification, making batteries and medicines. This leads to technological and scientific developments that shape the world.

An atom is composed of smaller particles called sub-atomic particles. These include *protons, neutrons,* and *electrons.* The arrangement of these particles within an atom is referred to as the atomic structure. The structures of atoms are understood through the atomic theory, which has been developed over time through experiments and scientific discoveries.

Task 1.2

Use reliable resources to analyse various atomic theory discoveries and come up with an idea about the structure of the atom.

About the year 400 BC, a Greek philosopher known as Democritus was the first to consider the idea that matter is made up of particles. Such an idea was not accepted because there was no experimental evidence to support it. About 2000 years later, an English man called John Dalton revived the discussion. He used experimental evidence to convince people that matter is made up of particles called atoms. Through that experiment, he deduced Dalton's spherical model of an atom shown in Figure 1.1. Dalton's discovery helped scientists understand chemical reactions and how substances combine.

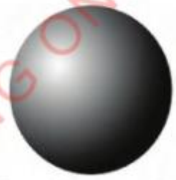


Figure 1.1: *Dalton's model of the atom*

Dalton atomic theory

In 1803, Dalton developed the theory about the atom. The four main points (assumptions) of Dalton's Atomic Theory are summarised as follows:

1. Matter is made up of tiny particles called atoms. (The word *atom* means 'unsplittable' in Greek).
2. Atoms can neither be created nor destroyed.
3. Atoms of the same element are identical and have the same mass and properties. Atoms of a given element are different from those of any other element. Atoms of different elements can be distinguished from one another by their respective relative weights.
4. Compounds are formed by a combination of two or more different kinds of atoms. Atoms always combine in simple whole number ratios.

Dalton never imagined that anyone would ever be able to see an atom. However, modern technology has provided direct evidence that shows the positions and patterns of individual atoms. The use of modern technology has enabled scientists to carry out experiments on the atom that Dalton could not. This has led to slight modifications to Dalton's Atomic Theory and thus formulated the so called modern concepts of Dalton's Atomic Theory.

These modifications include the following:

1. Atoms can be either created or destroyed by means of nuclear reactions. The atom can change form through special processes such as nuclear fusion (combining the atomic nuclei) or nuclear fission (splitting the atomic nucleus). For example, an atom of uranium-235 can be split into two separate atoms.
2. Some elements have atoms of more than one kind which differ slightly in mass. Such atoms are called isotopes. For example, carbon has three isotopes known as carbon-12, carbon-13, and carbon-14.
3. An atom is made up of smaller sub-atomic particles called protons, neutrons, and electrons.
4. Atoms of different elements may chemically combine in many different ratios to form compounds.

The modern atomic theory builds on Dalton's original ideas by recognising sub-atomic particles, isotopes, and nuclear reactions while retaining the ideas of chemical combinations and reactions. Dalton's discovery thus helped scientists understand chemical reactions and how substances combine.

Sub-atomic particles

In 1897, J. J. Thomson carried out experiments and described an atom as a sphere of positive charge, with negative particles called *electrons* spread throughout the sphere. This model of the atom was referred to as *plum pudding* model, as shown in Figure 1.2.

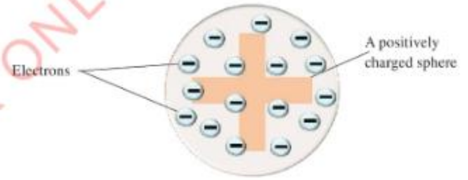


Figure 1.2: *Thomson's plum pudding model of the atom*

Thomson, therefore, managed to discover the electron among the three sub-atomic particles. His discovery led to inventions of electronic devices such as televisions, radios and computers. However, another scientist called Ernest Rutherford reasoned that if Thomson's model was correct, then the mass of the atom was evenly spread throughout the atom. He carried out experiments and discovered that most of the mass of an atom is concentrated in the nucleus (central

core) of the atom. Within the nucleus, there are positively charged particles called *protons*. This was the second sub-atomic particle to be discovered. Rutherford's discovery helped scientists discover nuclear energy used to produce electricity and in *radiotherapy*.

Rutherford's findings are summarised as follows:

1. Protons, the positively charged particles of an atom, are located in the nucleus.
2. Most of the mass of the atom is located in the nucleus.
3. The nucleus has a relatively smaller volume compared to the whole atom.
4. Electrons have very small masses compared to the protons.
5. Most of the space in an atom is empty.
6. Electrons are the negatively charged particles in an atom. They move around the nucleus in orbits.

Rutherford thus developed the *planetary model* of the atom shown in Figure 1.3.

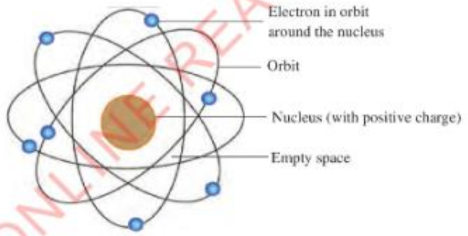


Figure 1.3: *Rutherford's planetary model of the atom*

In 1932, a scientist named James Chadwick discovered the *neutrons,* which also forms part of the nucleus. Figure 1.4 shows the locations of neutrons and other sub-atomic particles in an atom. Neutrons have the same mass as protons but no charge. They are located in the nucleus of an atom. They were the third sub-atomic particle to be discovered. The Chadwick discoveries have made nuclear power possible, helping to produce electricity and develop medical treatments such as cancer therapy.

The properties of neutrons are summarised as follows:

1. They have no charge (are neutral).
2. They have nearly the same mass as the corresponding protons.
3. They have a mass nearly 1840 times the mass of an electron.

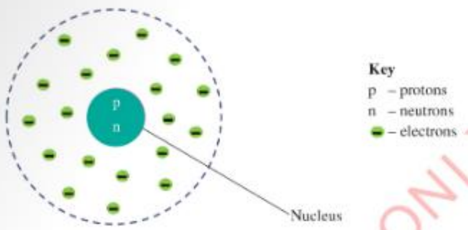


Figure 1.4: *Locations of sub-atomic particles in the atom*

Table 1.1 summaries the properties of sub-atomic particles of an atom.

Table 1.1: *Properties of sub-atomic particles*

Sub-atomic particle	Symbol	Location	Charge	Real mass (g)	Relative mass
Proton	p^+	In the nucleus	+1	1.6726×10^{-24}	1
Neutron	n	In the nucleus	0	1.6750×10^{-24}	1
Electron	e^-	Outside the nucleus	–1	9.109×10^{-28}	$\frac{1}{1840}$

Activity 1.1

Aim: To build a 3D atomic model of carbon

Requirements: Six medium-sized red beads and white beads, small-sized black beads, cardboard, clay or glue and string or wire

Procedure

1. Build the nucleus by randomly arranging six red beads (protons) and six white beads (neutrons). Then, stick them together using glue to form a tight cluster at the centre.

2. Create two circular loops of varying lengths of string or wire to represent the electron shells. Attach two black beads (electrons) on the first shell, and four black beads on the second shell. Ensure they are glued and evenly spaced on the string loops.
3. Label the model and attach the entire model to a cardboard base for stability.

Questions

1. Why are the electrons arranged in different shells?
2. If one more proton is added to the nucleus of this atomic model, would it still be a carbon atom? Explain.
3. Why were different colours used for protons, neutrons and electrons in the model?

Electronic arrangement

Task 1.3

Use a reliable interactive simulation to visualise the arrangements of electrons in an atom and explain how the electrons are arranged.

In 1913, Niels Bohr suggested that electrons rotate around the nucleus in special regions called *shells* or *orbits*. These shells (also known as *energy levels*) are at fixed distances from the nucleus. Each shell can only hold a specific number of electrons. The maximum number of electrons held within each shell can be determined by the formula $2n^2$, where n is the position of the shell from the nucleus. According to this formula, the:

first shell can hold $(2 \times 1^2) = 2$ electrons;

second shell can hold $(2 \times 2^2) = 8$ electrons; and

third shell can hold $(2 \times 3^2) = 18$ electrons.

The first four shells are represented by the letters K, L, M, and N, respectively as shown in Figure 1.5. Each electron in an atom is in a particular shell and the electrons must first occupy the lowest available shell nearest to the nucleus.

Figure 1.5: *First four shells of an atom*

For reasons beyond the scope of this book, the 3rd shell is more stable with 8 electrons. This is why even though the 3rd shell can hold up to 18 electrons, potassium element which has 19 electrons has only 8 electrons in its 3rd shell, and the last electron moves to the 4th shell. This is the same for calcium which has 20 electrons, where the last two electrons move to the 4th shell after the 3rd shell is completely filled with 8 electrons.

A shell which contains its maximum number of electrons is called a fully-filled shell. An atom with fully-filled outermost shell is said to be stable. Some atoms hold a maximum of 2 electrons (helium) or 8 electrons (neon) in their outermost shells. The elements with a maximum of 2 electrons in their outermost shells are said to exhibit a *duplet state,* while those with 8 electrons are said to exhibit an *octate state*. Electrons are arranged so that the lowest shells are filled first. This arrangement of electrons in different shells in an atom is called *electronic configuration*. Figure 1.6 shows the diagrammatic electronic configurations of hydrogen, helium, neon, potassium, and sulfur atoms.

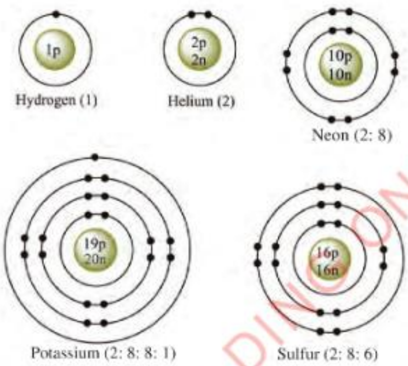


Figure 1.6: *Electronic configurations of some atoms*

Note that the hydrogen atom has no neutron in its nucleus. Helium and neon atoms have their outermost shells completely filled with electrons, and so they are stable atoms. The electrons are not fixed at particular positions within the shell. Instead, they move extremely fast and can be at any point within the shell.

Bohr's findings provided more information about elements that are summarised in a table called the *periodic table* (Appendix 1). The electronic arrangements of the first 20 elements are shown in Table 1.2.

Exercise 1.1

1. How did the discovery of the nucleus refine earlier atomic models?
2. Why are atoms considered electrically neutral under normal conditions?
3. Compare the Rutherford atomic model with the Bohr atomic model.
4. The atomic number is unique for each element in the periodic table. Explain.

Table 1.2: *Electronic arrangements of the first 20 elements*

Element	Chemical symbol	Number of electrons	Number of electrons in each shell				Electronic configuration (arrangement)
			1st	2nd	3rd	4th	
Hydrogen	H	1	1				1
Helium	He	2	2				2
Lithium	Li	3	2	1			2: 1
Beryllium	Be	4	2	2			2: 2
Boron	B	5	2	3			2: 3
Carbon	C	6	2	4			2: 4
Nitrogen	N	7	2	5			2: 5
Oxygen	O	8	2	6			2: 6
Fluorine	F	9	2	7			2: 7
Neon	Ne	10	2	8			2: 8
Sodium	Na	11	2	8	1		2: 8: 1
Magnesium	Mg	12	2	8	2		2: 8: 2
Aluminium	Al	13	2	8	3		2: 8: 3
Silicon	Si	14	2	8	4		2: 8: 4
Phosphorus	P	15	2	8	5		2: 8: 5
Sulfur	S	16	2	8	6		2: 8: 6
Chlorine	Cl	17	2	8	7		2: 8: 7
Argon	Ar	18	2	8	8		2: 8: 8
Potassium	K	19	2	8	8	1	2: 8: 8: 1
Calcium	Ca	20	2	8	8	2	2: 8: 8: 2

Task 1.4

Use drawings or cut-out pictures to represent different atomic models (Dalton's solid sphere, Thomson's plum pudding, Rutherford's nucleus and Bohr's orbits). Arrange the pictures in an order. Then, write one simple fact about each model.

Determination of atomic number and mass number

Task 1.5

Access an online simulation on how to build an atom. Explore the simulation by adding protons, neutrons, and electrons to form different elements. Observe changes in atomic number, atomic mass, and stability of the atoms.

The sub-atomic particles of an atom, namely protons, neutrons, and electrons, relate to the atomic number and mass number of the atom.

Atomic number

The *atomic number* is the number of protons in an atom. It is also known as the *proton number*. For example, the atomic number of hydrogen is 1 since it has only one proton. A sodium atom has 11 protons in the nucleus; therefore, its atomic number is 11. Since the number of protons is equal to the number of electrons in the neutral atom, its atomic number is not only the number of protons but it is also the number of electrons.

Therefore, for the neutral atoms:

Atomic number = Number of protons = Number of electrons

Mass number

Protons and neutrons are found in the nucleus of an atom and are called *nucleons*. The sum of the protons and neutrons in one atom of an element is called the *mass number or nucleon number* or *atomic mass*. This number is actually taken as the mass of the atom since the mass of the electron is negligible.

Thus,

Number of protons + Number of neutrons = Mass number

For example:

(i) Hydrogen has 1 proton and 0 neutrons. Therefore, its atomic number is 1, and mass number is $1 + 0 = 1$.

(ii) Boron has 5 protons and 6 neutrons. Its atomic number is 5 and mass number is $5 + 6 = 11$.

(iii) Nitrogen has 7 protons and 7 neutrons. Its atomic number is 7 and mass number is $7 + 7 = 14$.

It is also possible to calculate the number of neutrons and number of electrons of an atom if its mass number and atomic number are given.

Example 1.1

Atom Q has a mass number of 49 and an atomic number of 24. Calculate the number of neutrons and the number of electrons in atom Q.

Solution

Mass number = 49; atomic number = 24

(a) Neutron number = mass number – atomic number = 49 – 24 = 25

(b) Number of electrons = number of protons = atomic number = 24

Note: For the mass number with fractions, such as chlorine (35.5), calculating the number of neutrons and electrons involves only a whole number. In this case, for chlorine, 35 is used.

Nuclide notation

Atoms of different elements can be represented by chemical symbols that indicate their respective atomic numbers and mass numbers. Using an arbitrary element X, the mass number (A) is placed on its upper left end, while its atomic number (Z) is placed on the lower left end. Thus, element X is represented as $^{A}_{Z}X$. This is known as the *nuclide notation*. The following are examples of nuclide representations of different atoms:

(i) Hydrogen $^{1}_{1}H$ (ii) Boron $^{11}_{5}B$

(iii) Nitrogen $^{14}_{7}N$ (iv) Oxygen $^{16}_{8}O$

With this information, it is possible to deduce the number of neutrons and electrons in the atom, and to write the electronic configuration. For example, in the oxygen atom, the mass number is 16 and the atomic number is 8. Therefore, the number of electrons is 8 and the number of neutrons is 16 – 8 = 8. The nucleus of the oxygen atom can therefore be represented as shown in Figure 1.7.

Figure 1.7: *The nucleus of the oxygen atom*

Example 1.2

Potassium (K) atom has 19 electrons and the mass number of 39.

(a) Work out the:

(i) atomic number, and

(ii) number of neutrons.

(b) Give the nuclide notation.

(c) Show the representation of the nucleus of the potassium atom.

(d) Draw the electronic configuration of potassium.

Solution

(a) (i) Atomic number = number of protons = number of electrons = 19

(ii) Mass number = number of protons + number of neutrons

Number of neutrons = mass number – number of protons

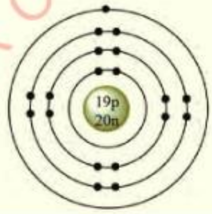
= 39 – 19

= 20

(b) $^{39}_{19}K$; where 39 is the mass number and 19 the atomic number.

(c)

(d)



Isotopes

Task 1.6

Watch educational video on isotopes and explain the uses of isotopes in carbon dating, medicine and agriculture.

Atoms of the same element have the same number of protons. However, the number of neutrons in the atoms of the same element may vary. This means that the atomic number of an element does not vary but the mass number can vary. Such atoms of an element are called *isotopes*. Isotopes are atoms of the same element with the same number of protons but different number of neutrons. Such an existence of the element is called *isotopy*. Isotopy is the existence of atoms of the same element having the same atomic number but different mass numbers. It is also possible to get the number of sub-atomic particles in a given isotope.

Example 1.3

State the number of protons, neutrons, and electrons in the following isotopes:

(a) $^{12}_{6}C$ and $^{14}_{6}C$

(b) $^{1}_{1}H$, $^{2}_{1}H$ and $^{3}_{1}H$

Solution

(a) $^{12}_{6}C$, Mass number = 12

Number of protons = atomic number = 6

Number of electrons = number of protons = 6

Number of neutrons = 12 – 6 = 6

$^{14}_{6}C$, Mass number = 14

Number of protons = 6

Number of electrons = 6

Number of neutrons = 14 – 6 = 8

(b) $^{1}_{1}H$, Mass number = 1

Number of proton = 1

Number of electron = 1

Number of neutron = 1 – 1 = 0

$^{2}_{1}H$, Mass number = 2

Number of proton = 1

Number of electron = 1

Number of neutron = 2 – 1 = 1

$^{3}_{1}H$, Mass number = 3

Number of proton = 1

Number of electron = 1

Number of neutrons = 3 – 1 = 2

Example 1.4

An isotope of carbon has a mass number of 13 and an atomic number of 6.

(a) Write its nuclide notation.

(b) How many neutrons does it have?

(c) How many electrons does it have?

Solution

(a) $^{13}_{6}C$

(b) Number of neutrons = 13 – 6 = 7

(c) Number of electrons = atomic number = 6

Many elements that occur naturally usually display isotopy. The most abundant isotope of an element is taken to be the representative of that element. This abundance is usually given in percentage. Examples of common elements that display isotopy are hydrogen, oxygen, carbon, chlorine, nitrogen, and neon (Table 1.3).

Table 1.3: *Examples of isotopes and their abundances*

Element	Chemical symbol	Atomic number	Isotopes	Abundance
Hydrogen	H	1	$^{1}_{1}H$ (protonium or hydrogen)	99.99%
			$^{2}_{1}H$ (deuterium)	0.01%
			$^{3}_{1}H$ (tritium)	Very rare
Carbon	C	6	$^{12}_{6}C$	98.9%
			$^{13}_{6}C$	1.1%
			$^{14}_{6}C$	Trace
Chlorine	Cl	17	$^{35}_{17}Cl$	75%
			$^{37}_{17}Cl$	25%
Oxygen	O	8	$^{16}_{8}O$	99.76%
			$^{17}_{8}O$	0.04%
			$^{18}_{8}O$	0.20%
Neon	Ne	10	$^{20}_{10}Ne$	90.5%
			$^{21}_{10}Ne$	0.3%
			$^{22}_{10}Ne$	9.2%
Nitrogen	N	7	$^{14}_{7}N$	99.6%
			$^{15}_{7}N$	0.4%
			$^{13}_{7}N$	Very rare

Applications of isotopes

Isotopes are special forms of elements that can be used in various areas, including research, medicine, industry, and agriculture. Some isotopes are *radioactive*. Radioactive isotopes are special types of atoms that give off energy called radiation. Even though the emitted radiations are dangerous if not handled properly, these isotopes are very useful in many ways. Scientists use these isotopes in various

applications, including determining the age of ancient objects using carbon-14, treating diseases in hospitals, tracking how plants absorb nutrients in the soil, and producing electricity for everyday uses.

Carbon dating: Finding the age of ancient items

Scientists use carbon-14 to determine the age of ancient objects. Carbon-14 is a radioactive isotope of carbon that slowly breaks down over time. When a plant or animal dies, it stops taking in carbon-14 from the air. Scientists measure how much carbon-14 is left in bones, wood, or fossils to estimate the number of years that have passed since the organism died. Carbon dating helps archaeologists and historians learn about the past, including the age of ancient human tools, animal fossils, and historical items.

Tracers in medicine, industry, and agriculture

Some radioactive isotopes are used as tracers that help scientists track movements or processes inside the body, in the environment, or in industrial systems.

(a) Medicine: Diagnosing and treating diseases

Iodine-131 is used in hospitals to check how the thyroid gland works. When a small dose of iodine-131 is administered to a patient, the radiation emitted by the iodine helps to create images of the thyroid. Special machines detect the radiation to help diagnose thyroid problems. The radiation also helps to shrink or destroy damaged thyroids cells leading to the treatment of thyroid related diseases.

(b) Agriculture: Studying how plants absorb nutrients

Phosphorus-32 is used to track how plants take in nutrients from the soil. Scientists use it to improve fertilisers and help farmers grow healthier crops.

(c) Industrial and environmental applications

Some isotopes, such as tritium -3 tracks how water moves underground. This allows scientists to understand the sources of drinking water. Chlorine isotopes are used to study the movements of chlorine in water sources such as rivers, lakes, and underground water. Sodium-24 is used to detect leaks in underground pipes. This isotope is made artificially.

Exercise 1.2

1. Complete the following table by filling in the number of protons, electrons, and neutrons of the atoms. The atomic numbers and mass numbers are given.

Atom	Atomic number	Mass number	Protons	Electrons	Neutrons
Sodium	11	23			
Oxygen	8	16			
Beryllium	4	9			
Fluorine	9	19			

2. An atom has 10 electrons.
   (a) Show how these electrons would be arranged in the shells around the nucleus.
   (b) Why is it important to know the arrangement of electrons around the nucleus?
3. Carbon has multiple atomic masses, carbon-12, carbon-13 and carbon-14. Why do these forms of carbon have different mass numbers and still belong to the same element?
4. How are isotopes identified based on their atomic number and mass number?
5. Provide two ways in which the knowledge of atomic numbers is useful in chemistry.
6. Write the nuclide notation of an atom with 7 protons, 7 neutrons, and 7 electrons and explain its significance.
7. Isotopes of an element are chemically identical but physically different. Explain.

Relative atomic mass

Task 1.7

Conduct a library search to find the differences between atomic mass and relative atomic mass.

An atom is very small and it would be difficult to measure its actual mass. To overcome this difficulty, chemists developed a simpler way to express the mass of an atom. This involved expressing the mass of an atom in relation to a chosen standard atomic mass. The carbon atom was chosen as the standard atom (reference atom) and its mass was arbitrarily chosen as 12 units (not actual value). Then, using an instrument called a mass spectrometer, all the other atoms were compared to this standard atom. This reference is called the *Carbon-12 scale*. For example, it was found that the:

(a) magnesium atom was twice as heavy as the reference atom; so its mass was put at 24.

(b) hydrogen atom was $\frac{1}{12}$ as heavy as the reference atom; so its mass was put at 1.

(c) helium atom was $\frac{1}{3}$ as heavy as the reference atom; so its mass was put at 4.

The mass of an atom obtained by comparing it with the arbitrary mass of a carbon-12 atom is called its *relative atomic mass* (R.A.M. or A_r). The relative atomic mass of an element is the average mass of one atom of the element relative to $\frac{1}{12^{th}}$ the mass of one atom of carbon-12. Therefore, R.A.M. may not necessarily be a whole number.

$$\text{That is, } A_r = \frac{\text{Average mass of atom of an element}}{\frac{1}{12^{th}}\text{ the mass of carbon-12 atom}}$$

Table 1.4 gives the atomic numbers and relative atomic masses of the first 20 elements in the periodic table. The relative atomic masses of such elements are obtained by calculating the average mass of all the isotopes of each element.

For isotopic elements, the relative atomic mass (R.A.M.) can be calculated using the following formula:

Relative atomic mass (R.A.M.) = $\sum$ isotopic mass × percentage abundance

Note: $\sum$ = Summation

Table 1.4: *Atomic numbers and relative atomic masses of some elements*

Element	Atomic number	Relative atomic mass
Hydrogen	1	1
Helium	2	4
Lithium	3	6.9
Beryllium	4	9
Boron	5	10.8
Carbon	6	12
Nitrogen	7	14
Oxygen	8	16
Fluorine	9	19
Neon	10	20.2
Sodium	11	23
Magnesium	12	24.3
Aluminium	13	27
Silicon	14	28.1
Phosphorus	15	31
Sulfur	16	32
Chlorine	17	35.5
Argon	18	39.9
Potassium	19	39.1
Calcium	20	40.1

Example 1.5

(a) Chlorine has two isotopes:

$^{35}_{17}Cl\ (75\%)$ and $^{37}_{17}Cl\ (25\%)$

The relative atomic mass of chlorine is: $\left(35 \times \frac{75}{100}\right) + \left(37 \times \frac{25}{100}\right)$

$$= \frac{2625 + 925}{100}$$

$$= \frac{3550}{100} = 35.5$$

(b) Neon has three isotopes:

$^{20}_{10}Ne\,(90.5\%),\ ^{21}_{10}Ne\,(0.3\%),\ ^{22}_{10}Ne\,(9.2\%)$

The relative atomic mass of neon is:

$$\left(20 \times \frac{90.5}{100}\right) + \left(21 \times \frac{0.3}{100}\right) + \left(22 \times \frac{9.2}{100}\right)$$

$$= \frac{1810 + 6.3 + 202.4}{100}$$

$$= \frac{2018.7}{100}$$

$$= 20.187$$

It can be noted that for both chlorine and neon, the R.A.M. is very close to the mass number of the isotope with the highest abundance, namely $^{35}_{17}Cl$ and $^{20}_{10}Ne$, respectively.

Chapter summary

1. An atom is the smallest particle of an element. It can only be split or destroyed by nuclear reactions.
2. There are three major sub-atomic particles, namely:
   (a) protons (positively charged),
   (b) neutrons (neutral), and
   (c) electrons (negatively charged).
3. Protons and neutrons are located in the nucleus of an atom while electrons are found in the shells or energy levels around the nucleus.

Next

4. The arrangement of electrons in different shells of an atom is known as electronic arrangement or electronic configuration.
5. Each shell can contain only a certain number of electrons, with the maximum being $2n^2$, where n is the position of the shell from the nucleus.
6. For any neutral atom of an element:
   Number of protons = atomic number
   Number of electrons = number of protons = atomic number
   Number of neutrons = mass number – atomic number
7. Isotopes are atoms of the same element with the same number of protons but different number of neutrons.
8. The relative atomic mass of an element is the average mass of one atom of the element relative to $\frac{1}{12^{th}}$ the mass of one carbon-12 atom.

Revision exercise 1

Choose the correct answer for Questions 1–8. For other questions, provide the answers as per the demands indicated.

1. Which description corresponds to a proton?
   (a) Relative mass = 1, charge = + 1
   (b) Relative mass = 1, charge = 0
   (c) Relative mass = $\frac{1}{1840}$, charge = –1
   (d) Relative mass = 4, charge = +2.
2. What are nucleons?
   (a) Neutrons and electrons
   (b) Neutrons and protons
   (c) Electrons and protons
   (d) Protons, neutrons and electrons
3. Which atomic property is the basis for the use of carbon-14 in archaeology for dating ancient objects?

   (a) Electronic configuration
   (b) Radioactive decay
   (c) Atomic mass
   (d) Relative atomic mass
4. Most atoms are neutral because they have
   (a) nucleus which is only made up of neutrons.
   (b) equal numbers of electrons and protons in the shells.
   (c) neutrons with zero charge.
   (d) the number of electrons which balance out the number of protons in the atom.
5. Which of these statements is true about isotopes of an element?
   (a) Contain the same number of protons, but different number of neutrons.
   (b) The number of neutrons is the same, but the number of protons is different.
   (c) Proton and neutron numbers are the same, but those of electrons are different.
   (d) Electrons are added to the nucleus,with the same number of protons.
6. Which sub-atomic particles are equal in number to protons in a neutral atom?
   (a) Electrons
   (b) Neutrons
   (c) Electrons of its ion
   (d) Neutrons of its ion
7. The atomic number for an element T is 9. What is its electronic configuration?
   (a) 2: 4: 3 (b) 2: 5: 2
   (c) 2: 7 (d) 2: 6: 1
8. In medical diagnosis and treatment, isotopes are used in
   (a) detecting malaria parasites by using special scans.
   (b) using radioactive tracers for imaging and treating cancer.
   (c) treating typhoid infections by using antibiotics.
   (d) determining a patient's age by checking body cells.

9. Copy the following table and fill the missing details of the arbitrary elements given.

Element	Atomic number	Mass number	Number of protons	Number of electrons	Electronic configuration
P	17	35			
Q		40	20		
R		12			2: 4
S		9		4	
T	3	7			
U		23			2: 8: 1
V		31	15		

10. State the number of protons, neutrons and electrons in the following atoms:
(a) $^{27}_{13}Al$ (b) $^{137}_{56}Ba$ (c) $^{90}_{38}Sr$ (d) $^{235}_{92}U$
11. An isotope of neon has a mass number of 21 and an atomic number of 10.
    (a) Write its nuclide notation.
    (b) How many neutrons does it have?
    (c) How many electrons does it have?
12. A sugar grain is made up of several atoms. Explain.
13. Form One students argue that atoms are visible to the naked eye because all objects are made of atoms. Explain this misconception using the atomic theory.
14. Medical doctors sometimes use X-rays to view patients' bodies. These rays interact mainly with atoms in the patient's bones.
    (a) Which sub-atomic particle is responsible for this interaction?
    (b) Why do you think this particle is important in medical imaging?
15. Describe the role of carbon-14 in determining the age of fossils.
16. Why are the relative atomic masses of elements rarely whole numbers?
17. Why is the relative atomic mass considered more practical than actual atomic mass?

Chapter Two
Periodic classification

Introduction

Many elements have been discovered. Therefore, studying each element individually is challenging. The elements are better understood and managed when they are classified. The most effective way to classify them is by using a table known as the periodic table. In this chapter, you will learn about the development of the periodic table, electronic configuration and element positioning in the periodic table, and changes in physical and chemical properties across periods and down the groups. The competencies developed will provide a foundation for studying elements and predicting how chemicals behave and their applications in everyday life.

Think

Periodic classification is the backbone of chemistry applications.

Development of the periodic table

Task 2.1

Utilise online or any reliable resources to search the dynamic changes that led to the development of the modern periodic table.

For a long time, chemists have used various ways of grouping elements with similar properties. The simplest of these has been classifying elements as either metals or non-metals. In 1866, a British chemist, John Newlands, thought of the idea of arranging elements in order of their increasing atomic masses. Newlands arranged the elements according to Table 2.1.

Table 2.1: *Newlands' first arrangement of elements*

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

Newlands noticed that an element tends to display characteristics similar to the 8th element in front of it. He arranged the elements in columns according to a law he called the *Law of Octaves* (Table 2.2). However, his classification was unfortunate since he grouped together certain elements which had different characteristics. For example, oxygen (O) was placed in the same group as iron (Fe) and sulfur (S). Newlands' ideas were therefore rejected by many scientists.

Table 2.2: *Newlands' octaves of elements*

H	Li	Be	B	C	N	O
F	Na	Mg	Al	Si	P	S
Cl	K	Ca	Cr	Ti	Mn	Fe

A Russian chemist, Dimitri Mendeleev, later improved Newlands' ideas and convinced other chemists to use them. He intended to illustrate recurring trends (periodic trends) in the properties of elements. In 1869, Mendeleev summarised his *Periodic Law* which states that *the properties of elements are a periodic function of their relative atomic masses.* He arranged elements according to their increasing atomic masses and by the similarity of properties. This resulted in an early version of the periodic table of elements (Table 2.3). A vertical column of elements is called a *group* and a horizontal row is called a *period.* He, however, left gaps in the table predicting that there were existing elements yet to be discovered.

Table 2.3: *Part of Mendeleev's periodic table*

Group / Period	1	2	3	4	5	6	7	8
1	H							
2	Li	Be	B	C	N	O	F	
3	Na	Mg	Al	Si	P	S	Cl	
4	K	Ca	–	Ti	V	Cr	Mn	Fe, Co, Ni
5	Cu	Zn	–	–	As	Se	Br	

Modern periodic table

The modern periodic table is a systematic arrangement of elements in order of increasing atomic numbers. It is a result of several modifications to Mendeleev's periodic table as new elements were discovered and scientific theories advanced to explain the chemical behaviour of elements. Unlike Mendeleev's Periodic Law, which was based on atomic masses, the modern periodic table is structured according to atomic numbers. Therefore, it led to the development of the modern Periodic Law which states that *the properties of elements are a periodic function of their atomic numbers.* This means that elements exhibit recurring (periodic) chemical properties when arranged in increasing atomic number, a phenomenon known as periodicity. For example, lithium, a very reactive metal, has one electron in its outermost shell. The eighth element after lithium is sodium, which shares many of the same characteristics as lithium, including a single electron in its outermost shell and high reactivity. Similarly, the eighth element after sodium is potassium, which also exhibits the same properties. This periodic recurrence of chemical behaviour is among the fundamental principles of the modern periodic table.

Table 2.4 shows the first twenty elements in the periodic table. The full periodic table is given in Appendix 1.

Note: In the periodic table, groups are shown using Roman numerals from I to VIII/0. These numbers indicate how many electrons are in the outermost shell of an element. Using Roman numerals helps make it easier to understand how elements react and form bonds. However, the modern International Union of Pure and Applied Chemistry (IUPAC) system, adopted worldwide, numbers groups from 1 to 18 consecutively. This standardised format, shown in Appendix 1, aims to facilitate group identification and promote consistency in educational and scientific contexts.

Table 2.4: *Position and electronic arrangements of the first twenty elements in the periodic table*

Groups	I	II	III	IV	V	VI	VII	VIII/0
Period 1	$_1H$ 1							$_2He$ 2
Period 2	$_3Li$ 2: 1	$_4Be$ 2: 2	$_5B$ 2: 3	$_6C$ 2: 4	$_7N$ 2: 5	$_8O$ 2: 6	$_9F$ 2: 7	$_{10}Ne$ 2: 8
Period 3	$_{11}Na$ 2: 8: 1	$_{12}Mg$ 2: 8: 2	$_{13}Al$ 2: 8: 3	$_{14}Si$ 2: 8: 4	$_{15}P$ 2: 8: 5	$_{16}S$ 2: 8: 6	$_{17}Cl$ 2: 8: 7	$_{18}Ar$ 2: 8: 8
Period 4	$_{19}K$ 2: 8: 8: 1	$_{20}Ca$ 2: 8: 8: 2						

Exercise 2.1

1. Why were the noble gases not included in Mendeleev's periodic table?
2. How does the electronic configuration of nitrogen compare to that of phosphorus?
3. Explain the significance of the Law of Octaves in understanding periodic trends.
4. Mendeleev left gaps in his periodic table. Predict how these gaps demonstrate his understanding of properties of elements.
5. Mendeleev used atomic masses to arrange elements. If you were to use atomic numbers, how would this impact the organisation of the periodic table?
6. How did the Mendeleev's and Newlands' works contribute to understanding the relationships between elements' physical and chemical properties.

Electronic configuration and element positioning in the periodic table

Task 2.2

Use reliable online resources to analyse the relationships between electronic configurations and the positions of elements in the periodic table.

The arrangement of electrons in an atom follows a specific pattern known as electronic configuration. This configuration determines the positioning of elements in the periodic table and their properties. Elements with similar valence electron configurations share common properties and belong to the same group in the periodic table. Elements are arranged in groups based on the number of electrons in their outermost shells. Elements in the same group exhibit similar chemical properties because they have the same number of *valence electrons*. Valence electrons are the outermost electrons of an atom that are involved in forming chemical bonds. The groups are labelled using Roman numerals (I to VIII), with the group number corresponding to the number of electrons in the outermost shell as shown in Figure 2.1.

The elements in Group I (except hydrogen) are called, *alkali metals* because they react with water to form an alkaline solution. These elements are lithium (Li), sodium (Na), and potassium (K), each having one electron in the outermost shell. Note that, even though hydrogen is placed in Group I, it is not an alkali metal. It carries some properties which are similar to those of Group I elements and others which are similar to those of Group VII elements. Therefore, some periodic tables such as Newlands' place it in the same group as Group VII elements, such as fluorine and chlorine.

Group II elements are called alkaline earth metals. They have properties similar to those of Group I elements, but they are less reactive. Group II elements also form alkaline solutions when they react with water. Their oxides are stable, insoluble solids that historically were referred to as "earths," giving the term earth

Group VII elements are called *halogens* ('salt formers' in Greek) because they react with metals to form compounds called salts. Group VIII elements are usually referred to as *Group 0* elements. They are called *noble gases* and all their shells are completely filled with electrons. They were formerly called *inert gases* because they do not readily react to form compounds. The elements in the block between Group II and Group III are *transition elements*. These elements have high densities and melting points, form coloured compounds, and often act as catalysts. Metalloid elements such as boron (B), silicon (Si), and germanium (Ge) exhibit both metallic and non-metallic properties.

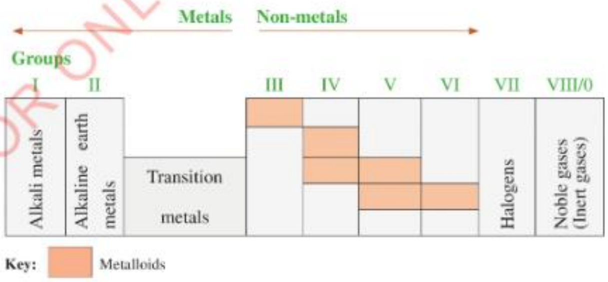


Figure 2.1: *Sketch of the periodic table showing groups of elements*

Elements are also arranged into periods, which indicate the number of electron shells in an atom. Table 2.4 shows that elements with the same number of shells belong to the same period. Periods are numbered from 1 to 7. The electronic configuration entails the number of shells for each element. Periods and number of shells of the first twenty elements are shown in Table 2.5.

Table 2.5: *Periods and number of shells of the first twenty elements*

Period	Elements	Number of shells
Period 1	Hydrogen and helium	1
Period 2	Lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine and neon	2
Period 3	Sodium, magnesium, aluminium, silicon, phosphorus, sulfur, chlorine and argon	3
Period 4	Potassium and calcium	4

Electronic configurations help to explain the positioning of elements in the periodic table. Elements in the same group share the same number of valence electrons, while elements in the same period have the same number of electron shells. This arrangement forms the basis for predicting chemical behaviour and classifying elements efficiently.

Exercise 2.2

1. Write the electronic configuration and indicate the *group* and *period* for each element in the table.

Element	Electronic configuration	Group	Period
Carbon			
Phosphorus			
Beryllium			
Oxygen			
Lithium			
Fluorine			
Silicon			

Element	Electronic configuration	Group	Period
Magnesium			
Aluminium			
Potassium			
Nitrogen			
Sulfur			

2. Argon and sodium are in the same period but behave differently. Use their electronic configurations to justify their differences.
3. How does an element's electronic configuration determine its position in the periodic table? Provide examples.
4. Why do elements in the same group of the periodic table have similar chemical properties?
5. Predict the position of an element with the electronic configuration 2: 8: 3 in the periodic table and justify your answer.

Task 2.3

Sketch a modern periodic table on manila paper using a pencil and ruler, ensuring clear rows for periods and columns for groups. Use different coloured marker pens to assign a distinct colour to each group. Then, fill each block with the chemical symbol, atomic number, and relative atomic mass of the corresponding elements. Select one element from each group and write its electronic configuration below its corresponding block to illustrate its relationship with its position in the periodic table.

Trends in physical and chemical properties across periods and down groups

Task 2.4

Utilise reliable online resources to search and analyse trends in the periodic table and describe how the properties of elements vary across the periods and down the groups.

The periodic table is a systematic arrangement of elements that helps in understanding patterns in their physical and chemical properties. Elements in the same row (period) and the same column (group) exhibit trends in *melting points, boiling points, density, electronegativity, ionisation energy, atomic size, and reactivity.* Recognising these trends is essential for applications in various settings, including home settings and industries such as material manufacturing, energy storage, and healthcare.

Melting point is the temperature at which a solid turns into a liquid. For example, ice melts into water when exposed to heat, such as on a hot day. On the other hand, boiling point is the temperature at which a liquid changes into a gas, for instance water boiling and turning into steam when heated on a stove. Density, another important property, refers to the mass per unit volume of a substance and reflects its compactness.

Electronegativity is the tendency of an atom to attract shared electrons in a chemical bond toward itself. This property plays a crucial role in interactions, such as those between the ions of salt and water molecules when the salt dissolves in water. On the other hand, ionisation energy is the energy required to remove an electron from an atom or ion in its gaseous state. Another fundamental property is the atomic size or radius, which measures the distance from an atom's nucleus to its outermost electron shell.

Reactivity refers to the ability of an atom of a given element to interact chemically with other substances. For instance, alkali metals (Group 1) are highly reactive due to their tendency to lose one electron, whereas noble gases (Group VIII) exhibit minimal reactivity because of their stable electronic configurations. Understanding these properties and their trends helps to explain the behaviour of elements in reactions and their industrial applications.

Trends in physical properties across periods

Task 2.5

Use an interactive simulation or any reliable resources to explore the trends in physical properties across periods of the periodic table. Analyse these trends and explain their practical applications in real-life scenarios.

The following are changes in physical properties across periods:

(a) Decrease in atomic radii

Atomic radii decrease from left to right across a period due to the increasing nuclear charge, which pulls the electrons closer to the nucleus. This trend contributes to designing strong and durable materials in construction and manufacturing, as smaller atomic sizes promote stronger atomic bonding and improved structural strength.

(b) Transition from metallic to non-metallic properties

Metals dominate the left side of the periodic table, while non-metals are positioned on the right. On moving across a period, metallic properties decrease, and non-metallic properties increase. This trend influences material selection, with metals such as iron and aluminium serving as structural materials due to their strengths, while non-metals such as carbon, silicon, sulfur, and chlorine in plastics and rubber function as insulating materials because of their poor conductivity.

(c) Variations in physical states

Elements exist in different states at room temperature (25 °C). Metals such as aluminium and iron remain solid, elements such as oxygen and nitrogen exist in gaseous form, and a few, including bromine, are liquids. This trend is essential for safe storage and transportation. For example, gases require pressurised cylinders, while solids remain stable in containers.

(d) Variations in melting and boiling points

Across a period, melting and boiling points generally increase at first due to stronger metallic or covalent bonding. However, they begin to decrease after Group IV as elements start to form molecules held together by weaker intermolecular forces. Typically, the melting and boiling points increase from Group I to Group IV, then decrease from Group V and stay relatively low to Group VIII.

This trend has several practical applications in everyday life and industry. For example, metals such as sodium (Na), magnesium (Mg), and aluminium (Al), which lie between Groups I and III, have relatively high melting points due to strong metallic bonding. These properties make them suitable for use in cooking utensils, building materials, and aircraft parts where heat resistance is important. Elements like silicon (Si) in Group IV, with a giant covalent structure and very high melting point, are essential in the manufacture of computer chips and solar panels. On the other hand,

elements found in Groups V to VIII, such as nitrogen (N_2), oxygen (O_2), and neon (Ne), exist as simple molecules with low melting and boiling points. These are commonly used in gas form; for example, oxygen is used in hospitals for respiration, nitrogen in food preservation, and neon in lighting systems.

(e) Increase in ionisation energy

Ionisation energy increases from left to right across a period due to stronger nuclear attraction, making it harder to remove electrons from an atom. Elements with high ionisation energy, such as noble gases, are used in lighting and insulation due to their chemical stability.

Trends in chemical properties across periods

The following are changes in chemical properties across periods:

(a) Increase in electronegativity

Electronegativity increases from left to right across a period, with non-metals attracting electrons more strongly. This is due to the stronger nuclear attraction resulting from the increasing number of protons in the nucleus. This trend is particularly crucial in various applications, such as semiconductor technology, where elements like silicon and germanium are widely used in electronic devices.

(b) Decrease in metallic character

Metallic characters decrease, and non-metallic characters increase across a period. Metals tend to lose electrons, while non-metals gain electrons. This trend is important in various applications such as battery production, where metals act as electron donors.

(c) Variations in chemical reactivity

Metals on the left are highly reactive and lose electrons easily, while non-metals on the right become more reactive in gaining electrons. For metals, reactivity decreases from left to right, while for non-metals, reactivity increases from left to right. This trend plays crucial roles in different activities such as drug formulation and material design, ensuring the stability and effectiveness of compounds in pharmaceuticals and engineering.

Trends in physical properties down a group

(a) Atomic size increases down a group as more electron shells are added. This expansion causes atoms to become larger, affecting their physical behaviour. Larger atomic size influences material performance under high-pressure

environments, making certain elements suitable for deep-sea applications and industrial machineries.

(b) Density also increases down a group as atomic mass increases more significantly than the atomic volume. Heavier elements tend to have stronger structural properties, making them valuable in industries that require durability and strength. These elements play crucial roles in construction and aerospace engineering, where materials must withstand extreme conditions.

(c) Melting points for metals generally decrease down a group due to weaker metallic bonding. As atomic sizes increase, the attraction between metal atoms weakens, reducing the energy required to melt the substances. This trend is significant in the design of alloys for safety devices such as fuses, which need to melt easily to prevent electrical hazards.

(d) Ionisation energy decreases down a group in the periodic table. This is because as atomic sizes increase, the outermost electrons are farther from the nucleus. As a result, the attractions between the nucleus and outer electrons become weaker, making it easier for metals to lose electrons. The trend in ionisation energy influences an element's reactivity, the types of compounds it forms, its electrical conductivity, and its role in biological and industrial processes such as metallurgy, battery design, and semiconductor manufacturing.

Trends in chemical properties down a group

(a) Electronegativity decreases down a group. This is because, as atomic size increases, the outer electrons are farther from the nucleus. This condition reduces the attractions between the nucleus and electrons, making atoms less able to attract electrons in a chemical bond.

(b) Metallic character increases down a group. These properties are crucial in selecting metals for use in catalysts and chemical processing industries. Reactivity trends differ between metals and non-metals as you move down a group. Metals become more reactive because they are more likely to lose electrons, which is advantageous in processes such as metal extraction from ores. In contrast, non-metals become less reactive because their ability to attract electrons weakens. This behaviour is applied in industries where reactive metals facilitate the breakdown of substances as in cleaning agents, while non-metals help stabilise compounds to prevent undesired reactions.

Trends in selected groups

Groups I, II, and VII elements exhibit distinct physical and chemical properties due to their unique positions in the periodic table. Group I (alkali metals) and

Group II (alkaline earth metals) are highly reactive metals, while Group VII (halogens) consists of reactive non-metals. Their physical properties, such as atomic radius, ionisation energy, density, and melting points, exhibit clear trends within each group, as shown in Tables 2.6 and 2.7. Chemically, they differ in reactivity and electronegativity, reflecting variations in their electronic configurations. Understanding these properties provides valuable insights into their behaviour and applications in chemistry and other fields.

Group I: Alkali metals

Group I consists of metals such as lithium (Li), sodium (Na), potassium (K), rubidium (Rb), and caesium (Cs) as shown in the Table 2.6. The data from this table show that the physical and chemical properties of these elements generally increase or decrease down the group. Each of these elements has one electron in its outermost shell. Lithium, sodium and potassium react very readily with water or air and are stored in oil.

Table 2.6: *Trends in Group I*

Name (Symbol)	Atomic number (z)	Electronic configuration	Atomic radius (picometres)	1^{st} ionisation energy (kJ /mol)	Melting point (°C)	Density (g/cm^3)	Electro-negativity
Lithium (Li)	3	2:1	152	526	180	0.54	1.0
Sodium (Na)	11	2:8:1	186	504	98	0.97	0.9
Potassium (K)	19	2:8:8:1	231	425	64	0.86	0.8
Rubidium (Rb)	37	2:8:18:8:1	244	410	39	1.5	0.8
Caesium (Cs)	55	2:8:18:18:8:1	262	380	29	1.9	0.7

Note: 1. Francium (Fr) is also an alkali metal but is rarely discussed or experimented in the discussion or experiments involving Group I elements due to its radioactive nature. It is also among the rarest naturally occurring elements.

2. The way electronic configurations of rubidium and caesium are written, is above the scope of this book. However, for the sake of understanding this section, the electronic configurations are written to show the period number, group number, and their relations to the trends.

The Group I elements have the following properties:

Physical properties

(a) They are good conductors of heat and electricity.

(b) They are soft metals.

(c) They have low density.

(d) They have shiny surfaces when freshly cut.

Chemical properties

(a) They burn in oxygen or air with a characteristic flame colour to form white solid oxides. These oxides dissolve in water to form alkaline solutions of the metal hydroxides.

$$\text{Metal + Oxygen} \longrightarrow \text{Metal oxide}$$

(b) They react vigorously with water to give alkaline solutions and hydrogen gas.

$$\text{Metal + Water} \longrightarrow \text{Metal hydroxide + Hydrogen gas}$$

Task 2.6

Utilise educational resources to explore the practical applications of Group I elements such as lithium in batteries or sodium in table salt. Prepare a poster or presentation linking their physical and chemical properties (for example, reactivity and conductivity) to everyday uses.

Activity 2.1

Aim: To investigate the reactivity of Group I elements (alkali metals) down the group

Requirements: Small pieces (equivalent to the head of a matchstick) of lithium (Li), sodium (Na), and potassium (K), three beakers, water, phenolphthalein or universal indicator, forceps or tongs, watch glass, dropper, safety goggles and gloves

Caution: Handle Group I elements cautiously by using small quantities, wearing protective gear, working in a controlled environment, avoiding direct contact, and keeping a fire extinguisher nearby to prevent accidents during their vigorous reactions with water.

Procedure

1. Fill three beakers halfway with water.
2. Add a few drops of phenolphthalein or universal indicator to each beaker.
3. Using forceps, carefully drop a small piece of lithium into the first beaker and observe the reaction.
4. Repeat Step 3 for sodium and potassium in separate beakers.
5. Record observations for each metal, focusing on the rate of reaction, colour change (due to the indicator), production of gas and sound.
6. Clean up all the apparatus safely under the supervision of the teacher.

Questions

1. Which element was the fastest to react with water?
2. How did the colour of water change and what did it indicate?
3. What trend did you observe in the reactivity of those elements down the group?
4. How does this experiment support the idea of periodic classification?

Group II: Alkaline earth metals

Group II consists of elements such as beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra) as shown in Table 2.7. The data from this table show that the physical and chemical properties of these elements generally increase or decrease down the group. These elements have two electrons in their outermost shells.

Table 2.7: *Trends in Group II elements*

Name (Symbol)	Atomic number (z)	Electronic configuration	Atomic radius (picometres)	1^{st} ionisation energy (kJ/mol)	Melting point (°C)	Density (g/cm^3)	Electro-negativity
Beryllium (Be)	4	2: 2	112	899	14849	1280	1.5
Magnesium (Mg)	12	2: 8: 2	160	738	7730	651	1.2
Calcium (Ca)	20	2: 8: 8: 2	197	590	4741	851	1.0
Strontium (Sr)	38	2: 8: 18: 8: 2	215	549	4207	800	1.0
Barium (Ba)	56	2: 8: 18: 18: 8: 2	217	503	3420	850	0.9

Note: The way electronic configurations of strontium and barium (Table 2.7) are written is beyond the scope of this book. However, for the sake of understanding this section, the electronic configurations are written to show the period number, group number, and their relations to the trends.

Group II elements have the following properties:

Physical properties

(a) They are harder metals than those in Group I.

(b) They are silvery-grey in colour when pure and clean. However, they tarnish quickly when left in the air due to the formation of the respective metal oxides.

(c) They are good conductors of heat and electricity.

Chemical properties

(a) They burn in oxygen or air with a characteristic flame colour to form a solid white oxide.

$$\text{Metal + Oxygen} \longrightarrow \text{Metal oxide}$$

(b) They react with water but much less vigorously than the elements in Group I.

$$\text{Metal + Water} \longrightarrow \text{Metal hydroxide + Hydrogen gas}$$

(c) The reactivity of metals increases down the group. For example, the reaction of calcium with water is vigorous, while that of magnesium with water is relatively slow.

Activity 2.2

Aim: To demonstrate the differences in reactivity of calcium and magnesium with water

Requirements: Two test tubes, magnesium ribbon, calcium, distilled water, and measuring cylinder

Procedure

1. Transfer about 5 cm^3 of distilled water in a test tube.
2. Add a small amount of calcium (half spatulaful) from the container in which it is stored. Record your observations.
3. Repeat Steps 1 and 2 using a clean piece of magnesium ribbon. Record your observations.

Questions

1. What happens when the powder of calcium and magnesium ribbon are dropped in water?
2. What is the difference in the reactivities of calcium and magnesium?

Group VII: Halogens

Task 2.7

Use online sources or any reliable resources to study trends in physical and chemical properties in Group VII.

Group VII elements include fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). These are highly reactive non-metals. They have the following properties:

Physical properties

(a) They exist in different physical states for example fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid.
(b) Their densities increase down the group.
(c) Their melting and boiling points increase down the group.

Chemical properties

(a) Their reactivity decreases down the group.
(b) They react with metals to form salts.
Metal + Halogen $\longrightarrow$ Metal halide
(c) They react with hydrogen to form acids (hydrogen halides).
Hydrogen + Halogen $\longrightarrow$ Hydrogen halide

Exercise 2.3

1. What happens to atomic size as you move down a group? Explain, giving reasons.
2. Predict how the densities of alkali metals change as you move down Group I.
3. The reactivity of alkali metals is directly related to their atomic size. Justify.
4. Investigate the use of lithium in rechargeable batteries. How do its properties

contribute to this application?

5. How do the electronic configurations of Group I elements determine their reactivity?
6. Compare the solubility of Group II hydroxides in water. How does this trend change down the group?
7. Calcium is essential in human bones and teeth. How do its chemical properties contribute to this biological role?
8. Why are Group II elements less reactive than Group I elements? Use their electronic configurations to support your answer.
9. Why do the reactivities of Group VII elements decrease down the group? Use fluorine and iodine as examples.
10. Explore the changes in physical states observed across Group VII elements. What trend is observed within the group?

Chapter summary

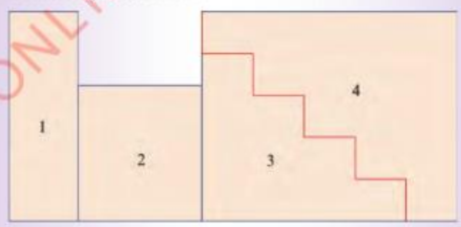
1. The periodic table of elements is a method of displaying chemical elements in a table format. It was developed after several modifications to the Mendeleev's periodic table.
2. Mendeleev's Periodic Law states that *the properties of elements are periodic functions of their relative atomic masses.*
3. The Modern Periodic Law states that *the properties of elements change systematically according to their atomic numbers.*
4. Periodicity refers to the regular periodic changes of properties of elements due to changes in atomic numbers.
5. Elements with the same number of electrons in their outermost shells belong to the same group.
6. The group number signifies the number of electrons in the outermost shell of an element.
7. Elements with the same number of shells belong to the same period.
8. The electronic configuration of an element entails the electronic arrangements in the shell(s), number of shells, and the group to which the element belongs.

Revision exercise 2

Choose the correct answer for Questions 1-7. For other questions, provide the answers as per the demands indicated.

1. Non-metals are generally better ________ than metals.
   (a) conductors of electricity
   (b) brittle materials
   (c) malleable materials
   (d) conductors of heat
2. The electronic arrangement of an element is 2: 3. This element is in _______ of the periodic table.
   (a) Group II
   (b) Group VIII
   (c) Period 3
   (d) Period 2
3. The following statements describe the alkaline earth metals except they
   (a) burn in oxygen to form solid white oxides.
   (b) become less reactive down the group.
   (c) are good conductors of heat and electricity.
   (d) react with water.
4. An element with electronic configuration of 2: 8: 3 belongs to Group ___ and period ____ of the periodic table.
   (a) III ... 2
   (b) III ... 3
   (c) II ... 4
   (d) II ... 5
5. Identify the electronic configuration of an element with 16 electrons.
   (a) 4: 8: 4
   (b) 3: 8: 5
   (c) 2: 8: 6
   (d) 2: 10: 4

6. Hydrogen is placed in Group I elements because it
   (a) is in period 1.
   (b) loses an electron.
   (c) is an inert gas.
   (d) has atomic mass of 1.
7. Group I elements burn in oxygen to form
   (a) metal oxides.
   (b) non-metal oxides.
   (c) hydroxides.
   (d) carbon dioxide.
8. (a) A particular metal reacts slowly with water to give a strong alkaline solution. In which group of the periodic table should it be placed? Give a reason.
   (b) What is common about the noble gases regarding the following properties?
      (i) Electronic arrangements
      (ii) Chemical reactions
9. Why are certain elements in the periodic table referred to as metalloids?
10. The following diagram represents the parts of the periodic table with four areas denoted by 1, 2, 3 and 4:



   (a) Which part contains mostly non-metals?
   (b) Which part contains most elements whose oxides dissolve in water?
   (c) Which part contains transition elements?
   (d) Which part mostly contains elements with both metallic and non-metallic characteristics?

11. (a) Given the elements calcium, sulfur, chlorine, helium and neon, write their:
    (i) period numbers.
    (ii) group numbers.
    (iii) atomic numbers.
    (iv) number of electrons in their atoms.
    (v) electronic configurations.

    (b) Which of the above elements would you expect to have similar properties? Give reasons.

12. During a science lesson, students added a small piece of sodium (Na) into water and observed fizzing and heat. They repeated the same with magnesium (Mg), and there was very little reaction. Using periodic classification, explain why the reaction of sodium with water was vigorous than that of magnesium.

13. Greenish Secondary School is building a simple application (App) that can help Form Two students to identify if an element is a metal, non-metal, or metalloid. What features of periodic classification should be included in the App to help students make the correct decision?

14. A cook wants to buy a new cooking pot. The shopkeeper offers two options: one made from aluminium and the other from a shiny, brittle material. Based on periodic classification, which material is better for cooking pots and why?

15. Form Two students are curious about the elements used in smartphones. They found that lithium (Li) is used in the battery and silicon (Si) in the screen. Based on their positions in the periodic table, explain why these elements are suitable for their specific roles in smartphones.

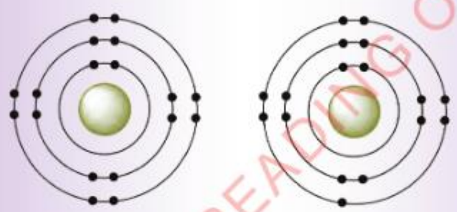
16. Magnesium has the electronic configuration 2: 8: 2. Explain how this configuration relates to its chemical reactivity and its position in the periodic table.

17. Two elements, A and B, have electronic configurations 2: 8: 8 and 2: 8: 2, respectively. Compare their positions in the periodic table and predict which is more reactive. Give reasons for your answers.

18. Predict the position of an element in the periodic table with the electronic configuration 2: 8: 8: 1. Justify your answer. What chemical properties would you expect this element to have?

19. Element R belongs to period 3 and Group VI in the periodic table.
    (a) Draw its atomic structure.
    (b) Determine its atomic number.

20. The following diagrams show the structures of atoms of elements X and Y:



X Y

    (a) Which of the elements is stable?
    (b) Which of the two elements is chemically more reactive? Explain.
    (c) Identify elements X and Y. Give the criteria you used.

21. How does the concept of periodicity in the periodic table aid modern chemists in organising information about chemical elements?

22. What happens when alkali earth metals burn in oxygen?

23. Consider elements $_{11}T$ and $_{17}Q$. Which of the two elements is more electronegative? Explain.

Chapter Three
Chemical bonding, formula and nomenclature

Introduction

Chemical bonding involves holding atoms together to form molecules or compounds. A chemical formula refers to symbols and numbers that represent the composition of a particular chemical substance. Nomenclature means naming. In this chapter, you will learn about the concept of chemical bonding, the concept of chemical formulas, the determination of empirical and molecular formulas of common compounds and the nomenclature of binary inorganic compounds using the IUPAC system. The competencies developed will enable you to explore the relationships between chemical substances and apply the acquired knowledge in understanding properties of different materials used in daily life.

Think

Contribution of chemical bonding, formulas, and nomenclature to global processes and activities

Concept of chemical bonding

Task 3.1

Use videos, interactive simulations, or other reliable resources to explore the concept of chemical bonding.

A *bond* is anything that holds two or more substances together. Many things used in our daily lives are constructed using different materials joined together by some bonds. For example, in a brick wall (Figure 3.1), each brick is joined to the other by a bond made of mortar. Similarly, chemical substances are made of atoms that are held together by chemical bonds.

Figure 3.1: *Bricks bonded using mortar to form a wall*

A *chemical bond* is a force of attraction that holds atoms or ions together to form molecules or compounds. The bond may result from forces of attraction between oppositely charged ions or through the sharing of electrons.

Note: A *molecule* is the smallest particle of an element or compound which can normally exist separately.

Chemical bonding

Chemical bonding involves electrons in the outermost shells of atoms. When the outermost shells are completely filled with electrons, the atoms are said to be stable, otherwise they are unstable. Table 3.1 shows the electronic arrangements of some elements and the stability of atoms of the elements. From Table 3.1, only helium, neon, and argon have stable electronic arrangements. Helium has a maximum number of 2 electrons in its outermost shell. Neon and argon have a maximum of 8 electrons in each of their respective outermost shells. These stable atoms are generally unreactive and can exist freely as single atoms. Unstable atoms cannot exist freely as single atoms. For unstable atoms to become stable, they should acquire electronic arrangements similar to those of noble gases. This means that they can either lose, gain or share electrons through chemical bonding.

Table 3.1: *Electronic arrangements and stability of the first twenty elements of the periodic table*

Element	Number of electrons	Electronic arrangement	Stability of an atom
Hydrogen	1	1	Unstable
Helium	2	2	Stable
Lithium	3	2: 1	Unstable
Beryllium	4	2: 2	Unstable
Boron	5	2: 3	Unstable
Carbon	6	2: 4	Unstable
Nitrogen	7	2: 5	Unstable
Oxygen	8	2: 6	Unstable
Fluorine	9	2: 7	Unstable
Neon	10	2: 8	Stable
Sodium	11	2: 8: 1	Unstable
Magnesium	12	2: 8: 2	Unstable
Aluminium	13	2: 8: 3	Unstable
Silicon	14	2: 8: 4	Unstable
Phosphorus	15	2: 8: 5	Unstable
Sulfur	16	2: 8: 6	Unstable
Chlorine	17	2: 8: 7	Unstable
Argon	18	2: 8: 8	Stable
Potassium	19	2: 8: 8: 1	Unstable
Calcium	20	2: 8: 8: 2	Unstable

Formation of ions

When an atom loses an electron to attain a stable electronic arrangement, the lost electron is transferred to another atom making it stable as well. The resulting species become *ions*. The one that loses an electron becomes a *positively charged ion*, while the one that gains an electron becomes a *negatively charged ion*. The number of electrons gained or lost will be equal to the charge of an ion. The positively charged ion is called a *cation* and the negatively charged ion is called an *anion*. Electrons are negatively charged while protons are positively charged; thus, the charge of the ion is due to the unbalanced number of electrons and protons. For example, a sodium atom has eleven protons and eleven electrons. Its charge is 0 because $(+11) + (-11) = 0$. After sodium loses one electron it will have 10 electrons. Its charge will become +1 because $(+11) + (-10) = +1$ (Figure 3.2).

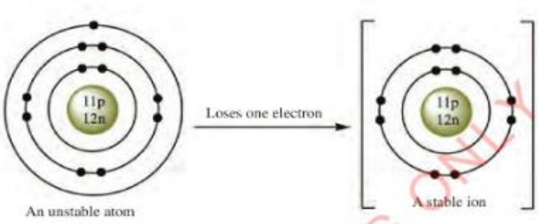


Figure 3.2: *Electron arrangement in sodium atom and its ion*

Electrovalent bonding

Electrovalent bonding, also known as ionic bonding, is a type of chemical bond formed when one atom transfers electrons to another, resulting in the formation of positively and negatively charged ions. These charged ions attract each other due to their opposite charges.

Electrovalent bonding usually occurs between *a metal* and *a non-metal*. The metal loses electron(s) and the non-metal gains electron(s). For example, when sodium and chlorine react to form sodium chloride, the sodium atom must lose an electron to acquire a stable noble gas structure of neon, while the chlorine atom must gain an electron to acquire a stable noble gas structure of argon. This results in a positively charged sodium ion and a negatively charged chloride ion which attract each other to form sodium chloride as shown in Figure 3.3. For this chapter, electrons are represented by dots with different colours.

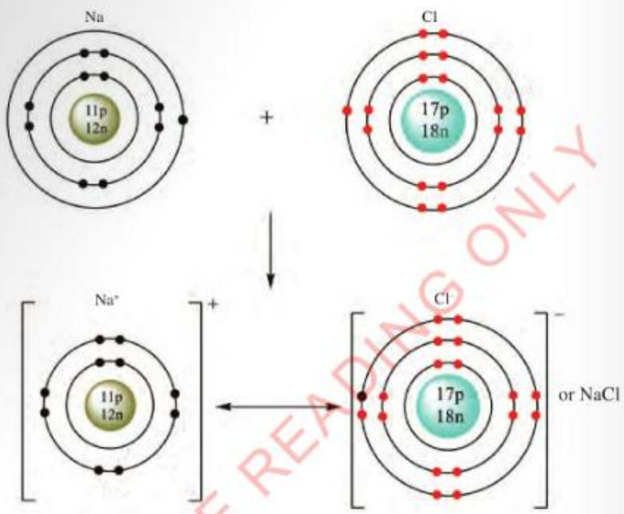


Figure 3.3: *Electrovalent bonding in sodium chloride*

Molecules of electrovalent compounds are *not discrete*. This means that the electrons are not localised or attached to particular ions of opposite charges to form pairs. Instead, a group of cations surround an anion and a group of anions surround a cation. This means that ions can move freely around each other, especially when in molten or in solution forms.

Properties of electrovalent compounds

Electrovalent (ionic) compounds have the following properties:

(a) They are generally soluble in water.

(b) They conduct electricity in solution or molten forms but not in their solid forms.

(c) They are usually crystalline solids at room temperature.

(d) They have high melting and boiling points.

(e) They are generally insoluble in non-polar solvents such as carbon tetrachloride and hexane. Non-polar solvents contain bonds between atoms with similar electronegativities.

(f) They are typically hard but brittle.

(g) They have high densities.

Activity 3.1

Aim: To investigate the electrovalent bonding in substances

Requirements: Table salt (sodium chloride), cooking oil, slaked lime, baking powder, sugar, three glass beakers, distilled water, a tablespoon, an electrical conductivity meter (or a small LED bulb with wires and a battery), and three stirring rods

Procedure

1. Fill a glass beaker with distilled water and add 2 tablespoons of table salt (sodium chloride). Stir with the stirring rod until the salt dissolves completely.
2. Attach the electrical conductivity meter (or set up the bulb circuit) to the solution. Observe whether the bulb lights up or the tester indicates conductivity.
3. Repeat Steps 1 and 2 for the baking powder, cooking oil, slaked lime, and sugar. Use a separate stirring rod for each substance.
4. Record your observations as shown in Table 3.2.

Table 3.2: *Experimental results*

S/N	Substance	Solubility	Conduct electricity	
			Yes	No
1	Table salt			
2	Baking powder			
3	Sugar			
4	Slaked lime			
5	Cooking oil			

Questions

1. Which substance(s) dissolved easily in water? Explain your answer.
2. Which substances are electrovalent? Write their constituent ions.

Exercise 3.1

1. Explain why cations are slightly smaller than their neutral atoms, whereas anions are slightly larger than their neutral atoms.
2. Magnesium and oxygen atoms combine to form magnesium oxide.
   (a) What is the charge on the magnesium ion in the oxide?
   (b) What is the charge on the oxide ion?
   (c) Illustrate the electron transfer using dots.
3. Why NaCl is soluble in water but not in oil, and how does chemical bonding explain this?

Covalent bonding

Covalent bonding occurs when atoms share electrons to achieve a stable electron configuration, typically similar to that of the nearest noble gas. This type of bond forms when the atoms involved have similar electronegativities (the ability to attract electrons). Atoms may gain the noble gas electronic structures without becoming ions but by sharing the outermost electrons. For example, a hydrogen atom needs one electron to acquire the helium electronic configuration. It can combine with another hydrogen atom and share the electrons so that each attain a stable helium configuration. This results in the formation of a hydrogen molecule (Figure 3.4). The kind of the bond formed between the two atoms is called a covalent bond. The *covalent bond* is a chemical bond formed by the sharing of one or more electrons between atoms.

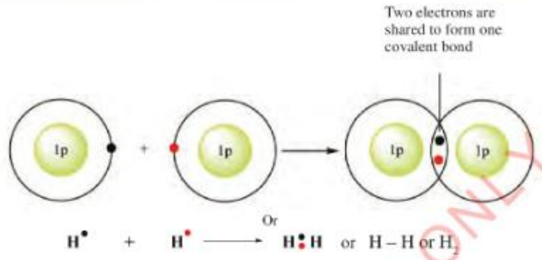


Figure 3.4: *Covalent bonding in a hydrogen molecule*

Note: A single line between two atoms indicates a single covalent bond carrying an electron pair. Two lines are used to represent bonding of two pairs, three lines represent three pairs, and so on.

Oxygen requires two electrons to acquire the stable atomic structure. Thus, two atoms of oxygen combine and share four electrons between them (Figure 3.5).

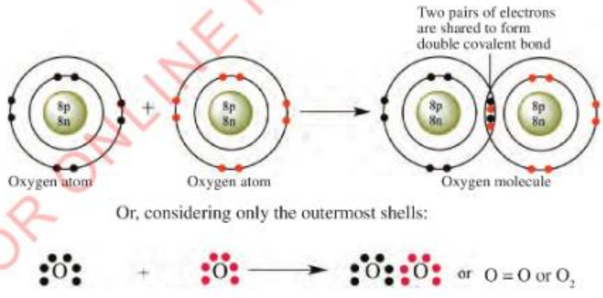


Figure 3.5: *Double covalent bond in oxygen molecule*

Covalent bonding also occurs between atoms of different elements such as between hydrogen and chlorine. Hydrogen requires one electron to attain a stable helium electronic configuration, while chlorine requires one electron to acquire the argon electronic configuration. The two atoms combine by sharing two electrons between them (Figure 3.6).

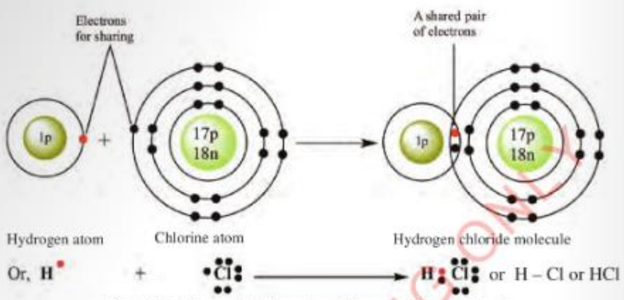


Figure 3.6: *Sharing of electrons in a hydrogen chloride molecule*

The same happens to the combination between oxygen and hydrogen. Oxygen requires two electrons to acquire a stable structure, while hydrogen requires one electron. Thus, one oxygen atom combines with two hydrogen atoms as shown in Figure 3.7, whereby each hydrogen provides one electron for sharing with oxygen (any pair from the outermost shell).

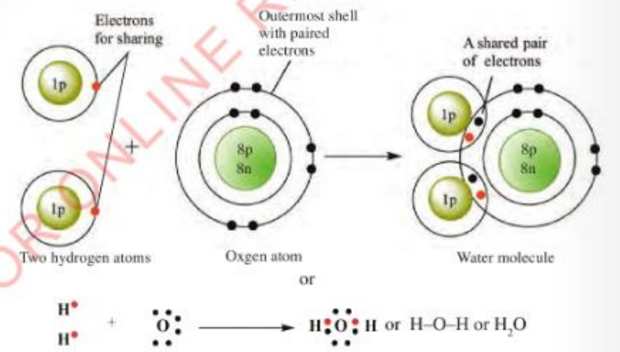


Figure 3.7: *Covalent bonding in a water molecule*

Covalent bonds are usually formed between non-metals. Molecules of covalent compounds are discrete or distinct. This means that the atoms forming the molecules cannot exist freely in the compounds. They remain bound together in molecules and their electrons are not free to form bonds with other atoms.

Properties of covalent compounds

The properties of covalent compounds include the following:

(a) Their melting and boiling points are usually low.
(b) They are usually liquids or gases at room temperatures.
(c) They do not conduct electricity.
(d) They are generally soluble in non-polar solvents such as oil, carbon tetrachloride and hexane.
(e) They are generally insoluble in polar solvents like water.
(f) They can easily vaporise at room temperature.
(g) They have distinct molecular structures.
(h) They are generally non-metallic compounds.

Task 3.2

Use chemical drawing software such as ChemDraw or ChemSketch to draw the chemical structures for sodium chloride and carbon dioxide based on their molecular formulas. Then, identify the type of bond present in each compound.

Activity 3.2

Aim: To demonstrate the formation of covalent bonds in molecules

Requirements: Ball-and-stick molecular model kit (or coloured playdough and toothpicks), manila paper, pens, coloured markers and a pair of compasses

Procedure

1. Use different coloured balls or playdough to make spheres representing atoms of hydrogen, oxygen, nitrogen and carbon.
2. Use two hydrogen atoms and one oxygen atom to model a water molecule (H_2O). Connect the hydrogen atoms to the oxygen atom using sticks or toothpicks.
3. For oxygen gas (O_2), connect two oxygen atoms.
4. Use similar steps to create models of methane (CH_4), carbon dioxide (CO_2), and NH_3.

Questions

1. What does a single stick represent?
2. How do atoms in the models share electrons to achieve stable molecules? Indicate the number of shared electron pairs in each molecular model.

Valency

Valency refers to the ability of an atom of a given element to combine with other atoms, and is measured by the number of electrons that the atom will donate, receive or share to form a chemical bond. It is the *combining power/capacity* of an element or a radical. The combining capacity of an atom of a given element is determined by the number of hydrogen atoms it combines with or displaces. For example, the valency of chlorine is 1 because one atom of hydrogen combines with one atom of chlorine to form hydrogen chloride (HCl). The valency of zinc is 2 because two atoms of hydrogen are displaced from dilute acids by one atom of zinc.

It is easy to predict the valencies of elements from the periodic table. Group I elements have one electron in their outermost shells, and so, their valency is 1. Group II elements have two electrons in their outermost shells, hence their valency is 2. Valencies are rarely above 4. For elements with more than four electrons in the outermost shells, the valency number is usually obtained by subtracting the number of electrons from eight. For example, sulfur with six electrons in the outermost shell has a valency of $8 - 6 = 2$. Those outermost electrons are called *valence electrons*. Some elements have more than one valency. For example, iron has valencies of 2 and 3, copper has valencies of 1 and 2, lead has valencies of 2 and 4, and manganese has valencies of 2, 4 and 7. Table 3.3 shows the valencies of some elements and radicals.

Task 3.3

Use ball and stick models or coloured beads to represent electrons in different shells. Build atoms and determine the valency by observing how many electrons are in the outer shell.

Radicals

Task 3.4

Use a chemistry simulation or software to identify, classify and write different formulas of radicals.

A *radical* is a group of atoms which behaves as a single unit and has a positive or negative charge. It contains at least one unpaired electron. Such a group maintains its identity throughout any chemical reaction. Most radicals form the non-metallic part of a compound, so their ions are negatively charged. Examples are CO_3^{2-} and SO_4^{2-} ions. An exception is for the ammonium radical, NH_4^+, which behaves like the metallic part of a compound and forms a positive ion. The valency of the radical is the same as the numerical value that the group acquires when it loses or gains an electron to form an ion. The common radicals with their formulas and valencies are shown in Table 3.3.

Table 3.3: *Valencies of some elements and radicals*

Category	Valency 1		Valency 2		Valency 3	
	Element	*Ion*	*Element*	*Ion*	*Element*	*Ion*
Metals	Potassium (K) Silver (Ag) Sodium (Na)	K^+ Ag^+ Na^+	Barium (Ba) Calcium (Ca) Iron (Fe) Lead (Pb) Magnesium (Mg) Mercury (Hg) Zinc (Zn)	Ba^{2+} Ca^{2+} Fe^{2+} Pb^{2+} Mg^{2+} Hg^{2+} Zn^{2+}	Aluminium (Al) Iron (Fe)	Al^{3+} Fe^{3+}
Non-metals	Chlorine (Cl) Fluorine (F) Iodine (I)	Cl^- F^- I^-	Oxygen (O) Sulfur (S)	O^{2-} S^{2-}	Nitrogen (N)	N^{3-}
Radicals	Ammonium radical Chlorate Cyanide Hydroxide Hydrogencarbonate Hydrogensulfate Nitrate Nitrite Permanganate	NH_4^- ClO_3^- CN^- OH^- HCO_3^- HSO_4^- NO_3^- NO_2^- MnO_4^-	Carbonate Dichromate Sulfate Sulfite Thiosulfate	CO_3^{2-} $Cr_2O_7^{2-}$ SO_4^{2-} SO_3^{2-} $S_2O_3^{2-}$	Phosphate	PO_4^{3-}

Note: Ammonium radical (NH_4^+) has a valency of 1 and can react like metals. Its compounds are similar to those of Group I elements.

Oxidation state

Oxidation state (also called *oxidation number*) is the total number of electrons that an atom either gains or loses in order to form a chemical bond with another atom. It is the measure of the electron control that an atom has in a compound compared to the atom in the pure element. The neutral atom has no charge.

Rules for assigning oxidation states

The following are the rules used to assign oxidation states of the elements:

1. The oxidation number of free elements is zero. For example, all elements in the periodic table have the oxidation number of zero.
2. The sum of the oxidation states of all atoms forming a molecule or ion is the net charge of that species. For example, nitrogen (N_2), hydrogen (H_2) and oxygen (O_2) molecules have the oxidation number of zero.
3. In simple ions that consist of only one atom, the oxidation number is equal to the charge on the ion. For example, the oxidation number of a sodium ion (Na^+) is +1, aluminium (Al^{3+}) is +3, iron(II) (Fe^{2+}) is +2, and iron(III) (Fe^{3+}) is +3. In an oxide ion (O^{2-}), the oxidation number of oxygen is –2.
4. In their compounds, Group I metals have an oxidation number of +1. Group II metals have an oxidation number of +2, while, Group III metals have an oxidation number of +3.
5. In their compounds, halogens always have an oxidation number of –1.
6. Hydrogen has an oxidation state of +1 in most compounds. The exception is in hydrides of active metals where the oxidation number is –1. For example, the hydrogen atom gains an electron from the lithium atom in lithium hydride (LiH).
7. Oxygen has an oxidation state of –2 when present in most compounds, except:
   (a) in peroxides, for example H_2O_2, where the oxidation number is –1.
   (b) when bonded with fluorine to form F_2O, the oxidation number is +2.

All oxidation numbers must be consistent with the conservation of charge. This means that for all neutral molecules, the oxidation number of all the atoms must add up to zero. For example, in H_2O, two hydrogen atoms each of charge +1 combine with one oxygen atom of charge –2. The charge of the H_2O molecule is $+2 - 2 = 0$.

Note: There is a close relationship between valency and oxidation state. However, they are not the same. Valency is a fixed value, but oxidation state is an arbitrary value (it may vary).

Example 3.1

Find the oxidation state of chlorine in $KClO_3$

Solution

The oxidation number of potassium is +1

The oxidation number for oxygen is –2

For the three oxygen atoms, the oxidation number is $(-2 \times 3) = -6$

$KClO_3$ is a neutral compound. Therefore, the oxidation number of the compound is zero.

Therefore, $+1 + Cl - 6 = 0$

$Cl = 6 - 1 = +5$

The oxidation number of chlorine in $KClO_3$ is +5.

Example 3.2

Find the oxidation number of sulfur in SO_4^{2-}

Solution

The total charge on the sulfate ion is –2

The oxidation number of oxygen is –2

Therefore, $S + (-2 \times 4) = -2$

$S - 8 = -2$

$S = 8 - 2$

$S = +6$

The oxidation state of sulfur in SO_4^{2-} is +6.

Example 3.3

Give the oxidation number of Cr in $Cr_2O_7^{2-}$

Solution

Total charge on the dichromate ion is –2

For oxygen, $-2 \times 7 = -14$

Therefore, $2Cr - 14 = -2$

$2Cr = +12$

$Cr = +6$

Therefore, the oxidation number of Cr is +6.

Exercise 3.2

Calculate the oxidation number of each underlined element in the following chemical substances:

1. (a) $Na_3\underline{P}O_4$ (b) $\underline{S}O_3^{2-}$ (c) $\underline{N}O_2^-$
2. Explain why ammonia (NH_3) is a covalent compound, while ammonium chloride (NH_4Cl) is an ionic compound.

Concept of chemical formulas

Task 3.5

Obtain containers such as bottles containing chemicals in the laboratory. Examine them and identify the chemical formulas of the substances on their labels.

A chemical formula is a representation that uses symbols to show the proportions of the elements present in a chemical compound. The number of atoms or groups of atoms are shown by number subscripts. For example, the chemical formula for sodium sulfide is Na_2S, which shows that two atoms of sodium combine with one atom of sulfur to form the molecule of sodium sulfide. For groups of atoms such as radicals, a bracket is used to show that they are being considered as a unit under one valency. For example, in calcium nitrate, $Ca(NO_3)_2$, the NO_3^- radical is in brackets. There are some points to remember when writing chemical formulas.

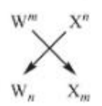
1. Positively charged ions (cations) are written before negatively charged ions (anions).
2. A radical must be treated as a unit.
3. Brackets are not used for single elements.
4. The valency 1 is simply assumed and not written in the formula.

The symbols and valencies of the atoms and radicals are important in writing a chemical formula. For example, for arbitrary elements W and X with valencies *m* and *n*, respectively, and where X can be a radical or an atom, the following steps can be used to come up with a chemical formula of their compounds:

Step 1: Write the symbols of the elements and radicals, in this case W and X.

Step 2: Write down the ions used, with their valencies as superscripts, that is $W^m X^n$.

Step 3: Interchange the valencies of W and X and write them as subscripts.



The formula of the chemical compound is W_nX_m.

Note: When *m* and *n* are equal, there is no need for the exchange and, therefore, are not written since they are in a ratio of 1:1.

Example 3.4

1. Give the formula of the compound of calcium and chlorine atoms.

   Step 1: Write down the symbols for calcium or chlorine atoms.
   Ca Cl

   Step 2: Write down the ions used in the compound with their charges.
   Ca^{2+} Cl^-

   Step 3: Write down the ions used with their valencies as superscripts.
   Ca^2 Cl^1

   Step 4: Interchange the valencies and write them as subscripts.

Ca^{2} Cl^{1}

1 2

$CaCl_2$

The formula of the chemical compound is $CaCl_2$. The compound contains two chlorine atoms and one calcium atom.

Example 3.5

Give the formula of the compound of aluminium and sulfate.

Step 1: Al SO_4

Step 2: Al^{3+} SO_4^{2-}

Step 3: Al^{3} SO_4^{2}

Step 4: Al^{3} SO_4^{2}

2 3

$Al_2(SO_4)_3$

The chemical formula is $Al_2(SO_4)_3$.

Exercise 3.3

1. Write the steps used to arrive to the following chemical formulas of the compounds:
   (a) $MgCl_2$ (b) Na_2SO_4 (c) NH_4NO_3
2. Write down the chemical formula of each of the compounds formed by the combination of the following elements:
   (a) Potassium and chlorine
   (b) Calcium and sulfur
   (c) Lithium and fluorine

Types of chemical formulas

Chemical formulas can be divided into three types, namely *empirical formula*, *molecular formula*, and *structural formula*.

An **empirical formula** is the formula which represents the simplest ratio of the atoms or ions in a compound. The simplest formula is usually determined by considering experimental data. That is why it is called 'empirical' which means 'based on experimentation'. For example, CH_2 shows there are twice as many hydrogen atoms as carbon atoms. It does not show the exact number of each atom of the element in the compound.

A **molecular formula** shows the actual number of each atom in a molecule. It is a multiple of the empirical formula. For example, if the empirical formula is CH_2, its molecular formula may be C_2H_4, C_4H_8, C_6H_{12}, and so on. Therefore, a molecular formula is equal to *n* multiplied by the empirical formula, where *n* is a whole number. Note that when *n* is 1, the empirical formula equals the molecular formula.

A **structural formula** is a graphic representation of molecular structure showing how the atoms are arranged. At this level, only the empirical and molecular formulas will be studied.

Formula calculations

When the percentage compositions of the elements that make up a compound are known, it is possible to obtain both the empirical and molecular formulas of such a compound. The following are the steps considered when calculating the empirical formula:

Step 1: Obtain the mass of each element in the sample compound. If expressed in percentages, convert the percentage of each of the elements to mass. If the mass or relative molecular mass (R.M.M.) of the compound is not given, an arbitrary mass of 100 g is usually used.

Note: The R.M.M. is the sum of the relative atomic masses (R.A.M.) of all the atoms in a molecule of the compound. The R.A.M. can also be obtained from the periodic table (See Appendix 1).

Step 2: Divide the mass of each element by its R.A.M.

Step 3: Divide each of the values obtained in Step 2 by the lowest value among them.

Step 4: Convert the ratios in Step 3 to whole numbers. These whole numbers give the ratio of each element in the compound.

Example 3.6

What is the empirical formula for a compound of mass 8.1 g if it consists of 4.9 g of magnesium and 3.2 g of oxygen?

Solution

Step 1: Obtain the mass of each element in the compound. These are already given:

Mass of magnesium = 4.9 g

Mass of oxygen = 3.2 g

Step 2: Divide the mass of each element by its R.A.M.

Magnesium, $\frac{4.9}{24} = 0.20$

Oxygen, $\frac{3.2}{16} = 0.20$

Step 3: Divide by the lowest quotient.

Mg : O

$\frac{0.20}{0.20} : \frac{0.20}{0.20}$

Step 4: Obtain their whole number ratios directly or by approximation.

Mg : O

1 : 1

The empirical formula is MgO.

Example 3.7

Given that a certain compound is 69.59% barium, 6.09% carbon and the rest is oxygen. Calculate the empirical formula of this compound.

Solution

Step 1: Assuming that you have 100 g of the compound, then the mass of each element will be:

Ba = 69.59 g

C = 6.09 g

The mass of oxygen will be 100 g – (mass of barium + mass of carbon)

= 100 g – (69.59 g + 6.09 g)= 24.32 g

Note that the 100 g is arbitrary for simplifying the calculations.

Step 2: Divide the mass of each element in the sample by its R.A.M.

$Ba = \frac{69.59}{137.3} = 0.51$

$C = \frac{6.09}{12} = 0.51$

$O = \frac{24.32}{16} = 1.52$

Step 3: Divide each value by the smallest number.

$\frac{0.51}{0.51} = 1$

$\frac{0.51}{0.51} = 1$

$\frac{1.52}{0.51} = 2.98$

Step 4: Obtain the whole number ratios.

Ba : C : O

1 : 1 : 3

The empirical formula is therefore $BaCO_3$.

Example 3.8

A compound contains 15.8% carbon and 84.2% sulfur. Calculate its empirical formula. If its relative molecular mass is 76, what is its molecular formula?

Solution

Step 1: Assume that you have 100 g of the compound, then find the mass of each element in grams.

C = 15.8 g

S = 84.2 g

Step 2: Divide the mass of each element in the sample by its R.A.M., i.e.

$$\frac{\text{Mass in sample}}{\text{R.A.M.}}$$

$$C = \frac{15.8\ g}{12\ g} = 1.32$$

$$S = \frac{84.2\ g}{32\ g} = 2.63$$

Step 3: Divide throughout by the smallest value, in this case 1.32:

$$\frac{1.32}{1.32} : \frac{2.63}{1.32} = 1:1.99$$

Step 4: Obtain whole number ratios:

C : S

1 : 2

The empirical formula is therefore CS_2.

The empirical formula together with the relative molecular mass can then be used to establish the molecular formula of the compound.

Molecular formula = $n \times$ empirical formula, where n is a whole number.

Hint: R.M.M. = $n \times$ sum of R.A.M.

Therefore, the molecular formula = $n \times$ empirical formula = $n(CS_2)$

Now, R.M.M. = $n \times$ sum of R.A.M.

$$76 = n \times [12 + (2 \times 32)]$$

$$76 = n \times (12 + 64)$$

$$76 = 76\,n$$

$$n = 1$$

Therefore, the molecular formula is CS_2.

Exercise 3.4

1. A compound has 1.121 g of nitrogen, 0.161 g of hydrogen, 0.480 g of carbon and 0.640 g of oxygen. Deduce its empirical formula.
2. A compound has an empirical formula CH. If it has a relative molecular mass of 78.11, deduce its molecular formula.
3. A compound is analysed and found to contain 52.17% carbon, 13.04% hydrogen, and 34.78% oxygen. Determine its
   (a) empirical formula.
   (b) molecular formula, if its relative molecular mass is 46.

Nomenclature of binary inorganic compounds

Everything in the universe bears a name to differentiate it from others. Chemical substances also bear names that range from those of elements to those of compounds. The name of a substance can originate from some factors such as the place of origin, founder, use, and type or classification. Items or substances that fall under a particular group or classification are named systematically. A systematic way of assigning names to items that belong to a particular group or classification is called *nomenclature*.

Scientists use IUPAC nomenclature in naming chemical compounds so that experts and other interested persons around the world understand exactly what the substance is. IUPAC stands for International Union of Pure and Applied Chemistry.

Binary inorganic compounds

While an *inorganic compound* is any substance in which two or more chemical elements (usually other than carbon) are combined, always in definite proportions, a *binary compound* is the one which is formed by two chemical substances. Examples of binary inorganic compounds are CaO, NaCl, and PCl_3.

Inorganic compounds are categorised into *ionic* and *covalent*. The nomenclature of ionic compounds differs slightly from that of covalent compounds.

Task 3.6

Use online resources such as Chemguide to explore detailed explanations and examples of the nomenclature of various binary inorganic compounds. Summarise the key rules for naming these compounds, including binary ionic compounds, covalent compounds and any exceptions.

Nomenclature of binary ionic compounds

Ionic compounds are formed when a metal combines with a non-metal. The following are the steps considered when naming binary ionic compounds:

1. Name the metallic ion that appears first in the formula using the name of the element itself.
2. The second part of the formula which is usually an anion in the compound will end with a suffix "ide". For example, oxygen becomes oxide, hydrogen becomes hydride and chlorine becomes chloride.

Note:

(a) Some metals always have fixed charges when they form ions, that is,

  (i) Group I metals have a charge of +1.

  (ii) Group II metals have a charge of +2.

  (iii) Group III metals have a charge of +3.

  (iv) Silver (Ag) has a charge of +1.

  (v) Zinc (Zn) has a charge of +2.

(b) Other metals are multivalent and can thus form more than one ion. For example; iron (Fe) is bivalent; it has valencies of 2 and 3, copper (Cu) is also bivalent; it has valencies of 1 and 2. Compounds formed from these metals must be distinguished by stating which valency has been used in the compound. The valency of the respective metal is indicated by capital Roman numbers in parentheses (brackets), for example cobalt(II) chloride, copper(II) oxide and iron(III) oxide. In chemical nomenclature, it is common practice not to leave a space between the name of a metal and its oxidation state written in Roman numbers.

Example 3.9

What is the name of the compound with the formula $FeCl_3$?

Solution

The total charge of the molecule is zero and Cl^- has a negative charge.

(i) Let x be the charge of Fe

(ii) $1(x) + 3(-1) = 0$

(iii) $x = +3$

(iv) So, the Fe is in the +3 oxidation state. Write the name 'iron' and place III in brackets beside it.

(v) Use the name 'chlorine' but change the last three letters to "ide": So the name is iron(III) chloride.

Example 3.10

What is the name of the compound with the formula CuS?

Solution

(i) Let x be the charge of Cu.

(ii) Sulfur has a charge of –2.

(iii) $1(x) + 1(-2) = 0$

$x = +2$ for Cu

(iv) Write the name copper and place II in brackets beside it.

(v) Use the name sulfur but change the last two letters to *ide*. The name of the compound is copper(II) sulfide.

Other examples of the names of binary ionic compounds are:

(a) MgO is named magnesium oxide.

(b) $AlCl_3$ is named aluminium chloride.

(c) MnO_2 is named manganese(IV) oxide.

Note: Manganese can have more than one charge, but each oxygen ion has a charge of –2. For the compound to be neutral, Mn must have a charge of +4.

Nomenclature of binary covalent compounds

Covalent compounds are formed between two non-metal elements. These compounds are named differently from ionic compounds. The number of atoms are presented by prefixes as shown in Table 3.4.

Table 3.4: *Examples of prefixes*

Number	Prefix	Number	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

The following are the steps to consider when naming binary covalent compounds:

1. Give the name of the first element.
2. Give the name of the second element with the ending changed to –*ide*.
3. If more than one compound can form between two elements, use prefixes to indicate the number of atoms of each element.

Example 3.11

Give the name for PCl_3.

Answer

(i) Since there is one phosphorus atom, use it as the first part of the name.

(ii) There are three chlorine atoms, so use *tri* in front of chlorine; then, drop the *ine* in chlorine and replace it with *ide*.

The name is phosphorus trichloride.

Example 3.12

What is the name for N_2O_4?

Answer

(i) Use the prefix *di* in front of nitrogen since there are two atoms.

(ii) Use the prefix *tetra* in front of the oxygen since there are four atoms.

(iii) Drop –*ygen* and replace it with *ide*:

(iv) The name is dinitrogen tetraoxide.

Table 3.5 gives the formulas and names of some binary covalent compounds.

Table 3.5: *Some binary covalent compounds*

Formula	Chemical name
CO_2	Carbon dioxide
CO	Carbon monoxide
N_2O_5	Dinitrogen pentoxide
HCl	Hydrogen chloride
NO	Nitrogen monoxide
SF_6	Sulfur hexafluoride

Note: IUPAC names of common compounds are given in Appendix 2.

Chemical names of common substances

Chemical names are typically used to provide precise descriptions of substance compositions, including those encountered in daily life. For example, requesting sodium chloride for use in food is uncommon; instead, the term "common salt" is used. This explains the existence of common names for certain substances. However, it is important to note that some common names are inaccurate and may vary from one place to another. Therefore, they cannot tell the chemical composition of a substance. The chemical names of common substances are provided in Appendix 2.

Activity 3.3

Aim: To identify and classify the chemical compounds in toothpaste

Requirements: Toothpaste tube (with ingredient list), notebook, pen/pencil, and chart or table for recording observations

Procedure

1. Read the ingredients on the toothpaste tube.
2. Write down the chemical compounds listed in the ingredients.
3. Explore the chemical formulas of the compounds and identify whether each of the compound contains a metal and a non-metal or only non-metals.
4. Record your findings as indicated in Table 3.6.

Table 3.6: *Chemical compounds present in toothpaste*

Compound	Formula	Ionic or covalent compound

Questions:

(a) In which ways do the ionic and covalent compounds found in toothpaste differ?

(b) Which compound is in the largest amount in the toothpaste?

Exercise 3.5

1. What is the name of NH_3? Why is it an exception in its nomenclature compared to other nitrogen compounds?
2. Write the IUPAC names of the following compounds:
   (a) $FeCl_2$
   (b) SO_2
   (c) P_4O_{10}
   (d) NO_2
3. Name the compound formed when zinc combines with oxygen.

Chapter summary

1. Chemical bonding involves electrons in the outermost shell of an atom. When the outermost shell is fully filled, the atom is said to be stable.
2. Ions are formed when an atom gains or loses electron(s). Cations are positively charged ions that result from atoms losing one or more electrons. Anions are negatively charged ions that result from atoms gaining one or more electrons.
3. Ionic (electrovalent) bonding usually occurs between a metal and a non-metal. It involves the transfer of electron(s) from the atoms of the metal to the atoms of the non-metal.

4. Covalent bonding takes place between two or more non-metals. It involves atoms of the non-metals sharing electrons that are in their outermost shells.
5. The ability of an atom to combine with other atoms according to the number of electrons it can give, take or share is known as valency.
6. The oxidation state (oxidation number) of an element is the number of electrons that need to be added, shared or removed by its atom, to make a neutral molecule. The oxidation number is arbitrary and may be positive, negative or zero.
7. A radical is a group of atoms which behaves as a single unit and has an overall positive or negative charge. A radical can also be an atom, molecule or ion that has unpaired valence electron. Such a group maintains its identity throughout any chemical reaction.
8. A chemical formula is a representation that uses chemical symbols to show the proportions of the elements present in a chemical compound.
9. An empirical formula is the simplest way of writing a chemical formula and indicates the ratio of the atoms in a compound.
10. A molecular formula is a chemical formula that shows the total number of atoms of each element in a molecule of a substance.
11. A systematic way of naming items or substances of a particular category is known as nomenclature.

Revision exercise 3

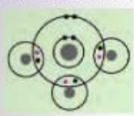
Choose the correct answer for Questions 1-11. For other questions, provide the answers as per the demands indicated.

1. What is the valency of Group I elements?
   (a) 1 (b) 2 (c) 3 (d) 4
2. What name is given to the force of attraction that holds atoms together to form a molecule?
   (a) Chemical change
   (b) Chemical bond
   (c) Friction
   (d) Centripetal force

3. What happens during the formation of ions?
   (a) Non-metal atoms gain protons.
   (b) Metal atoms lose their outermost electrons.
   (c) Metal atoms gain electrons in their outermost shells.
   (d) Non-metal atoms lose electrons of their outermost shells.
4. Electrovalent bonding occurs between
   (a) metals and metals.
   (b) metals and non-metals.
   (c) metals and inert gases.
   (d) non-metals and non-metals.
5. The following atoms cannot exist freely as single atoms, **except**
   (a) sodium.
   (b) chlorine.
   (c) argon.
   (d) magnesium.
6. A covalent bond is formed due to
   (a) opposite charges of atoms.
   (b) transfer of electrons of atoms.
   (c) forces of attraction between atoms.
   (d) sharing of electrons between atoms.
7. Which of the following is a property of covalent compounds?
   (a) They do not conduct electricity.
   (b) Atoms share only two pairs of electrons.
   (c) Mostly are soluble in water.
   (d) Their melting and boiling points are very high.
8. What is the oxidation number of nitrogen in NH_4^+?
   (a) +1 (b) –3 (c) +4 (d) +5
9. Which among the following is the correct valency of the phosphate radical?
   (a) 5 (b) –3 (c) 3 (d) –5

10. Which of the following sets of symbols represent cations?
    (a) K^+ and Mg
    (b) Mg and Al
    (c) Al^{3+} and Cl^-
    (d) K^+ and Al^{3+}
11. Identify the chemical formula for potassium chlorate(V).
    (a) K_2ClO_3 (b) $K(ClO_3)_2$ (c) K_3ClO_3 (d) $KClO_3$
12. Choose the correct formula for the combination of the following ions:
    (i) Mg^{2+} and PO_4^{3-}
    (a) $Mg_3(PO_4)_2$ (b) Mg_2PO_4 (c) $Mg_2(PO)_3$ (d) $MgPO_4$
    (ii) Ba^{2+} and N^{3-}
    (a) BaN (b) BaN_2 (c) Ba_2N_3 (d) Ba_3N_2
    (iii) Al^{3+} and SO_4^{2-}
    (a) $AlSO_4$ (b) $Al(SO_4)_2$ (c) $Al_2(SO_4)_3$ (d) $Al_3(SO_4)_2$
    (iv) Zn^{2+} and $C_2H_3O_2^-$
    (a) $ZnC_2H_3O_2$ (b) $Zn_2C_2H_3O_2$ (c) $Zn(C_2H_3O_2)_2$ (d) $Zn_2(C_2H_3O_2)_2$
13. (a) Identify the compounds with the incorrect IUPAC names.
    (i) KCl – Potassium chloride
    (ii) $Fe_2(SO_4)_3$ – Iron(II) sulfate(VI)
    (iii) $CaCl_2$ – Calcium dichloride
    (iv) Na_2SO_4 – Sodium tetraoxosulfate(VI)
    (b) Rename the incorrectly named compound(s) in (a) using the IUPAC rules.
14. You are provided with the following list of substances: Mg, MnO_4^{2-}, Ca^{2+}, Cl, N_2, $Cr_2O_7^{2-}$, Al^{3+}, and H_2. Which of these substances are:
    (a) atoms?
    (b) molecules?
    (c) ions?
    (d) radicals?

15. A molecule of a certain gas can be represented by the following diagram:



    (a) What is the name of the molecule?
    (b) What is the molecular formula of the gas?
    (c) What type of bonding holds the atoms of the molecule?
    (d) Name other five compounds with this type of bonding.
16. Write the electronic configuration of each of the following species:
    (a) Aluminum ion
    (b) Magnesium
    (c) Chloride ion
    (d) Neon
17. Name the following compounds:
    (a) MgI_2 (b) CCl_4 (c) $FeBr_2$ (d) CuI_2
    (e) H_2S (f) K_2O (g) PCl_5
18. Calculate the oxidation number of Cr in CrO_4^{2-} and $Cr_2O_7^{2-}$ radicals.
19. A compound of sulfur and oxygen is 40.1% sulfur by mass. What is the empirical formula for the compound? The R.A.M. are S = 32.07 and O = 16.00.
20. Write the chemical formulas of the following compounds:
    (a) Sodium hydrogen carbonate
    (b) Silver trioxonitrate(V)
    (c) Copper(I) oxide
    (d) Aluminium tetraoxosulfate(VI)
21. Write down any two different ionic states in each of the following elements:
    (a) Fe (b) Cu (c) Pb (d) Mn
22. An atom of element X (atomic number 11) and an atom of element Y (atomic number 9) combine to form a compound.
    (a) Write the formula of the compound.

    (b) State the type of the bond present in the compound. Explain your answer.
23. A hydrocarbon contains 88.88% carbon and 11.12% hydrogen by mass. Calculate the empirical formula of the molecule.
24. A compound consists of calcium 40%, carbon 12% and the rest is oxygen by mass. Determine the empirical formula of the compound.
25. (a) Calculate the oxidation number of the underlined elements in the following compounds:
    (i) $\underline{N}H_4Cl$ (ii) $Na_2\underline{S}O_4$ (iii) $\underline{Al}_2O_3$ (iv) $H_2\underline{O}_2$
26. Elements *T* and *Q* have atomic numbers 12 and 17, respectively. Use the two elements to answer the following questions:
    (a) Write the electronic configuration of element *Q*.
    (b) What is the valency of element *T*?
    (c) Write the chemical formula of a compound formed when *T* and *Q* combine.
    (d) Giving a reason predict the type of the bond formed by the combination of elements *T* and *Q*.
    (e) In which group and period in the periodic table do elements *Q* and T belong?
27. An organic compound contains 26.70% carbon, 2.20% hydrogen and 71.10% oxygen. If its mass is 90 g, determine its molecular formula.

Chapter Four

Chemical reactions

Introduction

Chemical reactions are an integral part of daily life, constantly occurring in and around us, often unnoticed. For example, everyday activities such as burning charcoal and wood, cooking food, respiration in living organisms, fuel combustion in engines, rusting of metals, and digestion involve transforming substances into new ones. These transformations result from chemical reactions. In this chapter, you will learn the concept of chemical reactions, including chemical equations, and types of chemical reactions. The competencies developed will enable you to accurately present chemical equations and analyse various chemical reactions that yield products essential to daily life.

Think

The impact of chemical reactions on natural and industrial processes

Concept of chemical reactions

Task 4.1

Use library books and reliable online resources to search for any five chemical reactions that occur in daily life. For each response, include the name of the chemical reaction, a brief explanation of how it happens, the chemical equation (if applicable) and an example of where this reaction is frequently observed.

A chemical reaction is a process in which one or more chemical substances are converted to one or more different substances. Chemical reactions take place when bonds between atoms in the reacting substance(s) are broken, atoms rearrange, and new bonds between the atoms are formed to make new substance(s). The chemicals that begin the reaction process are called *reactants* and the new substances formed are called *products*. The products have properties that are different from their respective reactants. Features that indicate a chemical reaction has taken place include one or more of the following: evolution of a gas; formation of precipitates; and change in colour, temperature or state.

A chemical reaction is either *reversible* or *irreversible*. A reversible reaction proceeds in both forward and backward directions. In the forward reaction, the reactants are converted into products, whereas in the backward reaction, the products become the reactants. An irreversible reaction proceeds in one direction from reactants to products.

Chemical equations

Task 4.2

1. Use library books and reliable online resources to explore various molecular equations.
2. Use chemistry software step by step to represent the molecular equations identified in 1.

A chemical reaction is expressed in the form of a *chemical equation*. A chemical equation is a symbolic or words representation of a chemical reaction. A chemical equation written in words is referred to as a *word equation*. For example, when calcium metal reacts with chlorine gas to form solid calcium chloride, the word equation is written as:

$$\text{Calcium} + \text{Chlorine gas} \longrightarrow \text{Calcium chloride}$$

A chemical equation written in symbols is referred to as a *formula equation*. For example, the formula equation for a reaction between calcium metal and chlorine gas is:

$$Ca(s) + Cl_2(g) \longrightarrow CaCl_2(s)$$

The formula equation is more useful than the word equation. However, it is advantageous to write the equation in word form first. Formula equations provide useful information, including compositions, amounts, formulas, and the physical states of substances that are involved. Formula equations may also state the conditions for the reactions to take place. The reactants and products in the formula equations are either solids, liquids or gases. These states of matter are represented in a chemical equation by using symbols, such as (s) for solid, (l) for liquid, and (g) for gas. When the substances that are involved in a reaction are dissolved in water, the word aqueous (aq) is used. These symbols should be included when writing the chemical equations and are stated in parentheses ( ), after the chemical symbol.

In a chemical equation, a headed arrow is used to separate the reactants and products.

A full-headed arrow is used to separate reactants and products for an *irreversible reaction*.

For example:

$$\text{Reactants} \longrightarrow \text{Products}$$

Double half-headed arrows pointing in the opposite directions are used to separate reactants and products for a reversible reaction. For example:

$$\text{Reactants} \underset{\text{backward reaction}}{\overset{\text{forward reaction}}{\rightleftharpoons}} \text{Products}$$

The arrows show the direction of a reaction; thus, it means 'produce' or 'yield'. Each individual substance is separated from the other by a plus sign (+). Note that the number of reactants and products are not necessarily the same. A chemical equation has the following key characteristics:

(a) Reactants and products: It lists the substances involved in the reaction, with reactants on the left side and products on the right side.

(b) Chemical formulas: Each substance is represented using its chemical formula (for example H_2O for water).

(c) Direction of reaction: An arrow points from the reactants to the products, indicating the direction of the reaction.

(d) Balanced equation: A chemical equation follows the law of conservation of mass, ensuring that the number of atoms for each element is the same on both sides of the equation.

(e) States of matter: The physical state of each substance is often indicated.

(f) Energy changes: If applicable, energy changes such as heat or light may be noted.

(g) Reaction conditions: Temperature, pressure or catalysts may be noted above or below the arrow.

Molecular equations

A molecular equation is an equation representing a reaction showing the reactants and products in undissociated form. In molecular equations, reactants and products are considered neutral regardless of their exact physical states. The following are examples of molecular equations:

$$PbO_2(s) + SO_2(g) \longrightarrow PbSO_4(s)$$

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

In principle, when writing molecular equations, the reactants and products should be balanced.

Balancing chemical equations

All chemical equations must be written in accordance with the *law of conservation of mass*. This law states that, *in a chemical reaction, the total mass of the products equals the total mass of the reactants.* This means, when balancing a chemical equation, the number of each atom on both sides of the equation must be equal because atoms do not varnish during a reaction, but are reorganised.

The following steps are followed when writing and balancing simple chemical equations:

1. Write the equation in a word form.
2. Write the unbalanced equation including correct chemical formulas for reactants and products.
3. List the number of atoms of each element on both sides of the equation.
4. Balance one element at a time. Usually start with metals or more complex elements, and complete by balancing hydrogen and oxygen if present.
5. Use coefficients (whole numbers) to balance atoms. Never change subscripts in a chemical formula.
6. Count atoms of all elements on both sides to make sure they are equal.
7. Simplify coefficients if necessary. The final equation should use the smallest whole-number coefficients and should include the state symbols.

Example 4.1

Hydrogen chloride gas is formed when hydrogen gas burns in chlorine gas. Write a balanced chemical equation for the reaction.

Step 1: Write the equation in word form

$$\text{Hydrogen gas} + \text{Chlorine gas} \longrightarrow \text{Hydrogen chloride gas}$$

Step 2: Write the unbalanced chemical equation using symbols

The reactants are hydrogen gas (H_2) and chlorine gas (Cl_2), and the product is hydrogen chloride gas (HCl):

$$H_2 + Cl_2 \longrightarrow HCl$$

Step 3: List the number of atoms on both sides

Reactants (left side):

H_2: 2 hydrogen atoms

Cl_2: 2 chlorine atoms

Products (right side):

HCl: 1 hydrogen atom and 1 chlorine atom per molecule.

There are two hydrogen atoms and two chlorine atoms on the left but only one of each on the right.

Step 4: Balance the equation

The HCl molecule contains one hydrogen atom and one chlorine atom on the right side. Thus, two HCl molecules are required to balance for both hydrogen atoms and chlorine atoms on the left.

$$H_2 + Cl_2 \longrightarrow 2HCl$$

Step 5: Verify the balancing

Count atoms of all elements on both sides to make sure they are equal.

Reactants (left side):

H_2: 2 hydrogen atoms

Cl_2: 2 chlorine atoms

Products (right side):

2HCl: 2 hydrogen atoms and 2 chlorine atoms.

Since the number of atoms is equal on both sides, the equation is balanced.

Steps 6: Write a balanced chemical equation

Write a balanced chemical equation with its state symbols.

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

Example 4.2

Zinc dissolves in dilute hydrochloric acid to form zinc chloride solution and hydrogen gas. Write a balanced chemical equation for this reaction.

Step 1: Write the equation in word form

$$\text{Zinc} + \text{Dilute hydrochloric acid} \longrightarrow \text{Zinc chloride} + \text{Hydrogen gas}$$

Step 2: Write the unbalanced chemical equation using symbols

The reactants are zinc (Zn) and hydrochloric acid (HCl), and the products are zinc chloride ($ZnCl_2$) and hydrogen gas (H_2):

$$Zn + HCl \longrightarrow ZnCl_2 + H_2$$

Step 3: List the number of atoms on both sides

Reactants (left side):

Zn: 1 zinc atom

HCl: 1 hydrogen atom and 1 chlorine atom per molecule

Products (right side):

$ZnCl_2$: 1 zinc atom and 2 chlorine atoms

H_2: 2 hydrogen atoms

There are two chlorine atoms in $ZnCl_2$ but only one chlorine atom from HCl on the left. Also, there are two hydrogen atoms in H_2 but only one hydrogen atom from HCl on the left.

Step 4: Balance the equation

To balance the equation, two HCl molecules are required to produce two chlorine atoms and two hydrogen atoms:

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

Step 5: Verify the balancing

Count atoms of all the elements on both sides to make sure they are equal.

Reactants (left side):

Zn: 1 zinc atom

2HCl: 2 hydrogen atoms and 2 chlorine atoms

Products (right side):

$ZnCl_2$: 1 zinc atom and 2 chlorine atoms

H_2: 2 hydrogen atoms

Since the number of the atoms is equal on both sides, the equation is balanced.

Step 6: Write the balanced chemical equation with its state symbols

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

Activity 4.1

Aim: To verify the law of conservation of matter (mass) and precipitation reaction

Requirements: Conical flasks, analytical balance, 100-mL measuring cylinders, 1 M barium chloride solution, and 1 M zinc sulfate solution

Procedure

1. Weigh the mass of two empty conical flasks and record the results.
2. Put about 50 cm^3 of 1 M barium chloride solution in one of the flasks and another 50 cm^3 of 1 M zinc sulfate in the second flask.
3. Weigh the flasks to get the mass of their solutions and flasks and record the results.
4. Pour the weighed solution of barium chloride into a flask containing zinc sulfate. Swirl the mixture.
5. Weigh the mixture after the reaction and record the results.

Questions

1. What is the total mass of the solutions before the reaction?
2. What is the mass of the mixture after the reaction?
3. What is the colour of the mixture after the reaction?
4. Write the balanced chemical equation for the reaction.

Ionic equations

An ionic equation is a chemical equation in which compounds in aqueous solutions or in molten state are written as dissociated ions. Ionic equations are commonly used in *displacement reactions* in aqueous solutions. In these equations, *spectator ions* are omitted to give a net ionic equation. *Spectator ions* are the ions that do not change their valence states in the reaction since they remain unchanged in a chemical reaction.

The following steps are followed when writing an ionic equation:

1. Write the balanced chemical equation in symbols; ensure all formulas are correct.
2. Split all soluble ionic compounds into individual ions to get the total ionic equation.

Note: Insoluble ionic compounds, gases, and liquids should not be split into ions.

3. Cancel out *'spectator'* ions (ions that appear unchanged on both sides).
4. Write the net ionic equation.

Example 4.3

Write a net ionic equation for the reaction between aqueous solutions of silver nitrate and calcium chloride.

Steps to write the net ionic equation

1. Write the balanced formula equation:

$$2AgNO_3(aq) + CaCl_2(aq) \longrightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$$

2. Write the complete ionic equation:

$$2Ag^+(aq) + 2NO_3^-(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \longrightarrow 2AgCl(s) + Ca^{2+}(aq) + 2NO_3^-(aq)$$

3. Identify and remove spectator ions:

$$2Ag^+(aq) + \cancel{2NO_3^-(aq)} + \cancel{Ca^{2+}(aq)} + 2Cl^-(aq) \longrightarrow 2AgCl(s) + \cancel{Ca^{2+}(aq)} + \cancel{2NO_3^-(aq)}$$

4. Write the net ionic equation:

$$2Ag^+(aq) + 2Cl^-(aq) \longrightarrow 2AgCl(s)$$

Example 4.4

Write a net ionic equation for the reaction of barium chloride with sodium sulfate aqueous solutions.

Steps:

1. Write the balanced formula equation

$$BaCl_2(aq) + Na_2SO_4(aq) \longrightarrow BaSO_4(s) + 2NaCl(aq)$$

2. Write the complete ionic equation:

$$Ba^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s) + 2Na^+(aq) + 2Cl^-(aq)$$

3. Identify and remove spectator ions:

$$Ba^{2+}(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + SO_4^{2-}(aq) \longrightarrow BaSO_4(s) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}$$

4. Write the net ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$$

Example 4.5

Write a net ionic equation for the reaction between dilute hydrochloric acid and an aqueous sodium hydroxide.

Steps:

1. Write the balanced formula equation:

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

2. Write the complete ionic equation:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \longrightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

3. Identify and remove spectator ions:

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \longrightarrow \cancel{Na^+(aq)} + \cancel{Cl^-(aq)} + H_2O(l)$$

4. Write the net ionic equation:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Exercise 4.1

1. Balance the following chemical equations:
   (a) $Na(s) + Cl_2(g) \longrightarrow NaCl(s)$
   (b) $P(s) + O_2(g) \longrightarrow P_2O_5(s)$
   (c) $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$
   (d) $C(s) + CO_2(g) \longrightarrow CO(g)$
   (e) $CaCO_3(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
2. With reasons, state whether the following chemical equations are balanced:
   (a) $Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$
   (b) $CaCl_2(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + Ca(NO_3)_2(aq)$
3. Balance each of the following chemical equations and write its net ionic equation:
   (a) $Na_2CO_3(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$
   (b) $H_2SO_4(aq) + KOH(aq) \longrightarrow K_2SO_4(aq) + H_2O(l)$
   (c) $Ca(s) + HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$
   (d) $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \longrightarrow PbSO_4(s) + NaNO_3(aq)$

Types of chemical reactions

Task 4.3

Use library books and reliable online resources to explore various chemical reactions which occur in everyday life, and classify each one based on its reaction type.

Chemical reactions drive numerous everyday processes, such as cooking, energy production, plant growth, and digestion. These reactions convert raw materials into essential products that support life and modern living. The following are the types of chemical reactions, along with their real-life applications.

Combination (synthesis) reactions

A combination reaction is a chemical reaction in which two or more chemical species combine to form a single product. Combination reactions are also referred to as synthesis reactions. This type of reaction is expressed in the form of:

$$A + B \longrightarrow AB$$

Generally, there are three types of combination reactions, namely a reaction between two or more elements, reaction between elements and compounds, and reaction between two compounds.

Combination reaction between two elements

This reaction occurs when two elements combine to give a single compound. An example of a combination reaction between two elements is when magnesium burns in air (oxygen) with a bright white flame to form white solids of magnesium oxide. Figure 4.1 shows the burning of magnesium in oxygen.

$$\text{Magnesium} + \text{Oxygen gas} \longrightarrow \text{Magnesium oxide}$$

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Activity 4.2

Aim: To burn magnesium in oxygen

Requirements: A pair of tongs, heat source (Bunsen burner), gas jar, lighter, magnesium ribbon, oxygen source, and steel wool

Procedure

1. Clean about 0.1 g of a piece of magnesium ribbon by using steel wool.

2. Hold the ribbon by using a pair of tongs and heat it over a Bunsen burner or any heat source flame.
3. When it starts to burn, lower it into the gas jar of oxygen as shown in Figure 4.1. Do not drop it into the jar.

Figure. 4.1: *Magnesium burning in oxygen*

Questions

1. What is the colour of flame when magnesium is burned?
2. What is the balanced chemical equation associated with this experiment?

Combination reaction between elements and compounds

This occurs when an element reacts with a compound to form another compound. For example, carbon monoxide gas reacts with oxygen gas to form carbon dioxide gas.

$$\text{Carbon monoxide gas} + \text{Oxygen gas} \longrightarrow \text{Carbon dioxide gas}$$

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$

Combination reaction between two compounds

Two compounds may react with each other to form a new compound. For example, calcium oxide (quicklime) reacts with carbon dioxide gas to form calcium carbonate (limestone).

$$\text{Calcium oxide} + \text{Carbon dioxide gas} \longrightarrow \text{Calcium carbonate}$$

$$CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$$

Another example of a combination reaction is the reaction between lead(IV) oxide and sulfur dioxide. If lead(IV) oxide is slightly heated and then lowered into a gas jar of sulfur dioxide, the two compounds combine to form one new compound, lead(II) sulfate.

$$\text{Lead(IV) oxide} + \text{Sulfur dioxide gas} \longrightarrow \text{Lead(II) sulfate}$$

$$PbO_2(s) + SO_2(g) \longrightarrow PbSO_4(s)$$

Combination reactions play crucial roles in various biological, environmental, and industrial processes. Some of these processes include:

(a) *Formation of water:* Formation of water relates to the combination reaction whereby hydrogen and oxygen gases combine to form water as shown in the following reaction:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

(b) *Rusting of iron:* Rusting is a slow chemical process in which iron reacts with oxygen and water to form hydrated iron(III) oxide (rust). It is an example of a combination reaction and a type of corrosion that weakens iron objects over time. The chemical reaction for rusting is:

$$4Fe(s) + 3O_2(g) + 6H_2O(l) \longrightarrow 4Fe(OH)_3(s)$$

The iron hydroxide ($Fe(OH)_3$) further transforms to a more stable form, hydrated iron(III) oxide which is reddish-brown (rust).

$$4Fe(OH)_3(s) \longrightarrow 2Fe_2O_3{\cdot}6H_2O(s)$$

Rusting is a major issue in construction, transportation, and machinery, leading to structural damage and economic losses.

(c) *Formation of calcium hydroxide (slaked lime):* Calcium oxide (quick lime) reacts with water to form calcium hydroxide (slaked lime) an essential material in construction that is frequently used for setting cement and plaster. The reaction equation is:

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

(d) *Photosynthesis:* In plants, carbon dioxide gas combines with water in the presence of sunlight to form glucose and oxygen.

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{sunlight}} C_6H_{12}O_6(s) + 6O_2(g)$$

(e) *Formation (synthesis) of ammonia*: In industry, ammonia gas is manufactured from a combination reaction between nitrogen and hydrogen gases as represented in the following reaction:

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Decomposition reactions

A decomposition reaction is a chemical reaction in which a compound breaks down (decomposes) into its components. This reaction can be expressed in the form of:

$$AB \longrightarrow A + B$$

The decomposition reaction is the opposite of a combination reaction. Generally, decomposition reactions are classified into three main types, namely catalytic, electrolytic and thermal reactions. In a *catalytic decomposition reaction*, an agent called a *catalyst* is introduced that alters the rate of a chemical reaction but remains unchanged at the end of the reaction. For example, potassium chlorate readily decomposes when heated in the presence of manganese(IV) oxide (catalyst) to produce oxygen gas and potassium chloride.

$$\text{Potassium chlorate} \xrightarrow[\text{heat}]{MnO_2} \text{Potassium chloride} + \text{Oxygen gas}$$

$$2KClO_3(s) \xrightarrow[\text{heat}]{MnO_2} 2KCl(s) + 3O_2(g)$$

An *electrolytic decomposition reaction* is achieved by exposing an aqueous solution or molten compound to an electric current. An example of an electrolytic decomposition reaction is the electrolysis of water, which is represented by the following chemical equation:

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g)$$

Thermal decomposition occurs when a compound is exposed to direct heat or radiation. For example, when lead(II) nitrate crystals are heated, they decompose with a cracking sound to produce lead(II) oxide, nitrogen dioxide gas, and oxygen gas.

$$\text{Lead(II) nitrate} \xrightarrow{\text{heat}} \text{Lead(II) oxide} + \text{Nitrogen dioxide gas} + \text{Oxygen gas}$$

$$2Pb(NO_3)_2(s) \xrightarrow{\text{heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$$

Some decomposition reactions are caused by light. For example, white silver chloride breaks down when exposed to light to give tiny black crystals of silver and chlorine gas.

$$\text{Silver chloride} \xrightarrow{\text{light}} \text{Silver} + \text{Chlorine gas}$$

$$2AgCl(s) \xrightarrow{\text{light}} 2Ag(s) + Cl_2(g)$$

Decomposition reactions are common in daily life. Some important processes that are related to decomposition reactions include:

(a) *Digestion of food:* This is a biological decomposition reaction in which Complex food molecules (carbohydrates, proteins, and fats) break down into simpler substances in the body through enzymatic reactions. The body absorbs and uses simpler substances for energy, growth, and repair. For example, starch in food is broken down into simple sugars by the amylase enzyme as shown in the following equation:

$$(C_6H_{10}O_5)_n + nH_2O \xrightarrow{\text{amylase}} n(C_6H_{12}O_6)$$

This reaction occurs in the stomach and intestines during digestion.

(b) *Spoilage of food:* Organic matters in food decompose over time because of activities of microorganisms such as bacteria and fungi, leading to decay and *fermentation*. For example, when milk spoils, lactose breaks down into lactic acid due to bacterial actions, resulting in a sour taste. The following equation shows its decomposition reaction.

$$C_{12}H_{22}O_{11}(aq) + H_2O(l) \xrightarrow{\text{bacteria}} 4C_3H_6O_3(aq)$$

(c) *Thermal decomposition of limestone:* In the cement industry, limestone (calcium carbonate) is heated to form lime (calcium oxide) and carbon dioxide as shown in the following reaction:

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

Combustion reactions

A combustion reaction is a chemical process in which a substance reacts with oxygen to produce heat and light. For example, burning of fuels such as wood, coal, diesel and petrol is a combustion reaction.

General formula: $\text{Fuel} + O_2(g) \longrightarrow CO_2(g) + H_2O(g) + \text{Energy}$

Combustion reactions are essential for energy production, heating, and transportation, making them indispensable in everyday life. These reactions release energy by burning fuels or materials, driving many practical applications. Some processes related to combustion reactions include:

(a) *Burning of fuel:* Burning fuel such as coal, oil, and natural gas is a combustion reaction, where hydrocarbons react with oxygen to produce carbon dioxide, water, and energy. For example, in the combustion of coal, carbon reacts with oxygen to form carbon dioxide and heat energy.

$$C(s) + O_2(g) \longrightarrow CO_2(g) + \text{Energy}$$

(b) *Cooking with gas stoves using natural gas (methane)*: Methane reacts with oxygen to produce carbon dioxide, water and heat energy. Its combustion reaction equation is:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) + \text{Energy}$$

The energy released heats the food during cooking.

(c) *Candle burning:* When a candle burns, the wax, usually made of paraffin (a hydrocarbon), reacts with oxygen from the air, resulting in the production of carbon dioxide, water, and heat.

(d) *Respiration in living organisms*: Cellular respiration is a slow combustion reaction where glucose reacts with oxygen to release energy needed for body functions.

$$C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) + \text{Energy}$$

Displacement reactions

A displacement reaction is a chemical reaction in which a more reactive element displaces a less reactive element from its compound (Appendix 3). This reaction is expressed in the general form as:

$$A + BC \longrightarrow AC + B$$

Two reactants yield two different products. For example, when zinc reacts with hydrochloric acid, it displaces hydrogen to produce zinc chloride and hydrogen gas.

$$\text{Zinc} + \text{Hydrochloric acid} \longrightarrow \text{Zinc chloride} + \text{Hydrogen gas}$$

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

Another example of a displacement reaction is when solid iron reacts with copper(II) sulfate. In this reaction, iron displaces copper from copper(II) sulfate.

$$\text{Iron} + \text{Copper(II) sulfate} \longrightarrow \text{Iron(II) sulfate} + \text{Copper}$$

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$

Displacement reactions have various practical applications in real life across different industries and fields. Here are a few examples:

(a) *Extraction of metals*: More reactive metals displace less reactive metals from their compounds. For example, iron is extracted from its ore (iron(III) oxide) using aluminium:

$$F_2O_3(s) + 2Al(s) \longrightarrow 2Fe(s) + Al_2O_3(s)$$

This is a highly exothermic reaction and is used in welding, like in railway tracks.

(b) *Purification of metals*: Less reactive metals are displaced from their solutions by more reactive metals. For example, impure silver is purified by displacing it with a more reactive metal, such as copper.

$$Cu(s) + 2AgNO_3(aq) \longrightarrow 2Ag(s) + Cu(NO_3)_2(aq)$$

(c) *Rust prevention*: In rust prevention treatments, displacement reactions are involved through a principle called *sacrificial protection*. A more reactive metal like zinc is coated onto iron or steel. When both are exposed to air and moisture, zinc reacts (oxidises) first, sacrificing itself to protect the iron from rust. During the process, zinc acts as a sacrificial anode, forming a protective layer of zinc. This is called *galvanisation*.

(d) *Treatment of wastewater*. Metals like aluminium or zinc displace harmful substances or ions from wastewater. For example, zinc is used to remove copper ions from industrial waste solutions.

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

The copper metal is easily recovered as solids from wastewater.

Precipitation reactions

A precipitation reaction is a chemical reaction in which two soluble substances (typically in aqueous solution) combine to give a soluble substance and an insoluble substance known as a precipitate. This reaction is expressed in the general form of:

$$AB(aq) + CD(aq) \longrightarrow AD(s) + CB(aq)$$

This type of reaction is also referred to as a *double displacement reaction.* An example of a precipitation reaction is the reaction between an aqueous solution of silver nitrate and an aqueous solution of sodium chloride to form white precipitates of silver chloride.

Silver nitrate + Sodium chloride $\longrightarrow$ Silver chloride + Sodium nitrate

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

Another example is when an aqueous solution of sodium sulfate is mixed with an aqueous solution of barium chloride to form solid barium sulfate.

$$Na_2SO_4(aq) + BaCl_2(aq) \longrightarrow BaSO_4(s) + 2NaCl(aq)$$

Precipitation reactions are common in various everyday activities. Some of the activities that relate to precipitation reactions include:

(a) *Formation of soap scum:* When soap is used in hard water (which contains calcium and magnesium ions), it reacts with these ions to form insoluble soap scum.

(b) *Curdling of milk:* When lemon juice or vinegar (acid) is added to milk, the casein proteins in the milk precipitate out, causing it to curdle.

(c) *Rainwater formation:* In cloud seeding, silver iodide (AgI) is introduced into clouds, where it reacts with water droplets to form solid ice crystals, leading to rainfall.

(d) *Treating wastewater:* In water treatment, chemicals like aluminium sulfate ($Al_2(SO_4)_3$) and lime ($Ca(OH)_2$) are added to wastewater to precipitate out harmful substances.

(e) *Formation of kidney stones (calcium oxalate crystals)*: In some individuals, excess calcium and oxalate in the urine combine to form calcium oxalate precipitates, which develop into kidney stones.

Redox reactions

A redox reaction is a chemical reaction in which the oxidation number or oxidation state of the participating chemical species changes by losing or gaining one or more electron(s). Redox is a short form for reduction-oxidation reaction. Redox reactions are common and important to some of the basic functions of life including photosynthesis, respiration, combustion, bleaching, digestion, and corrosion or rusting. In a redox reaction, oxidation and reduction reactions occur simultaneously.

A chemical reaction is said to be an oxidation reaction when its reacting substance combines with oxygen, or hydrogen is removed from it. It can also be a loss of electrons from that substance, or increase in oxidation state of that substance. On the other hand, a chemical reaction is said to be a reduction reaction when its reacting substance combines with hydrogen, or when oxygen is removed from that substance. It can also be a gain of electron(s) leading to the decrease in the oxidation state of that substance.

Oxidation and reduction reactions are opposite types of reactions. During a redox reaction, if one substance is oxidised by either gaining oxygen, losing hydrogen or losing one or more electrons, another substance must at the same time be reduced by losing oxygen, gaining hydrogen or gaining one or more electrons. For example, when copper(II) oxide is heated in hydrogen gas, it is reduced to copper metal, while the hydrogen gas is oxidised to water. In this reaction, copper(II) oxide gets reduced because hydrogen takes away the oxygen. Hydrogen is thus a reducing agent. Hydrogen gas gets oxidised because copper(II) oxide gives out its oxygen to it, thus copper(II) oxide is an oxidising agent.

Copper(II) oxide + Hydrogen gas $\xrightarrow{\text{heat}}$ Copper + Steam

addition of oxygen (oxidation)

$$CuO(s) + H_2(g) \xrightarrow{\text{heat}} Cu(s) + H_2O(g)$$

removal of oxygen (reduction)

Another example of a redox reaction is when iron metal is heated in chlorine gas to form iron(III) chloride.

$$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$$

In this reaction, iron metal is oxidised, while chlorine gas is reduced.

$$Fe(s) \longrightarrow Fe^{3+}(aq) + 3e^-$$

$$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$$

Iron is the *reducing agent* since it loses electrons to chlorine atoms. Chlorine gas is the *oxidising agent* since it accepts or gains electrons from iron.

Neutralisation reaction

A neutralisation reaction is chemical reaction between an acid and a base to give salt and water as the products. For example

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

Neutralisation reactions are common in everyday life, particularly in personal care, medicine and agriculture. For example, toothpaste neutralises acid produced by bacteria in the mouth. Antacids neutralise stomach acid.

Exercise 4.2

1. Differentiate oxidation from reduction in terms of electron transfer and changes in oxidation states.
2. Write a balanced chemical equation for each of the following reactions:
   (a) Zinc reacts with silver nitrate solution to produce zinc nitrate and silver.
   (b) Aqueous potassium iodide reacts with aqueous solution of lead(II) nitrate to produce potassium nitrate and lead(II) iodide.
3. Identify the type of reaction in each of the following chemical equations: Explain your answer.
   (a) $CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$
   (b) $CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq)$
   (c) $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$
   (d) $Zn(NO_3)_2(aq) + 2NaOH(aq) \longrightarrow Zn(OH)_2(s) + 2NaNO_3(aq)$
4. A student mixed lead metal with a solution of magnesium chloride but observed no reaction. However, when magnesium metal was added to lead(II) nitrate solution, a reaction occurred. Explain why one reaction occurred while the other did not.
5. The school laboratory has solutions of barium chloride ($BaCl_2$), sodium sulfate (Na_2SO_4), and potassium nitrate (KNO_3).
   (a) Which pair of solutions will produce precipitates when mixed?
   (b) Describe a simple experiment to confirm the presence of a precipitate.
6. Explain how corrosion is related to redox reactions. Suggest one way of preventing iron from rusting.

Chapter summary

1. Chemical equations are short forms of chemical reactions. The reactants are placed on the left hand side of the equation, while the products are placed on the right hand side of the equation.
2. Chemical equations are written by using the chemical symbols or formulas. The state of each of the reactants and products is indicated in parentheses and the equations are balanced.
3. An ionic equation is a chemical equation in which compounds in aqueous solutions or in molten states are written in terms of dissociated ions.
4. Spectator ions are ions that do not change their valence states in a reaction.
5. The net ionic equation is the result of an ionic equation without the spectator ions.
6. There are different types of chemical reactions. These include combination, decomposition, displacement, precipitation, neutralisation and redox reactions.

Revision exercise 4

Choose the correct answer for Questions 1–5. For other questions, provide the answers as per the demands indicated.

1. Which of the following statements describes chemical reactions?
   (a) Occur only in living organisms.
   (b) Occur in water only.
   (c) Produce new substances.
   (d) Only occur outside living organisms.
2. A certain metal hydroxide reacts with hydrochloric acid to produce salt and water only. What type of reaction is this?
   (a) Precipitation
   (b) Displacement
   (c) Neutralisation
   (d) Combination
3. Which of the following is **false** regarding the decomposition of a simple binary compound?
   (a) The products are uncertain.

(b) The elements of the compound become the products.
(c) The reactant is a single substance.
(d) Can have either an ionic or a molecular compound as the reactant.

4. The following processes are applications of neutralisation reactions, except?
   (a) Rubbing baking powder on the bee sting.
   (b) Adding quicklime in the acidic soil.
   (c) Applying vaseline to the burn wound.
   (d) Brushing teeth with toothpaste.

5. Why is it important to understand chemical equations in cooking?
   (a) To balance the temperature of the ingredients.
   (b) Chemical reactions prevent food overcooking.
   (c) It helps measure cooking time accurately.
   (d) To understand how ingredients behave when heated.

6. Analyse the following types of chemical reactions by writing simple balanced chemical equations and explaining at least three real-life applications:
   (a) Combination
   (b) Decomposition
   (c) Displacement
   (d) Precipitation
   (e) Redox

7. Write the word equation for each of the following reactions:
   (a) Burning calcium in oxygen gas.
   (b) Dissolving zinc in dilute sulfuric acid.
   (c) Reacting sodium with water.

8. Balance each of the following chemical equations that involve acids and hypothetical compounds:
   (a) $MCO_3(s) + HCl(aq) \longrightarrow MCl_2(aq) + H_2O(l) + CO_2(g)$
   (b) $MOH(aq) + H_2SO_4(aq) \longrightarrow M_2SO_4(aq) + H_2O(l)$
   (c) $MO(s) + HCl(aq) \longrightarrow MCl_2(aq) + H_2O(l)$

9. Complete and balance each of the following chemical equations:
   (a) $Zn(s) + Cl_2(g) \longrightarrow$
   (b) $Fe(s) + O_2(g) \longrightarrow$
   (c) $MgCl_2(aq) + AgNO_3(aq) \longrightarrow$

10. Write the net ionic equation for the reaction between hydrochloric acid and potassium hydroxide solutions.

11. (a) Differentiate between ionic equations and molecular equations and provide their significance in reaction analysis.
    (b) Compare total ionic equations and net ionic equations, illustrating their importance in predicting reaction outcomes and identifying spectator ions in chemical processes.

12. Study the following chemical equation of a reaction and answer the questions that follow.
$$3KOH(aq) + FeCl_3(aq) \longrightarrow 3KCl(aq) + Fe(OH)_3(s)$$
    (a) What type of the reaction is represented by the chemical equation?
    (b) Write a net ionic equation for the chemical reaction.

13. Explain the uses of balanced chemical equations in improving efficiency, safety, and cost-effectiveness in chemical manufacturing industries.

14. Describe the effect of heat on the decomposition of the following compounds and represent these reactions using balanced chemical equations:
    (a) Iron(II) sulfate
    (b) Calcium hydrogencarbonate
    (c) Ammonium nitrate
    (d) Copper(II) carbonate

15. Why combustion reactions are considered exothermic? Use an example to support your explanation.

16. A student placed a strip of zinc metal in copper(II) sulfate solution. After a while, the blue colour of the solution faded, and a reddish-brown deposit formed on the zinc strip.
    (a) Explain the observations made and identify the type of reaction that occurred.
    (b) What would happen if a silver plate replaced zinc in the same solution?

17. A farmer wants to remove calcium ions from hard water using a precipitation reaction.
    (a) Suggest the chemicals that can be added to precipitate calcium ions.
    (b) Write the balanced chemical equations for the reactions involved.

Chapter Five

Acids, bases, and salts

Introduction

Acids, bases, and salts are essential for the proper functioning of our bodies and the health of the environment. Medicines, foodstuffs, and fertilisers are among the substances that contain acids, bases, or salts. In this chapter, you will learn about acids, bases, acid-base indicators, salts, their properties, and uses. The competencies developed will enable you to prepare and use acids, bases, acid-base indicators and salts, as well as apply the concept of acid-base neutralisation in daily life.

Think

The usefulness of acids, bases and salts in daily life

Acids

Task 5.1

Use reliable online simulations, and other resources to explore the concepts of acids and their applications in daily life.

An *acid* is any substance which produces hydrogen ions (H^+) in water as the positively charged ions. In this case, acids are substances made of hydrogen ions and other ions.

Some acids occur naturally, but others are synthesised (do not occur naturally). Therefore, acids can be categorised as natural acids and synthetic acids. The substances which contain natural acids include sour milk and citrus fruits, such as oranges, lemons, and limes. Figure 5.1 and Table 5.1 show some natural sources of acids.

Citrus fruits Pineapples Plums Passion fruits

Figure 5.1: *Some fruits with natural sources of acids*

Table 5.1: *Examples of natural acids and their sources*

Name of natural acid	Natural sources
Citric acid	Citrus fruits such as lemons and oranges
Lactic acid	Sour milk, cheese, and other fermented dairy products
Acetic acid	Vinegar and rotten fruits such as grapes, pineapple and oranges
Tartaric acid	Grapes, bananas and tamarind
Oxalic acid	Spinach and tomatoes
Formic acid	Bee or ant stings
Ascorbic acid	Fruits and vegetables, for example, tomatoes, pawpaws, broccoli, oranges and grape fruits
Folic acid	Fruits, vegetables, nuts and grains
Malic acid	Apples, strawberries and plums

On the other hand, synthetic acids are artificially produced acids. Examples of synthetic acids include sulfuric acid, nitric acid and hydrochloric acid.

Acids can be classified as *mineral (inorganic) acids* and *organic acids*. Mineral acids are synthesised in industries from inorganic substances, while organic acids are synthesised in industries or obtained directly from natural organic materials. Examples of mineral acids include nitric acid (HNO_3), sulfuric acid (H_2SO_4), hydrochloric acid (HCl), carbonic acid (H_2CO_3), phosphoric acid (H_3PO_4), nitrous acid (HNO_2), and sulfurous acid (H_2SO_3). Examples of organic acids include ethanoic acid/acetic acid (CH_3COOH) and methanoic acid (HCOOH). The most common acids that are found in a chemistry laboratory include hydrochloric acid, sulfuric acid, nitric acid, and acetic acid. Common acids in the laboratory are stored in various containers such as plastic bottles or glass bottles as shown in Figure 5.2.

Figure 5.2: *Containers for storing acids in the laboratory*

Properties of acids

Acids have certain physical and chemical properties.

Physical properties of acids

The following are the physical properties of acids:

1. Acids have a sour taste: A sour taste of vinegar, sour milk, and many unripe fruits such as lemons, is caused by the acids present in them.
2. Most acids are water soluble: They can be diluted with water to reduce the intensity of their acidity. Diluted acids feel watery unlike concentrated ones.
3. Acids have the ability to conduct electricity: When acids are dissolved in water, they release ions which conduct electricity. The strength of acids as electrolytes varies; strong acids ionise completely, while weak acids ionise partially. This is why some acid solutions are commonly used in car and solar batteries.
4. Acids have varied boiling and melting points, depending on their molecular structures. Some acids have high boiling points, for example, sulfuric acid (337 °C), whereas others have low boiling points, for example formic acid (100.8 °C).
5. Acids are corrosive in nature: Acids can have a corrosive action on various substances such as paper as well as clothes. They can also damage or even destroy metals. Therefore, they should be handled with extreme care; otherwise they can corrode the skin.

6. Acids are mostly present in liquid or gaseous forms, but there are some which exist in solid form such as oxalic acid.
7. Acids change the colour of phenolphthalein indicator to colourless.
8. Acids change the colour of methyl orange indicator to red.

Chemical properties of acids

Acids have various chemical properties, including the following:

1. Acids have a pH less than 7 on a pH scale. Low pH values indicate high acidity, for example, hydrochloric acid (HCl) has a pH between 1.5 and 3.5, nitric acid (HNO_3) has pH around 3, and sulfuric acid (H_2SO_4) has pH between 2 and 3.
2. Solution of an acid turns blue litmus paper red (Figure 5.3). This is used in the identification of acids.

Figure 5.3: *Blue litmus paper immersed in an acid solution*

3. Acids react with carbonates and hydrogencarbonates to produce salt, water, and carbon dioxide gas as shown in the following chemical equations:

$$2HCl(aq) + CaCO_3(s) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$
$$H_2SO_4(aq) + Ca(HCO_3)_2(aq) \longrightarrow CaSO_4(s) + 2H_2O(l) + 2CO_2(g)$$

One test of the produced gas (carbon dioxide) is to bubble it through limewater. The limewater will turn milky or cloudy white. The laboratory experimental set-up for the reactions of acids and carbonates is shown in Figure 5.4.

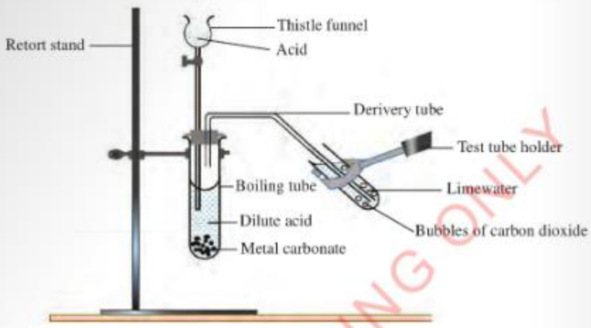


Figure 5.4: *Set-up for the reaction of acids with carbonates*

4. Acids react with metals above hydrogen in the reactivity series to form salt and hydrogen gas. Some examples are shown in the following equations:

$$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$$
$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$
$$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$$

Note: (a) The reaction between lead and dilute hydrochloric acid or sulfuric acid forms an insoluble coating of lead chloride and lead sulfate, respectively which prevents further reaction. The two reactions are very slow and the process may take hours before a few bubbles of hydrogen gas are observed. Copper does not react with dilute hydrochloric acid and sulfuric acid because it is below hydrogen in the reactivity series of metals. See Appendix 3.

(b) The reactions of dilute nitric acid and most metals produce water instead of hydrogen gas. This is because nitric acid is a strong oxidising agent. Its oxidising power depends on the concentration of the acid; the higher the concentration of acid, the higher the oxidising power of the acid. The following equations present the reaction between metals and concentrated nitric acid:

$$Pb(s) + 4HNO_3(l) \longrightarrow Pb(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$
$$Ca(s) + 4HNO_3(l) \longrightarrow Ca(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$$

5. Acids react with metal oxides or hydroxides to form salt and water only, as shown in the following examples:

$$CuO(s) + 2HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + H_2O(l)$$
$$Ca(OH)_2(aq) + 2HNO_3(aq) \longrightarrow Ca(NO_3)_2(aq) + 2H_2O(l)$$

Activity 5.1

Aim: To investigate the reactions of dilute acids with metals

Requirements: Test tubes, test tube rack, retort stand and clamp, wooden splints, lighter, 2 M sulfuric acid, 2 M hydrochloric acid, 2 M nitric acid, zinc, magnesium ribbon, lead, and copper

Procedure

1. Transfer about 5 cm³ of 2 M sulfuric acid, 2 M nitric acid, and 2 M hydrochloric acid in three separate test tubes.
2. Add a small piece of zinc metal to each test tube, one at a time, and record your observations.
3. Put a lighted splint at the mouth of each test tube, one at a time, as shown in Figure 5.5, and record the observations.
4. Repeat Steps 2–3 with magnesium, lead, and copper, each at a time instead of zinc metal.

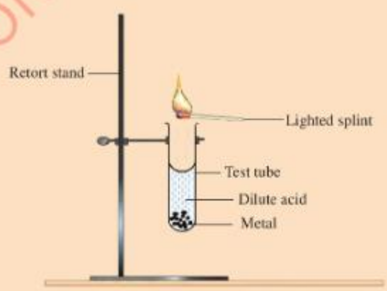


Figure 5.5: *Set-up for the reaction of dilute acids with metals*

Questions

1. What are the balanced chemical equations for the reaction between:
   (a) zinc and dilute hydrochloric acid?
   (b) zinc and dilute sulfuric acid?
   (c) magnesium and dilute hydrochloric acid?
   (d) magnesium and dilute sulfuric acid?
2. What causes the observations recorded in Step 3 of the experiments? Support with balanced chemical equations.
3. What happened in Steps 2 and 3 when lead and copper were used in the experiment? Write the balanced chemical equations for the reactions of these metals with sulfuric acid.

Task 5.2

Use virtual laboratories or interactive simulations and animations to explore the strength and the basicity of acids.

Strengths of acids

The strength of an acid is a measure of its ability to ionise (dissociate) in water to produce hydrogen ions (H^+). The more the acid ionises in water, the stronger it is. The strength of the acid solution can be determined by its pH value. *The pH is the measure of the degree of acidity or alkalinity of the solution.* Based on their strengths, acids are grouped into strong acids and weak acids.

Strong acids ionise completely in water to give large amounts of H^+. This complete ionisation makes them highly reactive. Common examples of strong acids are mineral acids such as hydrochloric acid, nitric acid, and sulfuric acid. For example, hydrochloric acid ionises completely in water to give H^+ and Cl^-.

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

Weak acids ionise partially in water to produce small amounts of H^+. This means that not all acid molecules release H^+. Common examples of weak acids are ethanoic acid, carbonic acid, and citric acid. For example, ethanoic acid (CH_3COOH) ionises partially in water to produce few CH_3COO^- and H^+.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

Note:

(a) The pH values of solutions increase with decrease in acidity; therefore, weak acids have higher pH values than strong acids.

(b) Solutions of strong acids are good conductors of electricity because they contain more free mobile ions to carry the charges than solutions of weak acids.

Basicity of an acid

Basicity of an acid is the number of ionisable hydrogen atoms per molecule of the acid that can be displaced by a metal in solution. For example, hydrochloric acid is *monobasic* as it has one ionisable hydrogen atom that can be displaced. Sulfuric acid and phosphoric acid are *dibasic* and *tribasic*, respectively. Sulfuric acid has two ionisable hydrogen atoms, while phosphoric acid has three ionisable hydrogen atoms that can be displaced (Table 5.2).

Table 5.2. *Basicity of some acids*

Acid	Basicity
$HCl \longrightarrow H^+ + Cl^-$	Monobasic
$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$	Dibasic
$H_3PO_4 \longrightarrow 3H^+ + PO_4^{3-}$	Tribasic

Exercises 5.1

1. Describe how the degree of ionisation leads to differences in acidic strength.
2. Should the strength of an acid be the only factor in determining its use in industry, home or laboratory? Explain your answer.
3. A particular skincare product contains glycolic acid. Why is it important to know the strength of the acid before using it on the skin?
4. Why do the physical properties of acids make them suitable for cleaning products?

Bases

A *base* is a substance that neutralises an acid by reacting with hydrogen ions. Bases include the oxides, hydroxides, and carbonates of metals.

Metal oxides are those compounds which contain a metal and an oxygen atom. Examples of metal oxides include sodium oxide, Na_2O; potassium oxide, K_2O; calcium oxide, CaO; magnesium oxide, MgO; and copper(II) oxide, CuO.

Hydroxides are those compounds which contain a hydroxy group (–OH). Examples of hydroxides include calcium hydroxide, $Ca(OH)_2$; magnesium hydroxide, $Mg(OH)_2$; sodium hydroxide, NaOH; potassium hydroxide, KOH; and ammonium hydroxide, NH_4OH.

Metal carbonates are made up of a metal ion and a carbonate ion. Examples of metal carbonates include sodium carbonate, Na_2CO_3; potassium carbonate, K_2CO_3; magnesium carbonate, $MgCO_3$; and calcium carbonate, $CaCO_3$.

Bases are abundant in nature. Examples of natural substances that contain bases are ashes, banana peels, and avocado. In addition, other bases occur naturally on land and in water bodies; such bases include soda ash (sodium carbonate), baking soda (sodium bicarbonate) and limestone (calcium carbonate) Figure 5.6 shows some sources of natural bases.

(a) Soda ash deposit (b) Limestone deposit

Figure 5.6: *Some sources of natural bases*

Many common household products such as detergents, deodorants, toothpaste, and baking powder contain bases. These products contain synthesised bases.

Some bases are readily soluble in water, while others are not. Soluble bases are called alkalis. An *alkali* is a soluble base which, when dissolved in water, forms hydroxide ions. Therefore, all alkalis are bases, but not all bases are alkalis. Examples of soluble and insoluble bases are shown in Table 5.3.

Table 5.3: *Examples of soluble and insoluble bases*

Insoluble bases	Soluble bases
Copper(II) hydroxide ($Cu(OH)_2$)	Sodium hydroxide (NaOH)
Lead(II) hydroxide ($Pb(OH)_2$)	Potassium hydroxide (KOH)
Calcium hydroxide ($Ca(OH)_2$)	Sodium oxide (Na_2O)
Copper(II) oxide (CuO)	Potassium oxide (K_2O)
Iron(III) oxide (Fe_2O_3)	Sodium carbonate (Na_2CO_3)
Lead(II) oxide (PbO)	Potassium carbonate (K_2CO_3)
Lead(II) carbonate ($PbCO_3$)	
Magnesium carbonate ($MgCO_3$)	

Note: Calcium hydroxide and lead hydroxide are moderately soluble in water. Oxides and hydroxides of other metals are insoluble except those of Group I elements.

Properties of bases

Bases have physical and chemical properties that differ from those of acids.

Physical properties of bases

Bases have the following physical properties:

1. Most bases have a bitter taste. For example, milk of magnesia, a common antacid.
2. Bases have a 'soapy' or slippery feel. For example, a rub of a drop of baking powder solution between fingers gives a slippery feel.
3. Most bases are insoluble in water.
4. Bases generally do not have odour except for ammonia which has a pungent smell.
5. Bases are corrosive depending on their pH and concentrations.
6. Soluble bases (alkalis) conduct electricity when dissociated into ions.

Chemical properties of bases

Bases have the following chemical properties:

1. The pH values of bases are greater than 7.
2. Bases turn red litmus paper blue as shown in Figure 5.7.

Figure 5.7: *Red litmus paper immersed in an alkali solution*

3. Bases change the colour of *phenolphthalein* (POP) indicator pink and that of *methyl orange* (MO) indicator yellow.
4. Soluble bases react with most cations to precipitate hydroxides. In this case, they precipitate insoluble metal hydroxides from their salt solutions. For example, potassium hydroxide reacts with iron(III) chloride to yield iron(III) hydroxide and potassium chloride.

$$3KOH(aq) + FeCl_3(aq) \longrightarrow Fe(OH)_3(s) + 3KCl(aq)$$

Ionically, the reaction is represented as:

$$3OH^-(aq) + Fe^{3+}(aq) \longrightarrow Fe(OH)_3(s)$$

5. Bases react with acids to form salt and water. For example, sodium hydroxide reacts with dilute hydrochloric acid to produce sodium chloride and water.

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

The ionic equation is

$$OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$$

6. Alkalis react with ammonium salts to produce ammonia gas.

$$KOH(aq) + NH_4Cl(s) \longrightarrow KCl(aq) + H_2O(l) + NH_3(g)$$

Strengths of bases

The strength of a base or alkali is its ability to ionise in aqueous solution to produce hydroxide ions (OH^-). Like acids, the strength of a solution of a base can be determined by the pH value of that particular solution. The more the base ionises in aqueous solution, the stronger it is. Alkalis are strong bases because they ionise completely in aqueous solutions to produce free ions. Examples of strong bases are potassium hydroxide (KOH) and sodium hydroxide (NaOH).

$$KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$$

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

Weak bases are those bases which do not completely dissociate into their ions in aqueous solutions. The most common example of a weak base is ammonium hydroxide.

$$NH_4OH(aq) \rightleftharpoons NH_4^-(aq) + OH^-(aq)$$

Acidity of a base

The *acidity* of a base is the number of hydroxide ions that a basic molecule can produce in the aqueous solution. On the basis of acidity, bases are classified into three types: monoacidic, diacidic, and triacidic bases. Table 5.4 shows some of the bases with their acidity.

Table 5.4: *Acidity of some bases*

Base	Acidity
$NaOH \longrightarrow Na^+ + OH^-$	Monoacidic
$Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$	Diacidic
$Al(OH)_3 \longrightarrow Al^{3+} + 3OH^-$	Triacidic

Activity 5.2

Aim: To identify acidic and basic substances

Requirements: Coconut, citrus fruits, ripe tomatoes, vinegar, sour milk, wood ash, pieces of chalk, yoghurt, blender, knife, water, red litmus papers, blue litmus papers, and beakers

Procedure

1. Prepare juices of coconut, citrus fruits, and ripe tomatoes.
2. Use litmus papers to test the acidity or basic nature of the juices obtained in Step 1.
3. Mix some wood ash and chalk powder, each with little water, decant to obtain a *supernatant* and use litmus papers to test the acidity or basic nature of the supernatant.
4. Take small amounts of vinegar, sour milk, and yoghurt into separate beakers and test for acidity or basic nature using litmus papers.

Questions

1. Which substances are acidic and which ones are basic?
2. How relevant is this activity to your daily life?

Task 5.3

Explore in books, internet sources and substances used at home or school to identify natural substances that contain acids or bases, then write a summary.

Neutralisation reaction

If an acid and a base are mixed together in correct amounts, a neutral solution is produced. A reaction of this type is called *neutralisation*. Neutralisation is therefore, a reaction between an acid and a base to produce salt and water. It is essentially the reaction between the hydroxide ions found in the basic solution and the hydrogen ions found in an acidic solution. The reaction is referred to as neutralisation because the resulting products are neither basic nor acidic.

$$\text{Acid} + \text{Base} \longrightarrow \text{Salt} + \text{Water}$$

Neutralisation reactions that involve the reaction of an acid with a carbonate or hydrogencarbonate produce carbon dioxide gas in addition to salt and water.

$$\text{Acid} + \text{Carbonate/hydrogencarbonate} \longrightarrow \text{Salt} + \text{Water} + \text{Carbon dioxide}$$

The following are examples of neutralisation reactions between bases and dilute acids:

Nitric acid reacts with potassium hydroxide to yield potassium nitrate and water.

$$HNO_3(aq) + KOH(aq) \longrightarrow KNO_3(aq) + H_2O(l)$$

Sulfuric acid reacts with calcium hydroxide to yield calcium sulfate and water

$$H_2SO_4(aq) + Ca(OH)_2(aq) \longrightarrow CaSO_4(s) + 2H_2O(l)$$

Hydrochloric acid reacts with sodium carbonate to yield sodium chloride, water, and carbon dioxide gas.

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

Hydrochloric acid reacts with sodium hydrogencarbonate to yield sodium chloride, water, and carbon dioxide gas.

$$HCl(aq) + NaHCO_3(aq) \longrightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

To identify the end point for an acid-base reaction, an indicator must be added. For example, when phenolphthalein indicator is used in a particular *titration* (by running the acid from a burette to a base in a flask), the pink colour of the content in the flask starts to change slowly to colourless. The *end point* is marked when the reaction mixture in the flask just turns colourless. During neutralisation of a base with an acid, the volume of the base in a flask is normally known. The volume of the acid that has neutralised the alkali is called a *titre value* or *titre volume*.

To get the exact volume (V) of the acid that neutralises a base, the initial burette reading, V_1 is subtracted from the final burette reading V_2.

$$V = V_2 - V_1$$

Activity 5.3

Aim: To investigate the neutralisation reaction between sodium hydroxide and dilute hydrochloric acid

Requirements: Pipette, burette, retort stand and clamp, 10-mL measuring cylinder, conical flasks, beakers, evaporating dish, white tile, heat source, tripod stand, wire gauze, 0.1 M sodium hydroxide, 0.1 M hydrochloric acid, and phenolphthalein indicator

Procedure

1. Transfer 25 cm^3 (or 20 cm^3) of sodium hydroxide solution into a conical flask by using a pipette.
2. Add 2 to 3 drops of phenolphthalein indicator to the solution.
3. Set the apparatus as shown in Figure 5.8.

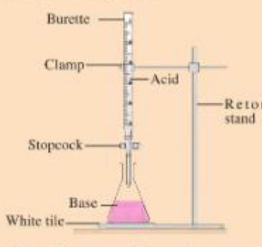


Figure 5.8: *Set-up of the acid-base titration*

4. Fill the burette with dilute hydrochloric acid and record the initial volume, V_1.
5. Run the acid into the alkali (base), a few drops at a time, until the indicator just changes colour. Swirl the components of the flask after every addition of the acid.
6. Record the new level of the remaining acid in the burette as V_2. Calculate the volume of the acid used in the reaction.
7. In a clean beaker, mix the same volume of the base used with the same volume of the acid obtained from the titration. Do not add any indicator.
8. Transfer 10 cm^3 of the resulting solution into an evaporating dish and carefully heat the solution until crystals start to form.
9. Stop heating and allow the remaining liquid to evaporate, leaving behind a white solid.

Questions

1. Why is the phenolphthalein indicator important in an acid-base experiment?
2. What is the balanced chemical equation representing the reaction in the experiment? Give the name of the white solids obtained in Step 9.
3. What is the ionic equation of the reaction in the experiment?

Applications of neutralisation reactions

Neutralisation reactions have many applications in daily life. Some of the applications are explained as follows:

Treating insect stings and bites

Insect bites or stings, and plants with stinging hair often cause small and red lumps on the skin which may be painful and itchy. Insects such as bees, have stings that inject an acidic liquid in the blood through the skin. The stings can be neutralised by rubbing baking soda on the affected area. Ant bites and nettle (a plant with stinging hair) stings contain methanoic acid which is neutralised by using baking soda or other alkaline substances such as cucumber and avocado. Wasp stings are alkaline and can be neutralised with vinegar which contains acetic acid. Figure 5.9 shows different sources of stings or bites.

Figure 5.9: *Organisms which produce acids and alkalis*

Relieving indigestion

Indigestion is a discomfort in the stomach that is associated with difficulty in digesting food. It is caused by the presence of excess acid like hydrochloric acid in the stomach. The excess acid can be neutralised by taking a liquid or tablets that contain magnesium or sodium hydrogencarbonate commonly known as antacids.

Soil treatment

Most plants grow well in soils that have optimal pH values. When soils are too acidic, for example, due to applications of ammonium based fertilisers, chemicals such as calcium oxide (quicklime) and calcium hydroxide (slaked lime) are added to adjust the soil pH (Figure 5.10). Such chemicals that neutralise the soil acidity are called *liming materials*.

Figure 5.10: *Soil liming*

Treating factory wastes

Liquid wastes from factories often contain acids and bases. If the wastes get into water bodies such as lakes, ponds and rivers, they can harm aquatic organisms like fish. Acidic wastes can be controlled by adding bases such as calcium hydroxide to neutralise them.

Preventing formation of acid rain

When the pH of rainwater falls below 5.6, it is called *acid rain*. Acid rain is caused by chemical reactions between rainwater and gases such as sulfur dioxide and nitrogen dioxide which are released into the atmosphere. Acid rain increases the acidity of soils, rivers and lakes and adversely affects vegetation and aquatic organisms. To reduce this problem, air pollution devices containing bases are fitted in exhaust pipes and chimneys to neutralise the acidic compounds before reaching the atmosphere.

Neutralising accidental spills

If an acid or an alkali spills on the floor or work surface in the laboratory, it can be neutralised. For example, sulfuric acid which is very corrosive, can be neutralised by adding sodium hydroxide.

Manufacturing fertilisers

The production of ammonium fertilisers is done through the neutralisation of ammonia with a mineral acid. Ammonium nitrate for example, is produced by the reaction of ammonia with nitric acid.

$$NH_3(g) + HNO_3(aq) \longrightarrow NH_4NO_3(aq)$$

Ammonia gas also reacts with sulfuric acid to give ammonium sulfate $(NH_4)_2SO_4$ fertiliser (Figure 5.11).

$$2NH_3(g) + H_2SO_4(aq) \longrightarrow (NH_4)_2SO_4(aq)$$

Figure 5.11: *Sack of ammonium sulfate fertiliser*

Uses of acids and bases

Task 5.4

Explore reliable books, online, or home resources to identify three everyday products containing acids and three containing bases.
For each product or material you find:

1. Write the trade name, common name and chemical name of the product.
2. Identify the acid or base it contains (if known).
3. Describe its use or purpose in daily life.

Acids and bases have many uses in daily life. For instance, some plants grow well in acidic soil, but others grow better in a neutral or alkaline soil. Farmers need to know the pH of the soil because the soil can be too acidic causing inability of plants to grow well and reduce the yields. Bases can be used by farmers to reduce the acidity of the soil, and acids can be used to correct the pH for the basic soils.

Amino acids and some vitamins such as vitamin C are important for our health. Sulfuric acid is used in car batteries. Citric acid is used in food preservation. Bases such as sodium hypochlorite are used in bleaching to remove stains on clothes. Ammonium hydroxide is used as a glass cleaner, and magnesium hydroxide is used as a laxative to regulate bowel movements. Magnesium hydroxide is also used as an antacid to relieve indigestion, sour stomach and heartburn. Tables 5.5 and 5.6 summarise the uses of some acids, and bases, respectively.

Table 5.5: *Uses of some acids in daily life*

Acids	Uses
Hydrochloric acid	(a) Cleaning metals before electroplating
	(b) Household cleaning
	(c) Leather processing
	(d) Swimming pool maintenance
	(e) Production of inorganic compounds such as chlorides
	(f) Production of organic compounds, such as vinyl chloride and dichloromethane during plastic manufacturing
	(g) Manufacture of fertilisers, textiles, and rubber

Acids	Uses
Sulfuric acid	(a) Removal of rust from iron and steel in industries (b) Manufacture of synthetic textiles such as nylon (c) Manufacture of fertilisers (d) Used as electrolyte in batteries (e) Production of sulfates
Nitric acid	(a) Production of nitrogen containing fertilisers, explosives, and rocket propellants (b) Manufacture of nylon (c) Used in testing the purity of precious stones
	(d) Used as an oxidising agent (e) Metal cleaning (f) Synthesis of dyes
Tartaric acid	(a) Manufacturing of soft drinks (b) Provides tartness (sourness) to food (c) Used as an emetic (a substance that induces vomiting)
Carbonic acid	(a) To make bubbly drinks like soft drinks and sparkling wine (b) Used in pharmaceuticals, cosmetics, fertilisers, and food processing
Ethanoic acid	(a) A main ingredient of vinegar (b) Making dyes and paints (c) Degreasing solvent and cleaning agent
Phosphoric acid	(a) Rust proofing (b) Used as a catalyst (c) Used as an ingredient in softening drinks and dental cements (d) Manufacturing of phosphates for softening water (e) Manufacturing of detergents and fertilisers (f) Used as a food additive lending acidic properties to soft drinks and other prepared foods
Benzoic acid	(a) Used to preserve food (b) Manufacture of a variety of products such as perfumes, dyes, medicines, and insect repellents

Table 5.6: *Uses of some bases in daily life*

Bases	Uses
Ammonia	(a) Production of fertilisers (b) Prevention of premature coagulation in natural or synthetic latex, and is used as a refrigerant (c) Purification of water supplies (d) Manufacture of plastics, explosives, textiles, pesticides, dyes, and other chemicals
Calcium hydroxide	(a) Neutralisation of the acidity of soil (b) Preparation of ammonia (c) Used in water and sewage treatments
Sodium hydroxide	(a) Production of soaps, dyes, detergents, and cleaners (b) Metal cleaning and processing (c) Processing cotton fabrics (d) Electroplating and electrolytic extraction
Magnesium hydroxide	(a) Pharmaceuticals as antacids (b) Neutralises acidic wastewater because it is a non-hazardous alkali (c) Used as pH control in the food industry (d) Used in chip fabrication for semiconductors
Aluminium hydroxide	(a) Manufacturing of other aluminium compounds, making gastric medicine (antacid), and fire extinguishers (b) Relief of heartburn, sour stomach, and peptic ulcer pain
Potassium hydroxide	(a) Used as an electrolyte in alkaline batteries and in electroplating (b) Making soap, paint, and varnish removers

Exercise 5.2

1. Group the following substances into either acids, bases or neutral solutions: Ash solution, fresh milk, sour milk, liquid soap, chalk solution, tomato juice, lemon juice, sugar solution, table salt solution, and cucumber juice.
2. Strong bases, such as sodium hydroxide, are used in industrial cleaning, while weak bases, such as baking soda, are safe for home uses. Why is this difference important in terms of safety and effectiveness?

3. A bee stung a gardener while working in the garden. What household material can be used to relieve the pain in the stung area? Explain your answer.

Acid-base indicators

Acid-base indicators are chemicals that are used to determine the chemical nature of a substance; whether it is acidic or basic. The acid-base indicators are also known as *pH indicators* because acidity and alkalinity relate to pH range. Indicators normally change colour to indicate the presence or absence of acids or bases. The pH indicators operate efficiently over a certain pH range.

pH scale

pH scale is a scale of numbers, from 0 to 14 which is used to express acidity, neutrality, and alkalinity. The pH 7 indicates neutrality of a substance. Any substance that has a pH below 7 is acidic, while a substance with pH above 7 is basic. The smaller the number below 7 in the scale, the more acidic is the substance. Similarly, the greater the number above 7, the more basic is the substance. This means that an acid with pH 1 is more acidic than that with pH 6, and the base with pH 14 is more basic than that with pH 8. Figure 5.12 shows the trends in acidity and alkalinity in the pH scale.

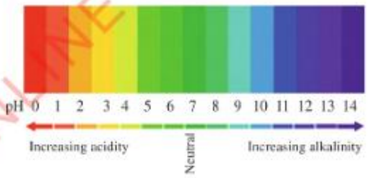


Figure 5.12: *Trends in acidity and alkalinity*

Occurrence of pH indicators

pH indicators can occur naturally or be synthesised.

Natural pH indicators

A natural pH indicator is a substance which is found naturally and can be used to determine whether the substance is acidic or basic. Examples of natural pH indicators are turmeric, grape juice, red cabbage, cassava leaves, onion, beetroot, and coloured flowers like that of rosella and hibiscus. For example, red cabbage juice is pink in acids and green in bases. An extract of hibiscus flowers has a

red colour in an acid and greenish yellow colour in a base. Onion is one of the olfactory indicators which diminishes its smell in a base and remains as it is in an acid. Vanilla essence has similar characteristics to onion. The smell of vanilla essence disappears when it is added to a base, but persists when it is added to an acid.

The most commonly used natural indicator is litmus which is extracted from lichens (Figure 5.13). Litmus has purple, red, and blue colours in distilled water, acidic solution, and basic solution, respectively. Litmus is available either in solution, or in the form of strips of paper.

Figure 5.13: *Lichen*

Activity 5.4

Aim: To prepare a pH indicator from hibiscus flowers

Requirements: Hibiscus flowers, water, beakers, droppers, mortar and pestle, knife, 50% ethanol solution, heat source, evaporating dish, dilute hydrochloric acid, and dilute sodium hydroxide

Procedure

1. Cut the hibiscus flowers into small pieces and then crush them by using mortar and pestle.
2. Add about 20 cm^3 of ethanol solution into the crushed hibiscus flowers and continue crushing. Decant the mixture obtained and then keep the solution obtained aside.
3. Boil the solution obtained to make sure that all the ethanol evaporates. Leave the residue to cool.

4. Add about 20 cm^3 of distilled water to the residue and stir to obtain an indicator solution.
5. Establish the colour of an indicator prepared by testing it with hydrochloric acid and sodium hydroxide, one at a time.

Questions

1. How can the colour changes observed with hibiscus flower extract help determine whether a substance is an acid or a base?
2. Why was ethanol used as a solvent in the extraction instead of water?
3. What other substances found in your environment can be used as acid-base indicators?

Activity 5.5

Aim: To prepare a pH indicator from cassava leaves

Requirements: Fresh cassava leaves, distilled water, ethanol, mortar and pestle or blender, filter paper, test tubes or small clear containers, dropper or pipette, known acidic and basic solutions (for example, vinegar, lemon juice, baking soda solution, soap solution), pH paper or a pH meter

Procedure

1. Collect fresh cassava leaves and rinse them thoroughly with distilled water to remove any dirt or impurities.
2. Chop the leaves into small pieces and then grind them thoroughly using a mortar and pestle. Alternatively, use a blender to blend the leaves into a fine pulp.
3. Add a small amount of distilled water to facilitate the grinding or blending process.
4. Add a small volume of ethanol to the mixture to enhance the extraction of pigments, especially *anthocyanins*.
5. Filter the homogenous mixture using filter paper to separate the liquid extract from the solid leaf residues. Collect the filtrate in a clean beaker.
6. Establish the colour of an indicator prepared by testing it with the acidic and basic solutions available, one at a time.

Questions

1. What colour changes were observed using cassava leaf extract indicator in the substances tested?
2. How can the knowledge gained from this activity be used in testing spoiled fruits?

Synthetic pH indicators

Synthetic pH indicators are chemical substances which are made from different substances in the laboratories used to determine the pH of substances. Examples of synthetic indicators include universal indicator, phenolphthalein (POP), and methyl orange (MO).

Universal indicator

The universal indicator is a chemical substance that displays several colour changes across a wide range of pH values as shown in Table 5.7. Basically, the universal indicator is a blend of pH indicator solutions which include phenolphthalein, methyl red, bromothymol blue, and thymol blue. When an acid or base solution is added to the universal indicator, a new colour is produced, which is used to determine the pH value of the acid or the base solution by matching the colour with the colours on the pH colour chart.

Table 5.7: *pH ranges of the universal indicator*

pH value	Colour of the universal indicator
1-2	Red
3	Pink
4	Brown
5	Yellow
6-8	Green
9-10	Blue
11-12	Indigo
13-14	Violet

Other indicators

The indicators such as methyl orange, methyl red, bromothymol blue, phenol red, thymol blue and phenolphthalein are used to test the acidity or alkalinity of acids and bases. Table 5.8 shows the colour changes and the pH ranges of selected indicators.

Table 5.8: *Colour changes of common indicators in acidic and basic solutions*

Indicator	Acidic	Basic	pH range
Phenolphthalein	Colourless	Pink	8–10
Methyl orange	Red	Yellow	3–5
Methyl red	Yellow	Red	5–8
Phenol red	Yellow	Red	7–8
Bromothymol blue	Yellow	Red	6–8

Activity 5.6

Aim: To test the acidity and alkalinity of substances by using synthetic indicators

Requirements: Dilute sulfuric acid, sodium hydroxide solution, red and blue litmus papers, litmus solution, beakers, phenolphthalein, bottled drinking water, and methyl orange

Procedure

1. Pour some dilute sulfuric acid and sodium hydroxide solution into four separate beakers each.
2. Insert a blue litmus paper into one of the beakers containing sulfuric acid. Record the observations.
3. Insert a red litmus paper into one of the beakers containing sodium hydroxide solution. Record the observations.
4. Add two to three drops of litmus solution, phenolphthalein, and methyl orange indicators to the other three beakers with the acidic solution and basic solution separately, and record your observations.
5. Test the acidity or alkalinity of bottled drinking water using all the above indicators and record your observations.

Questions

In Step 5, different results were obtained despite using the same indicators to test acidic and basic solutions.

(a) What do the results indicate?

(b) What other substances may give similar results?

Uses of pH indicators

pH indicators are used in a variety of ways, including the following:

1. They are used to give rough pH values of chemical solutions and drinks such as fruit juices and water.
2. They are used in acid-base titrations to show the completion of a reaction (end point).
3. They are used in agriculture to test the acidity or alkalinity of soil.
4. They are used to test the pH of water used in swimming pools to ensure that the water used is approximately neutral.
5. They are used in wastewater monitoring to avoid environmental pollution.

Exercise 5.3

1. Explain the importance of using indicators in activities involving acids and bases.
2. Explain one advantage and one limitation of using natural indicators.
3. How does the pH sensitivity of a synthetic indicator differ from that of a natural indicator, such as hibiscus extract?

Salts

Task 5.5

Use online simulation and interactive tools or any other reliable resources to investigate the properties, types, and everyday applications of salts.

A *salt* is an ionic compound which is made up of positively charged ions (cations) and negatively charged ions (anions). Salts occur naturally and are also synthesised in the laboratory. The main natural sources of salts include sea water and rocks. Different salts are used in our homes and industries for various purposes. These include table salt (NaCl), washing soda ($Na_2CO_3 \cdot 10H_2O$), Plaster of Paris (POP) ($CaSO_4 \cdot \frac{1}{2}H_2O$) and bleaching powder ($Ca(OCl)_2$).

Types of salts

Salts are categorised into neutral salts, acidic salts, and basic salts depending on their pH values when dissolved in water. Salts can also be categorised as normal salts, and double salts depending on their formation and compositions.

Neutral salts

These are formed by the reactions of strong acids and strong bases. They are called neutral salts because their aqueous solutions are neutral to litmus. Neutral salts, when categorised based on formation and composition, are termed as normal salts. A normal salt is formed by a complete replacement of replaceable hydrogen atoms from an acid molecule by a metal. Examples of these salts are $NaCl$, KCl, K_2SO_4, $KClO_3$, $MgCl_2$, $CuSO_4$, $MgSO_4$, and $NaNO_3$.

Acidic salts

Acidic salts are those salts that produce acidic solutions when dissolved in water. They are formed when a strong acid reacts with a weak base. Their aqueous solutions turn blue litmus paper red. Acidic salts may contain one or more replaceable hydrogen atom(s) present in an acid molecule or positive radical. Examples of acidic salts include $FeCl_3$, $ZnCl_2$, $HgCl_2$, $Fe_2(SO_4)_3$, $NaHSO_4$, $NaHPO_4$, $(NH_4)_2SO_4$, and NH_4Cl.

Basic salts

Basic salts are a class of salts that produce alkaline solutions when dissolved in water. Their aqueous solutions turn red litmus paper blue. A basic salt is a product of a reaction between a strong base and a weak acid. Examples of basic salts include CH_3COONa, Na_2CO_3, $NaHCO_3$, and $Na_2C_2O_4$.

Double salts

A double salt is a mixture of two salts, which dissolves in water to give two different cations or anions. Examples are potash alum, $[KAl(SO_4)_2\cdot 12H_2O]$, ferric alum $[NH_4Fe(SO_4)_2\cdot 12H_2O]$, and Mohr's salt $[(NH_4)_2Fe(SO_4)_2\cdot 6H_2O]$.

Preparation of salts

There are several methods of preparing salts. The selection of the method to be used depends on whether the salt required is soluble or insoluble in water.

Preparation of soluble salts

Soluble salts are prepared by reacting acids with metals, alkalis, insoluble bases or carbonates.

Reactions of acids with metals (displacement method)

The displacement method is suitable for preparing salts of moderately reactive metals such as zinc, iron, lead, tin, and copper. Highly reactive metals such as K, Na, and Ca react very rapidly with acids and can cause explosions. The displacement reaction occurs as shown in the general equation:

$$\text{Metal} + \text{Acid} \longrightarrow \text{Salt} + \text{Hydrogen gas}$$

For example, zinc displaces hydrogen from sulfuric acid as shown in the following reaction equation:

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

Activity 5.7

Aim: To prepare zinc sulfate

Requirements: Beaker, evaporating dish, glass rod, spatula, tripod stand, wire gauze, filter papers, filter funnel, heat source, zinc granules, and 1 M dilute sulfuric acid

Procedure

1. Put about 25 cm^3 of 1 M dilute sulfuric acid in a beaker, and then add some zinc granules into it. Mix the contents by using a glass rod.
2. Continue adding more zinc to the beaker to excess until no more effervescence is observed.
3. Filter out the excess zinc and pour the filtrate into an evaporating dish.
4. Heat the filtrate while ensuring that not all the water is evaporated.
5. Allow the heated filtrate to cool and then filter off any excess liquid.
6. Allow the crystals to dry by rubbing them between filter papers. Record the observations.

Questions

1. Why excess zinc is used in this experiment?
2. Why is it not allowed to dry up completely the filtrate when evaporating?

3. What will happen if magnesium ribbon is used instead of zinc granules. Support your answer with a chemical equation.

Reactions of acids with metal oxides and metal hydroxide (neutralisation method)

The neutralisation method is usually used to prepare soluble salts such as sodium, potassium, and ammonium salts. Metal oxides and metal hydroxides react with acids to produce salt and water only.

$$\text{Metal oxide} + \text{Acid} \longrightarrow \text{Salt} + \text{Water}$$
$$\text{Metal hydroxide} + \text{Acid} \longrightarrow \text{Salt} + \text{Water}$$

Activity 5.8

Aim: To prepare potassium chloride

Requirements: Pipette, burette, conical flasks, dropper, evaporating dish, tripod stand, wire gauze, heat source, 2 M potassium hydroxide solution, 2 M hydrochloric acid solution, phenolphthalein indicator, and distilled water

Procedure

1. Transfer 25 cm^3 of 2 M potassium hydroxide solution into a conical flask.
2. Add two drops of phenolphthalein indicator.
3. Fill the burette with 2 M of hydrochloric acid. Record the initial reading.
4. Slowly run the acid from the burette into the conical flask while gently swirling the flask to mix the acid and the base. Close the burette when the solution just changes its colour. Record the volume of the acid used.
5. Transfer the mixture obtained into an evaporating dish and heat until all the water evaporates.

Questions

1. What was the colour change of the reaction?
2. What volume of acid was used in the entire reaction?
3. What happened in Step 4? Explain with the aid of a chemical equation.

Reactions of acids with insoluble bases

Reactions of acids with insoluble bases result in the formation of soluble salts. For example, insoluble bases of zinc oxide and copper(II) oxide react with acids to form soluble salts like zinc sulfate, copper(II) sulfate, copper(II) chloride, and water.

$$\text{Insoluble base} + \text{Acid} \longrightarrow \text{Salt} + \text{Water}$$
$$ZnO(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2O$$

Activity 5.9

Aim: To prepare copper(II) chloride

Requirements: 100-mL measuring cylinder, tripod stand, 250-mL beaker, glass rod, evaporating dish, wire gauze, filter papers, funnel, heat source, 2 M hydrochloric acid solution, and copper(II) oxide

Procedure

1. Pour about 50 cm^3 of 2 M hydrochloric acid into a beaker and heat gently.
2. Dissolve some copper(II) oxide little by little in the hot acid and stir for some time until the metal oxide is in excess.
3. Filter off the solid and heat the filtrate until it approaches to dryness.

Questions

1. What chemical reaction took place in this experiment? Explain with a balanced chemical equation.
2. Why was it necessary to use hot hydrochloric acid in the reaction with copper(II) oxide?
3. Why did you heat the fitrate?
4. What is the colour of the product formed in Step 3?

Reactions of acids with metal carbonates

Most metal carbonates such as calcium carbonate, copper(II) carbonate and potassium carbonate react with acids to produce salt, water and carbon dioxide gas as shown in the following equation:

$$\text{Carbonate} + \text{Acid} \longrightarrow \text{Salt} + \text{Water} + \text{Carbon dioxide}$$

Activity 5.10

Aim: To prepare copper(II) sulfate

Requirements: Beaker, evaporating dish, glass rod, wire gauze, measuring cylinder, filter paper, filter funnel, tripod stand, heat source, dilute sulfuric acid, and copper(II) carbonate

Procedure

1. Transfer about 25 cm^3 of dilute sulfuric acid into a small beaker and heat gently.
2. Add solid copper(II) carbonate to the beaker containing the hot acid and stir the mixture by using a glass rod until it is saturated.
3. Filter the mixture and crystallise the filtrate by heating gently until almost all of the water evaporates. Do not allow evaporation to dryness.

Questions

1. Which reaction occurred in this experiment? Explain with a balanced chemical equation.
2. Why was the mixture filtered before evaporation?
3. How can the gas formed in this experiment be tested?

Reactions of acids with metal hydrogencarbonates

Metal hydrogencarbonates, or commonly known as bicarbonates, are compounds in the form of $MHCO_3$, where M represents a metal. These compounds react with acids in the same way as metal carbonates to form a salt, water, and carbon dioxide gas. For example, sodium hydrogencarbonate ($NaHCO_3$) reacts with dilute hydrochloric acid to produce sodium chloride, water, and carbon dioxide gas.

$$HCl(aq) + NaHCO_3(aq) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

Preparation of salts by direct combination

If salts are required in anhydrous forms, they cannot be prepared by wet methods. In this case, it is possible to prepare them by direct combination. For example, in the preparation of iron(III) chloride, dry chlorine gas is passed over hot iron wire or steel wool as shown in Figure 5.14. The iron(III) chloride vapour produced is condensed to give a solid salt. This is indicated in the following equation:

$$2Fe(s) + 3Cl_2(g) \xrightarrow{heat} 2FeCl_3(s)$$

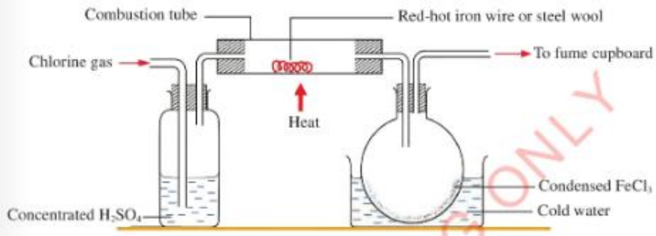


Figure 5.14: *Set-up of the direct preparation of iron(III) chloride*

Preparation of insoluble salts

Insoluble salts are prepared by using ionic precipitation or a double decomposition method. In this method, two soluble salts are used to form an insoluble salt and a soluble salt. The insoluble salt precipitates, while the soluble salt remains in solution. The precipitate is filtered and washed with distilled water and then dried. For example, to prepare insoluble lead(II) iodide, the solution containing lead ions and another containing iodide ions are mixed. Such solutions could be lead(II) nitrate and potassium iodide. A bright yellow precipitate of lead(II) iodide is formed after mixing. The chemical equation for this reaction is as follows:

$$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow PbI_2(s) + 2KNO_3(aq)$$

Other common insoluble salts which can be prepared by using this method include calcium sulfate ($CaSO_4$), magnesium carbonate ($MgCO_3$), silver chloride ($AgCl$), barium carbonate ($BaCO_3$), barium sulfate ($BaSO_4$), and lead(II) sulfate ($PbSO_4$).

Task 5.6

Use online search tools or other reliable resources to investigate and analyse various methods used in salt production in Tanzania.

Properties of salts

Salts exhibit different physical and chemical properties.

Physical properties of salts

The physical properties of salts include their appearance in colour, texture, and solubility.

Physical appearance of salts

Salts appear in crystalline or powder forms and exhibit different colours depending on their types. For example, hydrated copper(II) sulfate is a blue crystalline salt, while sodium chloride is a white crystalline salt. Ferrous chloride is a green crystalline salt, while calcium nitrate is a white crystalline salt. Calcium carbonate is a white powder. Figure 5.15 shows some examples of crystalline and powdered salts.

Figure 5.15: *Crystalline and powdered salts*

Solubility of salts

Solubility of a solute in a given solvent is the number of the grams of the solute required to saturate 100 grams of solvent at a given temperature, or it is the amount of grams or moles of a substance that dissolve in a solvent to form a *saturated solution* at a given temperature. The solubility of a solute in a solvent is not fixed but varies with temperature. Some solutes are more soluble in certain solvents than others. For example, 40 g of sodium chloride can dissolve in 100 g of water at 50 °C, while 80 g of potassium nitrate can dissolve in 100 g of water at the same temperature. In this case, potassium nitrate is more soluble than sodium chloride. Some salts such as sodium chloride and potassium nitrate are soluble in water, while others such as calcium carbonate and lead chloride are insoluble.

The solubility behaviour of some salts is summarised in Table 5.9.

Table 5.9: *Soluble and insoluble salts*

Salts	Soluble salts	Insoluble salts
Nitrates	All nitrates	None
Halides (fluorides, chlorides, bromides, and iodides)	All halides except those of Pb, Hg, and Ag	Silver, lead, and mercury halides such as AgCl and $PbCl_2$
Sulfates	All sulfates except $CaSO_4$, $BaSO_4$, and $PbSO_4$	$CaSO_4$, $BaSO_4$, and $PbSO_4$
Carbonates	All Group I metal carbonates and ammonium carbonate	All carbonates except Group I metal carbonates and ammonium carbonate

Activity 5.11

Aim: To investigate the solubility of salts

Requirements: Ten test tubes, test tube racks, 10-mL measuring cylinder, sodium carbonate, calcium nitrate, calcium carbonate, sodium sulfate, potassium chloride, barium sulfate, barium chloride, zinc chloride, lead chloride, ammonium carbonate, distilled water, and spatulas

Procedure

1. Put about 5 cm^3 of distilled water into each test tube.
2. Put 0.2 g of each salt in each test tube containing an equal amount of water, then shake gently. Record the observations in each test tube.

Questions

1. Which salts dissolved readily?
2. Which salts dissolved slowly?
3. Which salts did not dissolve?

Chemical properties of salts

Different salts exhibit different chemical properties. For example, they decompose upon heating to produce different substances depending on the type of their metals and positions in the reactivity series.

Action of heat on salts

Different salts behave differently when heated. Some salts will decompose after slight heating, yet others have to be heated strongly to decompose. Some salts do not decompose at all, even with strong heating.

Action of heat on carbonates

Upon heating, metal carbonates decompose to form metal oxides and carbon dioxide gas. The decomposition reaction can be represented by the following equation, where M represents a particular metal:

$$MCO_3(s) \xrightarrow{heat} MO(s) + CO_2(g)$$

Sodium and potassium carbonates do not decompose even when strongly heated. The carbonates of magnesium and calcium decompose to give magnesium oxide and calcium oxide, respectively, with the evolution of carbon dioxide gas.

$$MgCO_3(s) \xrightarrow{heat} MgO(s) + CO_2(g)$$
$$CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$$

Other metal carbonates decompose easily on heating. For example, green copper(II) carbonate decomposes easily upon heating to give black copper(II) oxide and carbon dioxide gas.

$$CuCO_3(s) \xrightarrow{heat} CuO(s) + CO_2(g)$$

Zinc carbonate decomposes easily on heating to form zinc oxide and carbon dioxide gas. The zinc oxide produced is yellow when hot and white when cold.

$$ZnCO_3(s) \xrightarrow{heat} ZnO(s) + CO_2(g)$$

Lead(II) carbonate decomposes when heated to give lead(II) oxide and carbon dioxide gas.

$$PbCO_3(s) \xrightarrow{heat} PbO(s) + CO_2(g)$$

Ammonium carbonate behaves so differently when heated. It decomposes slowly at room temperature to produce ammonia gas, water, and carbon dioxide.

$$(NH_4)_2CO_3(s) \longrightarrow 2NH_3(g) + H_2O(g) + CO_2(g)$$

A strong pungent smell is produced due to the presence of ammonia gas. The evolved gas turns limewater milky, indicating the presence of carbon dioxide. It also turns a blue cobalt(II) chloride paper pink indicating the presence of water

vapour. A red litmus paper turns blue due to the basic condition of ammonia gas in the solution. Table 5.10 summarises the behaviour of some carbonates on heating.

Table 5.10: *Effects of heat on some carbonate salts*

Carbonate salt	Behaviour on heating
Potassium carbonate and sodium carbonate	Do not decompose
Calcium carbonate, magnesium carbonate, zinc carbonate, iron(II) carbonate, lead(II) carbonate, and copper(II) carbonate	Decompose into corresponding metal oxides and carbon dioxide gas
Ammonium carbonate	Decomposes into ammonia gas, carbon dioxide gas, and water

Activity 5.12

Aim: To investigate the effects of heat on carbonate salts

Requirements: Sodium carbonate, magnesium carbonate, calcium carbonate, copper(II) carbonate, zinc carbonate, lead(II) carbonate, limewater, test tubes, spatulas, heat source, blue litmus paper, wire gauze, tripod stand, test tube holder and bent delivery tube

Procedure

1. Put a spatulaful of zinc carbonate into a dry test tube.
2. Hold the test tube with a test tube holder and heat over a non-luminous flame.
3. Pass the gas evolved into a test tube of limewater. Record the results.
4. Put a moist blue litmus paper at the mouth of the test tube, and record the observations.
5. Observe the colour change of the compound when hot and when cold.
6. Repeat Steps 1–5 for the other carbonates. Record all the observations.

Questions

1. What differences were observed when heating zinc carbonate and sodium carbonate?
2. What is common when heating magnesium carbonate and copper(II) carbonate?

Activity 5.13

Aim: To investigate the effect of heat on ammonium carbonate

Requirements: Spatula, 2 test tubes, test tube holder, heat source, cobalt(II) chloride paper, red litmus paper, ammonium carbonate, and limewater

Procedure

1. Put a spatulaful of ammonium carbonate salt into a dry test tube.
2. Hold the test tube with a test tube holder and heat the salt over a non-luminous flame.
3. Pass the gas evolved into a test tube containing limewater. Record the observations.
4. Put a moist red litmus paper and cobalt(II) chloride paper at the mouth of the test tube, and record the observations.

Questions

1. What differences did you observe between this activity and Activity 5.12?
2. What is the balanced chemical equation for the reaction in Step 3?

Action of heat on nitrates

Most nitrates decompose on heating to give the metal oxide, brown fumes of nitrogen dioxide, and oxygen gas. Nitrates of alkali metals such as sodium and potassium decompose slowly on heating to give oxygen and metal nitrite.

$$2NaNO_3(s) \xrightarrow{heat} 2NaNO_2(s) + O_2(g)$$

Nitrates of calcium and copper decompose into their respective oxides, nitrogen dioxide gas, and oxygen gas.

$$2Ca(NO_3)_2(s) \xrightarrow{heat} 2CaO(s) + 4NO_2(g) + O_2(g)$$
$$2Cu(NO_3)_2(s) \xrightarrow{heat} 2CuO(s) + 4NO_2(g) + O_2(g)$$

Exceptional behaviours are shown by ammonium nitrate and silver nitrate. Ammonium nitrate decomposes to give dinitrogen oxide gas and water, whereas silver nitrate decomposes to give silver metal, nitrogen dioxide, and oxygen gas.

$$NH_4NO_3(s) \xrightarrow{heat} N_2O(g) + 2H_2O(l)$$
$$2AgNO_3(s) \xrightarrow{heat} 2Ag(s) + 2NO_2(g) + O_2(g)$$

Table 5.11 summarises the effects of heat on nitrates.

Table 5.11: *Effects of heat on some nitrate salts*

Nitrate salts	Products
Nitrates of potassium and sodium	Metal nitrite + Oxygen
Nitrates of calcium, magnesium, aluminium, zinc, iron, lead, and copper	Metal oxide + Nitrogen dioxide + Oxygen
Nitrates of silver and mercury	Metal + Nitrogen dioxide + Oxygen
Ammonium nitrate	Dinitrogen oxide + Water

Action of heat on sulfates

Sulfates are generally more stable to heat than nitrates. Even sulfates of metals that are low in the reactivity series must be strongly heated to decompose. Sulfates of alkali metals and alkaline earth metals do not decompose when heated. However, few sulfates such as ammonium sulfate and copper(II) sulfate decompose when heated. For example, ammonium sulfate decomposes on heating to give sulfuric acid and ammonia gas. During heating, the sulfate initially decomposes to give ammonium hydrogen sulfate as shown in the following equations:

$$(NH_4)_2SO_4(s) \xrightarrow{heat} NH_4HSO_4(l) + NH_3(g)$$

On cooling the products, ammonia gas, and sulfuric acid are formed.

$$NH_4HSO_4(l) \xrightarrow{cool} NH_3(g) + H_2SO_4(l)$$

Table 5.12 summarises the effect of heat on some sulfate salts.

Table 5.12: *Effect of heat on some sulfate salts*

Sulfate salt	Products
Sulfate of potassium, sodium, magnesium and calcium	Do not decompose
Sulfate of zinc, copper(II) and iron (III)	Metal oxide + sulfur trioxide
Ammonium sulfate	Ammonia gas + sulfuric acid
Iron(II) sulfate	Iron(II) oxide + sulfur dioxide + sulfur trioxide

Activity 5.14

Aim: To investigate the action of heat on sodium sulfate and iron(II) sulfate

Requirements: Ignition tube, spatula, heat source, sodium sulfate, iron(II) sulfate, and red and blue litmus papers

Procedure

1. Put a spatulaful of sodium sulfate in the ignition tube and heat strongly.
2. Insert the wet blue and red litmus papers in the ignition tube while heating. Record the observations.
3. Repeat Steps 1 and 2 by using iron(II) sulfate in the place of sodium sulfate. Record the observations.

Questions

1. What were the colour changes in the two litmus papers?
2. What do the colour changes in the litmus papers imply?

Action of heat on chloride salts

All chloride salts except ammonium chloride do not decompose on heating. Ammonium chloride on heating, sublimes and decomposes to ammonia gas and hydrogen chloride gas.

$$NH_4Cl(s) \xrightarrow{heat} HCl(g) + NH_3(g)$$

Deliquescence, hygroscopy, and efflorescence in salts

Based on their behaviour when exposed to air, salts are categorised as deliquescent, hygroscopic or efflorescent.

Deliquescent salts

A salt that absorbs moisture from the air (environment) to form a solution is called a *deliquescent salt.* Examples of deliquescent salts are sodium hydroxide, phosphorus oxides, potassium hydroxide, and calcium chloride.

Hygroscopic salts

A *hygroscopic* salt is the one which absorbs water from air without necessarily forming a solution. Examples of hygroscopic salts are sodium chloride, sodium nitrate, calcium chloride, calcium oxide, and copper sulfate. The tendency of

calcium chloride to absorb water vapour explains its use as a drying agent for many gases, except ammonia gas because it reacts with it. The water absorbed by the salt may result in changes in the physical properties, such as boiling point and viscosity.

Efflorescent salts

An efflorescent salt is the one which when left in air loses all the water of crystallisation. Examples of efflorescent salts are hydrated sodium carbonate ($Na_2CO_3{\cdot}10H_2O$), hydrated sodium sulfate ($Na_2SO_4{\cdot}10H_2O$), and hydrated iron(II) sulfate ($FeSO_4{\cdot}7H_2O$).

Uses of salts

Salts have many uses in our daily lives. Some common uses include controlling soil pH, acting as antacids, serving as fertilisers, and helping to alleviate certain health disorders. They are also used for flavouring and preserving food, as well as in the production of pesticides.

Control of soil pH: Soils that are excessively acidic or alkaline are unsuitable for crop production. Calcium oxide (lime) is usually added to acidic soils to neutralise their acidity. When the soil is too alkaline, elemental sulfur or acidifying fertilisers like ammonium sulfate (SA) are added to remedy the condition.

Uses of salts as antacids: Antacids are substances that are in the forms of hydroxides or salts. These are used to relieve heartburn and acid build-up in the stomach. Some of the common antacids are magnesium sulfate (Epsom salt), aluminium hydroxide, and sodium hydrogencarbonate.

As inorganic fertilisers: Inorganic fertilisers are salts. Examples of these are ammonium sulfate, ammonium nitrate, and calcium phosphate.

Alleviation of health disorders: Various salts are used to alleviate problems that are associated with different health conditions. These salts include chlorides, phosphates, and sulfates of sodium, potassium, and calcium.

Production of pesticides: Certain salts are used as pesticides to kill or control pests such as insects, rodents, weeds and fungi. For example, copper(II) sulfate, iron(II) sulfate, and sodium chlorate(V) are used as pesticides for controlling fungi in crops.

Other uses of salts: There are many uses that are specific to some salts as summarised in Table 5.13.

Table 5.13: *Uses of some salts in daily life*

S/N	Salt	Uses
1.	Sodium chloride (NaCl)	(a) An essential ingredient in food (b) Preservation of food (c) Manufacturing of soaps (d) Making a freezing mixture which is used by ice-cream vendors
2.	Sodium carbonate (Na_2CO_3) (Washing soda)	(a) To remove stain from clothes (b) Production of detergents and soaps (c) Manufacture of glass, paper, and textiles (d) Softening hard water (e) In fire extinguishers
3.	Sodium hydrogencarbonate ($NaHCO_3$)	(a) Used as baking soda (b) In fire extinguishers (c) As an antacid in medicine
4.	Ammonium chloride (NH_4Cl)	(a) Used as an electrolyte in the manufacture of dry batteries (b) Used as an acidifier
5.	Copper(II) sulfate ($CuSO_4$)	(a) Used as a fungicide to kill certain germs (b) In electroplating (c) In dyeing (d) In correction of copper deficiency in soil and animals (e) Preparation of catalysts
6.	Silver bromide (AgBr)	Making photographic films
7.	Calcium chloride ($CaCl_2$)	(a) Used as a drying agent (b) In freezing mixtures (c) Used as food additive (d) In water treatment plants (e) As a food preservative
8.	Calcium carbonate ($CaCO_3$)	(a) Plays an important role as a construction and building material (marble) or as an ingredient in cement (b) In medicinal industries to manufacture antacids and tablets (c) Manufacture of paints, paper, and plastics

S/N	Salt	Uses
9.	Potash alum [$KAl(SO_4)_2{\cdot}12H_2O$]	(a) Purifying water (b) Used as an antiseptic (c) In dyeing
10.	Ammonium salts	(a) Manufacture of plastics, synthetic fibres, dyes, explosives, and pharmaceuticals (b) Used as fertilisers

Activity 5.15

Aim: To investigate the preservative effects of table salt

Requirements: Table salt (sodium chloride-NaCl), pieces of fresh meat, beaker, petri dishes

Procedure

1. Take two pieces of fresh meat.
2. Sprinkle a generous amount of salt on one piece and leave the other without salt.
3. Leave them at room temperature for 24 hours.
4. Observe and record any appearance, texture, and odour.

Questions

1. What happened to the salted and unsalted pieces of meat?
2. What can you conclude on the effects of the salt in the two pieces of meat?

Exercise 5.4

1. Classify the following substances as neutral, acidic, or basic salts: NaCl, $NaHSO_4$ and Na_2CO_3. Explain your classification.
2. Describe how acidic and basic salts are formed using an example for each.
3. Differentiate between a double salt and a neutral salt, giving one example in each.
4. A salt is prepared by mixing ammonia solution and hydrochloric acids. Will the resulting salt be acidic, basic, or neutral? Explain your answer.

5. Explain the advantages of the following types of salts in real-life applications by providing two examples for each:
   (a) Deliquescent salts
   (b) Efflorescent salts
   (c) Hygroscopic salts

Project: Determination of pH values of natural and synthetic substances

Collect samples of various substances such as tap water, bottled water, river water, fruits juices, and soaps and detergents. Determine the pH values using natural and synthetic acid-base indicators, litmus papers, or pH meter. Write a detailed report outlining the methodology, materials used, observations, results, discussions, and conclusions regarding the pH properties of the substances. Then, prepare a brief presentation summarising the experiments, results and key learnings. Enhance the presentation with visuals such as photos of the tests or a pH chart derived from the collected data.

Chapter summary

1. An acid is a chemical substance which produces hydrogen ions (H^+) in water as the positively charged ions.
2. A base refers to a chemical substance which, when dissolved in water, produces hydroxide ions (OH^-) as negatively charged ions. It is also defined as a metal oxide or hydroxide which neutralises an acid. An alkali is a soluble base.
3. A strong acid or strong base dissociates completely in water. A weak acid or weak base dissociates partially in water.
4. A pH indicator is a chemical substance that exhibits different colours in solutions of different acidities or alkalinities.
5. The basicity of an acid is the number of hydrogen atoms per molecule of the acid that can be displaced by a metal in a solution. An acid that has only one hydrogen atom that can be displaced is said to be monobasic. An acid with a basicity of two is dibasic, while the one with basicity of three is tribasic.

6. Neutralisation is a chemical reaction in which an acid reacts with a base to form salt and water.
7. Neutralisation is applied in various useful situations such as treating insect stings, relieving indigestion, neutralising harmful substances in the environment and manufacturing fertilisers.
8. Acids and alkalis are used in the manufacture of various industrial products.
9. Salt is an ionic compound which is made up of the positively charged ion (cation) and negatively charged ion (anion).
10. The solubility of a solute in a given solvent is the number of grams of the solute that will saturate 100 grams of a solvent at a given temperature.
11. All sodium, potassium, and ammonium salts are soluble in water.
12. All nitrates are soluble in water.
13. All chlorides are soluble in water except those of silver, mercury, and lead.
14. All sulfates are soluble in water except those of barium, calcium and lead.
15. All carbonates are insoluble in water except those of sodium, potassium, and ammonium carbonates.
16. Salts have various uses. These include: controlling soil pH, as antacids, as inorganic fertilisers, alleviating health disorders, and in the preservation and seasoning of food.

Revision exercise 5

Choose the correct answer for Questions 1–6. For other questions, provide the answers as per the demands indicated.

1. Which of the following is a natural acid?
   (a) Nitric acid
   (b) Phosphoric acid
   (c) Citric acid
   (d) Sulfuric acid
2. An acid is any substance that;
   (a) contains hydrogen in its formula.
   (b) dissolves in water to produce hydrogen ions.
   (c) can react with NaOH.
   (d) contains oxygen in its formula.

3. What name is given to the process of dissolving sodium chloride in water?
   (a) Precipitation process
   (b) Neutralisation reaction
   (c) Hydration process
   (d) Dissolution process
4. Which of the following salts is soluble in water?
   (a) Calcium carbonate
   (b) Copper carbonate
   (c) Magnesium carbonate
   (d) Sodium carbonate
5. Which of the following substances is produced when calcium carbonate is strongly heated?
   (a) Metal oxide and gas
   (b) Metal oxide and smoke
   (c) High heat and gas
   (d) A bicarbonate and gas
6. Which of the following salts decomposes to its own metal?
   (a) Sodium chloride
   (b) Potassium carbonate
   (c) Zinc sulfate
   (d) Silver nitrate
7. Match the premises on properties and uses of acids and bases in **Column A** with their correct responses in **Column B**.

Column A	Column B
(i) The reaction of hydrogen ions with hydroxide ions	A. Weak acid B. Alkali
(ii) Ion responsible for acidic properties of substances	C. Ant-pain D. Hydrogen ion
(iii) Alleviates excess stomach acidity	E. Antacid
(iv) Instrument commonly used to measure the acidity or alkalinity of a solution	F. Titration G. Acid
(v) Produces hydroxide ions when dissolved in water	H. Neutralisation I. pH meter

8. Explain the following terms and illustrate their significance by providing an example for each:
   (a) pH indicators
   (b) Double salt
   (c) Titration
   (d) Solubility
9. Sort out the following salts into groups of soluble and insoluble salts: Sodium carbonate, lead nitrate, lead sulfate, calcium carbonate, copper(II) sulfate, sodium nitrate, zinc chloride, silver chloride, barium carbonate, sodium sulfate, sodium chloride, potassium chloride, ammonium carbonate, and barium sulfate.
10. With reasons, suggest the best methods for preparing the following compounds:
    (a) Silver chloride
    (b) Lead(II) sulfate
    (c) Copper(II) nitrate
    (d) Calcium carbonate
    (e) Anhydrous zinc chloride
11. What are the thermal behaviours of metals carbonates, nitrates, chlorides, and sulfates when heated? Give an example in each case.
12. How are salts useful in our life in the following aspects?
    (a) Food preservation and preparation
    (b) Agricultural practices
    (c) Medical applications
13. Describe acids and bases, citing five examples in each.
14. Write the balanced chemical equations and net ionic equations for the reactions between sodium hydroxide and;
    (a) sulfuric acid
    (b) nitric acid
15. A chemistry student tested five solutions A, B, C, D, and E with a universal indicator solution to determine their pH values. The results obtained are shown in the following table:

Beaker of solution	A	B	C	D	E
pH	1	5	7	9	14

Which solution had a;
(i) strong alkali?
(ii) strong acid?
(iii) weak acid?
(iv) neutral solution?
(v) weak base?

16. During a neutralisation reaction, a student placed 25 cm^3 of sodium hydroxide in a flask and added few drops of phenolphthalein. The base required 22 cm^3 of dilute hydrochloric acid for complete neutralisation.
 (a) What apparatus should have been used to measure the exact volume of sodium hydroxide solution?
 (b) What reagent was more concentrated than the other? Explain.
 (c) Name the salt that was formed.
 (d) How would you obtain pure crystals of the salt resulting from the solution?
 (e) Write a balanced net ionic equation for the reaction.

Appendices

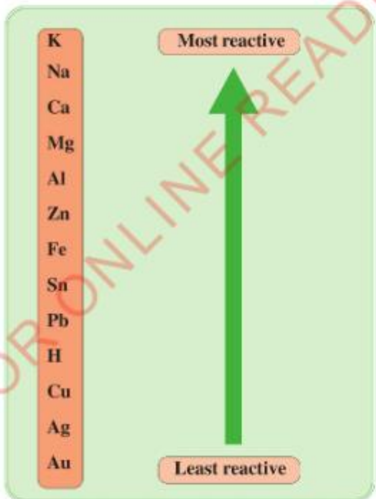
Appendix 1: *The periodic table*

Appendix 2: *List of chemical substances with IUPAC names*

S/N	Common name	Chemical formula	IUPAC name
1	Aspirin	$C_9H_8O_4$	Acetylsalicylic acid
2	Baking soda	$NaHCO_3$	Sodium hydrogen carbonate
3	Bleach	$NaOCl$	Sodium hypochlorite
4	Carbon dioxide	CO_2	Carbon dioxide
5	Carbon monoxide	CO	Carbon monoxide
6	Caustic potash	KOH	Potassium hydroxide
7	Caustic soda	$NaOH$	Sodium hydroxide
8	Chalk	$CaCO_3$	Calcium carbonate
9	Chloroform	$CHCl_3$	Trichloromethane
10	Common salt	$NaCl$	Sodium chloride
11	Copper sulfate (Cupric sulfate)	$CuSO_4$	Copper(II) sulfate
12	Dinitrogen oxide (Nitrous oxide)	N_2O	Dinitrogen monoxide
13	Dolomite	$CaMg(CO_3)_2$	Calcium magnesium carbonate
14	Fluospar/fluorite	CaF_2	Calcium fluoride
15	Gypsum	$CaSO_4{\cdot}2H_2O$	Calcium sulfate dihydrate
16	Limestone	$CaCO_3$	Calcium carbonate
17	Limewater	$Ca(OH)_2$	Calcium hydroxide solution
18	Marble	$CaCO_3$	Calcium carbonate
19	Milk of magnesia	$Mg(OH)_2$	Magnesium hydroxide
20	Plaster of Paris (POP)	$CaSO_4{\cdot}\frac{1}{2}H_2O$	Calcium sulfate hemihydrate
21	Quick lime	CaO	Calcium oxide
22	Saltpetre	KNO_3	Potassium nitrate
23	Slaked lime	$Ca(OH)_2$	Calcium hydroxide
24	Soda ash	Na_2CO_3	Sodium carbonate
25	Sodium nitrate	$NaNO_3$	Sodium nitrate

26	Sodium sulfate	Na_2SO_4	Sodium sulfate
27	Sulfuric acid	H_2SO_4	Sulfuric acid or tetraoxosulfate
28	Sulfate of ammonia (SA)	$(NH_4)_2SO_4$	Ammonium sulfate
29	Sulfur dioxide	SO_2	Sulfur(IV) oxide
30	Sulfur trioxide	SO_3	Sulfur(VI) oxide
31	Urea	$CO(NH_2)_2$	Carbamide or Diaminomethanal
32	Water	H_2O	Dihydrogen monoxide

Appendix 3: *Reactivity series of metals*



Glossary

Acid	a substance which produces hydrogen ions (H^+) in water as the positively charged ions
Anhydrous	contains no water
Base	a substance that neutralises an acid by reacting with hydrogen ions
Basic oxides	oxides which react with acids to give salt and water as the only products
Basicity of an acid	the number of ionisable hydrogen ions per molecule of an acid
Catalyst	a substance that alters the rate of a chemical reaction but it remains unchanged at the end of the reaction
Cations	are positively charged particles
Corrosion	a natural process that weakens metal structures, causing damage and reducing their lifespan
Decomposition	breaking down of a chemical compound into elements or simpler compounds
End point	the point in titration process whereby reactants have reacted completely and is indicated by a physical change such as colour change of an indicator used or precipitates formation
Galvanisation	a process of applying a protective zinc coating to iron or steel to prevent rusting and corrosion
Hydrocarbons	organic compounds made up of only carbon (C) and hydrogen (H) atoms.
Irreversible reaction	a reaction that proceeds in only one direction
Mineral acids	also called inorganic acid is an acid derived from one or more inorganic compounds
Neutralisation	a process where an oxide or hydroxide ions of a base combine with an acid in the correct ratio to form salt and water
Organic acid	an acidic organic compound that contains primarily carbon, hydrogen, and oxygen.
Oxidation	addition of oxygen to a substance or removal of hydrogen from a substance or loss of electrons from a substance

pH	the negative logarithm of hydrogen ion concentration
Redox reaction	a reaction in which both oxidation and reduction occur simultaneously
Reduction	addition of hydrogen or removal of oxygen from a substance or gain of electrons by a substance
Relative atomic mass	the atomic masses of all the atoms of a particular element compared to the atomic mass of carbon-12 isotope
Relative molecular mass	is the average mass of a molecule relative to the mass of a carbon-12 atom
Respiration	a process in living organisms involving the production of energy, typically with the intake of oxygen and the release of carbon dioxide from the oxidation of complex organic substances
Reversible reaction	a reaction that proceeds in both the forward and backward directions
Sacrificial anode	a highly reactive metal used to protect another metal from corrosion by preventing metal deterioration by making the protected metal act as a cathode in an electrochemical cell
Salt	an ionic compound which is made up of cation(s) and anion(s)
Saturated solution	a solution that holds the maximum dissolved solute at a given temperature and pressure.
Thermal decomposition	a breakdown reaction that occurs when a compound is exposed to heat
Thermite	a mixture of finely powdered aluminium and iron oxide that produces a very high temperature on combustion, used in welding and for incendiary bombs
Titrant	a standard solution usually added to the burette during titration
Titration	the process of determining the quantity of analyte by adding measured volumes of a titrant

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