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Textbook of Chemistry

Grade

9



National Book Foundation
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Chemistry

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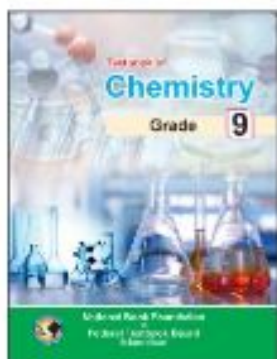
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OUR MOTTO

• Standards • Outcomes • Access • Style

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Textbook of
Chemistry Grade - 9



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First Edition	:	2012 Qty. 10,000
New Developed Edition	:	Mar. 2015 Qty. 25,000
6th Print	:	Mar. 2019 Qty. 60,000
7th Print	:	July 2019 Qty. 11,000
8th Print	:	Sep. 2019 Qty. 8000
9th Print	:	Feb. 2020 Qty. 81,000
Price	:	Rs. 180/-
Code	:	STE-498
ISBN	:	978-969-37-0820-2
Printer	:	Malik House Printers, Lahore

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PREFACE

CHEMISTRY GRADE - 9 is developed according to the National Curriculum 2006 and National Style Guide. It is being published since 2012 and in 2015 it was presented under the new management and supervision of textbook development principles and guidelines with new design and layout.

Chemistry is one of the most exciting and useful subjects. It may not be an easy subject for you. If you develop a sound approach to thinking through an idea and solving problem, you will find it easy. This textbook provides several ways to develop this approach.

This textbook has been made friendly by giving problem solving strategy. This strategy is to facilitate reasoning, not memorizing toward a solution. With each skill or concept, self-assessment exercises have been given. These exercises would encourage students to think, develop skills and use information for a variety of purpose.

It is caused of development of skills such as questioning, observing, inferring, and predicting, measuring, hypothesizing, classifying, designing experiments, collecting data, analyzing data and interpreting data. At the end of each chapter summary of the key points is given. This provides quick reviews of the salient features of each chapter.

Science, Technology and Society connections have also been given in each chapter. These will help to understand the impact of science and technology on society. Students will learn how scientists decide what constitute scientific knowledge, how science is related to other ways of knowledge and how people have contributed to and influenced developments in science.

This textbook has been developed according to Student-Centered Inquiry Based (SCIB) criteria. It will encourage students to develop skills and think scientifically rather than simply memorize and study scientific facts. It is also expected from teachers that they will engage students in scientific inquiry activities to develop such skills.

Our efforts are to make textbooks teachable with quality, i.e. maintaining of standards. It is a continuous effort and we will get feedback of the yearly feasibility reports and redesign the textbook every year. The test items given in the exercises are for learning reinforcement. The examination questions should be prepared according to the SLO's and Bloom's Taxonomy.

Quality of Standards, Pedagogical Outcomes, Taxonomy Access and Actualization of Style is our motto. With these elaborations, this series of new development was presented for use. After educational feedback, research, NCC Standards 2016, the book is being published again.

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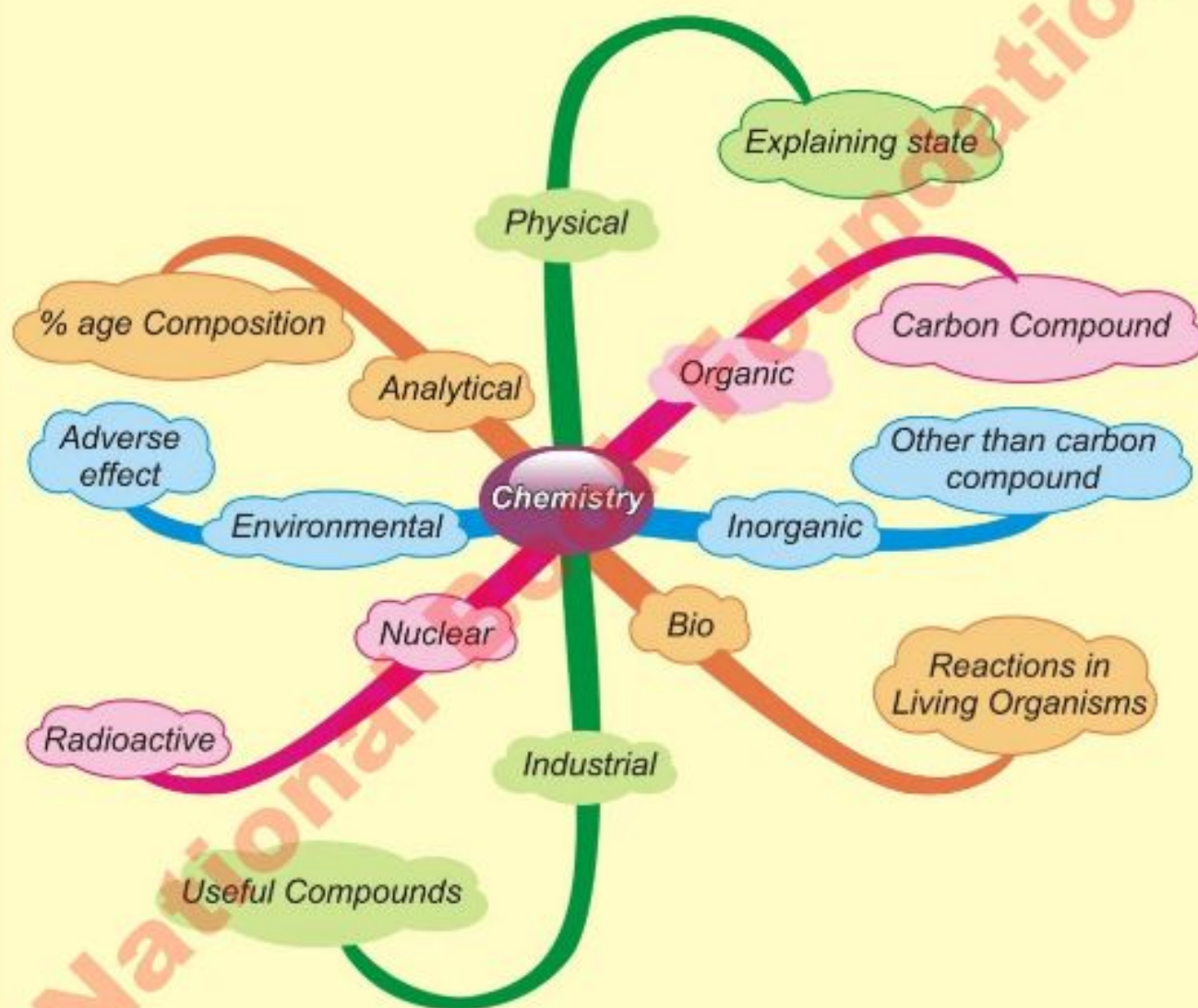
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اللہ کے نام سے شروع جو بڑا مہربان، نہایت رحم والا ہے

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MIND MAP





1

FUNDAMENTALS OF CHEMISTRY



This is a 15 days lesson

After completing this lesson, you will be able to:

- Change atomic mass, molecular mass and formula mass into gram atomic mass, gram molecular mass and gram formula mass.
- Differentiate between branches of chemistry.
- Differentiate between empirical and molecular formula.
- Differentiate among elements, compounds and mixtures.
- Differentiate between molecules and molecular ions
- Define ions, molecular ions, formula units and free radicals.
- Define atomic number, atomic mass, atomic mass unit.
- Define relative atomic mass based on C-12 scale.
- Distinguish between atoms and ions.
- Distinguish between ion and free radical. Classify the chemical species from given examples.
- Distinguish between matter and a substance.
- Describe how Avogadro's number is related to a mole of any substance. Distinguish among the terms gram atomic mass, gram molecular mass and gram formula mass.
- Identify the representative particles of elements and compounds.
- Identify and provide example of different branches of chemistry.
- Relate gram atomic mass, gram molecular mass and gram formula mass to mole.



Pre- Reading

What are the simplest components of wood, rocks and living organisms? This is an age-old question. Ancient Greek Philosophers believed that everything was made of an elemental substance. Some believed that substance to be water, other thought it was air. Some other believed that there were four elemental substances.

As 19th century began, John Dalton proposed an atomic theory. This theory led to rapid progress in chemistry. By the end of the century however, further observations exposed the need for a different atomic theory. 20th century led to a picture of an atom with a complex internal structure.

A major goal of this chapter is to acquaint you with the fundamental concepts about matter. In this chapter you will

learn some basic definitions to understand matter. This knowledge will help you in grade XI.

**Reading**

1.1 BRANCHES OF CHEMISTRY

Chemistry is defined as the science that examines the materials of the universe and changes that these materials undergo.

The study of chemistry is commonly divided into eight major branches:

1. Physical Chemistry

The branch of Chemistry that deals with laws and theories to understand the structure and changes of matter is called Physical Chemistry.

2. Organic Chemistry:

The branch of Chemistry that deals with substances containing carbon (Except carbonates, bicarbonates, oxides and carbides).

3. Inorganic Chemistry:

The branch of Chemistry that deals with elements and their compounds except organic compounds is called Inorganic Chemistry.

4. Biochemistry:

The branch of Chemistry that deals with physical and chemical changes that occur in living organisms is called Biochemistry.

5. Industrial Chemistry:

The branch of Chemistry that deals with the methods and use of technology in the large-scale production of useful substances is called industrial chemistry.

Do You Know?

Do you know the debate going on for centuries about the corpuscular nature of matter? An ancient Greek philosopher, Empedocles thought that all materials are made up of four things called elements:

1. Earth
2. Air
3. Water
4. Fire

Plato adopted Empedocles theory and coined the term element to describe these four substances. His successor, Aristotle also adopted the concept of four elements. He introduced the idea that elements can be differentiated on the basis of properties such as hot versus cold and wet versus dry. For example, heating clay in an oven could be thought of as driving off water and adding fire, transforming clay into a pot. Similarly water (cold & wet) falls from the sky as rain, when air (hot and wet) cools down. The Greek concept of four elements existed for more than two thousand years.

Society, Technology and Science

Archimedes was a Greek philosopher and mathematician and inventor of many war machines. Greek emperor gave him the task to determine whether his crown was made of pure gold or impure gold. Archimedes took the task and started thinking on it. He knew that the volume of an object determines the volume of the liquid it displaces, when submerged in the liquid. One day when he was taking bath, he observed that more water overflowed the bath tank as he sank deeper into the water. He also noticed that he felt weightless as he submerged deeper in the bath tank. From these observations he concluded that the loss in weight is equal to the weight of water overflowed. Thinking this he at once designed an experiment in his mind to check the purity of crown. He thought, he should weigh the crown and equal weight of the pure gold. Both should be dipped in water in separate containers, since every substance has different mass to volume ratio. If the crown was made of pure gold, it would displace same weight of water as an equal weight of pure gold. If the crown is impure, it would displace different mass of water than the pure gold. Thinking this, he was so excited that he ran from the bath shouting "Eureka" which means I have found it. Like Archimedes discovery, science developed through observations and experiments rather than by speculation alone.

**Teacher's Point**

Teacher may give examples of branches of Chemistry applied in daily life.



6. Nuclear Chemistry:

The branch of Chemistry that deals with the changes that occur in atomic nuclei is called nuclear chemistry.

7. Environmental Chemistry:

The branch of Chemistry that deals with the chemicals and toxic substances that pollute the environment and their adverse effects on human beings is called environmental chemistry.

8. Analytical Chemistry:

The branch of Chemistry that deals with the methods and instruments for determining the composition of matter is called Analytical Chemistry.

1.1.1 Differentiation between Branches of Chemistry

Vinegar contains 5% acetic acid. Acetic acid (CH_3COOH) is a colourless liquid that has characteristic vinegar like smell. It is used to flavour food. Various types of studies on this compound can help you to differentiate between various branches of chemistry.

1. Explanation of its transformation into gaseous state or solid state, applications of laws and theories to understand its structure is **physical chemistry**.
2. Since this is a carbon compound, its method of preparations and study of its physical and chemical characteristics is **organic chemistry**.
3. But the study of its component elements, carbon, hydrogen and oxygen is **inorganic chemistry**. This is because inorganic chemistry deals with elements and their compounds except carbon compounds. However, some carbon compounds such as CO_2 , CO, metal carbonates, hydrogen carbonates and carbides are studied in inorganic chemistry.
4. The study of chemical reactions that acetic acid undergoes in the bodies of human beings is **biochemistry**.

Society, Technology and Science

Theories are tentative. They may change if they do not adequately provide explanation of the observed facts. The work of scientists help to change existing theories of the time. In 1803, the British chemist John Dalton presented a scientific theory on the existence and nature of matter. This theory is called Dalton's atomic theory. Main postulates of his theory are as follows:

1. All elements are composed of tiny indivisible particles called atoms.
2. Atoms of a particular element are identical. They have same mass and same volume.
3. During chemical reaction atoms combine or separate or re-arrange. They combine in simple ratios.
4. Atoms can neither be created nor destroyed.

Dalton was able to explain quantitative results that scientists of his time had obtained in their experiments. He nicely explained the law of chemical combinations. His brilliant work became the main stimulus for the rapid progress of the chemistry during nineteenth century. However, series of experiment that were performed in 1850's and beginning of twentieth century clearly demonstrated that atom is divisible and consists of subatomic particles, electrons, protons and neutrons. Also the atoms of an element may differ in masses (such atoms are called isotopes). Thus some of the postulates of Dalton's atomic theory were found defective and were changed.



5. Use of technology and ways to obtain acetic acid on the large scale is **industrial chemistry**.
6. The study of the effect of radioactive radiations or neutron on this compound or its component elements is **nuclear chemistry**.
7. The study of any adverse effects of this compound or the compounds that are derived from it, on the humans is **environmental chemistry**.
8. The method and instruments used to determine its percentage composition, melting point, boiling point etc is **analytical chemistry**.

Example 1.1: Identifying examples of different branches of chemistry

Identify the branch of chemistry in each of the following examples:

1. Photosynthesis produces glucose and oxygen from carbon dioxide and water in presence of chlorophyll and sunlight.
2. Plantation helps in overcoming green house effect.
3. Haber's process converts large quantities of hydrogen and nitrogen into ammonia (NH_3).
4. Ammonia is a colourless gas with pungent irritating odour. It is highly soluble in water.
5. A chemist performed an experiment to check the percentage purity of a sample of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$).
6. An analyst determines that NO_2 is responsible for acid rain.
7. Chlorofluorocarbon compounds are responsible for the depletion of ozone layer.
8. α -particles (He^{++}) when bombard on nitrogen atom, a proton is emitted.

Problem Solving strategy:

Concentrate on the basic definition of each branch of chemistry and identify branch of chemistry in each example.

Solution:

1. Biochemistry, since photosynthesis is a chemical reaction that occurs in plants (living organism).
2. Environmental chemistry, since green house effect is an environmental problem.
3. Industrial chemistry, since large scale production of any substance is the subject of industrial chemistry.
4. Inorganic chemistry, since it deals with properties of inorganic compounds.
5. Analytical chemistry, since it deals with analysis of a compound, whether organic or inorganic in nature.
6. Environmental chemistry, since acid rain is an environmental problem.
7. Environmental chemistry, since depletion of ozone layer is environmental problem.
8. Nuclear chemistry, since nuclear change can emit protons.



Self-Assessment Exercise 1.1

Identify the branch of chemistry that is related to the following information:

1. Hair contain a special class of proteins called keratins, which are present in nails and wool. Proteins contain chains of C-atoms.
2. Acetylene is the simplest hydrocarbon that contains carbon-carbon triple bond. Hydrocarbons are the compounds of carbon and hydrogen.
3. White lead is a pigment used by artists for centuries, the metal Pb in the compound is extracted from its ore, galena (PbS).
4. Sulphuric acid (H_2SO_4) is extremely corrosive to skin.
5. Gases can be compressed by applying pressure.
6. Meat, Milk and eggs contain long chains of carbon compounds.
7. Element radium decays by emitting α -particles and is converted into another element radon.
8. Calorimeter is a device that measures the amount of heat, a substance absorbs on heating or emits on cooling.

Society, Technology and Science Molecularity of the physical world

World is composed of a few more than a hundred elements. Elements are building blocks of all the substances that make up all living and non-living things. This means elements are building blocks for everything in the universe. The same elements that make up earth also make up moon. A careful observation of the physical world reveals that matter usually occurs as mixtures. Most of the components of these mixtures are elements and compounds that exist as molecules. Only noble gases exist as monoatomic molecules, other substances exist as polyatomic molecules. Air consists of many elements and compounds all existing in molecular form. For instance O_2 , N_2 , CO_2 , H_2O and the noble gases. Water a molecular substance cover 70% of the earth's crust. It also fills the empty under the earth. Rocks and earth are mixtures of numerous compounds. Clay and sand consists of long chains of atoms called giant molecules. Petroleum and coal that are complex mixtures also contain hundreds of thousands of molecular compounds. Living things contain thousands of different substances such as carbohydrates, proteins, fats, lipids, DNA, RNA etc. All these substances are molecular in nature.



Reading

1.2 BASIC DEFINITIONS

Some of the important definitions used to understand matter are given below:

1.2.1 Elements, Compounds and Mixtures

Anything that occupies space and has mass is called matter. Any matter that has a particular set of characteristics that differ from the characteristics of another kind of matter is called a substance. For instance, oxygen, water, carbon dioxide, urea, glucose, common salt etc are different substances.



A substance that cannot be converted to other simpler substances is called an element. For example substances like oxygen, hydrogen, iron, copper, aluminium etc are elements. An element is now defined as a substance whose all the atoms have the same atomic number.

A compound is a pure substance that consists of two or more elements held together in fixed proportions by natural forces called chemical bonds. The properties of compounds are different from the properties of the elements from which they are formed. For example, water, carbon dioxide, copper sulphate, sodium chloride etc are compounds. Elements and compounds have uniform composition throughout.

An impure substance that contains two or more pure substances that retain their individual chemical characteristics is called a mixture. A mixture can be converted into two or more pure substances by a physical method. Examples of mixture are air, water containing dissolved oxygen, table salt dissolved in water, salt + sand etc. A mixture that consists of two or more visibly different components is called a heterogeneous mixture. For example sand + salt, oil floating on water etc. A mixture that consists of only one phase is called a homogeneous mixture. For example, sugar mixed in water, salt dissolved in water. Homogeneous mixtures also have uniform composition throughout.

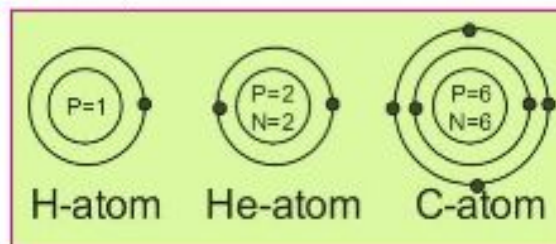
Science Tit Bits

Bad breath may be good for you. The chemistry of garlic is not simple. Garlic contains more than 200 compounds. People who eat a lot of garlic have a lower chance of getting stomach cancer, suffering from heart disease or having a stroke than people who eat little or no garlic.

In fact the entire physical world is made up of mixture of elements and compounds. Most of its components are made up of molecules.

1.2.2 Atomic Number, Mass Number

The number of protons in the nucleus of an atom is known as its atomic number. For example, there is only one proton in the nucleus of H-atom; therefore its atomic number is 1. All the atoms of a given element have the same number of protons and therefore the same atomic number. Do you think atomic number of He is 2? What is the mass number of C-atom? **The total number of protons and neutrons in an atom is known as its mass number.**



Some atoms of an element have different number of neutrons, such atoms are called isotopes. We will discuss isotopes in section 2.2.

$$\text{No. of neutrons} = \text{mass number} - \text{atomic number}$$

Example 1.2: Determining the number of protons and neutrons in an atom

Atomic number of an element is 17 and mass number is 35. How many protons and neutrons are in the nucleus of an atom of this element?

Problem Solving Strategy:

Number of protons are equal to atomic number and

Number of neutrons = mass number – atomic number

**Solution:**

Number of protons = atomic number = 17

Number of neutrons = mass number – atomic number
= 35-17 = 18

1.2.3 Relative Atomic Mass and Atomic Mass Unit

The first quantitative information about atomic masses came from the work of Dalton, Gay Lussac, Lavoisier, Avogadro and Berzelius. By observing the proportions in which elements combine to form various compounds, nineteenth century chemists calculated relative atomic masses. An atom is extremely small particle; therefore we cannot determine the mass of a single atom. However, it is possible to determine the mass of one atom of an element relative to another experimentally. This can be done by assigning a value to the mass of one atom of a given element, so that it can be used as standard. By international agreement in 1961, light isotope of carbon C-12 has been chosen as a standard. This isotope of carbon (C-12) has been assigned a mass of exactly 12 atomic mass unit. This value has been determined accurately using mass spectrometer. The mass of atoms of all other elements are compared to the mass C-12. Thus **“the mass of an atom of an element relative to the mass of an atom of C-12 is called its relative atomic mass”**.

One atomic mass unit (amu) is defined as a mass exactly equal to one-twelfth the mass of one C-12 atom.

Mass of one C-12 atom = 12 amu

$$1 \text{ amu} = \frac{\text{mass of one C-12 atom}}{12}$$

A hydrogen atom is 8.40% as massive as the standard C-12 atom. Therefore, relative atomic mass of hydrogen.

$$\begin{aligned} &= \frac{8.40}{100} \times 12 \text{ amu} \\ &= 1.008 \text{ amu} \end{aligned}$$

Similarly, relative atomic masses of O, Na, Al are 15.9994 amu, 22.9898 amu, 26.9815 amu respectively. Table 1.1 shows the relative atomic masses of some elements.

Table 1.1 relative atomic masses of some elements

Element	Relative atomic mass	Element	Relative atomic mass
H	1.008 amu	Al	26.9815 amu
N	14.0067amu	S	32.06 amu
O	15.9994amu	Cl	35.453 amu
Na	22.9898 amu	Fe	55.847 amu



1.2.4 Empirical Formula, Molecular Formula

Recall that the chemical formula of a compound tells us which elements are present in it and the whole number ratio of their atoms. In a chemical formula element's symbol and numerical subscripts show the type and the number of each atom in a compound. There are several types of chemical formulas for a compound. Here you will learn about two types of chemical formulas.

1. Empirical Formula

The empirical formula of a compound is the chemical formula that gives the simplest whole-number ratio of atoms of each element. For example in compound hydrogen peroxide there is one H atom for every O atom. Therefore, simplest ratio of hydrogen to oxygen is 1 : 1. So the empirical formula of hydrogen peroxide is written as **HO**.

The simplest ratio between C, H and O atoms in glucose is 1 : 2 : 1. What is the empirical formula of glucose?

2. Molecular Formula

A molecular formula gives the actual whole number ratio of atoms of each element present in a compound. For example there are actually two H atoms and two O atoms in each molecule of hydrogen peroxide. Therefore, actual ratio of hydrogen to oxygen atoms is 2 : 2. So molecular formula of hydrogen peroxide is **H₂O₂**.

The actual ratio between C, H and O atoms in a glucose molecule is 6 : 12 : 6. What is the molecular formula of glucose?

An empirical formula shows the simplest number of atoms of each element in a compound whereas the molecular formula shows the actual number of atoms of each element in the molecule of a compound.

Benzene is a compound of carbon and hydrogen. It contains one C atom for every H atom. There are actually six C atoms and six hydrogen atoms in each molecule of benzene. Identify empirical and molecular formula for benzene from the following formulas.



Molecular formulas for water and carbon dioxide are H₂O and CO₂ respectively. What are empirical formulas for these compounds?



Self-Assessment Exercise 1.2

Write the empirical formulas for the compound containing carbon to hydrogen in the following ratios:

- (a) 1:4 (b) 2:6 (c) 2:2 (d) 6:6

For many compounds, empirical and molecular formulas are same. For example water (H₂O), carbon dioxide (CO₂), ammonia (NH₃), methane (CH₄), sulphur dioxide (SO₂) etc. Can you show it why?



Teacher's Point

A teacher may give numerical example to determine E.F of a Compound.



Self-Assessment Exercise 1.3

1. Aspirin is used as a mild pain killer. There are nine carbon atoms, eight hydrogen atoms and four oxygen atoms, in this compound. Write its empirical and molecular formulas.
2. Vinegar is 5% acetic acid. It contains 2 carbon atoms, four hydrogen atoms and 2 oxygen atoms. Write its empirical and molecular formulas.
3. Caffeine ($\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$) is found in tea and coffee. Write the empirical formula for caffeine.



Reading

1.2.5 Molecular Mass and Formula Mass

Molecular mass is the sum of atomic masses of all the atoms present in the molecule. All you have to do is to add up the atomic masses of all the atoms in the compound. For example,

$$\begin{aligned}
 \text{Molecular mass of water } \text{H}_2\text{O} &= 2(\text{atomic mass of H}) + \text{atomic mass of oxygen} \\
 &= 2(1.008) + 16.00 \\
 &= 2.016 + 16.00 \\
 &= 18.016 \text{ amu}
 \end{aligned}$$

Example 1.3: Determining molecular mass

1. Determine the molecular mass of glucose $\text{C}_6\text{H}_{12}\text{O}_6$ which is also known as blood sugar.
2. Determine the molecular mass of naphthalene C_{10}H_8 , which is used in mothballs.

Problem solving strategy:

Multiply atomic masses of carbon, hydrogen and oxygen by their subscripts and add.

Solution:

1. Molecular mass of $\text{C}_6\text{H}_{12}\text{O}_6$

$$\begin{aligned}
 &= 6(12.00) + 12(1.008) + 6(16.00) \\
 &= 180.096 \text{ amu}
 \end{aligned}$$
2. Molecular mass of C_{10}H_8

$$\begin{aligned}
 &= 12 \times 10 + 1 \times 8 \\
 &= 120 + 8 = 128 \text{ amu}
 \end{aligned}$$

The term molecular mass is used for molecular compounds. Whereas, the term formula mass is used for ionic compounds. Ionic compounds consist of arrays of oppositely charged ions rather than separate molecules. So we represent an ionic compound by its formula unit. A formula unit indicates the simplest ratio between cations and anions in an ionic compound. For example, the common salt consists of Na^+ and Cl^- ions. It has one Na^+ ion for every Cl^- ion. So formula unit for common salt is NaCl .

The sum of the atomic masses of all the atoms in the formula unit of a substance is called formula mass.

**Example 1.4: Determining formula mass**

1. Sodium Chloride, also called as table salt is used to flavour food, preserve meat, and in the preparation of large number of compounds. Determine its formula mass.
2. Milk of magnesia which contains $\text{Mg}(\text{OH})_2$, is used to treat acidity. Determine its formula mass.

Problem solving strategy:

Add the atomic masses of all the atoms in the formula unit.

Solution:

1. Formula mass of NaCl = 1 x Atomic mass of Na + 1 x Atomic mass of Cl
 = 1 x 23 + 1 x 35.5
 = 58.5 amu
2. Formula mass of $\text{Mg}(\text{OH})_2$ = 24 + 16x2 + 1x2
 = 24 + 32 + 2
 = 58 amu

**Self-Assessment Exercise 1.4**

1. Potassium Chlorate (KClO_3) is used commonly for the laboratory preparation of oxygen gas. Calculate its formula mass.
2. When baking soda, NaHCO_3 is heated carbon dioxide is released, which is responsible for the rising of cookies and bread. Determine the formula masses of baking soda and carbon dioxide.
3. Following compounds are used as fertilizers. Determine their formula masses.
 - (i) Urea, $(\text{NH}_2)_2\text{CO}$
 - (ii) Ammonium nitrate, NH_4NO_3 .

**Reading****Important information**

Many scientists regarded atom as merely a convenient mental construct and nothing more. This is because atom is so small that it cannot be seen with the naked eye. Today, however, we have sophisticated instruments to weigh atoms and even visualize them. Figure 1.1 shows an image of gold atoms on the surface.

**Figure 1.1 view of surface atoms of gold**

The image has been drawn by computer from signal sent to it by an instrument called a scanning tunneling microscope. The computer has drawn gold atoms as topped peaks.

1.3 CHEMICAL SPECIES

Most of the matter is composed of molecules or ions, formed by atoms.



1.3.1 Ions (Cation, Anions), Molecular Ions and Free Radicals

Atom is the smallest particle of an element that cannot exist in free state. It is electrically neutral. On the other hand an ion is a charged species formed from an atom or chemically bonded groups of atoms by adding or removing electrons. **Positively charged ions are called cations, whereas, the negatively charged ions are called anions.** An ionic compound

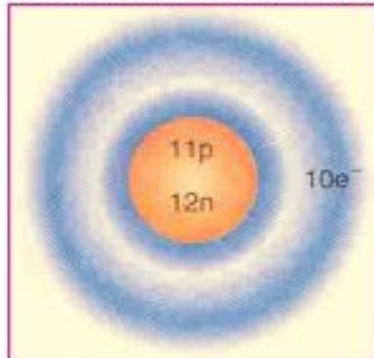


Fig.1.2 Na⁺ ion

contains anions and cations in such number that the compound is electrically neutral. Metal atoms generally lose one or more electrons and form cations. For example Na forms Na⁺ by losing one electron, Ca forms Ca²⁺ by losing two electrons. The Non-metal atoms usually gain one or more electrons and form anions. For example a chlorine atom gains one electron and forms Cl⁻ ion. An O-atom gains two electrons and forms O²⁻ ion.

Let us understand why an ion acquires a net positive or negative charge. Consider the formation of Na⁺ ion.

Figure 1.2 shows the sodium ion. Note that sodium has a nucleus of 11 protons and 12 neutrons. Thus its nucleus has a total charge of +11. Around the nucleus, there are 10 electrons, with a total charge of -10. The charge on the ion is $+11 + (-10) = +1$



Self-Assessment Exercise 1.5

Explain Why?

1. An oxide ion O²⁻ has -2 charge.
2. Magnesium ion, Mg²⁺ has +2 charge.
3. Sulphide ion, S²⁻ has -2 charge.



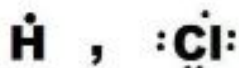
Reading

Molecular Ion

When a molecule loses or gains one or more electrons, the resulting species is called a molecular ion. For example O₂ when loses one electron it forms O₂⁺ ion, but when it absorbs an electrons it forms O₂⁻ ion. These ions are called molecular ions. Similarly N₂⁺, N₂⁻ etc are examples of molecular ions. These are short lived species and only exist at high temperature. Molecular ions do not form ionic compounds.

Free Radicals

A free radical is an atom which has an unpaired electron and bears no electrical charge. For example



are free radicals



When substances like halogens are exposed to sun light, their molecules split up into free radicals.

Difference between Ion and Free Radical



Chlorine free radical



Chloride ion

Which species has even number of electrons? Which species has odd number of electrons?

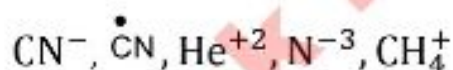
A free radical has an unpaired electron, so it has odd number of electrons. A free radical is an electrically neutral species. Whereas an ion has even number of electrons, so it has no unpaired electrons.

Dot (.) indicates an unpaired electron.



Self-Assessment Exercise 1.6

Identify ions, molecular ions and free radicals from the following species.



Reading

1.3.2 Representative Particles of Elements and Compounds

The term representative particles refer to species present in a substance. These species are atoms, molecules or formula units. For instance water exists as molecules, carbon exists as atoms.

Example 1.5: Identifying representative particles of elements and compounds

Figure 1.3 shows some molecules. Identify particles of elements and compounds.

Problem Solving Strategy:

Elements have atoms of same sizes and compounds have atoms of different sizes.

Solution:

Particles of elements are A, C, D and E. Particles of compounds are B and F.

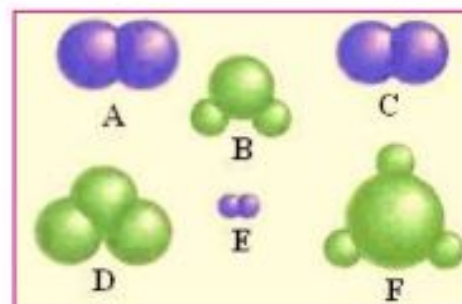


Fig 1.3 Some common molecules



Molecules can also be classified as monoatomic or polyatomic. Inert gases consist of monoatomic molecules such as He, Ne, Ar, Kr, Rn. **A molecule that contains only one atom is called monoatomic. Molecules that contain two or more similar or different atoms are called polyatomic molecules.** For example, H_2 , O_2 , HCl, NH_3 etc are polyatomic molecules.



Self-Assessment Exercise 1.7

- Observe the given figure and identify the diagrams that represents the particles of :
 - An element whose particles are atoms.
 - An element whose particles are molecules.
 - A compound.
 - A mixture of an element and a compound.
 - A mixture of two elements.
 - A mixture of two compounds.
- Observe the given figure and decide which diagram represents particles in an element, a compound or a mixture.



Reading

1.4 AVOGADRO'S NUMBER AND MOLE

How do you count shoes? As shoes come in pairs, so you would most likely count them by pairs rather than individually. Similarly eggs, oranges etc. are counted in dozens, but paper by ream. Thus, the counting unit depends on what you are counting.

Chemists also use a practical unit for counting atoms, molecules and ions. They use a counting unit called mole to measure the amount of a substance.

A mole is an amount of a substance that contains 6.022×10^{23} particles of that substance. This experimentally determined number is known as Avogadro's number. It is represented by N_A . Just as a dozen eggs represent twelve eggs, a ream of paper represent 500 papers, a mole of a substance represents 6.022×10^{23} representative particles of a substance..

Society, Technology and Science

During 600 – 1600 AD, Chemical history was dominated by a pseudo-science called alchemy. Earlier alchemists were obsessed with the idea of turning cheap metals into gold. They searched for ways to change less valued metals such as lead into gold. They could not succeed and wasted their time and money. Therefore, the works of earlier alchemists handicapped progress of science. However, during that period they discovered many new processes such as distillation, sublimation and extraction. These processes are still in use today. Such processes are contributing a lot in the progress of science. This means the works of different scientists at the same time handicap or promote the growth of science.



For example a mole of carbon is 6.022×10^{23} atoms. A mole of sulphur is 6.022×10^{23} atoms. A mole of water is 6.022×10^{23} molecules.

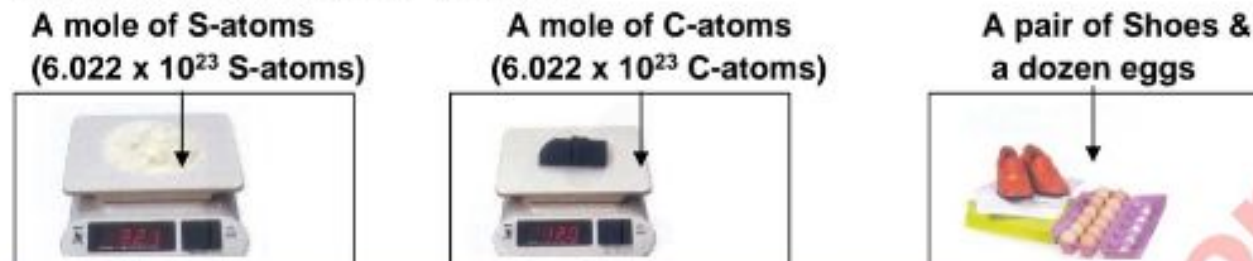


Figure 1.4 A mole of S-atoms, a mole of C-atoms & pair of Shoes & a dozen eggs

What is the mass of one mole C-atoms?

How many atoms are there in 32.1 g of S-atoms?

Does a dozen eggs have same mass as a dozen bananas? Does a mole of carbon atoms have a different mass than a mole of sulphur atoms?

The mass of one mole of substance is called as molar mass. What are the molar masses of carbon and sulphur? The term representative particles in a substance are atoms, molecules, formula units or ions. For instance water exists as molecules, therefore, one mole of water contains 6.022×10^{23} molecules of water. Hydrogen exists as H_2 molecules, so one mole of hydrogen contain 6.022×10^{23} molecules. Carbon exists as atoms so 1 mole of carbon contains 6.022×10^{23} atoms.

Society, Technology and Science Size of the Mole

Entire population cannot count 1 mole of coins in a year. They need about one million year to count them. So, when counting a pile of coins, it would not be convenient to count them one by one. The concept of mole has given a very simple method to count large number of items. Mole is not only a number but also represents definite amount of a substance. Just as 6.02×10^{23} carbon atoms weigh 12 g, 6.02×10^{23} coins will also have a definite mass. So an easy way is to weigh them. If you know the mass of one coin, you can count them by weighing.

1.4.3 Gram Atomic Mass, Gram Molecular Mass and Gram Formula Mass

A mole of S-atoms
(6.022×10^{23} S-atoms)



What is the mass of 6.022×10^{23} S-atoms?
Is this mass of S-atoms equal to its atomic mass?

A mole of C-atoms
(6.022×10^{23} C-atoms)



What is the mass of one mole of C-atoms?
Is this mass of C-atoms equal to its atomic mass?



Teacher's Point

Teacher may give more examples of mole.



Atomic mass of an element expressed in grams is called gram atomic mass.

Is the gram atomic mass of C-atoms 12 g?

What is the gram atomic mass of S-atoms?

If each of the carbon and sulphur sample shown above contains one mole of atoms, why do the samples have different masses?

Atomic mass of C = 12amu $\therefore$ gram atomic mass of C = 12g

Atomic mass of Na = 23amu $\therefore$ gram atomic mass of Na = 23g

Atomic mass of Zn = 63.54amu $\therefore$ gram atomic mass of Zn = 63.54g

Gram atomic mass of an element contains 1 mole of atoms.

Therefore,

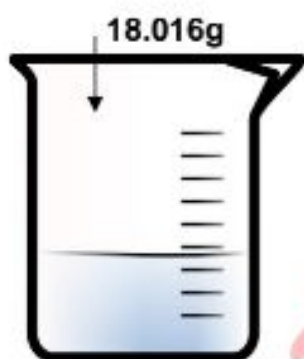
Mass of 1 mole of C-atoms = 12g

Mass of 1 mole of Na-atoms = 23g

Mass of 1 mole of Zn-atoms = 63.54g

A mole of H_2O -molecules

$(6.022 \times 10^{23} \text{ H}_2\text{O-molecules})$



What is the mass of one mole of water molecules?

Is this mass of water molecules equal to molecular mass of water?

A mole of $\text{C}_6\text{H}_{12}\text{O}_6$ - molecules

$(6.022 \times 10^{23} \text{ C}_6\text{H}_{12}\text{O}_6 \text{-molecules})$



What is the mass of 6.022×10^{23} molecules of glucose?

Is this mass of glucose molecules equal to molecular mass of glucose?

Molecular mass of a substance expressed in grams is called gram molecular mass.

Molecular mass of H_2O = $2 \times 1.008 + 16$

= 18.016amu

So, gram molecular mass of H_2O

= 18.016g

Molecular mass of $\text{C}_6\text{H}_{12}\text{O}_6$

= $6 \times 12 + 12 \times 1.008 + 16 \times 6$

= 180.096amu

So, gram molecular mass of $\text{C}_6\text{H}_{12}\text{O}_6$

= 180.096g



Formula mass of a substance expressed in gram is called gram formula mass.

An ionic compound is represented by the formula unit that represents the simplest ratio between the ions of a compound. For example NaCl, KCl, CuSO_4 etc.

$$\begin{aligned}\text{Formula mass of NaCl} &= 23 + 35.5 \\ &= 58.5\text{amu}\end{aligned}$$

Therefore, gram formula mass of NaCl = 58.5g = mole of NaCl formula units.

$$\begin{aligned}\text{Formula mass of KCl} &= 39 + 35.5 \\ &= 74.5\text{amu}\end{aligned}$$

So, gram formula mass of KCl = 74.5g

Difference between the Terms Gram Atomic Mass, Gram Molecular Mass And Gram Formula Mass

- Gram atomic mass represents one mole of atoms of an element, gram molecular mass represents one mole of molecules of a compound or an element that exists in molecular state whereas gram formula mass represents one mole of ionic formula units of a compound.
- Gram atomic mass contains 6.022×10^{23} atoms, gram molecular mass contains 6.022×10^{23} molecules whereas gram formula mass contain 6.022×10^{23} formula units.
- All of these quantities represent molar mass. Mass of one mole of a substance expressed in grams is called molar mass. **"Therefore, mole can be defined as atomic mass, molecular mass or formula mass expressed in grams".**

1.5 CHEMICAL CALCULATIONS

In this section, you will learn about the chemical calculations based on the concept of mole and Avogadro's number.

1.5.1 Mole-Mass Calculations

Example 1.5: Calculating mass of one mole of a substance

Calculate the molar masses of (a) Na (b) Nitrogen (c) Sucrose $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

Problem solving strategy:

If an element is a metal then its molar mass is its atomic mass expressed in grams (gram atomic mass). If an element exists as molecule, its molar mass is its molecular mass expressed in grams (gram molecular mass).

Solution:

a) 1 mole of Na = 23g

b) Nitrogen occurs as diatomic molecules.

$$\begin{aligned}\text{Molecular mass of N}_2 &= 14 \times 2 \\ &= 28\text{amu}\end{aligned}$$

$$\text{Therefore, mass of 1 mole of N}_2 = 28 \text{ g}$$



$$\begin{aligned} \text{c) Molecular mass of } \text{C}_{12}\text{H}_{22}\text{O}_{11} &= 12 \times 12 + 1 \times 22 + 16 \times 11 \\ &= 144 + 22 + 176 \end{aligned}$$

$$\text{Therefore, mass of 1 mole of sucrose} = 342\text{g}$$



Self-Assessment Exercise 1.8

Calculate the mass of one mole of (a) Copper (b) Iodine (c) Potassium (d) Oxygen

Example 1.5(a): Calculating the mass of a given number of moles of a substance

Oxygen is converted to ozone (O_3) during thunder storms. Calculate the mass of ozone if 9.05 moles of ozone is formed in a storm?

Problem solving strategy:

Ozone is a molecular substance. Determine its molar mass and use it to convert moles to mass in grams.

$$9.05 \text{ moles of } \text{O}_3 \longrightarrow ? \text{ g of } \text{O}_3$$

Solution:

$$\begin{aligned} 1 \text{ mole of } \text{O}_3 &= 16 \times 3 \\ &= 48 \text{ g} \end{aligned}$$

$$1 \text{ mole of } \text{O}_3 = 48 \text{ g}$$

$$\begin{aligned} \text{So, } 9.05 \text{ moles of } \text{O}_3 &= 48 \text{ g} \times 9.05 \\ &= 434.4\text{g of } \text{O}_3 \end{aligned}$$

Example 1.6: When natural gas burns CO_2 is formed. If 0.25 moles of CO_2 is formed, what mass of CO_2 is produced?

Problem solving strategy:

Carbon dioxide is a molecular substance. Determine its molar mass and use it to convert moles to mass in grams

$$0.25 \text{ moles of } \text{CO}_2 \longrightarrow ? \text{ g of } \text{CO}_2$$

Solution:

$$\begin{aligned} \text{Molar mass of } \text{CO}_2 &= 12 + 16 \times 2 \\ &= 44\text{g} \end{aligned}$$

$$1 \text{ mole of } \text{CO}_2 = 44\text{g of } \text{CO}_2$$

$$\begin{aligned} \text{So, } 0.25 \text{ moles of } \text{CO}_2 &= 44 \times 0.25 \\ &= 11\text{g of } \text{CO}_2 \end{aligned}$$

**Example 1.7: Converting grams to moles**

How many moles of each of the following substance are present?

- (a) A balloon filled with 5g of hydrogen.
 (b) A block of ice that weighs 100g.

Problem solving strategy:

Hydrogen and ice both are molecular substances. Determine their molar masses. Use the molar mass of each to convert masses in grams to moles.

mass $\longrightarrow$? moles

Solution:

- a) Molar mass of H_2 $= 1.008 \times 2$
 $= 2.016g$
 So, 1 mole of H_2 $= 2.016g$
 2.016g of H_2 $= 1$ mole of H_2
 1g of H_2 $= \frac{1}{2.016}$ moles of H_2
 5g of H_2 $= \frac{1}{2.016} \times 5$ moles of H_2
 $= 2.48$ moles of H_2
- b) 1 mole of H_2O $= 2 \times 1.008 + 16$
 $= 2.016 + 16$
 $= 18.016g$
 So, 18.016g of H_2O $= 1$ mole
 1g of H_2O $= \frac{1}{18.016}$ moles
 100g of H_2O $= \frac{1}{18.016} \times 100$ moles
 $= 5.55$ moles of H_2O

**Self-Assessment Exercise 1.9**

- The molecular formula of a compound used for bleaching hair is H_2O_2 . Calculate (a) Mass of this compound that would contain 2.5 moles. (b) No. of moles of this compound that would exactly weigh 30g.
- A spoon of table salt, NaCl contains 12.5grams of this salt. Calculate the number of moles it contains.
- Before the digestive systems X-rayed, people are required to swallow suspensions of barium sulphate $BaSO_4$. Calculate mass of one mole of $BaSO_4$.

**Teacher's Point**

A teacher may encourage students solve self-assessment exercise themselves.



Reading

1.5.2 Mole-Particles Calculations

Example 1.8: Calculating the number of atoms in given moles

1. Zn is a silvery metal that is used to galvanize steel to prevent corrosion. How many atoms are there in 1.25 moles of Zn.
2. A thin foil of aluminium (Al) is used as wrapper in food industries. How many atoms are present in a foil that contains 0.2 moles of aluminium?

Problem solving strategy:

Remember that symbols Zn and Al stand for one mole of Zn and Al atoms respectively.

Solution:

1. 1 mole of Zn contains $= 6.022 \times 10^{23}$ atoms
 1.25 moles of Zn contain $= 6.022 \times 10^{23} \times 1.25$
 $= 7.53 \times 10^{23}$ Zn atoms
2. 1 mole of Al contains $= 6.022 \times 10^{23}$ atoms
 So 0.2 moles of Al will contain $= 6.022 \times 10^{23} \times 0.2$
 $= 1.2044 \times 10^{23}$ atoms

Example 1.9: Calculating the number of molecules in given moles of a substance

1. Methane (CH_4) is the major component of natural gas. How many molecules are present in 0.5 moles of a pure sample of methane?
2. At high temperature hydrogen sulphide (H_2S) gas given off by a volcano is oxidized by air to sulphur dioxide (SO_2). Sulphur dioxide reacts with water to form acid rain. How many molecules are there in 0.25 moles of SO_2 .

Problem solving strategy:

Remember that CH_4 is a molecular compound, thus 1 mole of methane will have 6.022×10^{23} molecules. Similarly, SO_2 is a molecular compound, its one mole will also have 6.022×10^{23} molecules.

Solution:

1. 1 mole of CH_4 contains $= 6.022 \times 10^{23}$ molecules
 So, 0.5 moles of CH_4 will contain $= 6.022 \times 10^{23} \times 0.5$
 $= 3.011 \times 10^{23}$ molecules
2. 1 mole of SO_2 contains $= 6.022 \times 10^{23}$ molecules
 So, 0.25 moles of SO_2 will contain $= 6.022 \times 10^{23} \times 0.25$
 $= 1.5055 \times 10^{23}$ molecules

**Example 1.10: Calculating the number of moles in the given number of atoms**

Titanium is corrosion resistant metal that is used in rockets, aircrafts and jet engines. Calculate the number of moles of this metal in a sample containing 3.011×10^{23} Ti-atoms.

Problem solving strategy:

Remember that 1 mole of an element contains 6.022×10^{23} atoms.

Thus,

$$6.022 \times 10^{23} \text{ atoms} = 1 \text{ mole}$$

$$3.011 \times 10^{23} \text{ atoms} \longrightarrow ? \text{ moles}$$

Solution:

$$6.022 \times 10^{23} \text{ Ti atoms} = 1 \text{ mole of Ti}$$

$$1 \text{ Ti atom} = \frac{1}{6.022 \times 10^{23}} \text{ moles of Ti}$$

$$\begin{aligned} 3.011 \times 10^{23} \text{ Ti atoms} &= \frac{1}{6.022 \times 10^{23}} \times 3.011 \times 10^{23} \text{ moles of Ti} \\ &= 0.5 \text{ moles of Ti} \end{aligned}$$

Example 1.11: Calculating number of moles in the given number of molecules

Formaldehyde is used to preserve dead animals. Its molecular formula is CH_2O . Calculate the number of moles that would contain 3.011×10^{22} molecules of this compound.

Problem Solving Strategy:

Remember that 1 mole of any compound contains 6.022×10^{23} molecules.

Thus,

$$6.022 \times 10^{23} \text{ molecules} = 1 \text{ mole of compound}$$

$$3.011 \times 10^{22} \text{ molecules} \longrightarrow ? \text{ moles}$$

Solution:

$$6.022 \times 10^{23} \text{ molecules} = 1 \text{ mole of formaldehyde}$$

$$1 \text{ molecule} = \frac{1}{6.022 \times 10^{23}} \text{ moles of formaldehyde}$$

$$\begin{aligned} 3.011 \times 10^{22} \text{ molecules} &= \frac{1}{6.022 \times 10^{23}} \times 3.011 \times 10^{22} \text{ moles of formaldehyde} \\ &= 0.05 \text{ moles of formaldehyde} \end{aligned}$$

**Self-Assessment Exercise 1.10**

1. Aspirin is a compound that contains carbon, hydrogen and oxygen. It is used as a painkiller. An aspirin tablet contains 1.25×10^{30} molecules. How many moles of this compound are present in the tablet?
2. A method used to prevent rusting in ships and underground pipelines involves connecting the iron to a block of a more active metal such as magnesium. This method is called cathodic protection. How many moles of magnesium are present in 1 billion (1×10^9) atoms of magnesium.

**Key Points**

- Chemistry is the science of materials of the universe.
- The branch of Chemistry that deals with laws and theories to understand the structure and changes of matter is called Physical Chemistry. Organic chemistry deals with carbon compounds (except bicarbonates carbonates oxides and carbides).
- The branch of Chemistry that deals with elements and their compounds except organic compounds is called Inorganic Chemistry.
- Physical and chemical changes that occur in living organisms are studied in biochemistry.
- Industrial chemistry is concerned with the large scale production of chemical substances.
- An element is a substance all the atoms of which have the same atomic number.
- A compound consists of two or more elements held together in fixed proportions by chemical bonds.
- An impure substance that contains two or more pure substances that retain their individual chemical characteristics is called a mixture.
- The number of protons in the nucleus of an atom is known as its atomic number.
- The total number of protons and neutrons in an atom is called its mass number.
- Atoms of an element that have different number of neutrons are called isotopes.
- The mass of an atom of an element relative to the mass of an atom of C-12 is called relative atomic mass.
- One atomic mass unit is defined as the mass exactly equal to one-twelfth the mass of one C-12 atom.
- Chemical formula of a compound that gives the simplest whole-number ratio between atoms is called empirical formula.
- Molecular formula of a compound gives the exact number of atoms present in a molecule.
- Molecular mass is the sum of atomic masses of all the atoms present in the molecule.
- Positively charged ions are called cations and negatively charged ions are called anions.
- When a molecule loses or gains electrons, the resulting species is called molecular ion.



- Free radical is an atom or group of atoms that contains an unpaired electron.
- The number of representative particles in one mole of the substance is known as Avogadro's number.
- The amount of matter that contains as many atoms, ions or molecules as the number of atoms in exactly 12g of C-12 is called mole. Mole can also be defined as atomic mass, molecular mass or formula mass expressed in grams.
- Atomic mass of an element expressed in grams is called gram atomic mass.
- Molecular mass of an element or a compound expressed in grams is its gram molecular mass.
- Gram formula mass is the formula mass of a substance in grams.

REFERENCES FOR ADDITIONAL INFORMATION

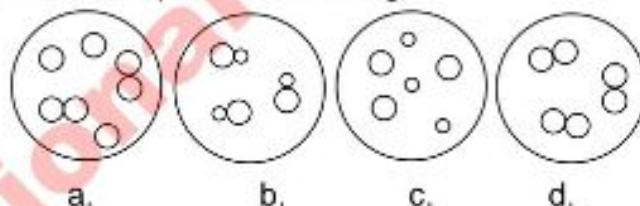
- Zumdahl, Introductory Chemistry.
- Raymond Chang, Essential Chemistry.

**Activity**

Take two balloons. Fill one balloon full of air and other half. Then weigh both and compare the masses of both. It confirms that masses of gases increase with increase in number of molecules in it.

**Review Questions****1. Encircle the correct answer:**

- (i) Which of the following lists contains only elements?
 (a) Air, water, oxygen (b) Hydrogen, oxygen, brass
 (c) Air, water, fire, earth (d) Calcium, sulphur, carbon
- (ii) The diagrams below represent particles in four substances, which box represent the particles in nitrogen.



- (iii) What is the formula mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. (Atomic masses: Cu=63.5, S=32, O=16, H=1)
 (a) 159.5 (b) 185.5 (c) 249.5 (d) 149.5
- (iv) A compound with chemical formula Na_2CX_3 has formula mass 106amu. Atomic mass of the element X is
 (a) 106 (b) 23 (c) 12 (d) 16
- (v) How many moles of molecules are there in 16g oxygen.
 (a) 1 (b) 0.5 (c) 0.1 (d) 0.05
- (vi) What is the mass of 4 moles of hydrogen gas.
 (a) 8.064g (b) 4.032g (c) 1g (d) 1.008g



- (vii) What is the mass of carbon present in 44g of carbon dioxide.
(a) 12g (b) 6g (c) 24g (d) 44g
- (viii) The electron configuration of an element is $1s^2 2s^2$. An atom of this element will form an ion that will have charge.
(a) +1 (b) +2 (c) +3 (d) -1
- (ix) Which term is the same for one mole of oxygen and one mole of water?
(a) volume (b) (c) mass (d) atoms (e) molecules
- (x) If one mole of carbon contains x atoms, what is the number of atoms contained in 12g of Mg.
(a) x (b) $0.5x$ (c) $2x$ (d) $1.5x$
2. Give short answers.
- (i) Differentiate between an ion and a free radical
- (ii) What do you know about corpuscular nature of matter?
- (iii) Differentiate between analytical chemistry and environmental chemistry.
- (iv) What is mole?
- (v) Differentiate between empirical formula and molecular formula.
- (vi) What is the number of molecules in 9.0 g of steam?
- (vii) What are the molar masses of uranium -238 and uranium -235?
- (viii) Why one mole of hydrogen molecules and one mole of H-atoms have different masses?
3. Define ion, molecular ion, formula unit, free radical, atomic number, mass number, atomic mass unit.
4. Differentiate between (a) atom and ion (b) molecular ion and free radical.
5. Describe how Avogadro's number is related to a mole of any substance.
6. Calculate the number of moles of each substance in samples with the following masses:
- (a) 2.4 g of He (b) 250mg of carbon (c) 15g of sodium chloride
(d) 40g of sulphur (e) 1.5kg of MgO
7. Calculate the mass in grams of each of the following samples:
- (a) 1.2 moles of K (b) 75moles of H_2 (c) 0.25 moles of steam
(d) 1.05 moles of $CuSO_4 \cdot 5H_2O$ (e) 0.15moles of H_2SO_4
8. Calculate the number of molecules present in each of the following samples:
- (a) 2.5 moles of carbon dioxide (b) 3.4 moles of ammonia, NH_3
(c) 1.09 moles of benzene, C_6H_6 (d) 0.01 moles of acetic acid, CH_3COOH
9. Decide whether or not each of the following is an example of empirical formula:
- (a) Al_2Cl_6 (b) Hg_2Cl_2 (c) NaCl (d) C_2H_6O
10. TNT or trinitrotoluene is an explosive compound used in bombs. It contains 7 C-atoms, 5 H-atoms, 3 -N atoms and 6 O-atoms. Write its empirical formula.
11. A molecule contains four phosphorus atoms and ten oxygen atoms. Write the empirical formula of this compound. Also determine the molar mass of this molecule.



12. **Indigo ($C_{16}H_{10}N_2O_2$)**, the dye used to colour blue jeans is derived from a compound known as **indoxyl (C_8H_7ON)**. Calculate the molar masses of these compounds. Also write their empirical formulas.
13. **Identify the substance that has formula mass of 133.5amu.**
(a) $MgCl_2$ (b) S_2Cl_2 (c) BCl_3 (d) $AlCl_3$
14. **Calculate the number of atoms in each of the following samples:**
(a) 3.4 moles of nitrogen atoms (b) 23g of Na (d) 5g of H atoms
15. **Calculate the mass of the following:**
(a) 3.24×10^{18} atoms of iron (b) 2×10^{10} molecules of nitrogen gas
(c) 1×10^{25} molecules of water (d) 3×10^6 atoms of Al
16. **Identify the branch of chemistry that deals with the following examples:**
1. A cornstalk grows from a seed.
 2. Dynamite ($C_3H_5N_3O_9$) explodes to form a mixture of gases.
 3. Purple iodine vapour appears when solid iodine is warmed.
 4. Gasoline (a mixture of hydrocarbons) fumes are ignited in an auto mobile engine.
 5. A silver article tarnishes in air.
 6. Ice floats on water.
 7. Sulphur dioxide is the major source of acid rain.
 8. Many other light chlorinated hydrocarbons in drinking water are carcinogens.
 9. In Pakistan most of the factories use wet process for the production of cement.
 10. Carbon-14 is continuously produced in the atmosphere when high energy neutrons from space collide with nitrogen-14.



Think-Tank

1. What mass of sodium metal contains the same number of atoms as 12.00g of carbon.
2. What mass of oxygen contains the same number of molecules as 42g of nitrogen.
3. Calculate the mass of one hydrogen atom in grams.
4. Observe the given figure. It shows particles in a sample of air.
 - a) Count the substances shown in the sample
 - b) Is air a mixture or pure substance? Explain?
 - c) Identify the formula of each substance in air.
 - d) Argue whether each substance in air is an element or a compound.
 - e) What is the most common substance in air?
5. Calculate the number of H-atoms present in 18g H_2O .
6. Calculate the total number of atoms present in 18g H_2O .





2

STRUCTURE OF ATOMS



This is a 14 days lesson

After completing this lesson, you will be able to:

- Describe the contribution the Rutherford Atomic Theory in the structure of the atoms.
- Explain how Bohr's Atomic Theory differed from Rutherford's Atomic theory.
- Compare isotopes of an element.
- Describe the structure of an atom including the location of the proton, electron and neutron.
- Describe the contribution that Rutherford made to the development of the atomic theory
- Describe the presence of sub shells in a shell.
- Define isotopes.
- Distinguish between shells and sub shells.
- Discuss the properties of the isotopes of H, C, Cl, U.
- Draw the structures of different isotopes from mass number and atomic number.
- State the importance and uses of isotopes in various fields of life.
- Write the electronic configuration of the first 18 elements in the Periodic Table.



Pre- Reading

The structure of atom tells how the sub-atomic particles are arranged. Atoms are so small that they can only be visualized with a scanning tunneling microscope. See figure 2.1 that shows an image of gold atoms on the surface. In grade VII you have learned about the structure of atom and sub-atomic particles, electrons, protons and neutrons. In this chapter you will learn about the arrangement of these particles in an atom. Atomic structure was formulated from a series of experiments during the later part of nineteenth century and the beginning of the twentieth

century. We will discuss contributions of British physicist Rutherford and Danish physicist Neil Bohr for determining the structure of an atom. Bohr's structure of an atom nicely explains the arrangement of elements in the periodic table and periodicity of properties. How? To understand this you should know the structure of atoms.



Reading

2.1 THEORIES AND EXPERIMENTS RELATED TO ATOMIC STRUCTURE

In grade VII you have learned about the structure of atom and sub-atomic particles. You have also learned the distribution of electrons in shells (KLM only) using $2n^2$ formula. In this section you will learn about theories and experiments related to atomic structure.



2.1.1 Rutherford's Atomic Model

In 1911 Rutherford performed an experiment in order to know the arrangement of electrons and protons in atoms.

2.1.1.1 Rutherford's Experiment

Rutherford bombarded a very thin gold foil about 0.00004cm thickness with α -particles. (figure 2.1). He used α -particles obtained from the disintegration of polonium. α -particles are helium nuclei that are doubly positively charged (He^{++}). Most of these particles passed straight through the foil. Only few particles were slightly deflected. But one in 1 million was deflected through an angle greater than 90° from their straight paths. Rutherford performed a series of experiments using thin foils of other elements. He observed similar results from these experiments.

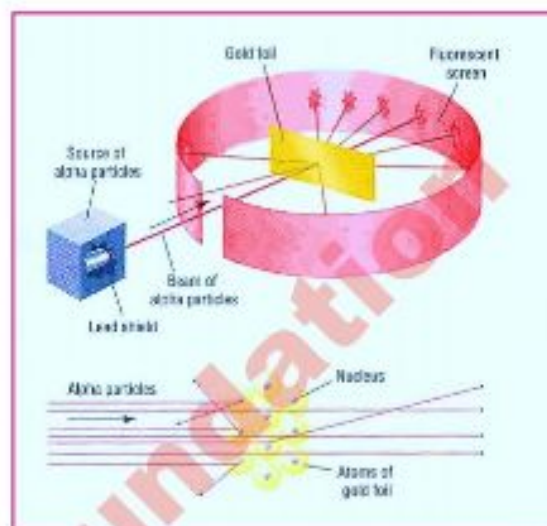


Figure 2.1 Rutherford's Experiment

Rutherford drew the following conclusions:

1. Since majority of the α -particles passed through the foil undeflected, most of the space occupied by an atom must be empty.
2. The deflection of a few α -particles through angles greater than 90° shows that these particles are deflected by electrostatic repulsion between the positively charged α -particles and the positively charged part of atom.
3. Massive α -particles are not deflected by electrons.

On the basis of conclusions drawn from these experiments, Rutherford proposed a new model for an atom. He proposed a planetary model (similar to the solar system) for an atom. An atom is a neutral particle. The mass of an atom is concentrated in a very small dense positively charged region. He named this region as nucleus. The electrons are revolving around the nucleus in circles. These circles are called orbits. The centripetal force due to the revolution of electrons balances the electrostatic force of attraction between the nucleus and the electrons.

2.1.1.2 Defects in Rutherford's Atomic Model

Rutherford's model of an atom resembles our solar system. It has following defects:

1. Classical physics suggests that electron being charged particle will emit energy continuously while revolving around the nucleus. Thus the orbit of the revolving electron becomes smaller and smaller until it would fall into the nucleus. This would collapse the atomic structure.
2. If revolving electron emits energy continuously it should form a continuous spectrum for an atom but a line spectrum is obtained.



Teacher's Point

Tube light is a discharge tube



Bohr formulated new explanation and a new theory to remove defects from the Rutherford's atomic model.

2.1.2 Bohr's Atomic Theory

In 1913 Neil Bohr, proposed a model for an atom that was consistent with Rutherford's model. But it also explains the observed line spectrum of the hydrogen atom. Main postulates of Bohr's atomic theory are as follows:

1. The electron in an atom revolves around the nucleus in one of the circular orbits. Each orbit has a fixed energy. So each orbit is also called energy level.
2. The energy of the electron in an orbit is proportional to its distance from the nucleus. The farther the electron is from the nucleus, the more energy it has.
3. The electron revolves only in those orbits for which the angular momentum of the electron is an integral multiple of $\frac{h}{2\pi}$ where h is Planck's constant (its value is 6.626×10^{-34} J.s).
4. Light is absorbed when an electron jumps to a higher energy orbit and emitted when an electron falls into a lower energy orbit. Electron present in a particular orbit does not radiate energy.
5. The energy of the light emitted is exactly equal to the difference between the energies of the orbits.

Society, Technology and Science

Rutherford was the first scientist who proposed first atomic model of an atom. He suggested that all of the positive charge and most of the mass of the atom is concentrated in the nucleus. The remaining volume of the atom is occupied by electrons that revolve around the nucleus in circles called orbits. These suggestions remained unchallenged. But his model could not explain the stability of an atom and line spectrum for an atom. Bohr leaped over difficulty by using Quantum Theory of Radiation that was proposed by Max Planck. Bohr proposed that an electron moves around the nucleus in well defined circular paths called orbits. An orbit has fixed energy. Electron present in an orbit does not emit energy. Bohr atomic theory explains nicely the stability of an atom and also explains why an atom gives line spectrum. Development of Bohr's atomic model explains how interpretations of experimental results of other scientists help chemists to formulate new explanations and new theories.

$$\Delta E = E_2 - E_1$$

Where ΔE is the energy difference between any two orbits with energies E_1 and E_2

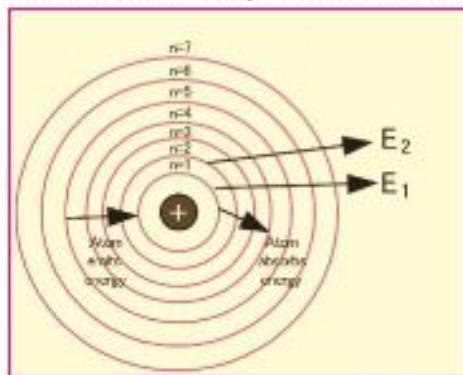


Figure 2.2 Bohr's model of the atom

**Self-Assessment Exercise 2.1**

Draw Bohr's Model for the following atoms indicating the location for electrons, protons and neutrons,

- (a) Carbon (Atomic No. 6, Mass No. 12)
- (b) Chlorine (Atomic No. 17, Mass No. 35)

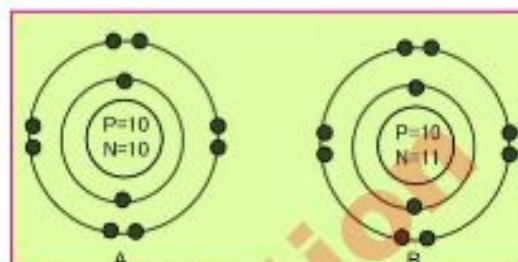
**Reading**

Figure 2.3 Bohr's model for atom A and atom B

2.2 ISOTOPES

Figure 2.3 shows Bohr's Model for two atoms A and B

Can you identify three similarities and two differences in these atoms?

You will find,

- (a) Both the atoms have same number of protons.
- (b) Both the atoms have same number of electrons.
- (c) Both have same atomic number.
- (d) Both have different number of neutrons.
- (e) Both differ in total number of protons and neutron. This means they have different mass numbers.

Since both the atoms have the same atomic number, they must be the atoms of same element and are called isotopes. The word isotope was first used by Soddy. It is a Greek word "isos" means same and "tope" means place.

Isotopes are atoms of an element whose nuclei have the same atomic number but different mass number. This is because atoms of an element can differ in the number of neutrons. Isotopes are chemically alike and differ in their physical properties.

How does the discovery of isotopes contradicted Dalton's atomic theory?

Isotopes of Hydrogen

Hydrogen has three isotopes. Hydrogen -1 (Protium) has no neutron. Almost all the hydrogen is Hydrogen -1. Its symbol is ${}^1_1\text{H}$. Hydrogen -2 (deuterium) has one neutron and hydrogen -3 (Tritium) has two neutrons. Their symbols are ${}^2_1\text{H}$ and ${}^3_1\text{H}$ respectively. Because hydrogen -1 also known as protium has only one proton, adding a neutron doubles its mass.

Society, Technology and Science

Dalton's atomic theory explained data from many experiments. So it was widely accepted. Discovery of sub-atomic particles and isotopes proved that some of the Dalton's ideas about atoms were not correct. Scientists did not discard his theory. Instead, they revised the theory to take into account new discoveries. This shows how prevailing theories bring about changes in them.



Protium / Hydrogen is a colourless, odourless, and tasteless gas. It is insoluble in water and is highly inflammable gas. Water that contain hydrogen-2 atoms in place of hydrogen-1 is called heavy water. Table 2.1 Shows some physical properties of ordinary water and heavy water.

Table 2.1 - Comparison of ordinary water and heavy water.

Property	Ordinary water	Heavy water
Melting Point	0.00°C	3.81°C
Boiling point	100°C	101.2°C
Density at 25°C	0.99701 g/cm ³	1.1044 g/cm ³

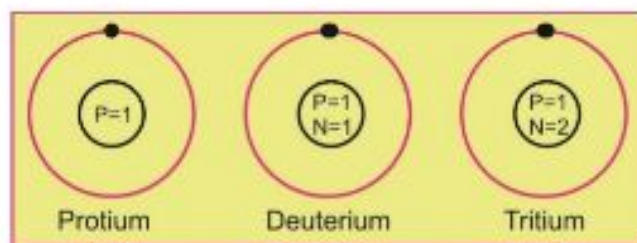


Figure 2.5 Isotopes of hydrogen

At what temperature would a sample of heavy water freeze?

Naturally occurring hydrogen contains 99.99% protium, 0.0015% Deuterium. Tritium is radioactive and is rare. Tritium is not found in naturally occurring hydrogen because its nucleus is unstable.

Isotopes of Carbon

Carbon has three isotopes. Carbon-12, carbon-13 and carbon -14. Almost all the carbon is carbon-12. Its symbol is ${}^{12}_6\text{C}$. It has six neutrons and six protons. Carbon-13 has symbol ${}^{13}_6\text{C}$. It has seven neutrons and six protons. Carbon-14 has eight neutrons and six protons. Its symbol is ${}^{14}_6\text{C}$. Different forms of carbon are black or greyish black solids except diamond. They are odourless and tasteless. They have high melting and boiling points and are insoluble in water.



Activity 2.1

Carbon has three isotopes ${}^{12}_6\text{C}$, ${}^{13}_6\text{C}$, ${}^{14}_6\text{C}$ Figure 2.6 shows incomplete structure of isotopes of carbon. Can you complete it?

Natural abundance of isotopes of carbon is as follows

$${}^{12}_6\text{C} = 98.8\%, \quad {}^{13}_6\text{C} = 1.1\%, \quad {}^{14}_6\text{C} = 0.009\%$$

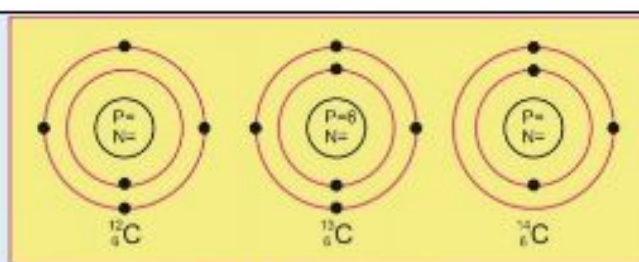


Figure 2.6 Isotopes of Carbon

Isotopes of Chlorine

There are two natural isotopes of chlorine, chlorine-35 and chlorine-37. An atom of chlorine-35 has 17 protons and 18 neutrons. An atom of chlorine-37 has 17 protons and 20 neutrons. Chlorine-35 occurs in nature about 75% and chlorine-37 about 25%. Chlorine is a greyish yellow gas with sharp pungent irritating smell. It is fairly soluble in water.

**Activity 2.2**

Chlorine has two isotopes. Figure 2.7 shows the structure of isotopes of chlorine. Can you write isotope symbol for each?

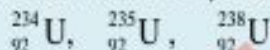
Isotope symbols: _____
 Natural abundance 75.77% 24.23%



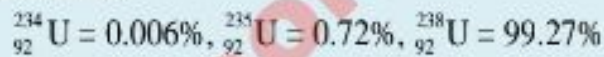
Figure 2.7 Isotopes of chlorine

Isotopes of Uranium**Activity 2.3**

Uranium has three isotopes with mass number 234, 235 and 238 respectively.



The ${}_{92}^{235}\text{U}$ isotope is used in nuclear reactors and atomic bombs, whereas the ${}_{92}^{238}\text{U}$ isotope lacks the properties necessary for these applications. ${}_{92}^{234}\text{U}$ is rare. Natural abundance of Uranium isotopes is as follows:



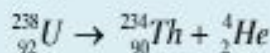
Fill in the blanks?

${}_{92}^{234}\text{U}$ has ____ protons, ____ electrons and ____ neutrons

${}_{92}^{235}\text{U}$ has ____ protons, ____ electrons and ____ neutrons

${}_{92}^{238}\text{U}$ has ____ protons, ____ electrons and ____ neutrons

When uranium-238 decays into thorium-234, it emits alpha particle. An alpha particle is doubly positively charged helium nucleus.



The fission of uranium-235 yields smaller nuclei, neutron and energy. The nuclear energy released by the fission of one kilogram of uranium-235 is equivalent to chemical energy produced by burning more than 17000 kg of coal.

Chemical properties of an element depend upon the number of protons and electrons. Neutrons do not take part in ordinary chemical reactions. Therefore, isotopes of an element have similar chemical properties.

2.2.2 Uses of Isotopes

Stable and radioactive isotopes have many applications in science and medicines. Some of these are as follows:

- (i) Radioactive iodine -131 is used as a tracer in diagnosing thyroid problem.



- (ii) Na-24 is used to trace the flow of blood and detect possible constrictions or obstructions in the circulatory system.
- (iii) Iodine-123 is used to image the brain.

Important information

Carbon-14 is used to estimate the age of carbon-containing substances. Carbon atoms circulate between the oceans, and living organism at a rate very much faster than they decay. As a result the concentration of C-14 in all living things keep on increasing. After death organisms no longer pick up C-14. By comparing the activity of a sample of skull or jaw bones, with the activity of living tissues, we can estimate how long it has been since the organism died. This process is called dating.

- (iv) Cobalt-60 is commonly used to irradiate cancer cells in the hope of killing or shrinking the tumors.
- (v) Carbon-14 is used to trace the path of carbon in photosynthesis.
- (vi) Radioactive isotopes are used to determine the molecular structure e.g. sulphur-35 has been used in the structure determination of thiosulphate, $S_2O_3^{2-}$ ion.
- (vii) Radioactive isotopes are also used to study the mechanism of chemical reactions.
- (viii) Radioactive isotopes are used to date rocks, soils, archaeological objects, and mummies.

2.3 ELECTRONIC CONFIGURATION

To understand electronic configuration, you should know about shells and sub-shells.

2.3.1 Shells

According to Bohr's atomic theory, the electron in an atom revolves around the nucleus in one of the circular paths called shells or orbits. Each shell has a fixed energy. So each shell is also called energy level. Each shell is described by an n value. n can have values 1,2,3,....

When,

$n = 1$, it is K shell

$n = 2$, it is L shell

$n = 3$, it is M shell etc.

As the value of n increases the distance of electron from the nucleus and energy of the shell increases.

2.3.2 Sub-Shells

A shell or energy level is sub divided into sub-shells or sub-energy levels. n value of a shell is placed before the symbol for a sub-shell. For instance

$n = 1$, for K shell. It has only one sub-shell which is represented by 1s.

For L shell $n = 2$, L shell has two sub-shells, these are designated as 2s and 2p.

For M shell $n = 3$ So M shell has 3 sub-shells called 3s, 3p and 3d. While N shell has 4s, 4p, 4d and 4f sub-shells.

s sub-shell can accommodate maximum 2 electrons.

p sub-shell can accommodate maximum 6 electrons.



d sub-shell can accommodate maximum 10 electrons.

f sub-shell can accommodate maximum 14 electrons.

The increasing order of energy of the sub-shells belonging to different shells is given below.

$$1s < 2s < 2p < 3s < 3p < 4s < 3d \dots$$

The arrangement of electrons in sub-shells is called as the electronic configuration.

We can fill the electrons present in various elements by using Auf Bau Principle. **According to this principle, electrons fill the lowest energy sub-shell that is available first.** This means electron will fill first 1s, then 2s, then 2p and so on.

2.3.3 Electronic Configuration of First 18 Elements.

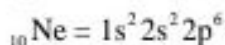
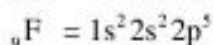
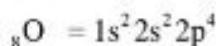
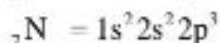
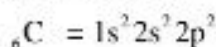
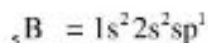
Electronic configuration is the distribution of electrons among the different sub-shells of an atom. We can do this by listing the symbol for the occupied sub-shells one after another. Show the number of electrons in the sub-shell as a superscript to each symbol. Because the energies of sub-shells increase in the order, 1s, 2s, 2p, 3s, 3p (as indicated in section 2.2.1), the first five sub-shells fill in that order. Hydrogen has atomic number 1. So it has only one electron that will occupy lowest energy sub-shell 1s. The electronic configuration of H is $1s^1$.

Helium has atomic number 2, so it has two electrons. Since s sub-shell can accommodate two electrons, so electronic configuration of He is $1s^2$.

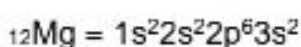
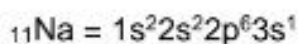
Lithium has atomic number 3, so it has three electrons, two will fill 1s sub-shell and one 2s sub-shell. So electronic configuration of Li is $1s^2 2s^1$.

Beryllium has atomic number 4, so it has four electrons. Two of these electrons go into 1s sub-shell and two will go to 2s sub-shell. Thus electronic configuration of Be is $1s^2 2s^2$.

Once 2s sub-shell is filled, the 2p sub-shell begins to fill. 2p sub-shell can hold maximum 6 electrons. So next six elements will have configuration in which 2p sub-shell will be progressively filled. Therefore, these elements will have following electronic configuration.

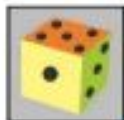


After 2p has completely filled, the additional electrons will fill 3s sub-shell, so electronic configuration of Na & Mg would be





After 3s has completely filled 3p sub-shell begins to fill. So next six elements have electronic configuration by filling 3p sub-shell.



Self-Assessment Exercise 2.3

Write the complete electronic configuration for the following elements;

- Al (atomic number 13)
 Si (atomic number 14)
 P (atomic number 15)
 S (atomic number 16)
 Cl (atomic number 17)
 Ar (atomic number 18)

Figure 2.10 shows the electronic configuration in the sub-shell last occupied for the first eighteen elements.

H 1s ¹							He 1s ²
Li 2s ¹	Be 2s ²	B 2s ² 2p ¹	C 2s ² 2p ²	N 2s ² 2p ³	O 2s ² 2p ⁴	F 2s ² 2p ⁵	Ne 2s ² 2p ⁶
Na 3s ¹	Mg 3s ²	Al 3s ² 3p ¹	Si 3s ² 3p ²	P 3s ² 3p ³	S 3s ² 3p ⁴	Cl 3s ² 3p ⁵	Ar 3s ² 3p ⁶

Figure 2.10 Valence shell configuration of first 18 elements



Self-Assessment Exercise 2.4

Write the electronic configuration for the following isotopes

- (a) $^{14}_6\text{C}$, (b) $^{35}_{17}\text{Cl}$, (c) $^{37}_{17}\text{Cl}$



Key Points

- Rutherford proposed a planetary model for an atom. The nucleus of an atom is composed of protons. The electrons surround the nucleus and occupy most of the volume of the atom.
- According to Bohr's atomic model, the electron in an atom revolves around the nucleus in fixed circular orbits called shells. Energy is absorbed when an electron jumps to a higher energy orbit and emitted when an electron falls into a lower energy orbit.



- Isotopes are atoms of an element that differ in the number of neutrons.
- $^{235}_{92}\text{U}$ isotope is used in nuclear reactors and atomic bombs.
- Radioactive isotopes have many applications in science and medicines such as killing cancer cells, diagnosing thyroid problem, to image the brain, to detect obstruction in the circulatory system, to date rocks, soils, mummies etc.
- A shell or energy level is divided into sub-shells.
- The arrangement of electrons in sub-shells is called as the electronic configuration.
- According to the Auf Bau Principle, electrons fill the lowest energy levels first.

REFERENCES FOR ADDITIONAL INFORMATION

- B.Earl and LDR Wilford, Introduction to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).



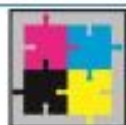
Review Questions

1. Encircle the correct answer:

- According to Bohr atomic model:
 - Each orbit has fixed energy, so each orbit is called sub-energy level.
 - The energy of the electron is inversely proportional to its distance from the nucleus.
 - Light is absorbed when an electron jumps a lower energy orbit.
 - The further the electron is from the nucleus, the more energy it has.
- Chlorine has two isotopes, both of which have
 - same mass number.
 - same number of neutrons.
 - different number of protons.
 - same number of electrons.
- Number of neutrons in $^{27}_{13}\text{M}$ are
 - 13
 - 14
 - 27
 - 15
- Which isotope is commonly used to irradiate cancer cells?
 - Iodine-123
 - Carbon-14
 - Cobalt-60
 - Iodine-131
- M shell has sub-shells:
 - 1s, 2s
 - 2s, 2p
 - 3s, 3p, 3d
 - 1s, 2s, 3s
- A sub-shell that can accommodate 6 electrons is
 - s
 - d
 - p
 - f
- $_{11}\text{Na}$ has electronic configuration:
 - $1s^2 2s^2 3s^1$
 - $1s^2 2s^2 2p^7$
 - $1s^2 2s^2 2p^5 3s^2$
 - $1s^2 2s^2 2p^6 3s^1$
- Rutherford used _____ particles in his experiments.
 - He atoms
 - He^+
 - He^{+2}
 - He^{-2}



- (ix) Which of the following statement is not correct about isotopes?
 (a) they have same atomic number (b) they have same number of protons
 (c) they have same chemical properties (d) they have same physical properties
- (x) Which isotope is used in nuclear reactors?
 (a) U-234 (b) U-238 (c) U-235 (d) All of these
2. **Give short answers**
- Distinguish between shell and sub-shell
 - Why an atom is electrically neutral?
 - How many sub-shells are there in N shell.
 - Give notation for sub-shells of M shell.
 - List the sub-shells of M Shell in order of increasing energy
 - Can you identify an atom without knowing number of neutrons in it?
3. The electronic configurations listed are incorrect. Explain what mistake have been made in each and write correct electronic configurations.
 $x = 1s^2 2s^2 2p^4 3p^2$, $y = 1s^2 2s^1 2p^1$, $z = 1s^2 2s^2 2p^5 3s^1$
4. **Which orbital in each of the following pairs is lower in energy?**
 (a) 2s, 2p (b) 3p, 2p (c) 3s, 4s
5. **Draw Bohr's Model for the following atoms indicating the location for electron, protons and neutrons:**
 (a) Potassium (Atomic No 19, Mass No. 39) (b) Silicon (Atomic No. 14 Mass No. 28)
 (c) Argon (Atomic No. 18 Mass No. 39)
6. **Write electronic configuration for the following elements:**
 (a) $^{28}_{14}\text{Si}$ (b) $^{24}_{12}\text{Mg}$ (c) $^{27}_{13}\text{Al}$ (d) $^{40}_{18}\text{Ar}$
7. **Describe the contribution that Rutherford made to the development of the atomic theory.**
8. **Explain how Bohr's atomic theory differed from Rutherford's atomic theory.**
9. **Describe the presence of sub shells in a shell.**
10. **State the importance and uses of isotopes in various fields of life.**
11. **The atomic number of an element is 23 and its mass number is 56.**
 a. How many protons and electrons does an atom of this element have?
 b. How many neutrons does this atom have?
12. **The atomic symbol of aluminum is written as $^{27}_{13}\text{Al}$. What information do you get from it?**
13. **How the prevailing theories about testing bring changes in atomic theories?**
14. **How experimental results of some scientists help chemist to formulate new theories and new explanation.**

**Think-Tank**

15. M-24 is a radioactive isotope used to diagnose restricted blood circulation, for example in legs. How many electrons, protons and neutrons are there in this isotope. Valence shell electronic configuration of M is $3s^1$.
- 16: Two isotopes of chlorine are $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$. How do these isotopes differ? How are they alike?
- 17: How many electrons can be placed in all of the sub-Shells in the $n=2$ shell?
- 18: Mass number of an atom indicates total number of protons and neutrons in the nucleus. Can you identify an atom without any neutron?
- 19: The table shows the nuclei of five different atoms.

Name of atom	Number of Protons	Number of neutrons
A	5	6
B	6	6
C	6	7
D	7	7
E	8	8

- Which atom has the highest mass number?
 - Which two atoms are isotopes?
 - Which atom has the least number of electrons?
 - Which atom will have electronic configuration $1s^2 2s^2 2p^3$.
 - Which of the atom contains the most number of electrons?
- 20: Naturally occurring nitrogen has two isotopes N-14 and N-15 select isotope that has greater number of electrons.



3

PERIODIC TABLE AND PERIODICITY OF PROPERTIES



This is a 20 days lesson

After completing this lesson, you will be able to:

- Classify the element (into two categories: groups and periods) according to the configuration of their outer most shell.
- Distinguish between a period and a group in the periodic table. (Understanding)
- State the periodic law.
- Determine the demarcation of the periodic table into an s block and p block.
- Determine the location of families in the periodic table. Recognize the similarity in the chemical and physical properties of elements in the same family of elements.
- Describe how electronegativities change within a group and within a period in the periodic table.
- Explain the shape of the periodic table.
- Explain how shielding effect influences periodic trends.
- Identify the relationship between electronic configuration and the position of an element in the periodic table.



Pre- Reading

By the end of 18th century, 23 elements were known, by 1870, 65, by 1925, 88, today there are 109. These elements combine to form millions of compounds. It is very difficult rather impossible to remember information concerning reactions, properties and atomic masses of elements. So we clearly need some way to organize our knowledge about them.

The periodic table is one of the most important tools in chemistry. It is very useful for understanding and predicting the properties of elements. For instance if you know physical and

chemical properties of one element in a group, you can predict about the physical and chemical properties of any other element present in the same group. You can use periodic table to relate trends in the reactivity of elements with their atomic structure. You can also predict which elements can form ionic or covalent bonds.



Reading

3.1 PERIODIC TABLE

One of the most important activities is the search for order. A large number of observations or objects can be arranged into groups according to common features they share,



it becomes easier to describe them. After the discovery of atomic number by Moseley in 1913, it was noticed that atomic number could serve as a base for systematic arrangement of elements. Thus elements are arranged in the order of increasing atomic number. A table showing systematic arrangement of elements is called periodic table. It is based on the Periodic law that states **if the elements are arranged in the order of their increasing atomic numbers, their properties are repeated in a periodic manner.**

3.1.1 Periods and Groups of Elements.

The most commonly used form of the periodic table is shown in figure 3.1. Note that the elements are listed in order of increasing atomic numbers, from left to right and from top to bottom. Hydrogen (H) is in the top left corner. Helium (He), atomic number 2, is at the top right corner. Lithium (Li), atomic number 3, is at the left end of the second row.

The horizontal rows of the periodic table are called periods. There are varying number of elements in periods. How many periods you find in the periodic table? There are seven periods. The number of elements per period range from 2 in period 1 to 32 in period 6. First three periods are called **short periods** and the remaining periods are called **long periods**. The properties of elements within a period change gradually as you move from left to right in it. But when you move from one period to the next, the pattern of properties within a period repeats. This is in accordance to the periodic law.



Activity 3.1

Look at the periodic table and write number of elements present in the relevant period in the table 3.1

Table 3.1 Number of elements in the periods of the periodic table

Period No.	No. of elements
First	
Second	
Third	
Fourth	
Fifth	
Sixth	
Seventh	



Teacher's Point

Teacher may ask students to memorise elements of 1st two groups and 1st two periods.

Representative Elements (*p* Series)

Representative Elements (*s* Series)

[illegible]



Elements that have similar properties lie in the same column in the periodic table. **Each vertical column of elements in the periodic table is called a group or family.**

Elements with similar valance shell electronic configuration are placed in the same group. Each group is identified by a number and the letter A or B. Group A elements are called **normal** or **representative elements**. They are also called main group elements. Group B elements are called **transition elements**.

Some groups of elements in the periodic table have been given group names. For example metallic elements in Group 1A are called **alkali metals**. Group IIA elements are called **alkaline earth metals**. The elements in Group VIIA are **halogens**. The Group VIIIA elements are called **noble gases** because they do not readily undergo chemical reactions.



Self-Assessment Exercise 3.1

In which period and group the following elements are present in the periodic table. (a) Mg
(b) Ne (c) Si (d) B

Example 3.1: Identifying the group and period of an element

Identify the group and period of ${}_{13}^{27}\text{Al}$, ${}_{5}^9\text{B}$, ${}_{12}^{24}\text{Mg}$ on the basis of electronic configuration.

Problem Solving Strategy:

Write electronic configuration of element. Identify its valence shell. Remember that n value of the valence shell indicates period. Total number of electrons in the valence shells represents group number.

Solution:

$$\text{a) } {}_{13}^{27}\text{Al} = \frac{1s^2}{K} \frac{2s^2, 2p^6}{L} \frac{3s^2, 3p^1}{M} \quad \begin{array}{ccc} K & L & M \\ 1 & 2 & 3 \end{array}$$

Valence shells is M

As $n = 3$, Al is present in the 3rd period. Since total number of electrons in the valence sub-shells are $2+1=3$, it must be present in Group IIIA.

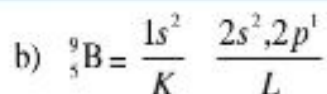
Society, Technology, Science

In 1864, John Newland, an English chemist arranged 24 elements in order of increasing atomic masses. He noticed that every eighth element, starting from any point, has similar properties. Few rows of his arrangement are shown below:

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

His scheme however, failed because many elements were found out of place in his arrangement. For instance Ti does not resemble C and Si, Mn does not resemble N and P and Fe does not resemble O and S. However his arrangement of elements in order of increasing atomic masses formed basis for later classification of elements.

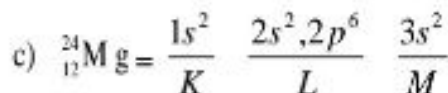
In 1869, Mendeleev, a Russian chemist developed a classification scheme of elements. He recognized that if elements were placed in order of increasing atomic masses, the properties of elements repeated at regular intervals. He arranged 65 elements in periods and groups. Development of the periodic table nicely explains how scientist can build on one another's work.



K	L
1	2

Valence shell is L

So $n = 2$, B is present in the 2nd period. Since total number of electrons in the valence shell are $2+1=3$, it must be present in Group IIIA.



K	L	M
1	2	3

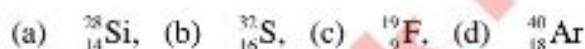
Valence shell is M

So $n = 3$, Mg is present in the 3rd period. Since total number of electrons in the valence shell are 2, it must be present in Group IIA.



Self-Assessment Exercise 3.2

Identify the group and period of the following elements on the basis of electronic configurations.



Example 3.2: Classifying or dividing elements into groups and periods

Electronic configuration of atoms of some elements are given below. Classify them in groups and periods.

- A. $1s^2 2s^2$
- B. $1s^2 2s^2 2p^3$
- C. $1s^2 2s^2 2p^5$
- D. $1s^2 2s^2 2p^6 3s^2$
- E. $1s^2 2s^2 2p^6 3s^2 3p^5$
- F. $1s^2 2s^2 2p^6 3s^2 3p^3$

Problem solving Strategy:

Remember that:

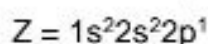
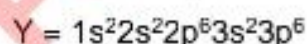
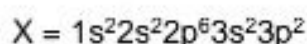
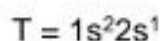
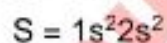
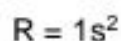
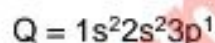
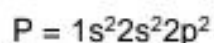
1. The elements whose atoms have similar valence shell electronic configuration belong to the same group.
2. The n value of the valence shell indicates period.
3. The elements whose atoms have same value of n for the valence shell lie in the same period.

**Solution:**

		IIA	VA	VIIA
Period	2	A	B	C
Period		$2s^2$	$2s^22p^3$	$2s^22p^5$
	3	D	F	E
		$3s^2$	$3s^23p^3$	$3s^23p^5$

**Self-Assessment Exercise 3.3**

Electronic configuration of atoms of some elements are given below. Place them into groups and periods.



	IA								VIIIA
		IIA		IIIA	IVA	VA	VIA	VIIA	

3.1.5 s and p Blocks in the Periodic Table

On the basis of outer most valence sub shell, elements in the periodic table can also be classified into four blocks. Elements of Group IA and Group IIA contain their valence electrons in s sub-shell. Therefore, these elements are called **s-block** elements. The elements of Group IIIA to VIIIA (except He) are known as **p – block** elements, because their valence electrons lie in p sub-shell. Figure 3.2 shows blocks in the periodic table



3.1.6 Valence Shell Electronic Configuration and Position of an Element in the Periodic Table

You can determine the valence shell electronic configuration of an element from its position in the periodic table. Period number of element indicates n value of the valence shell. Whereas group number of element indicates the number of electrons in the valence shell.

Example 3.3: Obtaining the valence shell configuration

Write the valence shell electronic configuration of the following elements from their position in the periodic table.

- (a) Phosphorus (b) Neon

Problem Solving Strategy:

Remember that

Period number = n value of valence shell

Group number = number of valence electrons

Distribute the electron in the sub-shells of valence shell.

Solution:

- a) Period number of phosphorus is 3,

As $n = 3$ therefore, valence shell is M

So valence electrons will be present in 3s and 3p sub-shells

The group number is 5, so there are 5 electrons in the valence shell

Two electrons will fill 3s sub-shell and remaining 3p sub-shell. Thus, the valence shell electronic configuration is $3s^2 3p^3$

- b) Period number of Ne is 2. So, $n = 2$ and valence shell is L. Valence electrons will be present in 2s and 2p sub-shells.

Group number for Ne is 8,

This means there are 8 electrons in the valence shell. Two electrons will fill 2s sub-shell and remaining six 2p sub-shell. Thus the valence shell electronic configuration for Ne is $2s^2 2p^6$.

Example 3.4: Obtaining the position of element in the periodic table from electronic configuration

Find out the position of the following elements in the periodic table from the electronic configuration:

- (a) Nitrogen (atomic number: 7) (b) Oxygen (atomic number: 8)



Teacher's Point

A teacher may ask student to draw Valence shells electronic configuration of some common elements.

Period

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1A	2A	3B	4B	5B	6B	7B	8	9	10	11B	12B	13A	14A	15A	16A	17A	18A
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
H 1s ¹	He 1s ²	Li 2s ¹	Be 2s ²	B 2s ² 2p ¹	C 2s ² 2p ²	N 2s ² 2p ³	O 2s ² 2p ⁴	F 2s ² 2p ⁵	Ne 2s ² 2p ⁶	Na 3s ¹	Mg 3s ²	Al 3s ² 3p ¹	Si 3s ² 3p ²	P 3s ² 3p ³	S 3s ² 3p ⁴	Cl 3s ² 3p ⁵	Ar 3s ² 3p ⁶
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K 4s ¹	Ca 4s ²	Sc 3d ¹ 4s ²	Ti 3d ² 4s ²	V 3d ³ 4s ²	Cr 3d ⁵ 4s ¹	Mn 3d ⁵ 4s ²	Fe 3d ⁶ 4s ²	Co 3d ⁷ 4s ²	Ni 3d ⁸ 4s ²	Cu 3d ¹⁰ 4s ¹	Zn 3d ¹⁰ 4s ²	Ga 4s ² 4p ¹	Ge 4s ² 4p ²	As 4s ² 4p ³	Se 4s ² 4p ⁴	Br 4s ² 4p ⁵	Kr 4s ² 4p ⁶
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb 5s ¹	Sr 5s ²	Y 4d ¹ 5s ²	Zr 4d ² 5s ²	Nb 4d ⁴ 5s ¹	Mo 4d ⁵ 5s ¹	Tc 4d ⁵ 5s ²	Ru 4d ⁷ 5s ¹	Rh 4d ⁸ 5s ¹	Pd 4d ¹⁰	Ag 4d ¹⁰ 5s ¹	Cd 4d ¹⁰ 5s ²	In 5s ² 5p ¹	Sn 5s ² 5p ²	Sb 5s ² 5p ³	Te 5s ² 5p ⁴	I 5s ² 5p ⁵	Xe 5s ² 5p ⁶
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs 6s ¹	Ba 6s ²	La* 5d ¹ 6s ²	Hf 5d ² 6s ²	Ta 5d ³ 6s ²	W 5d ⁴ 6s ²	Re 5d ⁵ 6s ²	Os 5d ⁶ 6s ²	Ir 5d ⁷ 6s ²	Pt 5d ⁹ 6s ¹	Au 5d ¹⁰ 6s ¹	Hg 5d ¹⁰ 6s ²	Tl 6s ² 6p ¹	Pb 6s ² 6p ²	Bi 6s ² 6p ³	Po 6s ² 6p ⁴	At 6s ² 6p ⁵	Rn 6s ² 6p ⁶
87	88	89	104	105	106	107	108	109									
Fr 7s ¹	Ra 7s ²	Ac** 6d ¹ 7s ²	Db 6d ² 7s ²	Jl 6d ³ 7s ²	Rf 6d ⁴ 7s ²	Bh 6d ⁵ 7s ²	Hn 6d ⁶ 7s ²	Mt 6d ⁷ 7s ²									

f-Block

*** Lanthanides**

58	59	60	61	62	63	64	65	66	67	68	69	70	71
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
4f ¹ 5d ¹ 6s ²	4f ² 6s ²	4f ³ 6s ²	4f ⁴ 6s ²	4f ⁶ 6s ²	4f ⁷ 6s ²	4f ⁷ 5d ¹ 6s ²	4f ⁹ 6s ²	4f ¹⁰ 6s ²	4f ¹¹ 6s ²	4f ¹² 6s ²	4f ¹³ 6s ²	4f ¹⁴ 6s ²	4f ¹⁴ 5d ¹ 6s ²

**** Actinides**

90	91	92	93	94	95	96	97	98	99	100	101	102	103
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr
6d ² 7s ²	5f ¹ 6d ¹ 7s ²	5f ² 6d ¹ 7s ²	5f ⁴ 6d ¹ 7s ²	5f ⁶ 7s ²	5f ⁷ 7s ²	5f ⁷ 6d ¹ 7s ²	5f ⁹ 7s ²	5f ¹⁰ 7s ²	5f ¹¹ 7s ²	5f ¹² 7s ²	5f ¹³ 7s ²	5f ¹⁴ 7s ²	5f ¹⁴ 6d ¹ 7s ²

Metal

Metalloid

Nonmetal

Figure 3.2: Blocks in the periodic table



Problem Solving Strategy:

Write electronic configuration of the element. Identify the valence shell configuration, coefficient of s or p sub-shell represents period number and total number of electrons in valence shell is equal to the group number.

Solution:

- a) Electronic configuration of N = $1s^2 2s^2 2p^3$
 Valence shell has configuration = $2s^2 2p^3$
 Period number = 2
 Group number = 2 + 3 = 5
 $\therefore$ Nitrogen is present in the 2nd period of Group V-A
- b) Electronic configuration of oxygen = $1s^2 2s^2 2p^4$
 Valence shell has configuration = $2s^2 2p^4$
 So, Period number = 2
 Group number = 2 + 4 = 6
 $\therefore$ Oxygen is present in the 2nd period of Group VI-A



Self-Assessment Exercise 3.4

1. Obtain the valence shell configuration of Al and S from their position in the periodic table.
2. Find out the position of Ne (At. No 10) and Cl (At. No. 17) in the periodic table.

3.1.7 Shape of the Periodic Table

Recall that the horizontal rows in the period table are called periods. How are these periods formed?

Elements are arranged in order of increasing atomic number. First period contains only two elements, H and He. Both these elements have valence electron in K shell. K shell can not have more than two electrons. As K shell is completed at He, so the period also ends at He, Lithium (Li) atomic number 3 has one electron in L shell, so second period begins with Li. Since L shell can accommodate 8 electrons, so eight elements come in the 2nd period. Second period ends at Ne which has eight electron, in L shell ($2s^2 2p^6$).

Next elements Na has valence electron in the third shell (M – shell), in Na valence electron is present, in 3s sub – shell, which has similar electronic configuration as Li ($2s^1$), So it comes under Li. Mg with $3s^2$ valence shell electronic configuration come under Be ($2s^2$), Similarly next six elements Al, Si, P, S Cl and Ar on the bases of similar valence shell electronic configuration come under B, C, N, O, F, and Ne respectively. Ar has $3s^2 3p^6$ valence shell configuration similar to Ne ($2s^2 2p^6$). Next element K has $4s^1$ electronic configuration in the valence shell, which is similar to Na($3s^1$). So K comes under Na and a new period (4th) begins with K. In this way elements having similar valence shell configuration come in the same group. The arrangement of the elements into periods has an important consequence. The elements that have similar properties end up in the same group in the periodic table.



3.2 PERIODICITY OF PROPERTIES

In section 3.1.4 you learned that, the electronic configuration of elements show a periodic variation with the increasing atomic number. Therefore, the elements also show periodic variation in their physical and chemical properties. Elements having similar valence shell electronic configuration have been placed in the same group, one below the other. Chemical properties depend on the valence shell electronic configuration. Because all the elements of a particular group have similar valence shell electronic configuration, they possess similar chemical characteristics. Physical properties depend on the sizes of atoms. Since sizes of atoms change gradually from top to bottom in a group. Therefore, elements show gradation in physical properties in the same group. In a period of periodic table the number of electrons present in the valence shell increase gradually from left to right. Their chemical and physical properties also show variation in the same manner. In this section you will learn variation in some of the physical properties of elements in a group and across a period.

3.2.1 Shielding Effect

Figure 3.2 shows electronic configuration of Li, Be and Mg.

Which atom has more shells, Be or Mg? Which atom has more electrons between the nucleus and the valence electrons, Be or Mg?

Electrons present in the inner shells cut off attractive force between the nucleus and the valence electrons.

The reduction in force of attraction between nucleus and the valence electrons by the electrons present in the inner sub-shells is called shielding effect.

Which atom has greater shielding effect, Be or Mg?

As you move from top to bottom in a group the number of electronic shells increase. So the number of electrons in the inner shell also increase. As a result shielding effect increases.

Which atom, Li or Be has greater number of shells? Which atom, Li or Be has greater number of electrons between nucleus and valence electrons?

As you move from left to right in a period the number of electrons in the inner shells remains constant, therefore, shielding effect remains constant.

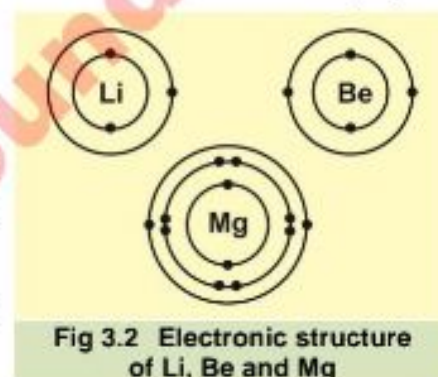


Fig 3.2 Electronic structure of Li, Be and Mg

Example 3.5: Identifying the element whose atoms have greater shielding effect, using periodic table

Choose the elements whose atoms you expect to have greater shielding effect.

- (a) Be or Mg (b) C or Si



Teacher's Point

A teacher may ask student to draw shell of first four elements group A to show shielding effect.

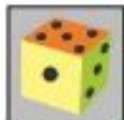


Problem Solving Strategy:

Look at the periodic table and find the relative position of given elements in the periodic table. Apply the trend of increasing shielding effect in a group.

Solution:

- (a) Mg atoms will have greater shielding effect.
- (b) Si atoms will have greater shielding effect.



Self-Assessment Exercise 3.5

Choose the element whose atoms you expect to have smaller shielding effect.

- (a) F or Cl
- (b) Li or Na
- (c) B or Al

All the physical and chemical properties of the elements depend on the electronic configurations of their atoms. Now we will discuss four properties of atoms that are influenced by the electronic configuration: atomic size, ionization energy, electron affinity and electronegativity. These properties are periodic. They generally increase and decrease in a recurring or repeating manner through the periodic table. This means they show consistent changes or trends, within a group or a period. These trends correlate with the behaviour.

3.2.2 Atomic Size

The size of an atom depends on its electronic configuration. **The size of an atom is the average distance between the nucleus of an atom and the outer electronic shell.** Figure 3.3 shows atomic radii of the main group elements.

Figure 3.3 shows the variation in atomic radii in a period and within a group. You can see two general trends in atomic radii.

- (1) The atomic radius decreases in any given period as you move across the period. This is because as you move from one element to the next on its right in a period. Another electron is added to the same valence shell. At the same time positive charge on the nucleus also increases by 1. The attractive force of the nucleus for the valence shell electron increases. Therefore, the shell size and atomic radius decreases. For example, in going from lithium to beryllium, atomic size decreases. This you can understand from the valence shell electronic configuration of Li ($2s^1$) and B ($2s^2$). In going from Li to Be, there is no change in the shell number n , but atomic number increases from 3 to 4. Due to this the force of the nucleus for the valence shell electron increases. Therefore, atomic radius decreases.

The atomic radius increases in any given main group as you move down the group of elements. This is because the size of an atom is determined by the size of its valence shell. As you move to the next lower element in the group, the atom has an additional shell of electrons. This increases atomic radius. For example, in going from Li to Na atomic radius increases.



Period	I-A		II-A		III-A	IV-A	V-A	VI-A	VII-A	VIII-A
	1	H 37 •								He 31 •
	2	Li 152 ●	Be 112 ●		B 85 ●	C 77 ●	N 75 ●	O 73 ●	F 72 ●	Ne 71 ●
	3	Na 186 ●	Mg 160 ●		Al 143 ●	Si 118 ●	P 110 ●	S 103 ●	Cl 100 ●	Ar 98 ●
	4	K 227 ●	Ca 197 ●		Ga 135 ●	Ge 122 ●	As 120 ●	Se 119 ●	Br 114 ●	Kr 112 ●
	5	Rb 248 ●	Sr 215 ●		In 167 ●	Sn 140 ●	Sb 140 ●	Te 142 ●	I 133 ●	Xe 131 ●
	6	Cs 265 ●	Ba 222 ●		Tl 170 ●	Pb 146 ●	Bi 150 ●	Po 168 ●	At (140) ●	Rn (140) ●
	7	Fr (270) ●	Ra (220) ●							

Figure 3.3: Atomic radii of the main group elements (in picometer)

Consider electronic configuration of Li ($1s^2 2s^1$) and Na ($1s^2, 2s^2, 2p^6, 3s^1$). A new electronic shell has been added that increases atomic size.

Example 3.6: Identifying the element that has greater atomic radius

Choose the element whose atom you expect to have larger atomic radius in each of the following pairs.

- (a) Mg, Al (b) C, Si

Problem Solving Strategy:

Remember that the larger atom in any:

- Period lies further to the left in the periodic table.
- Group lies closer to the bottom in the periodic table.
- Check the periodic table and choose the element.

**Solution:**

- (a) The larger atom is Mg
(b) The larger atom is Si

**Self-Assessment Exercise 3.6**

Using the periodic table but without looking at the figure 3.3, choose the element whose atom you expect to have smaller atomic radius in each of the following pairs.

- (a) O or S (b) O or F

3.2.3 Ionization Energy

	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	H 1312							He 2372
2	Li 520	Be 899	B 801	C 1086	N 1402	O 1314	F 1681	Ne 2081
3	Na 496	Mg 738	Al 578	Si 786	P 1012	S 1000	Cl 1251	Ar 1521
4	K 419	Ca 590	Ga 579	Ge 762	As 947	Se 941	Br 1140	Kr 1351
5	Rb 403	Sr 549	In 558	Sn 709	Sb 834	Te 869	I 1008	Xe 1170
6	Cs 376	Ba 503	Tl 589	Pb 716	Bi 703	Po 812	At 926	Rn 1037

Figure 3.4 Ionization energies of the main group elements

You have learned in section 1.3.1 how cations are formed. Ionization energy is an important property of atoms that explains cation formation. **“Ionization energy is defined as the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom”.**



Ionization energy is a measure of the extent to which the nucleus attracts the outermost electron. A high value of ionization energy means stronger attraction between the nucleus and the outermost electron. Whereas a low

ionization energy indicates a weaker force of attraction between the nucleus and the outermost electron. Figure 3.4 shows the ionization energies of the main group elements. Values are given in units of kJ/mole^{-1} or kJ/mole .

Trends in the values of ionization energies.

The ionization energy value decreases from top to bottom in a group. This is because the shielding effect in atoms increases as you descend. Greater shielding effects results in a weaker attraction of the nucleus for the valence electrons. So, they are easier to remove. This leads to decrease in ionization energy from top to bottom in a group.

Which atom has greater shielding effect, Li or Na?

As you move from left to right in a period, the shielding effect remains constant. But progressively nuclear charge increases. A stronger force of attraction between nucleus and the



valence electron increases. This leads to increase in ionization energy from left to right in a period.

Which atom has higher ionization energy, Li or Be?

Example 3.7: Identifying the element that has smaller ionization energy

Choose the element whose atom you expect to have smaller ionization energy in each of the following pairs.

- (a) B, C (b) N, P

Problem Solving Strategy

Remember that ionization energy:

- (a) Increases across a period. The element that has smaller ionization energy will be further to the left in the periodic table.
- (b) Decreases from top to bottom in a group. The element that has smaller ionization energy will correspond to the element closer to the bottom.
- (c) Check the periodic table to choose the element.

Solution:

- (a) The atom with the smaller ionization energy is B
- (b) The atom with the smaller ionization energy is P.



Self-Assessment Exercise 3.7

Which atom has the smaller ionization energy?

- (a) B or N (b) Be or Mg (c) C or Si

3.2.4 Electron Affinity

Electron affinity explains the anion formation. **Electron affinity is defined as the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uninegative gaseous ion.**

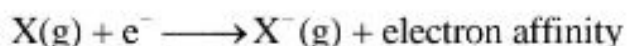


Figure 3.5 shows electron affinities of main group elements.

As you move from left to right across a period, the electron affinity generally increases. This is due to increase in nuclear charge and decrease in atomic radius, which binds the extra electron more tightly to the nucleus. But shielding effect remains constant in each period. Therefore, alkali metals have lowest and halogens have the highest electron affinities in each period.



The electron affinity decreases from top to bottom in a group. This is due to increase in shielding effect. Due to increase in shielding effect added electron binds less tightly to the nucleus. As a result less energy is released.

H -73								He 0
Li -60	Be 0		B -27	C -122	N +7	O -141	F -328	Ne 0
Na -53	Mg 0		Al -44	Si -134	P -71.7	S -200	Cl -349	Ar 0
K -48	Ca 0		Ga -29	Ge -120	As -77	Se -195	Br -325	Kr 0
Rb -47	Sr 0		In -29	Sn -121	Sb -101	Te -190	I -295	Xe 0
Cs -45	Ba 0		Tl -30	Pb -110	Bi -110	Po -180	At -270	Rn 0

Figure 3.5 electron affinities of main group elements

There are several exceptions to the general trend of election affinity values. You will learn reasons for it in grade XI.

3.2.5 Electronegativity

Electronegativity is the ability of an atom to attract the electrons towards itself in a chemical bond. Figure 3.6 shows as scale of electronegativities of the elements devised by Linus Pauling. The American chemist Linus Pauling devised a method for calculating relative electronegativities of elements.

H 2.1						He	
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne 2.1
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar 3.0
K 0.5	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 2.1
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9						

Figure 3.6 the electronegativities of elements.

**Activity 3.3****Determining the general trends in the electronegativities****You will need:**

- Figure 3.6

Carry out the following:

1. Move across the second period from left to right and note down the variation in electronegativity values.
2. Move across the 3rd period from left to right and note down the variation in electronegativity values.
3. Make generalization about the variation in electronegativities across a period and write reason.
4. Move from top to bottom in Groups IA and IIA and note down the variation in electronegativities value.
5. Move from top to bottom in Groups VIA and VIIA and note down the variation in electronegativities value.
6. Make generalization about the trend in electronegativity values in a group. Give reason.

**Key Points**

- When elements are arranged in the order of their increasing atomic number, their properties are repeated in a periodic manner.
- A horizontal row of elements in the periodic table is called a period.
- A column of elements in the periodic table is called a group or a family.
- Group IA and IIA elements are called s-block elements, since s sub-shell fills in these elements.
- Elements in group IIIA to VIIIA are called p-block elements, because filling of valence p sub-shell occurs in these elements.
- The length of a period in the periodic table depends on the type of sub-shell that fills.
- The decrease in force of attraction between nucleus and the valence electron by the electrons present in the inner sub-shells is called shielding effect.
- The size of atom is the average distance between the nucleus of an atom and the outer electronic shell.
- The atomic radii decrease from left to right in a period. Whereas these increase from top to bottom in a group.
- Ionization energy is the minimum amount of energy required to remove the outermost electron from an isolated gaseous atom.



- Electron affinity is the amount of energy released when an electron adds up in the valence shell of an isolated atom to form a uninegative gaseous ion.

REFERENCES FOR ADDITIONAL INFORMATION

- B.Earl and LDR Wilford, Introduction to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).
- Lawarie Ryan, Chemistry for you.



Review Questions

1. Encircle the correct answer:

- Number of periods in the periodic table are:
(a) 8 (b) 7 (c) 16 (d) 5
- Which of the following groups contain alkaline earth metals?
(a) 1A (b) IIA (c) VIIA (d) VIIIA
- Which of the following elements belongs to VIIIA?
(a) Na (b) Mg (c) Br (d) Xe
- Main group elements are arranged in _____ groups.
(a) 6 (b) 7 (c) 8 (d) 10
- Period number of $^{27}_{13}\text{Al}$ is:
(a) 1 (b) 2 (c) 3 (d) 4
- Valence shell electronic configuration of an element M (atomic no. 14) is:
(a) $2s^2 2p^1$ (b) $2s^2 2p^2$ (c) $2s^2 2p^3$ (d) $3s^2 3p^2$
- Which of the following elements you expect to have greater shielding effect?
(a) Li (b) Na (c) K (d) Rb
- As you move from right to left across a period, which of the following does not increase:
(a) electron affinity (b) ionization energy
(c) nuclear charge (d) shielding effect
- All the elements of Group IIA are less reactive than alkali metals. This is because these elements have:
(a) high ionization energies (b) relatively greater atomic sizes
(c) similar electronic configuration (d) decreased nuclear charge

2. Give short answers

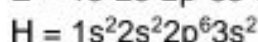
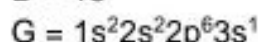
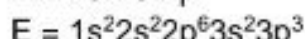
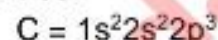
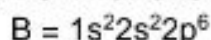
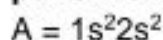
- Write the valence shell electronic configuration of an element present in the 3rd period and Group IIIA.
- Write two ways in which isotopes of an element differ.
- Which atom has higher shielding effect, Li or Na?
- Explain why, Na has higher ionization energy than K?
- Alkali metals belong to S-block in the periodic table, why?



3. Arrange the elements in each of the following groups in order of increasing ionization energy:
 (a) Li, Na, K (b) Cl, Br, I
4. Arrange the elements in each of the following in order of decreasing shielding effect.
 (a) Li, Na, K (b) Cl, Br, I (c) Cl, Br
5. Specify which of the following elements you would expect to have the greatest electron affinity.

S, P, Cl

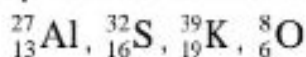
6. Electronic configuration of some elements are given below, group the elements in pairs that would represent similar chemical properties.



7. Arrange the elements in groups and periods in Q. No. 6.

IA		IIA		IIIA	IVA	VA	VIA	VIIA	VIIIA

8. For normal elements, the number of valence electrons of an element is equal to the group number. Find the group number of the following elements.



9. Write the valence shell electronic configuration for the following groups:

- Alkali metals
- Alkaline earth metals
- Halogens
- Noble gases

10. Write electron dot symbols for an atom of the following elements

- (a) Be (b) K (c) N (d) I

11. Write the valence shell electronic configuration of the atoms of the following elements.

- An element present in period 3 of Group VA
- An element present in period 2 of Group VIA



12. Copy and complete the following table:

Atomic number	Mass number	No. of protons	No. of neutrons	No. of electrons
11			12	
		14	15	
	47		25	
	27			13

13. Imagine you are standing on the top of Neon-20 nucleus. How many kinds of sub-atomic particles you would see looking down into the nucleus and those you would see looking out from the nucleus.
14. Chlorine is a reactive element used to disinfect swimming pools. It is made up of two isotopes Cl-35 and Cl-37. Because Cl-35 is more than Cl-37, the atomic mass of chlorine is 35.5amu. is closer to 35 than 37. Write electronic configuration of each isotope of chlorine. Also write symbol for these isotopes (atomic number for chlorine is 17).
15. In which block, group and period in the periodic table where would you place each of the following elements with the following electronic configurations?
 (a) $1s^2 2s^1$ (b) $1s^2 2s^2 2p^5$ (c) $1s^2 2s^2 2p^6 3s^2$ (d) $1s^2$



Think-Tank

- What types of elements have the highest ionization energies and what types of elements have the lowest ionization energies. Argue.
- Two atoms have electronic configuration $1s^2 2s^2 2p^6$ and $1s^2 2s^2 2p^6 3s^1$. The ionization energy of one is 2080kJ/mole and that of the other is 496kJ/mole. Match each ionization energy with one of the given electronic configuration. Give reason for your choice.
- Use the second member of each group from Group IA, IIA and VIIA to judge that the number of valence electron in an atom of the element is the same as its group number.
- Letter A, B, C, D, E, F indicates elements in the following figure:

						C		
A				B				
	D				E			
							F	



- a. Which elements are in the same periods?
 - b. Write valence shell electronic configuration of element D.
 - c. Which elements are metals?
 - d. Which element can lose two electrons?
 - e. In which group E is present?
 - f. Which of the element is halogen?
 - g. Which element will form dipositive cation?
 - h. Write electronic configuration of element E
 - i. Which two elements can form ionic bond?
 - j. Can element C form C_2 molecule? Interpret.
 - k. Which element can form covalent bonds?
 - l. Is element F a metal or non-metal?
5. Electronic configurations of four elements are given below:
(a) $1s^2 2s^1$ (b) $1s^2 2s^2 2p^5$ (c) $1s^2 2s^2 2p^6 3s^2$ (d) $1s^2$
- Which of these elements is
- i) An alkali metal
 - ii) An alkaline earth metal
 - iii) A noble gas
 - iv) A halogen
6. Argue in what region of the periodic table you will find elements with relatively
- a) high ionization energies
 - b) low ionization energies



4

STRUCTURE OF MOLECULES



This is a 20 days lesson

After completing this lesson, you will be able to:

- Find the number of valence electrons in an atom using the Periodic Table
- Describe the importance of noble gas electronic configurations.
- Describe the formation of cations from an atom of a metallic element. Describe the formation of anions from an atom of a non-metallic element.
- Describe the ways in which bonds may be formed.
- Describe the formation of a covalent bond between two nonmetallic elements.
- Describe with examples single, double and triple covalent bonds.
- Draw electron cross and dot structures for simple covalent molecules containing single, double and triple covalent bonds.
- Explain how element attain stability.
- Recognize a compound as having ionic bonds. Identify characteristic of ionic compounds.
- State the octet and duplet rules
- State the importance of the noble gas electronic configurations in the formation of ion.



Pre- Reading

All the matter in this world is composed of almost entirely compounds and their mixtures. Human, animal and plant bodies, rocks, soil, petroleum, coal etc. are all complex mixtures of compounds. In compounds different kinds of atom are bounded together. Few elements also consist of unbounded atoms. For instance helium, neon, argon, xenon and krypton present in the atmosphere consist of unbounded atoms. The manner in which various atoms are bonded together has a profound effect on the properties of substances.

Some substances are hard and tough, others are soft and

flexible why? Resins are widely used to paint dams, bridges, buildings and automobiles. What makes them sticky? How do adhesives such as glue bind two surfaces together? What is the nature of such linkages? The answer lies in the nature of bonding and structure of their molecules. Therefore, to understand the behaviour of various substances, you must understand the nature of chemical bonding and structure of molecules.



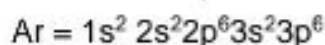
Reading

4.1 WHY DO ATOMS REACT?

In chapter 3, you have learned the arrangement of elements on the periodic table. You noticed that each period on the periodic table starts with alkali metal (except period I that starts



with hydrogen) and ends at a noble gas. The noble gases have ns^2np^6 electronic configuration in the outer most shell. These elements are sometimes called the inert gases. This is because they do not participate in chemical reactions. Electronic configurations of first three noble gases are shown below:



Note that these elements have completely filled outer most s and p sub-shells. Helium contains two electrons and remaining noble gases contain 8 electrons in the valence shell. Because of these configuration noble gases are stable and not active. In 1916 a chemist G.N. Lewis used this fact to explain why atoms undergo chemical reactions. He called his explanations as Octet Rule. An octet is a set of eight. In forming compounds, atoms tend to gain electronic configuration of a noble gas. Remember that each noble gas (except He) has eight electrons configuration in the valence shell. Thus the octet rule takes its name from this fact about noble gases.

The tendency of atoms to acquire eight electron configuration in their valence shell, when bonding, is called octet rule.

Helium has two electrons in its valence shell and is also chemically inert. Some elements that are close to He on the periodic table tend to achieve two electron configuration in their valence shell. For example hydrogen, lithium and beryllium etc tend to achieve two electron configuration in the valence shell.

The tendency of some atoms to acquire two electron configuration in their valence shell, when bonding, is called duplet rule.

Example 4.1: Obtaining the number of valence electrons in an atom using the periodic table.

Find the number of valence electrons in the following atoms using the periodic table.

- (a) Carbon (b) Magnesium (c) Phosphorus

Problem Solving Strategy:

Remember that the group number of main group elements indicates the number of valence electrons in an atom. Check the group number of the elements in the periodic table and find the number of valence electrons.

Solution:

- (a) Carbon belongs to Group IVA, so it contains four electrons in the valence shell.
(b) Magnesium belongs to Group IIA, so it contains two electrons in the valence shell.
(c) Phosphorus is present in Group VA, so it has five electrons in the valence shell.

**Self-Assessment Exercise 4.1**

Find the number of electrons in valence shell of the following atoms using the periodic table. (a) Silicon (b) Sulphur (c) Bromine (d) Argon (e) Potassium (f) Nitrogen

**Reading**

4.2. CHEMICAL BONDS

Atoms combine to form various types of substances. But what holds them together? Fundamentally, some forces of attraction hold atoms together in substances. These forces are called chemical bonds. Basically the forces of attraction that lead to chemical bonding between atoms are electrical in nature. Electronic structure of an atom helps us to understand how atoms are held together to form substances. Atoms other than the noble gases have a tendency to react with other elements. These elements are reactive because they tend to gain stability by losing or gaining electrons. When atoms gain or lose electron they acquire the configuration of next noble gas element.

Atoms can also acquire the configuration of next noble gas element by sharing electrons.

4.3 TYPES OF BONDS

Depending on the tendency of an atom to lose or gain or share electrons, there are two types of bonds:

1. Ionic bonds
2. Covalent bonds

4.3.1 Ionic Bonds

Ionic bonds are formed between two atoms, when one atom loses electron to form cation and the other atom gains this electron to form anion.

Example 4.2: Describing the formation of cations.

Describe the formation of Na^+ and Mg^{+2} cations.

Problem Solving Strategy:

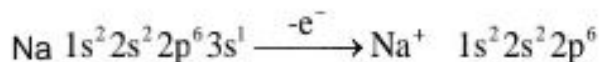
1. Sodium belongs to Group IA on the periodic table. It has only one electron in the valence shell. Sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.
2. Magnesium belongs to Group IIA in the periodic table. It has two valence electrons. Magnesium atom loses these electrons to achieve noble gas configuration. Represent



this by drawing the complete electronic configuration or using an electron dot structure. This number also corresponds to the Group number in the periodic table.

Solution:

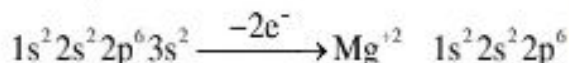
(a) Formation of Na^+ ion



You can also represent this by following electron dot structure,

(b) Formation of Mg^{+2} ion

Mg



You can also represent this by electron dot structure,



Self-Assessment Exercise 4.2

- Describe the formation of cations for the following metal atoms:
 - Li(atomic no 3)
 - Al(atomic no. 13)
- Represent the formation of cations for the following metal atoms using electron dot structures.
 - K
 - Ca

Example 4.3: Describing the formation of anions.

Describe the formation of anions for the following non-metal atoms:

- Oxygen(atomic no.8)
- Fluorine (atomic no. 9)

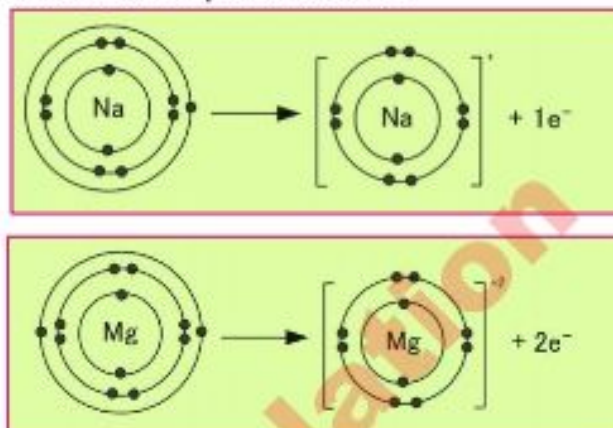
Problem Solving Strategy:

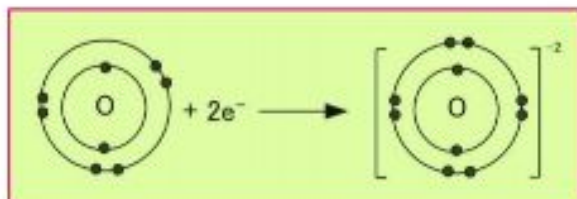
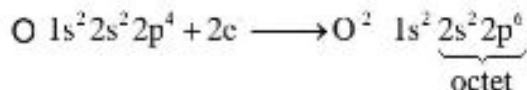
- Write electronic configuration or dot structure.
- Find the number of electrons needed to acquire eight electron configuration.
- Represent addition of electrons.

Solution:

(a) Formation of anion by oxygen atom.

Oxygen belongs to Group VIA on the periodic table. So it has six electrons in its valence shell. It needs two electrons to achieve noble gas configuration.

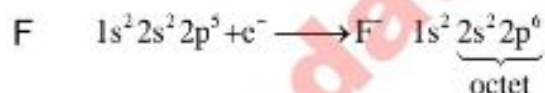
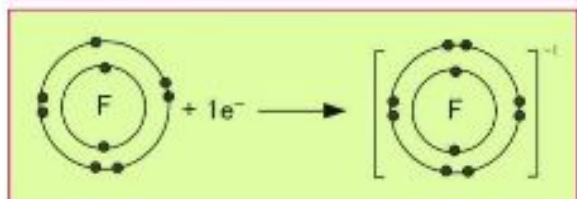




You can also represent this by electron dot structure,

(b) Formation of anion by fluorine atom

Fluorine belongs to Group VIIA on the periodic table. So it has seven electrons in the valence shell. A fluorine atom therefore, requires only one electron to complete octet.



You can also represent this by electron dot structure,



Self-Assessment Exercise 4.3

- Describe the formation of anions by the following non-metals.
 - Sulphur (atomic No. 16)
 - Chlorine (atomic No. 17)
- Represent the formation of anions by the following non-metals using electron dot structures.
 - N
 - P
 - Br
 - H

Anions and cations have opposite charges. They attract one another by electrostatic forces. **"The forces of attraction that bind oppositely charged ions are called ionic bonds"**. Compounds that consist of ions joined by electrostatic forces are called ionic compounds. The total positive charge of the cations must be equal to the total negative charge of the anions. This is because ionic compounds as a whole are electrically neutral.

Example 4.4: Representing ionic bond formation.

For each of the following pairs of atoms, use electron dot & electron cross structures to write the equation for the formation of ionic compound.

- Na and Cl
- Mg and F

Problem Solving Strategy:

- The metal atoms form cations and non-metal atoms form anions.
- The number of electrons lost by metal atoms of group IA, IIA and IIIA equals the group number.

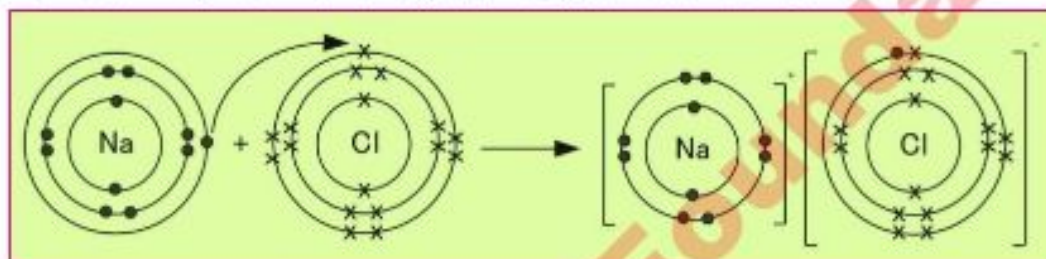


3. To write the final form of the equation, you need to know the simplest ratio of cations to anions that you require for the neutral compound.
4. Write equation using electron dot & electron cross structures.

Solution:

(a) Na is metal and Cl is non-metal.

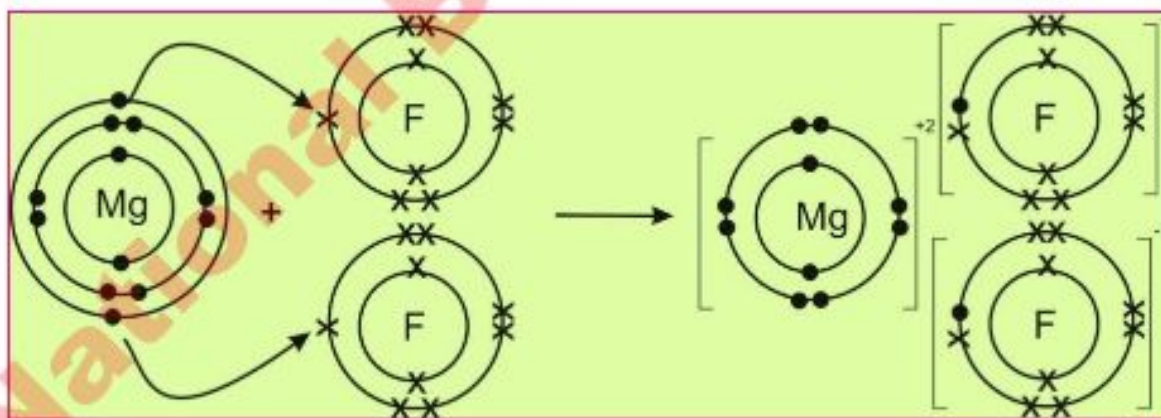
Metal atom tends to lose electrons and non – metal atoms tends to gain electrons to acquire electronic configuration of nearest noble gas. Since Na atom has one electron in the outer most shell. It losses one electron to form Na^+ ion. Since Cl atom has seven electrons in outermost shell, it needs one electron to complete octet. So it gains one electron to form Cl^- ion. For every Na^+ ion, you need one Cl^- ion.



(b) Mg is metal and F is non-metal.

Mg atom has two electrons in the outermost shell. It losses two electrons to form Mg^{+2} ion. Since F atom has seven electrons in the outermost shell, so it gains one electron to form F^- ion.

For every Mg^{+2} ion you need two F^- ions.

**Self-Assessment Exercise 4.4**

For each of the following pairs of atoms, use electron dot and electron cross structures to write the equation for the formation of ionic compound.

- (a) Mg and O (b) Al and Cl

**Example 4.5: Recognizing a compound as having ionic bonds.**

Recognize the following compounds as having ionic bonds.

(a) MgO (b) NaF

Problem Solving Strategy:

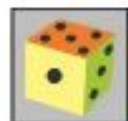
1. The metal atom loses electrons to form cations and non-metal atom gains electrons to form anions.
2. The number of electrons lost by metal atoms of group IA, IIA and IIIA equals the group number. The number of electrons gained by the non-metal atoms is equal to 8 minus group number.
3. Find the simplest ratio of cations to anions, to identify the compound.

Solution:

(a) MgO

Mg is metal and O is non-metal. Mg atom has two electrons in outermost shell. So it loses two electrons to form Mg^{+2} ion. Since O atom has six electrons in outermost shell, so it gains two electrons to form O^{-2} ion. In this way both the atoms acquire nearest noble gas configuration. For every Mg^{+2} ion you need one O^{-2} ion. Chemical formula of resulting compound is MgO. Therefore MgO is an ionic compound.

(b) Na is metal and F is non-metal. Na atom has one electron in outmost shell. So it loses one electron to form Na^{+} ion. Since F atom has seven electrons in outermost shell, so it gains one electron to form F^{-} ion. Na atom by losing one electron and F atom by gaining one electron acquire nearest noble gas electronic configuration. You need one F^{-} ion for each Na^{+} ion. Therefore, NaF is an ionic compound.

**Self-Assessment Exercise 4.5**

Recognize the following compounds as having ionic bonds:

(a) KCl (b) AlCl_3 (c) MgF_2

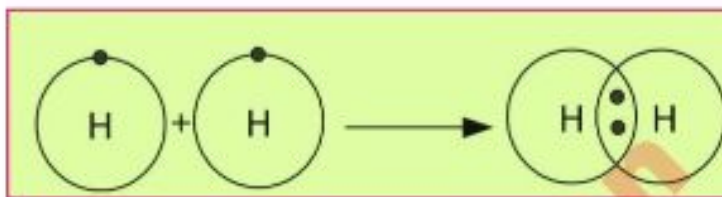
**Reading****4.3.2 Covalent Bonds**

In the preceding section you have learned the formation of ionic bonds between two atoms. Ionic compounds such as NaCl, are crystalline solids with high melting points. However, some compounds have very different properties. HCl is a gas at room temperature. Water (H_2O) is a liquid. Such compounds are not ionic. These compounds are made up of non-metal atoms. Non-metals have high ionization energies, therefore, they do not lose electrons. Non-metal



atoms tend to share electrons among themselves or with other non-metal atoms to form a chemical bond called covalent bond. **A Covalent bond is formed by mutual sharing of electrons between two atoms.**

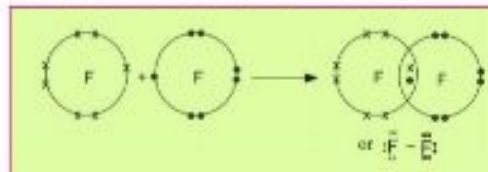
Consider the formation of covalent bond in hydrogen molecule. A hydrogen atom has a single valence electron. Two hydrogen atoms share their valence electrons to form a diatomic molecule.



In the formation of this molecule, each hydrogen atom achieves the electron configuration of the noble gas, helium which has two valence electrons. An electron pair in the region between the two atoms is attracted to both hydrogen nuclei. This means it is a more stable situation than that exists in separate atoms. Because of this stability two atoms form a covalent bond.

We can represent the formation of a covalent bond between two atoms using electron-dot and electron-cross symbols for the atoms and the resulting molecule. As already discussed that valence electrons are represented by dots. Just to understand sharing, we represent valence electrons in one atom by dots and in the other atom by crosses. However, remember that all the electrons are identical and cannot be differentiated. A shared pair of electrons is also represented by a dash (-) in a molecule.

Consider the formation of a bond between two fluorine atoms. Fluorine belongs to Group VIIA, so it has seven electrons in the valence shell. It needs one more electron to attain the electron configuration of a noble gas. Thus two F-atoms share an electron pair and achieve electron configuration of Ne. For sharing each F-atom contributes one electron to complete the octet.

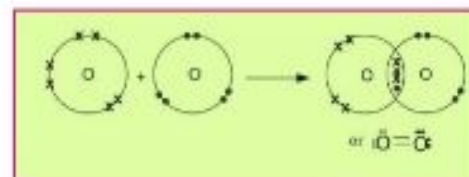


Pairs of valence electrons which are not shared between atoms are called unshared pairs or lone pairs.

Covalent bond that is formed by the sharing of one electron pair is called single covalent bond. So H_2 and F_2 molecules contain a single covalent bond.

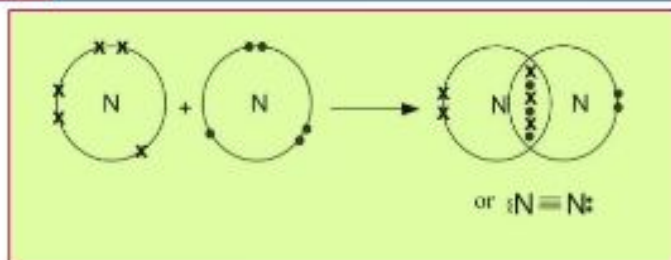
Can you explain the formation of covalent bond between H-atom and a F-atom?

Sometimes atoms may share two or three electron pairs to complete octet. **Double covalent bonds are the bonds that are formed by sharing of two electron pairs.** **Triple covalent bonds are the bonds that involve three shared pairs of electrons.**



Teacher's Point

A teacher may show bonds to students by balls and springs.



Consider the formation of O_2 molecule. Oxygen is in Group VI A, so it has 6 electrons in the valence shell. It needs two electrons to complete its octet. So for sharing each O-atom contributes two electrons.

Can you explain the formation of N_2 molecule?

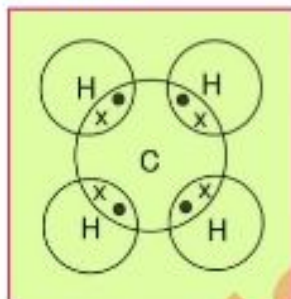
Example 4.6: Drawing electron cross and dot structures for simple covalent molecules containing single covalent bonds

Draw electron cross and dot structures for (a) CH_4 that is a major component of natural gas (b) H_2O that covers about 80% of the earth crust.

Problem Solving Strategy:

1. Decide from the chemical formula which atom is the central atom. An atom that contributes more electrons for sharing is the central atom. Show its valence electrons by dots. Note the number of electrons it needs to complete octet. If the number of electrons needed equals the other atoms, each atom will form a single covalent bond.
2. Arrange other atoms around the central atom. Connect the central atom by single bonds. Use cross to represent electrons of the other atoms.
3. Check whether the arrangement of electron satisfies the octet rule.

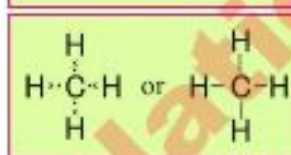
Solution:



(a) CH_4

(i) C has four electrons in the valence shell and needs four electrons to complete its octet. H has only one valence electron and needs one electron to complete duplet. So C can form four single bonds with four H-atoms. C is the central element.

(ii) Connect the atoms with a dot and a cross

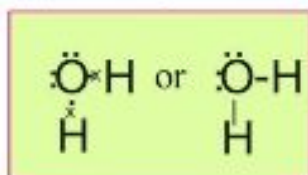
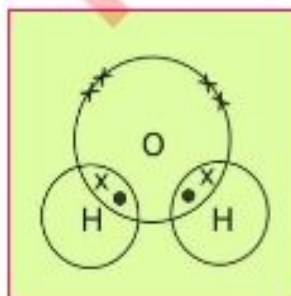


H_2O

(i) O has six valence electrons $\ddot{O}:$ and each hydrogen atom has one valence electron $\overset{x}{H}$. So O-atom needs two electrons to complete octet. Each H needs one electron to complete duplet.

(ii) O is central atom and will form two single bonds with H-atoms.

(iii) Arrange H-atoms around O and connect them by a pair of electrons (one dot and one cross)



**Self-Assessment Exercise 4.6**

Draw electron cross and dot structures for the following molecules:

- NH_3 , that is used to manufacture urea.
- CCl_4 , a dry cleansing agent.
- SiCl_4 , used to make smoke screens.
- H_2S , a poisonous gas.

Example 4.7: Drawing electron cross and dot structures for molecules containing multiple bonds

Draw electron cross and dot structures for the following molecules:

- CO_2 , a component of air and is responsible for green house effect.
- HCN , used as insecticide.

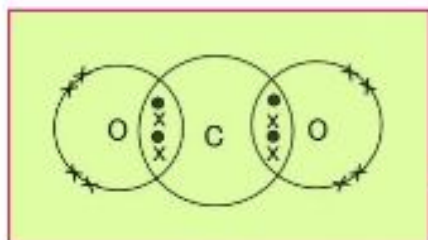
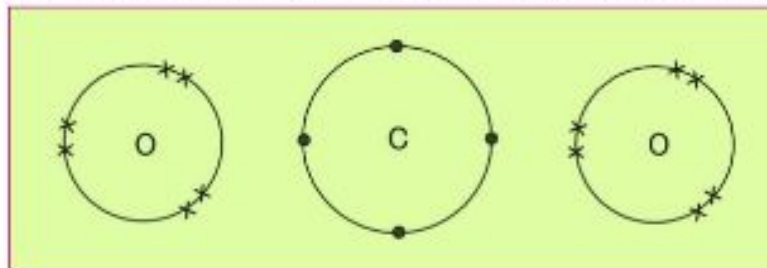
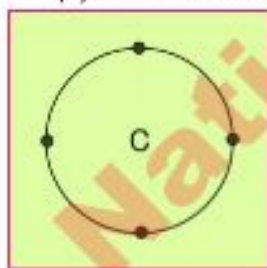
Problem Solving Strategy:

- Decide from the formula which atom is to be in the center. Show its valence electrons by dots. Note the number of electrons it needs to complete octet.
- Show valence electron of the other atoms by cross and find the number of electrons each of the atoms needs to complete octet or duplet.
- Connect central atom with the other atoms by electron pair or pairs to satisfy the octet rule.

Solution:

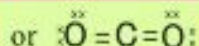
(a) CO_2

- C has four electrons in the valence shell. It needs four electrons to complete octet.
- Each oxygen atom has six valence electrons and needs two electrons to have an octet.



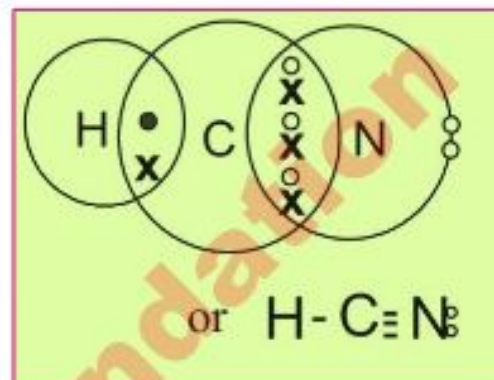
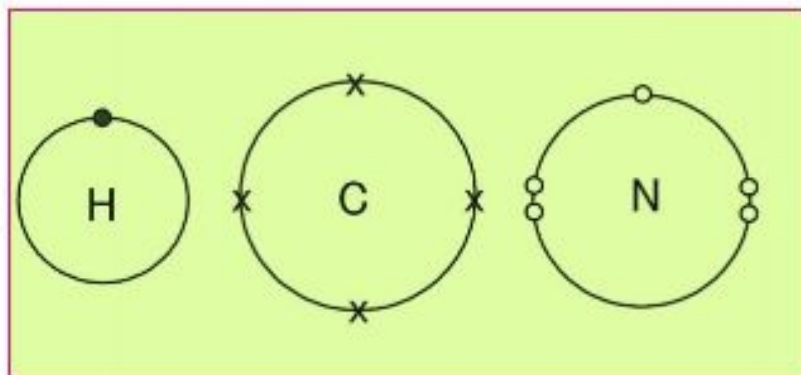
(iii) C is central atom, arrange O-atoms around it.

(iv) Since C needs four electrons and there are only two oxygen atoms. So it will share its two electrons with each oxygen atom.





- (b) HCN
- (i) H has one, C has four and N has five electrons.
- (ii) C needs four and N needs three electrons. So C shares one electron with H to form a single bond and three electrons with N to form a triple bond. This will satisfy octet rule.



Self-Assessment Exercise 4.7

Draw electron cross and electron dot structures for the following molecules:

- (a) CS_2 , an organic solvent that dissolves sulphur, phosphorus etc
- (b) N_2 , a component of air.

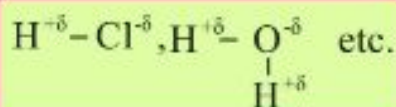


Reading

4.4 INTERMOLECULAR FORCES

A covalent bond can occur between two similar atoms such as in H_2 , N_2 , O_2 , Cl_2 , etc. It can also occur between two unlike atoms, such as in HCl , H_2O , NH_3 , HCN , CO_2 , etc.

When two identical atoms share electron pairs, both the atoms exert same force on the shared electron pairs. Such a covalent bond is called non-polar covalent bond. For example, bond in $\text{H}-\text{H}$, $\text{O}=\text{O}$, $\text{N}=\text{N}$ etc are non-polar covalent bonds. On the other hand, when two different atoms share electron pair, both the atoms exert different forces on the shared electron pair. More electronegative atom pulls shared electrons pairs with greater force towards itself than the other. So more electronegative atom partially draws electron density towards itself. This makes it partially negatively charged and other atom partially positively charged. Such a covalent



bond is called polar covalent bond. The forces of attraction thus created between the molecules are called intermolecular forces. For example,



These intermolecular forces are weaker than an ionic or a covalent bond. There are several types of intermolecular forces. We will discuss two of these. Dipole-dipole interactions occur between polar molecules. Figure 4.1 shows these interactions

You know that paints and dyes are used to protect solid surfaces from the atmospheric effects. They also give visual appeal. Resins are used to coat materials that give toughness, flexibility, adhesion and chemical resistance. For example dams, bridges, floors, trains, buses, cars etc are painted with resins. The synthetic resins are used where water resistance is required. Chemically, resins are either adhesive or they form bond linkages with the material being bonded together. What is the nature of these linkages?

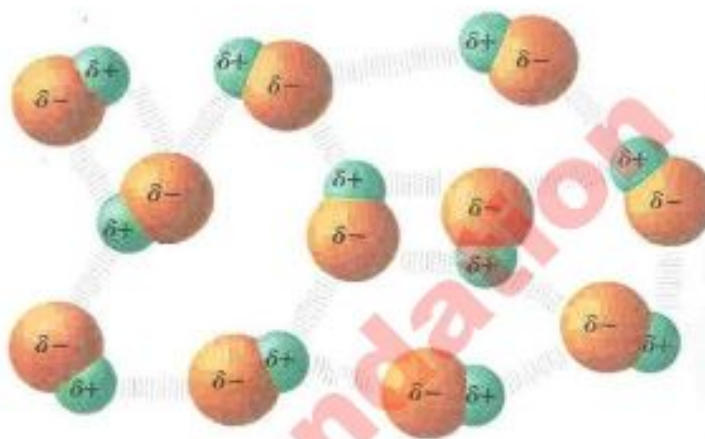


Fig 4.1: Dipole-Dipole interactions

Notice that slightly negative end of polar molecule is weakly attracted to the slightly positive end of another molecule. Such attracting forces are called dipole-dipole interactions.

Molecules in which hydrogen is covalently bonded to a very electronegative atom such as oxygen, nitrogen or fluorine is also weakly bonded to a lone pair of electron of another electronegative atom. This other atom may occur in the same molecule or in a nearby molecule. This intermolecular interaction is called hydrogen bonding. Oxygen, nitrogen or fluorine makes hydrogen very electron-deficient. Thus interaction of such a highly electron deficient hydrogen and lone pair on a nearby electronegative atom compensates for the deficiency. Figure 4.2 shows hydrogen bonding in water molecules.

The interaction of a highly electron deficient hydrogen and lone pair on a nearby highly electronegative atom such as N, O or F is called hydrogen bond. This phenomenon is called hydrogen bonding.

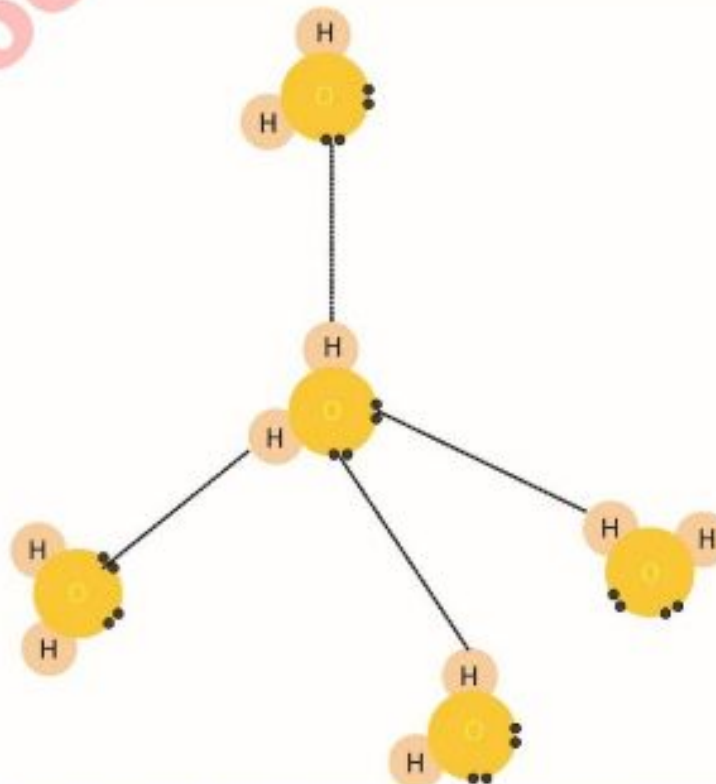


Fig 4.2: Hydrogen bonding in water



Society, Technology and Science

Epoxy adhesives have excellent chemical resistance, good adhesion properties, good heat resistance and they form strong and tough coating. Therefore, propellers and parts of aircraft, boats, cars, trucks etc are held together by epoxy adhesives. Epoxy adhesives contain partially positively charged H-atoms and oxygen atoms containing lone pairs in their molecules. Epoxy adhesives are, therefore, sticky and can make H-bonds with other substances. Modern aircraft, boats and automobiles such as cars, trucks etc and even in space craft epoxy adhesives are used for assembling, saving money and reducing weight. This means glues and adhesives have become an essential item in our daily life.

These intermolecular forces are extremely important in determining properties of water, biological molecules, such as proteins, DNA etc and synthetic materials such as glue, paints, resins etc. The adhesive action of paints and dyes is developed due to hydrogen bonding. Synthetic resins bind two surfaces together by hydrogen bonding or dipole-dipole interactions.

4.5 NATURE OF BONDING AND PROPERTIES

Compounds that consist of ions joined by electrostatic forces are called ionic compounds. At room temperature most of the ionic compounds are crystalline solids. Figure 4.3 shows arrangement of ions in NaCl and CsCl crystals.

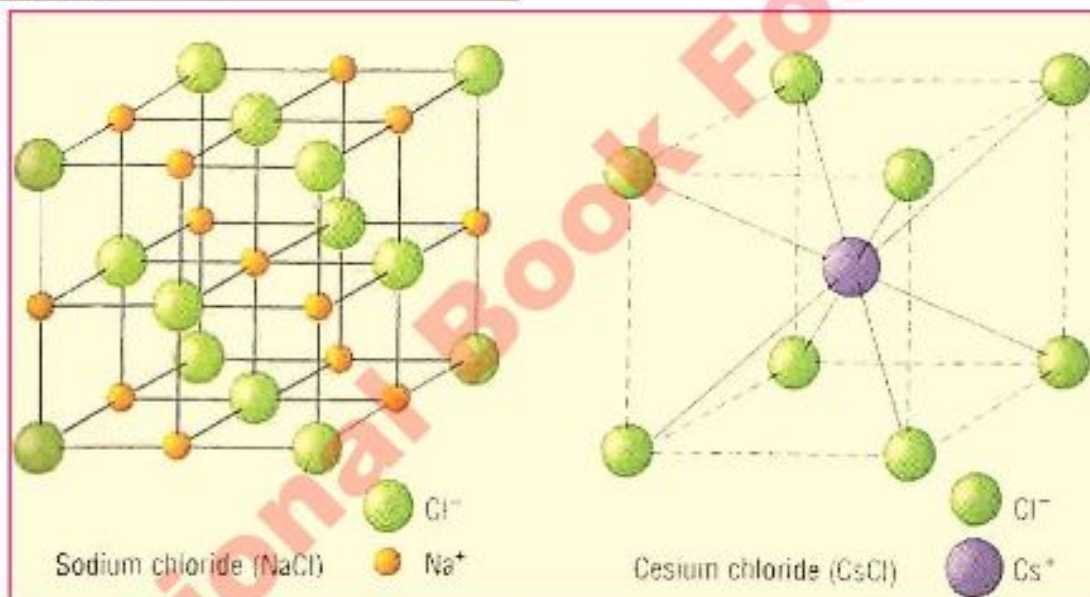


Fig 4.3: Arrangement of ions in NaCl and CsCl crystals

Note that both NaCl and CsCl form colorless cubic crystals. Each Na^+ ion is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions. Internal structure of CsCl is different from NaCl. In CsCl each Cs^+ ion is surrounded by eight Cl^- ions and each Cl^- ion is surrounded by eight Cs^+ ions. Thus in crystals each ion is attracted strongly to each of its neighbours. The large attracting forces result in a very stable structure. So ionic compounds have high melting points. For example, melting point of NaCl is 801°C .



Teacher's Point

A teacher may ask the students to make a model of NaCl



When melted ionic compounds conduct electricity, figure 4.4 shows that NaCl melts and the Na^+ and Cl^- ions are free to move throughout the molten salt. When voltage is applied, Na^+ ions move towards negative electrode. At the same time, Cl^- ions move towards positive electrode. This movement of ions inside a cell is responsible for flow of electricity between the electrodes in the external wire.

Aqueous solutions of ionic compounds also conduct electricity. This is because when an ionic compound dissolves in water, the ions are free to move about in the aqueous solution.

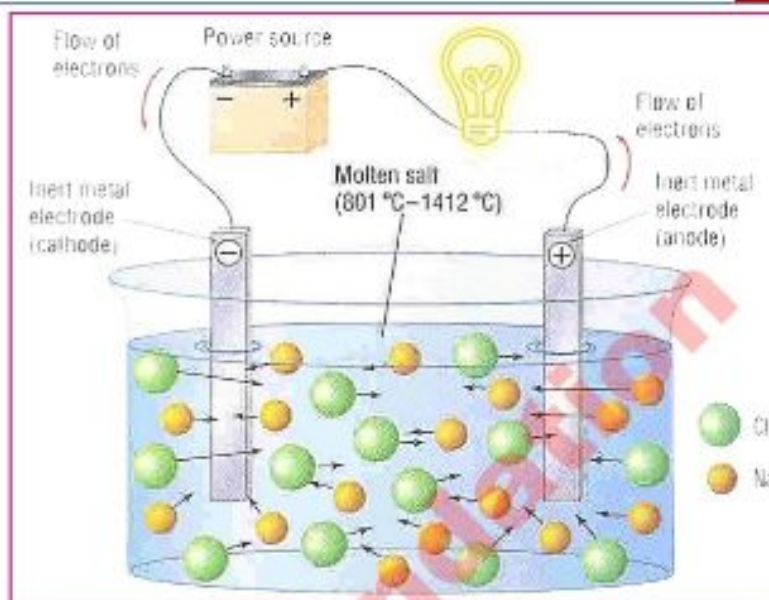


Figure 4.4: conduction of electricity through molten NaCl

Society, Technology and Science

Synthetic adhesives such as glues and epoxy resins are used in large scale industrial applications. Glues are less costly than synthetic adhesives. Glues are extensively used as an adhesive for veneer, plywood, corrugated cartons and laminated boards. Glues for glass and metal are also available. Epoxy resins are used where water resistance is required. They form strong and tough coating. They also give flexibility and chemical resistance. For these reasons, dams, bridges, thermal power stations are coated with epoxy resins. In modern air craft and automobile epoxy resin adhesives are used for assembling, saving money and reducing weight. Hence there is a need for more adhesives.



Key Points

- An octet is a set of eight. In order to gain stability atoms tend to gain electron configuration of nearest noble gas.
- The tendency of atoms to acquire eight electron configuration in their valence shell, when binding is called octet rule.
- Ionic bonds are formed between two atoms, when one atom loses electrons and other atom gains these electrons. The force of attraction that binds oppositely charged ions is called ionic bonds.
- Ionic compounds have high melting points. They conduct electricity in molten state.
- A bond that is formed by the sharing of electrons between two atoms is called a covalent bond. A covalent bond can be single, double or triple.
- The interaction of a highly electron deficient hydrogen and lone pair on a nearby electronegative atom is called hydrogen-bond.



- The adhesive action of paints and dyes is developed due to hydrogen bonding.

REFERENCES FOR ADDITIONAL INFORMATION

- Lawarie Ryan, Chemistry for you.
- Iain Brand and Richard Grime, Chemistry (11-14).
- Silberg, Chemistry.
- Raymond Chang, Essential Chemistry.



Review Questions

1. Encircle the correct answer:

- (i) Which of the following atoms will form an ion of charge -2?

<u>Atomic Number</u>	<u>Mass Number</u>	<u>Atomic Number</u>	<u>Mass Number</u>
(a) 12	24	(b) 14	28
(c) 8	8	(d) 10	20

- (ii) Which of the following atoms will not form cation or anion.

- A (Atomic No. 16)
- B (Atomic No. 17)
- C (Atomic No. 18)
- D (Atomic No. 19)

- (iii) Which of the following atoms will form cation.

<u>Atomic Number</u>	<u>Atomic Number</u>
(a) 20	(b) 18
(c) 17	(d) 15

- (vi) Which of the following atoms obey duplet rule?

- O₂
- F₂
- F₂
- N₂

- (v) Silicon belongs to Group IVA. It has _____ electrons in the valence shell

- 2
- 3
- 4
- 6

- (vi) Phosphorus belongs to third period of Group VA. How many electrons it needs to complete its valence shell.

- 2
- 3
- 4
- 5

- (vii) In the formation of AlF₃, aluminum atom loses _____ electrons.

- 1
- 2
- 3
- 4

- (viii) Which of the following is not true about the formation of Na₂S:

- Each sodium atom loses one electron
- Sodium forms cation
- Sulphur forms anion
- Each sulphur atom gains one electron

- (ix) Identify the covalent compound

- NaCl
- MgO
- H₂O
- KF



2. Give short answers
 - i. State octet and duplet rules.
 - ii. Explain formation of covalent bond between two nitrogen atoms
 - iii. How does Al form cation?
 - iv. How does O form anion?
 - v. Draw electron cross and dot structure for H_2O molecule.
3. Describe the importance of noble gas electronic configuration.
4. Explain how elements attain stability?
5. Describe the ways in which bonds may be formed.
6. Describe the formation of covalent bond between two non-metallic elements.
7. Explain with examples single, double and triple covalent bond.
8. Find the number of valence electrons in the following atoms using the periodic table:
(a) Boron (b) Neon (c) Rubidium (d) Barium (e) Arsenic
9. Represent the formation of cations for the following metal atoms using electron dot structures.
(a) Al (b) Sr (c) Ba
10. Describe the formation of anions for the following non-metal atoms:
(a) P (b) Br (c) H
11. Represent the formation of cations for the following metal atoms using electron dot structures.
(a) Mg (b) Li (c) Be
12. For each of the following pairs of atoms, use electron dot and electron cross structures to write the equation for the formation of ionic compound.
(a) K and Cl (b) Ca and S (c) Al and N
13. Recognize the following compounds as having ionic bonds.
(a) $MgCl_2$ (b) KBr (c) NaI
14. An atom of an element has atomic number 9 and mass number 19.
 - (a) State the number of protons and neutrons in the nucleus of this atom.
 - (b) State the number of electrons in this atom.
 - (c) Show with electron cross-dot diagrams, the formation of ions in the reaction of this atom with sodium atom.
15. Is there a need for more adhesives?
16. What is the importance of glues and adhesives in our society?



Think-Tank

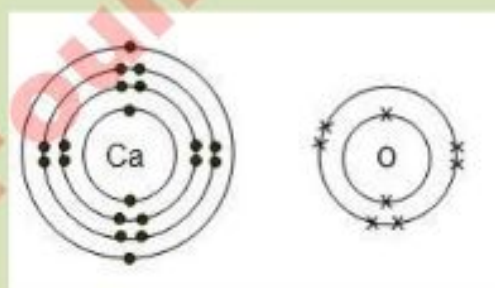
1: Magnesium oxide is a compound made up of magnesium ions and oxide ions.



- What is the charge on these ions.
- How these ions get these charges.
- Show with electron cross-dot diagrams the formation of these ions.

2: The diagrams below show the electronic structures of an atom of calcium and an atom of oxygen.

Draw structures of the ions that are formed when these atoms react.



3: Draw electron cross and dot structure for the following molecules:

- COCl_2 , a poisonous gas called phosgene that has been used in World War-II.
- HOCl , hypochlorous acid is unstable, decomposes to liberate atomic oxygen that makes HOCl a strong oxidizing agent.

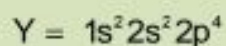
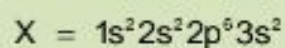
4: The table below shows the properties of four substances:

Substance	Melting point	Electrical Conductivity	
		In solid state	In molten state
A	High	NIL	NIL
B	High	NIL	Good
C	Low	NIL	NIL
D	High	Good	Good

- Which substance is a metal?
- Which substance is an ionic compound?
- Which substance is a covalent compound?
- Which substance is a non-metal?

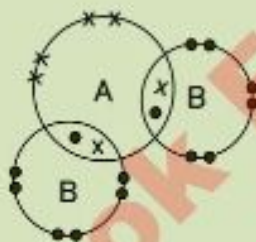


5: Electronic configuration of two elements X and Y are given below:



Predict which of the following compounds is likely to form when X and Y react? Explain.

- (a) A covalent compound of formula XY_2
 - (b) An ionic compound of formula XY_2
 - (c) An ionic compound of formula XY
 - (d) An ionic compound of formula X_2Y
- 6: The following figure shows the electron dot and cross diagram of molecule AB_2 . Which of the elements could be A and B? Apatise.



7: Illustrate the total number of shared electrons in a molecule of CO_2 ?

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In 1980, he launched himself into the teaching profession, when he joined the Chemistry Department of Army Burn Hall College Abbottabad. In 1982, he joined the Islamabad College for Boys, G-6/3, Islamabad and also served as Principal in various colleges in Islamabad. He retired as Principal from Islamabad College for Boys (ICB) G-6/3, Islamabad. On the occasion of the Silver Jubilee celebrations of I.C.B, G-6/3 he was declared as the best teacher in Chemistry and awarded a gold medal.

His rich experience in teaching Chemistry, has enabled him to serve as member, National Review Committee from 2002-2006 for finalizing textbook manuscripts in the subject of Chemistry for Secondary and Higher Secondary classes, developed by various Textbook Boards. He is also Managing / Co-Author of several textbooks. For queries about book in question, please contact at: prof.iqtedar@gmail.com.



Approved by the Capital Administration & Development Division
(Curriculum Development & Textbook Production Wing),
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vide letter No. F.1-1/2007 Chem Dated: 16-03-2015

قومی ترانہ

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تو نشانِ عزمِ عالی شان ارضِ پاکستان
مدرکزِ یقین شاد باد!

پاک سرزمین کا نظام قوتِ اخوتِ عوام
قوم، ملک، سلطنت پائندہ تابندہ باد!
شاد باد منزلِ مسرور!

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ترجمانِ ماضی، شانِ حالِ جہانِ استقبال
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5

PHYSICAL STATES OF MATTER



This is a 18 days lesson

After completing this lesson, you will be able to:

- Account for pressure-volume changes in a gas using Boyle's Law.
- Account for temperature-volume changes in a gas using Charles's Law.
- Compare the physical states of matter with regard to intermolecular forces present between them.
- Describe physical properties of solids (melting and boiling points).
- Differentiate between amorphous and crystalline solids.
- Explain the effect on the pressure of a gas by a change in the **a. volume b. temperature**.
- Explain the properties of gases (diffusion, effusion and pressure).
- Explain the effect of temperature and external pressure on vapour pressure and boiling point.
- Explain the allotropic forms of solids.
- Summarize the properties of liquids like evaporation, vapour pressure, boiling point.



Pre- Reading

At room temperature most of the pure substances exist in any of the three states of matter, gas, liquid or solid. The study of gases provides excellent example of the scientific method in action. It illustrates how observations lead to natural laws. The earth's atmosphere is a mixture of gases. It not only supports life but also accommodates exhaust gases that accompany many industrial processes. Therefore, it is important to understand the behaviour of a gas.

The most common liquid on earth is water. Apart from other uses, water is used for cooking food. Why do mountaineers carry with them pressure cooker? Do they need to cook food quickly? Food cooked in ordinary kettles at higher altitude remains semi-cooked why? How does pressure cooker help in proper cooking of food? Why pressure cooker saves time in the kitchen?

You will learn about phase changes such as melting, vaporization and sublimation. Phase changes play many important roles in our daily lives. For instance freeze-dried foods are light-weight and can be conveniently re-constituted by adding water. How? This chapter will enable to understand all this.



Reading

GASEOUS STATE

5.1 TYPICAL PROPERTIES

The following properties are common to all gases:

1. Diffusion

The characteristic smell of rotten egg spreads throughout the room. This is because of the concentration of hydrogen sulphide molecules at a particular place is higher, they start moving towards a place where their concentration is lower. This movement of molecules from a higher concentration to a lower concentration is called diffusion. We can also say that the diffusion of a gas is its spreading tendency from one place to another. Molecules in a gas are widely separated and they do not have any attractive or repulsive forces. Thus molecules of one gas can move easily between the molecules of another gas.

2. Effusion

When gas is allowed to pass through a hole of the size of a molecule, only one molecule at a time can pass through the hole. **The escape of gas molecules through the hole one after the other without collision is called effusion.** Figure 5.1 shows diffusion and effusion of gases.

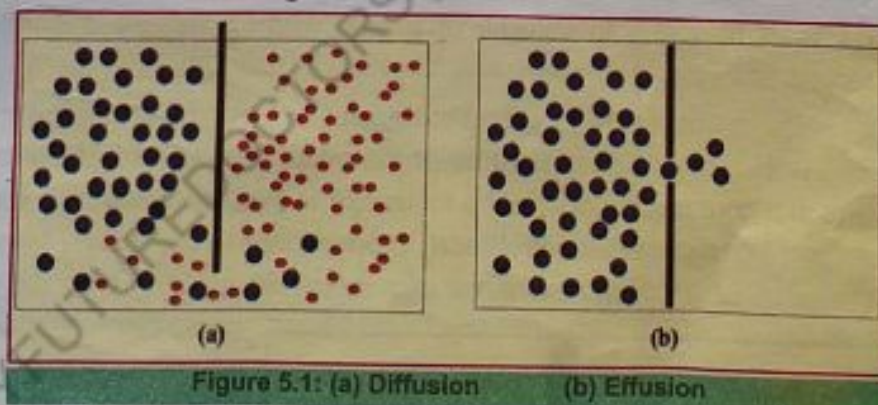


Figure 5.1: (a) Diffusion (b) Effusion

3. Pressure

A gas exerts pressure on the walls of its container, as you can see when you blow a balloon. Force exerted by a gas on unit area of a container is called its pressure. All gases exert pressure. At any point a gas exerts an equal pressure in all directions.

At sea level at 0°C the atmospheric pressure is 760mm of Hg or 760 torr. This pressure is referred as one atmosphere.

$$\begin{aligned}\text{So,} \quad 1\text{atm} &= 760\text{mm Hg} \\ &= 760\text{ torr}\end{aligned}$$



The SI unit of pressure is pascal (Pa) which is very small unit, so kilopascal (kPa) is used for atmospheric pressure under ordinary conditions. One pascal equals the force of one Newton exerted on an area of one square meter.

$$1 \text{ atm} = 101.325 \text{ kPa} \\ = 1.01325 \times 10^5 \text{ Pa.}$$

4. Compressibility

Q5(a)

(Gases are highly compressible. This is due to the presence of large empty spaces between the gas molecules. On applying pressure distances between the gas molecules decrease, therefore, its volume decreases.)

5. Mobility

Gas can flow and can be transported through pipes over long distances but it can also leak more rapidly out of small hole. This is due to tendency of a gas to expand and fill the entire available space.

6. Density

Q5(b)

Q33

(Gases have relatively low densities under normal conditions. This is because the molecules are much farther apart in the gas. When a gas is cooled its density increases because its volume decreases.) For example, density of oxygen at 20°C is 1.4 g/dm^3 and at 0°C is 1.5 g/dm^3 .

5.1.1 Effect on the Volume of a Gas by a Change in Pressure and Temperature

Figure 5.2 shows what happens to volume of a gas, when pressure over the gas is increased at constant temperature.

Q26

(According to the kinetic molecular theory, gas molecules are in constant random motion. They move in straight line until they collide with another molecule or the walls of the container. The pressure a gas exerts in a container is due to the force exerted on the walls of the container. The average kinetic energy of gas molecules is directly proportional to the Kelvin temperature. So the average kinetic energy of a collision when a gas molecule collides with the wall of a

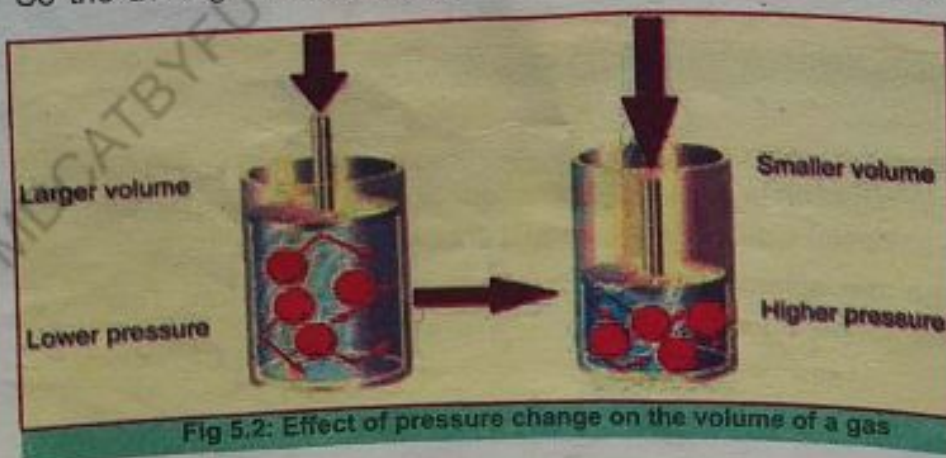


Fig 5.2: Effect of pressure change on the volume of a gas

container will not change when the pressure decreases at constant temperature. (There are large empty spaces between the molecules. On increasing pressure on the gas, the distances between molecules decrease. So the volume of the gas decreases.)



On the other hand when pressure of the gas is kept constant and temperature is increased, the average kinetic energy of gas molecules increases. So, the molecules of the gas hit the wall of container more frequently and energetically. This increases internal pressure. As a result, volume of the gas increases to restore constant pressure. (Fig 5.3)

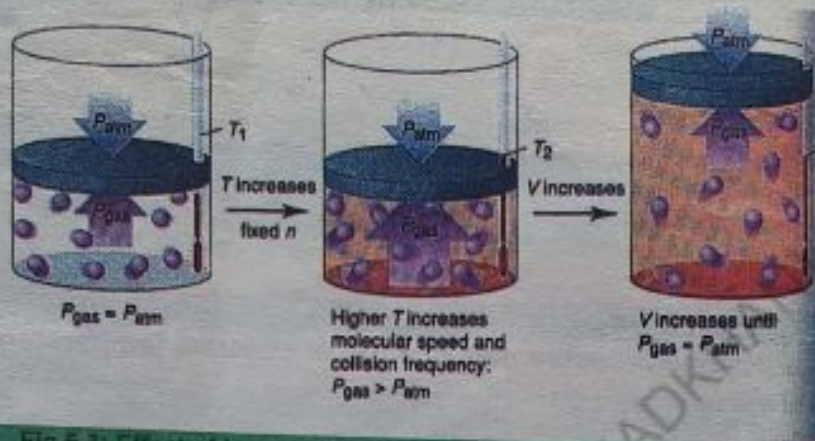


Fig 5.3: Effect of temperature change on the volume of a gas

5.2 LAWS RELATED TO GASES

A sample of a gas can be characterised by four variables:

1. Pressure (P)
2. Volume (V)
3. Temperature (T)
4. The number of moles of gas (n)

The relationship that expresses the influence of one variable on another with the two variable constant are called gas laws.

5.2.1 Boyle's Law

Figure 5.4 shows a syringe filled with air.

Seal needle's tip with alfe. What happens when its piston is moved from position A to B?

A student obtained data of volume and pressure for a given mass of a gas at constant temperature. It is shown in table 5.1

Society, Technology and Science

Scientists use the power of reasoning to explain their observation. For instance when a balloon is filled with air, it expands. A scientist would explain it by saying that air molecules are free to move inside their container. There is no attractive or repulsive force between the molecules. As a result gas expands until it takes the shape of its container. Therefore, air expands to fill the interior of the balloon evenly.

Do You Know?

Earth's atmosphere consists of gases that provide the oxygen and pressure necessary to support life as we know it. Atmospheric gases moderate temperature extremes and one of them (ozone) shields living things on the surface from harmful ultraviolet radiations. In contrast the moon has no atmosphere.

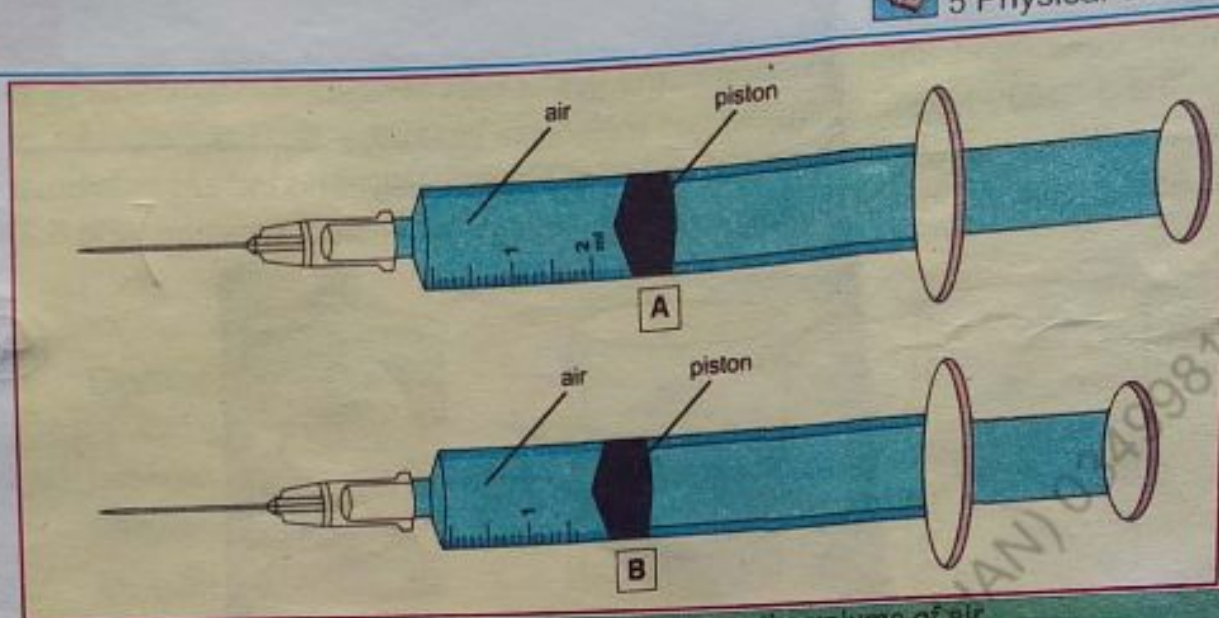


Fig 5.4: Effect of pressure change on the volume of air

Table 5.1: Volume-pressure data of a gas at constant temperature

Experiment No.	Pressure (atm)	Volume (dm ³)
1.	0.500	4.00
2.	1.00	2.00
3.	2.00	1.00
4.	4.00	0.500

How can you interpret this data? If you carefully look at the data, you will find that as the volume of the gas is decreasing, the pressure is increasing. This means an inverse relationship may exist between volume and pressure

So,

$$V \propto \frac{1}{P}$$

$$V = \frac{1}{P} \times \text{Constant}$$

$$V \times P = \text{constant}$$

This means product of volume and pressure must be constant if inverse relationship exists between them. So this table may be expanded to table 5.2 to show this relationship:

Table 5.2: Relationship between volume and pressure at constant temperature

Experiment No.	Pressure (atm)	Volume (dm ³)	V x P (dm ³ .atm)
1.	0.500	4.00	2.00
2.	1.00	2.00	2.00
3.	2.00	1.00	2.00
4.	4.00	0.500	2.00



Table 5.2 clearly shows that the product of volume and pressure is keeping constant.

Therefore, we can write **$V \times P = \text{constant}$**

$$V = \frac{1}{P} \times \text{Constant}$$

$$V \propto \frac{1}{P}$$

This means the volume of a given mass of a gas at constant temperature is inversely proportional to the pressure on the gas. This relationship is known as Boyle's Law.

Boyle's law states that the volume of a fixed amount of a gas at a given temperature is inversely proportional to the applied pressure.

Example 5.1: Accounting for pressure-volume changes in a gas using Boyle's Law

Q 20 Ethene is used as anaesthetic gas. The pressure on 2.5 dm^3 of ethene changes from 1.05 to 2.10 atm. The volume of ethene becomes 1.25 dm^3 if the temperature remains constant. Explain this change using Boyle's law.

Problem Solving Strategy:

1. According to the Boyle's law, product of pressure and volume is constant at any two sets of conditions.
2. Calculate $P \times V$ for the two sets of condition and compare.

Solution:

$$\begin{aligned} P_1 \times V_1 \text{ before change} &= 1.05 \times \text{atm} \times 2.5 \text{ dm}^3 \\ &= 2.625 \text{ atm. dm}^3 \end{aligned}$$

$$\begin{aligned} P_2 \times V_2 \text{ after change} &= 2.1 \text{ atm} \times 1.25 \text{ dm}^3 \\ &= 2.625 \text{ atm. dm}^3 \end{aligned}$$

$$\therefore P_1 V_1 = P_2 V_2$$

(Thus the calculated result agrees with the pressure-volume relationship according to the Boyle's Law.)

Note: Relationship $P_1 V_1 = P_2 V_2$ can be used to determine any variable knowing the other three.

Society, Technology and Science

In early 1600s, Galileo argued that suction pumps were able to draw water from a well because of the "force of vacuum" inside the pump. After Galileo's death, the Italian mathematician and physicist E. Torricelli proposed another explanation. He suggested that the air in the atmosphere has weight. The force of atmosphere pushing down on the surface of water drives the water into the suction pump when it is evacuated. In 1643 Torricelli invented a device called barometer. He measured atmospheric pressure by the barometer and found that at sea level and 0°C , the atmospheric pressure is 760mm Hg. Torricelli's work soon caught the attention of British scientist Robert Boyle. He modified barometric tube into a J-shaped tube. By adding mercury to the open end of the tube, he trapped a small volume of air in the sealed end. He studied what happened to the volume of the air as he added more mercury to the open end. Boyle's from the studies discovered the pressure-volume relationship. J-tube was further modified and another device known as manometer was developed that can measure the pressure of any gas. This means instrumentation improves as science progresses.



Self-Assessment Exercise 5.1

1. A student obtained following data in an experiment at 20°C.

P(atm)	V(dm ³)	P(atm)	V(dm ³)
0.350	0.707	0.951	0.261
0.551	0.450	1.210	0.205
0.762	0.325		

Explain pressure-volume relationship using this data and the Boyle's Law.

2. Ammonia gas is used as refrigerant 0.474 atm. pressure is required to change 2000cm³ sample of ammonia initially at 1.0 atm. to 4.22dm³ at constant temperature. Show that this data satisfies Boyle's law.



Fig 5.5



Reading

5.2.5 Charles's Law

Figure 5.5 shows study of the effect of temperature on volume of a gas. A known mass of gas is trapped in graduated glass tube by placing small amount of mercury. The tube is immersed in a water bath that can be warmed with the heater. After each change of water temperature you can measure the volume of a gas. The pressure exerted on the gas sample is constant because the amount of mercury and atmospheric pressure above it do not change.

A student studying effect of temperature on volume of a given mass of a gas obtained data as given in Table 5.3.

Table 5.3 volume-temperature data of a gas at constant pressure.

Sr. No	Temperature 0°C	Volume (dm ³)
1.	0	1.00
2.	100	1.37
3.	200	1.73
4.	300	2.10

How can you interpret this data? You can interpret this data by plotting a graph. Place values of volume on the y-axis and value of absolute temperature on x-axis. You can convert centigrade temperature to absolute or Kelvin temperature by adding 273 to it. Absolute temperature is represented by T.

$$T = t^{\circ}\text{C} + 273$$

You will get the following graph (figure 5.6).

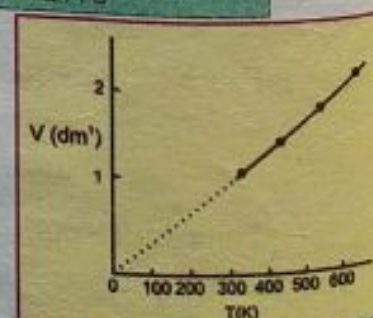


Figure 5.6: Dependence of volume on absolute temperature



Straight line between the temperature and volume suggests that there is a direct relationship between volume and temperature measured at Kelvin scale at constant pressure.

Q23

$$V \propto T$$

$$V = \text{constant} \times T$$

Q29

$$\frac{V}{T} = \text{constant}$$

$$V_1 T_1 = V_2 T_2$$

for two sets of conditions

This relationship is known as the Charles's Law.

The law states that **the volume of given mass of a gas varies directly with absolute temperature at constant pressure.** Since the ratio of $\frac{V}{T}$ remains constant at constant pressure, you can also observe this relationship by expanding table 5.3 to 5.4

Table 5.4: Volume-Temperature relationship

Sr. No.	Volume (dm ³)	Temperature(K)	$\frac{V}{T}$
1.	1.00	273	0.0037
2.	1.37	373	0.0037
3.	1.73	473	0.0037
4.	2.10	573	0.0037

Since ratio of $\frac{V}{T}$ is constant,

$$\frac{V}{T} = \text{constant}$$

$$V = T \times \text{constant}$$

$$V \propto T$$

and

Example 5.2: Accounting for temperature-volume changes in a gas using Charles's

Table 5.5 shows data of volume of a gas and its temperature for the given mass of a gas at 900 mm Hg.

Table 5.5: Temperature-volume data for a gas at 900mmHg

Temperature(°C)	Volume(cm ³)
0	107.9
5	109.7
10	111.7
15	113.6
20	115.5



Explain volume-temperature relationship using Charles's Law.

Problem Solving Strategy:

1. According to the Charles's law, ratio of volume to absolute temperature is constant for any set of conditions.
2. Convert $^{\circ}\text{C}$ temperature to Kelvin temperature by adding 273.
3. Find $\frac{V}{T}$ for each set of conditions and compare.

Solution:

Table 5.5 can be expanded into another table 5.6

Table 5.6 Temperature-volume relationship

Temperature ($^{\circ}\text{C}$)	Volume (cm^3)	Temperature (K)	$\frac{V}{T}$
0	107.9	273	$\frac{107.9}{273} = 0.3952$
5	109.7	278	$\frac{109.7}{278} = 0.3948$
10	111.7	283	$\frac{111.7}{283} = 0.3947$
15	113.6	288	$\frac{113.6}{288} = 0.3944$
20	115.5	293	$\frac{115.5}{293} = 0.3942$

The ratio $\frac{V}{T}$ is fairly constant. Thus volume of the gas varies directly with the absolute temperature as stated by the Charles's law.



Self-Assessment Exercise 5.2

1. A chemist obtained data shown in table 5.7 in an experiment at 1 atm.

Table 5.7 Temperature-volume data of a gas at 1 atm

Temperature($^{\circ}\text{C}$)	Volume(cm^3)	Temperature($^{\circ}\text{C}$)	Volume(cm^3)
25	117.5	35	121.3
30	119.4	40	123.2

Explain volume-temperature relationship using Charles's law.



- Q4 2. A bacterial culture isolated from sewage produces 36.4 cm^3 of methane (CH_4) gas at 27°C and 760 mm Hg . This gas occupies 33.124 cm^3 at 0°C and same pressure. Explain volume-temperature relationship from this data.
- Q8 3. A perfect elastic balloon filled with helium gas has a volume of $1.25 \times 10^3 \text{ dm}^3$ at 1.00 atm and 25°C on ascending to a certain altitude where temperature is 15°C the volume of balloon becomes $1.208 \times 10^3 \text{ dm}^3$. Show that this data satisfies the Charles's law.



Reading

LIQUID STATE

5.3 TYPICAL PROPERTIES

Some typical properties of liquids are given below:

5.3.1 Evaporation

Place some liquid such as ether or acetone in an open container and observe. You will notice that the volume of the liquid gradually decreases and finally no more of the liquid is left. This is because liquids constantly change into gas or vapour even when the temperature is less than the boiling point of a liquid. **Conversion of a liquid to a gas or vapour at all temperatures is called vaporization or evaporation.** (figure 5.7)

Q64 [Reason (In evaporation, some molecules in the liquid break away and enter the gas or vapour state. Only those molecules which have greater kinetic energy than average can break away from the surface. This means the molecules with the highest kinetic energy escape first. The molecules in the liquid have a lower average kinetic energy than the molecules that have escaped. Therefore liquid's temperature will decrease. Is evaporation a cooling process? Liquids evaporate faster when heated. This is because added heat increases the kinetic energy of the molecules. This enables more molecules to overcome attractive forces keeping them in the liquid state and escape as vapours.) Q34



Figure 5.7:
Evaporation in an
open container



Self-Assessment Exercise 5.3

Give reason:

- Q34 1. When you put nail polish remover on your palm, you feel a sensation of coldness.
- Q34 2. Wet clothes dry quickly in summer than in winter.



Reading

5.3.2 Vapour Pressure

The evaporation of a liquid in a closed container is somewhat different. Figure 5.8.

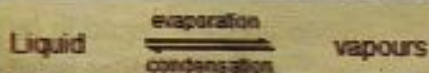


Figure 5.8:
Evaporation in a
closed container

Notice that in a closed container no molecules can escape into the outside air, when a partially filled container is sealed (as shown in figure 5.8). Some of the liquid molecules vapourize. As the time passes, the number of molecules changing into vapour increases. Some of these molecules because of their random motions will collide with the liquid surface. Such molecules are recaptured by the molecules at the surface of the liquid. This process is called condensation. These two opposing processes will continue. After some time, the number of molecules evaporating will become equal to the number of molecules condensing. At this stage equilibrium is established between the liquid and its vapours.

At equilibrium liquid continues to evaporate and condense but at equal rates.

The pressure exerted by the vapours of a liquid in equilibrium with its liquid is called vapour pressure. At equilibrium the space above the liquid is saturated with vapours.



Effect of Temperature on Vapour Pressure

Table 5.8 shows vapour pressures of some liquids at various temperatures.

Table 5.8: Vapour pressures of some liquids at various temperatures

Vapour Pressure (kPa) of Several Substances at Various Temperatures						
	0°C	20°C	40°C	60°C	80°C	100°C
Water	0.61	2.33	7.37	19.92	47.34	101.33
Ethanol	1.63	5.85	18.04	47.02	108.34	225.75
Diethyl ether	24.70	58.96	122.80	230.65	399.11	647.87

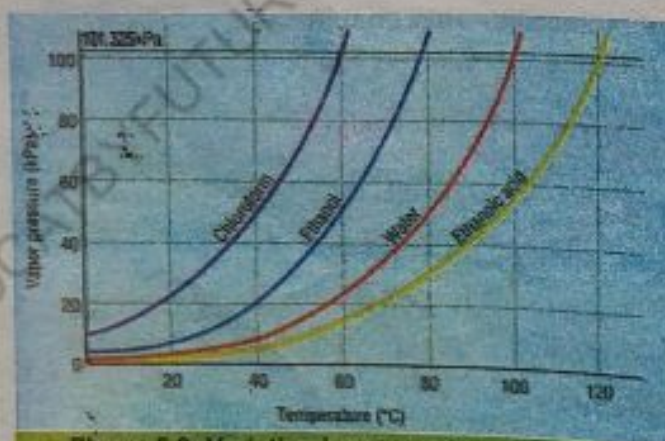


Figure 5.9: Variation in vapour pressures with temperature for some liquids

Look at the table 5.8 and notice something. You will find that the vapour pressure of liquids changes with temperature. This is because an increase in temperature of a liquid increases the kinetic energy of its molecules. As a result more of the molecules will have minimum kinetic energy needed to escape the surface of the liquid. *(and so vapour pressure also changes)*

5.3.3 Boiling Point

In an open container atmosphere exerts pressure on the liquid surface. When a liquid is heated in an open container, average



kinetic energy of the molecules increases. Therefore, the temperature of the liquid also increases. Heating allows larger number of molecules at the liquid's surface to overcome the attractive forces responsible keeping them in the liquid. This increases the vapour pressure of the liquid. Vapour pressure of the liquid keeps on increasing with the increase in temperature. At a certain temperature, the vapour pressure of the liquid becomes equal to the atmospheric pressure or external pressure. At this stage liquid starts boiling. Bubbles of vapours form throughout the liquid, rise to surface and escape into the air. **The temperature at which vapour pressure of a liquid becomes equal to the external or atmospheric pressure is called boiling point.** Figure 5.9 shows the variation in vapour pressures of four liquids with temperature. It also indicates the boiling points of these liquids.

Remember that a liquid boils when its vapour pressure becomes equal to the atmospheric pressure i.e., 760mmHg or 101.325 kPa at sea level.



Self-Assessment Exercise 5.4

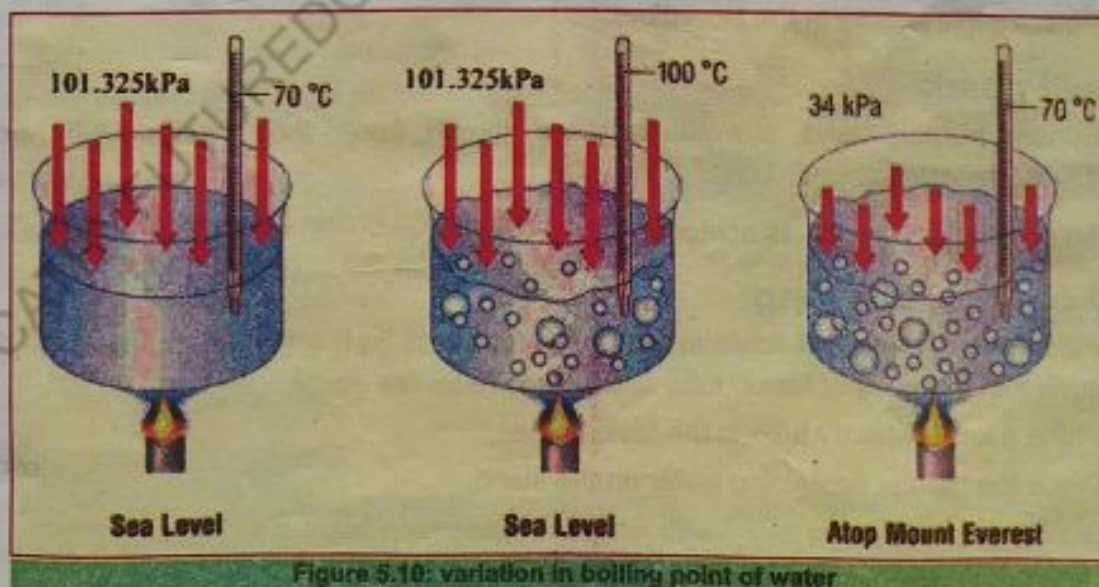
Look at figure 5.9 and predict the boiling points of the four liquids at normal atmospheric pressure.



Reading

5.3.3.1 Effect of External Pressure on Boiling Point

Liquids boil when their vapour pressure is equal to the pressure exerted on the liquid by its surroundings. The normal boiling point of water is 100°C . In the mountains the atmospheric pressure is less than 1 atm. so water boils below 100°C . (In a pressure cooker at 2atm, water does not boil until the temperature reaches 120°C .) Figure 5.10 shows this variation in boiling point of water.





To understand why water boils at 70°C on the top of Mount Everest and 120°C in a pressure cooker, although the normal boiling point of water is 100°C . Concentrate on the definition of boiling point "**a liquid boils when its vapour pressure is equal to the pressure exerted on the liquid by its surroundings**". (See figure 5.10). *Do not use*

When the pressure of atmosphere is 1atm or 101.325 kPa water boils at 100°C at sea level. This is because at this temperature vapour pressure of water is 1atm or 101.325 kPa. At Mount Everest at about 8850m above sea level atmospheric pressure is only 34kPa. So water boils at this height above sea level, when its vapour pressure is 34kPa at 70°C . So water boils at 70°C .

Q252 Pressure cooker is equipped with a valve that controls the pressure inside the pot. This valve generally exerts a pressure of 2 atm. Therefore, the valve does not allow water vapours to escape until the pressure inside the pot reaches 2 atm. Because vapour pressure of water becomes 2atm when the temperature reaches 120°C . So water boils at 120°C in a pressure cooker.



Self-Assessment Exercise 5.5

- Q132
1. The boiling point of water on the top of Mount Everest is 70°C , while at Murree 98°C . Explain this difference.
 2. If you try to cook an egg in boiling water while camping at an elevation of 0.5km in the mountain, you will find that it takes longer than it does at home. Explain why?



Activity 5.1

Determining boiling point of an organic liquid

Perform this activity in the laboratory.

You will need:

- Bunsen burner, stand, wire gauze, stirrer, thread, fusion tube, capillary tube, beaker, thermometer and match box.
- An organic liquid such as acetone.

Carry out the following:

1. Place small amount of liquid in the fusion tube and tie it with thermometer with thread such that the ends of fusion tube and thermometer are equal.
2. Place a long capillary tube in the fusion tube.
3. Place the beaker containing water on the stand.
4. Suspend thermometer along with fusion in water: (see figure 5.11)
5. Heat the beaker and stir water with the stirrer.



5 Physical States of Matter

- Continue heating and stirring until bubbles start rising from the lower end of the capillary tube.
- See the temperature on the thermometer. This is the boiling point of the liquid. (Q92) (1033)

A process called distillation is used to purify liquids. **Distillation is the process in which a liquid is heated to vapourize it and the vapours are cooled to condense them back to the liquid in a different container.**



Activity 5.2

Carrying out distillation process

Perform this activity in the laboratory.

You will need:

- A round bottom flask with side arm, thermometer, a glass condenser, a conical flask.
- Tripod stand, plastic tubes, stand, Bunsen burner.
- Aqueous solution of NaCl.

Carry out the following:

- Place NaCl solution in the round bottom flask and place it on the tripod stand.
- Fix thermometer on the mouth of the flask.

Connect one end of condenser with the round bottom and other end with the conical flask. See figure 5.12 for the assembly.

- Connect lower side arm of condenser with the plastic tube to the water tap. Connect upper side arm with another plastic tube to drain circulating water.
- Heat NaCl solution and open tap so that water circulates through the condenser.
- Observe what happens.



Teacher's Point

Teacher may show apparatus of B.P and distillation to students.

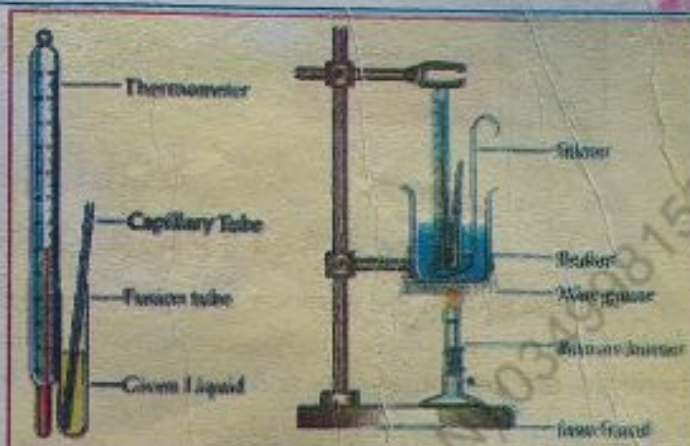


Figure 5.11: Determination of the boiling point of an organic compound.

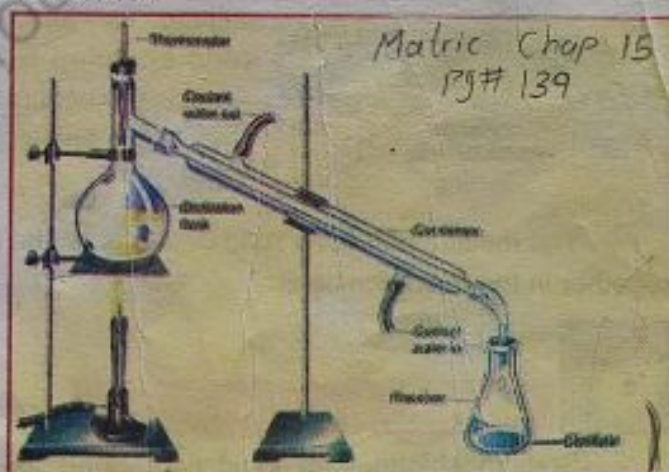


Figure 5.12: Distillation apparatus



Reading

SOLID STATE

Q15(a)

5.4 TYPICAL PROPERTIES

Unlike liquids, the particles in solids are not free to move. In solids particles are closely packed together and have fixed positions. They can only vibrate about their fixed positions. Solids, therefore, are incompressible and do not flow. See figure 5.13.

5.4.1 Melting and Freezing

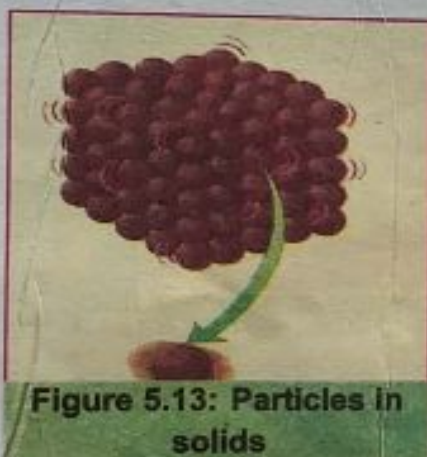


Figure 5.13: Particles in solids



The particles in a solid vibrate about their fixed positions. When you heat a solid, its particles vibrate more rapidly. This is because their kinetic energy increases on heating. When its temperature is raised to a level at which the vibrations of the particles become so intense that the particles start leaving fixed positions. Thus arrangement of particles within the solid breaks down. Eventually the solid melts. **The temperature at which a solid turns into a liquid is called melting point.** On cooling, the liquid freezes. Thus freezing of a liquid is the reverse of melting. **The temperature at which a liquid changes into the solid is called freezing point.** The melting and freezing of a substance occurs at the same temperature. At this temperature, the liquid and solid substances are in equilibrium with each other.

(Q15(b)) The melting point of a solid depends on the strength of attractive forces that hold particles together in the fixed positions.



Self-Assessment Exercise 5.6

(Q15(b)) Sodium chloride, an ionic compound, has a high melting point of 801°C . Whereas molecular solid such as ice has relatively low melting point of 0°C . Explain this difference.



Reading

Sublimation

(Q33) Some solids on heating change to vapours without passing through the liquid state. This process is called sublimation. Substances like iodine, benzoic acid, ammonium chloride etc sublime. Figure 5.14 shows sublimation of violet-black crystals of iodine.



Figure 5.14: Sublimation of violet-black crystals of iodine

Observe that dark crystals of iodine deposit on the under side of the glass watch placed on the top of a beaker containing solid iodine that is being heated. The iodine vapour sublimates from iodine crystals in the bottom of the beaker and condenses to form crystals on the glass watch.



Activity 5.3

Determining melting point of organic solids

Perform this activity in the laboratory.

You will need:

- Beaker, thermometer, capillary tube, Bunsen burner, glass rod and wire gauze.
- Organic solid such as naphthalene.

Carry out the following:

1. Close one end of the capillary tube by heating.
2. Fill one third of this capillary tube with small amount of powdered naphthalene.
3. Attach this capillary tube with the thermometer with the thread.
4. Suspend thermometer along with capillary tube in a beaker containing water.
5. Heat beaker slowly with constant stirring water.
6. Note the temperature when the naphthalene in the capillary tube just starts melting.

This is melting point of the naphthalene. Figure 5.15 shows schematic assembly for determining melting point of the solid.

Society, Technology and Science

Freeze-dried foods are light-weight and conveniently re-constituted by adding water. When salt is applied on meat it draws out considerable amount of water from the meat. After this meat is frozen and placed in a chamber attached to a vacuum pump. By lowering the pressure below the vapour pressure of ice, the ice crystals sublime and the meat is dried without the loss of its flavour. Dried meat needs no refrigerator because bacteria such as salmonella which causes food poisoning cannot grow on salt and in the absence of moisture. Thus curing with the salt helps preserve meat.

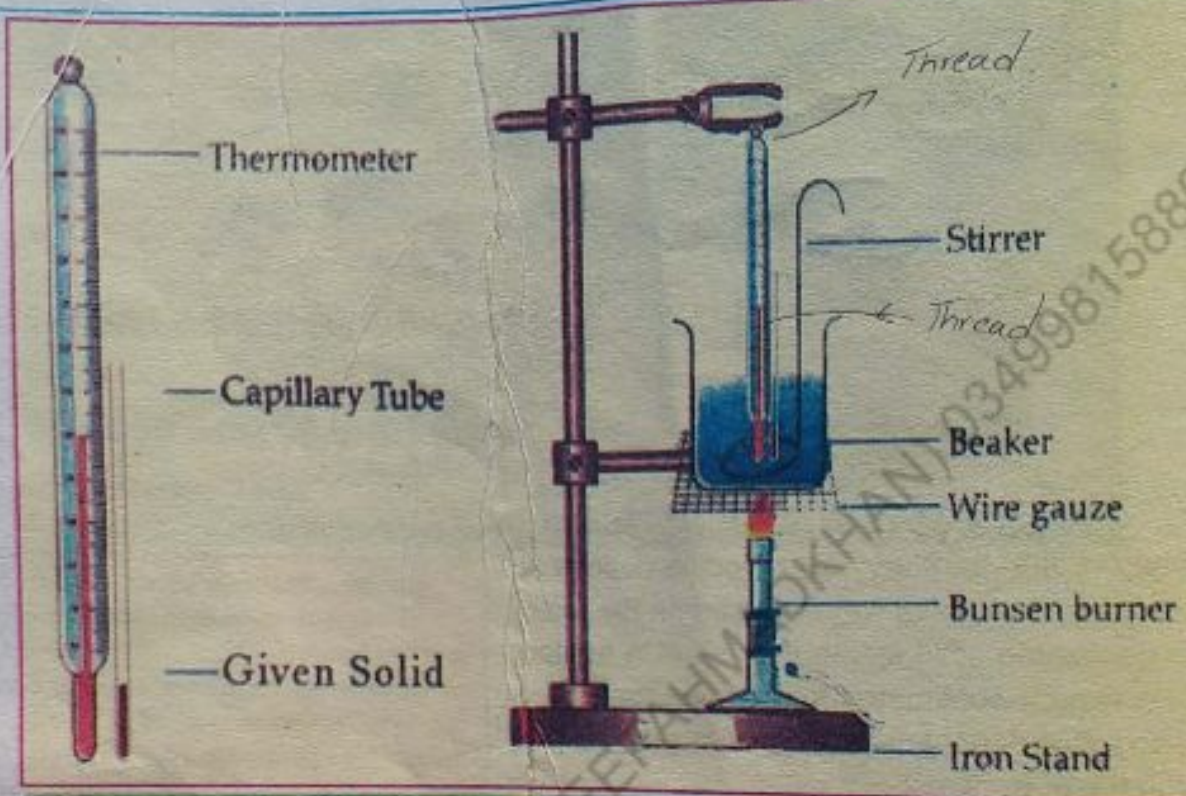


Figure 5.15: Determination of melting point of a solid



Activity 5.4

Carrying out sublimation of air freshener

Perform this activity in the laboratory.

You will need:

- Small pieces of solid air-freshener.
- A shallow container.
- Two clean plastic cups.
- Hot water and ice.
- Cardboard strips.

Carry out the following:

1. Place some pieces of solid air-freshener in one of the cups.
2. Bend the cardboard strips and place over the rim of the cup containing air freshener pieces.
3. Place the other cup inside the first cup. It should not touch the pieces of the air freshener. Adjust the cardboards if necessary. This assembly is your sublimate (figure 5.16).
4. Fill the upper cup with pieces of ice cubes.



5. Fill the shallow container about one third with the hot water.
6. Carefully place sublimate in hot water. Observe what happens.

Note: you can perform this activity with naphthalene as well.



Reading

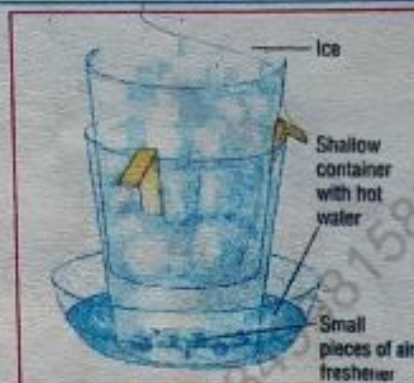


Figure 5.16: Sublimation

5.5 TYPES OF SOLIDS

A solid consists of an array of particles which may be atoms, ions or molecules. These particles are held together by strong attractive forces between them. The strength of force that exists between the particles of a solid determines the melting point of the solid. You can recognize solids into two major groups based on their macroscopic appearance. This is based on the arrangement of particles of the solid.

A crystalline solid is a solid that is composed of orderly, repeating three-dimensional arrangement of particles. This means a crystalline solid has well defined shape because of the orderly arrangement of its particles. Crystalline solids have sharp melting points. Figure 5.17 shows shape of sodium chloride crystal.

Notice that the Na^+ and Cl^- ions are close and packed together in a regular array. **Whereas an amorphous solid is one that lacks ordered arrangement of its particles.** This means it does not have a well defined arrangement of its particles. Amorphous solid do not melt at a definite temperature but gradually soften when heated. For example, glass, plastic etc are amorphous solids.)

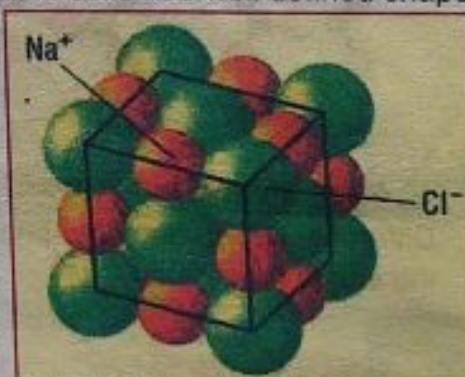


Figure 5.17: Sodium chloride crystal



Self-Assessment Exercise 5.7

1. Quartz is the crystalline form of silicon dioxide (SiO_2). It is a hard, brittle and colourless solid. When quartz is heated above its melting point (about 1600°C) and cooled rapidly, an amorphous solid called quartz glass results. Figure 5.18 shows two-dimensional representation of quartz glass and quartz crystal. Identify each? Give reasons.



Teacher's Point

A teacher may show model of balls of NaCl to students.

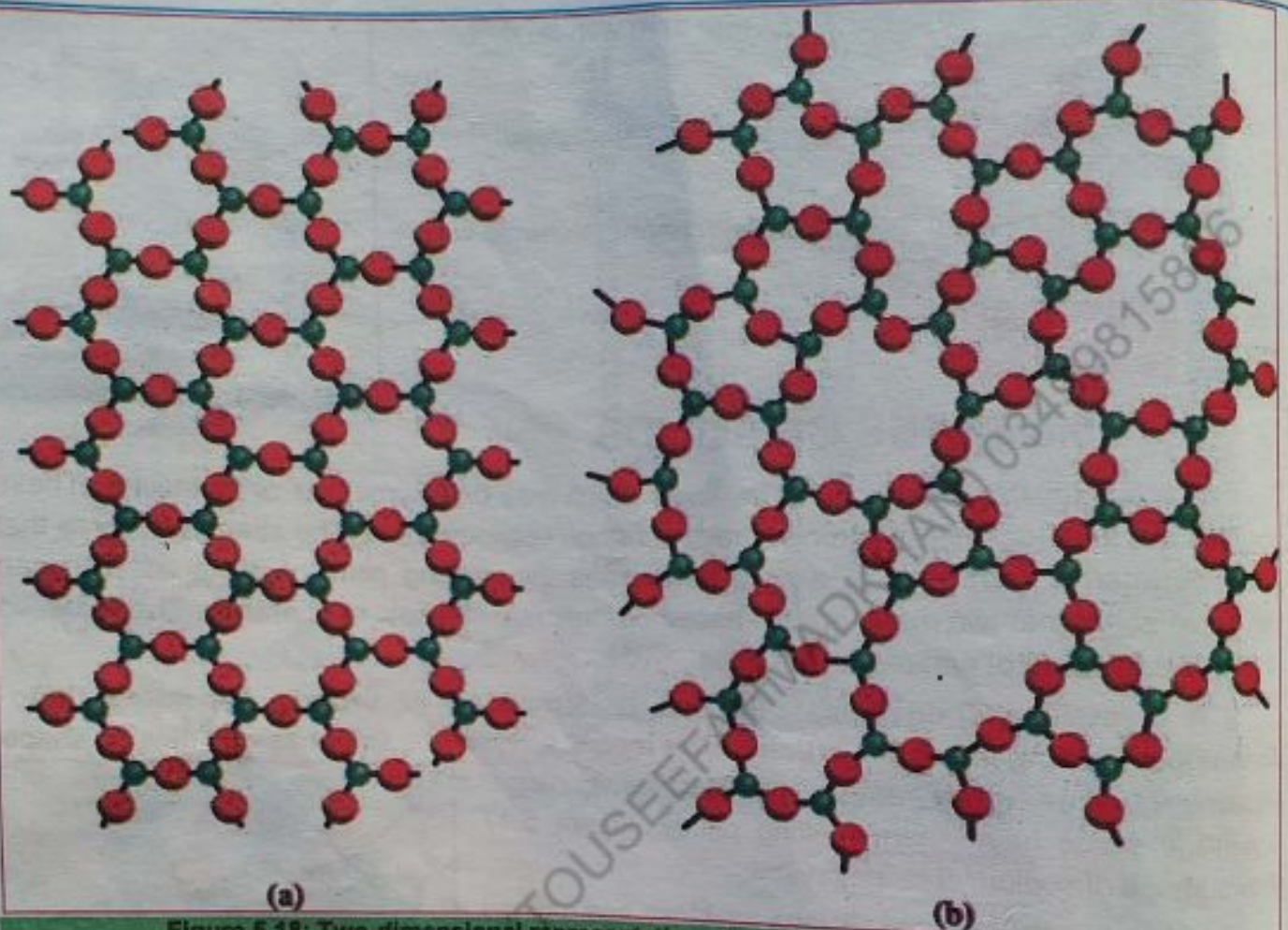


Figure 5.18: Two-dimensional representation of quartz glass and quartz crystal

2. Differentiate between amorphous and crystalline solids by writing appropriate sentences in the empty boxes.



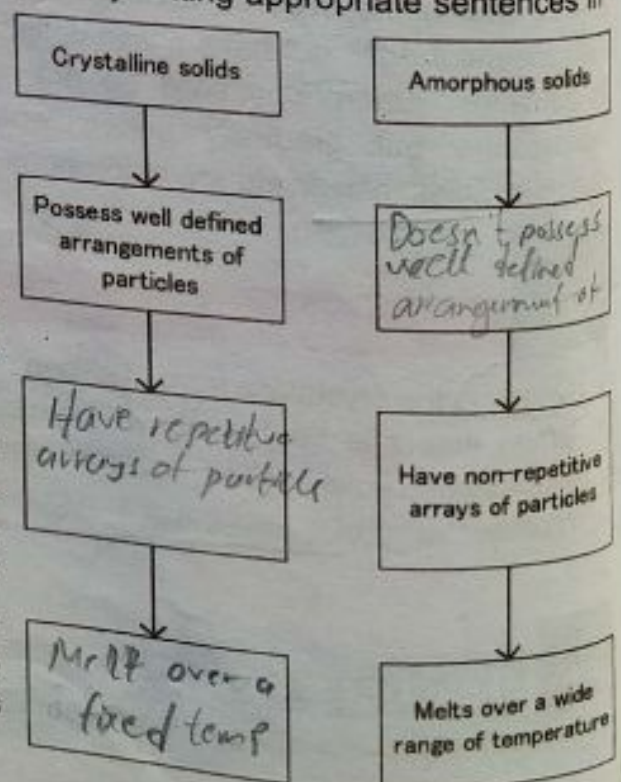
Reading

5.6 ALLOTROPES

Some elements can exist in two or more different arrangements of atoms. Figure 5.19 shows three different arrangements of carbon atoms.

Notice that in diamond each carbon atom is bonded to four others, creating a rigid compact array. This makes diamond the hardest known substance. In graphite carbon atoms are arranged in layers of hexagonal arrays. Weak bonds exist between the layers that allow them to slide over one another. This makes graphite soft. In bucky ball, 60 carbon atoms

Q73





form a hollow sphere. (Carbon atoms are arranged in Diamond, graphite and bucky ball are called allotropes of carbon.) pentagons and hexagons just like a soccer ball. (because all are made of carbon atoms and all are solids.) **The different forms of an element in the same physical state**

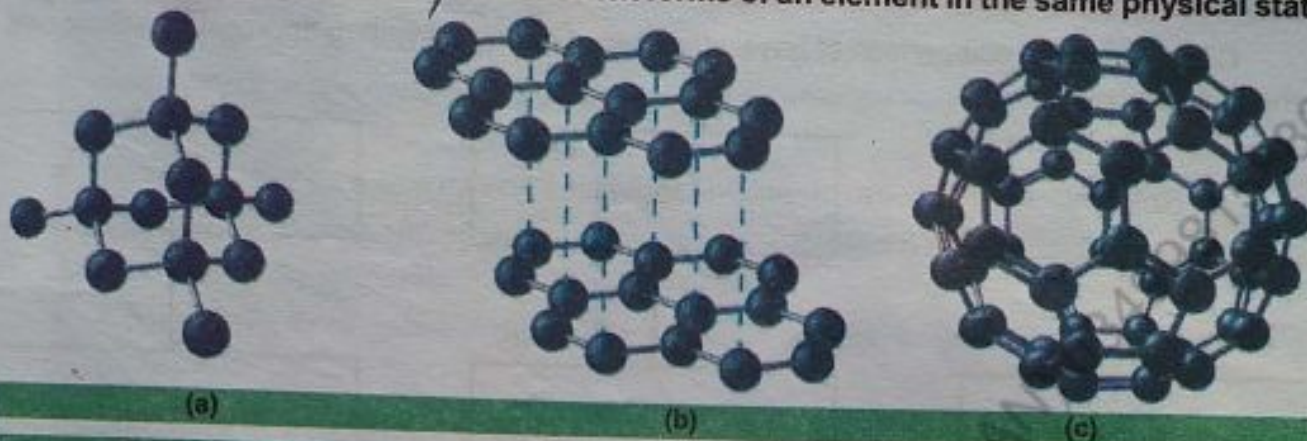


Figure 5.19: Arrangements of carbon atoms in (a) diamond (b) graphite (c) buckyball

and phase are called **allotropic form or allotropes**. Another common non-metal phosphorus is found in two common allotropic forms, known as white phosphorus and red phosphorus. The white phosphorus consists of tetraatomic molecules, P_4 that are not bonded to each other. Whereas in red phosphorus, P_4 molecules are bonded to one another in long chains. Figure 5.20.

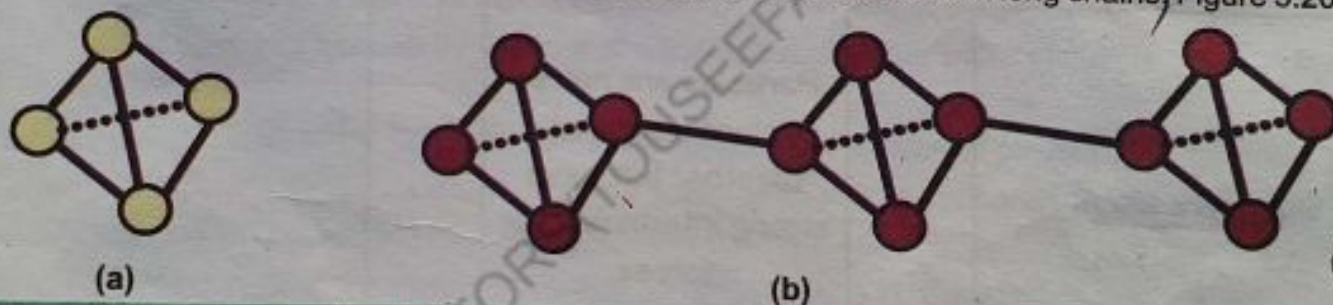


Figure 5.20: (a) white phosphorus (b) red phosphorus

Sulphur consists of molecules that contain eight atoms, S_8 . Sulphur also exists in several crystalline and an amorphous forms. (Rhombic and monoclinic forms are its important crystalline allotropes. Plastic sulphur is the amorphous form.)

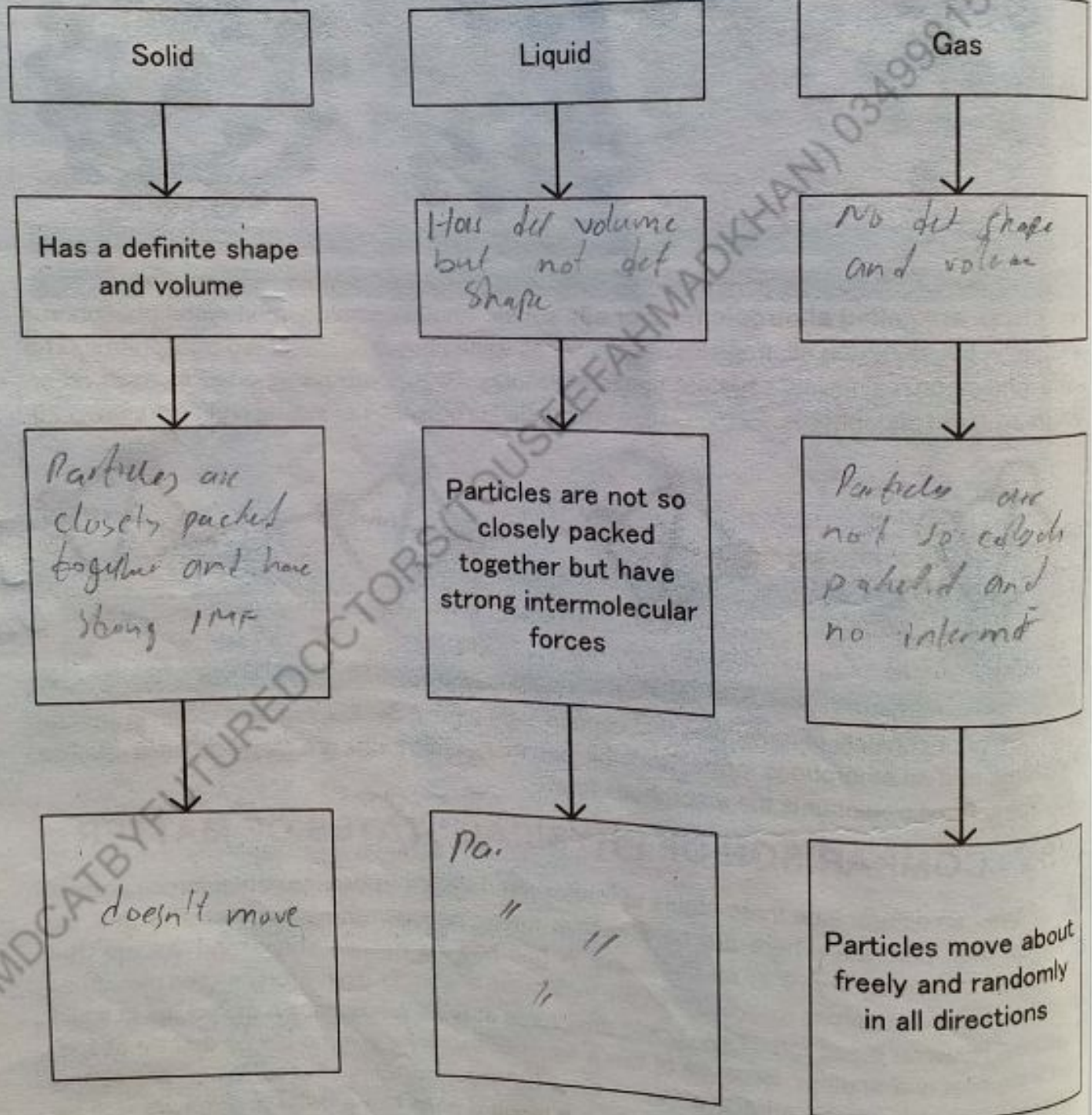
5.7 COMPARISON OF PHYSICAL STATES OF MATTER

We can differentiate three states of matter with regards to intermolecular forces present between their particles. (There are no attractive forces between the molecules of a gas. So molecules can freely move in all directions. A gas has no definite shape and volume.) But (significant attractive forces exist between the molecules of a liquid. These forces keep molecules together. However these forces can't hold molecules in fixed positions. So molecules in liquid can slide past one another. Because of this a liquid has definite volume but no definite shape.) Whereas (in solids strong attractive forces exist between particles. These forces are strong enough to hold particles in fixed positions. As a result a solid has a definite shape as well as definite volume.)



Self-Assessment Exercise 5.8

Copy and make comparison of physical states of matter by writing appropriate sentences in the empty boxes.



**Key Points**

- Diffusion is the movement of molecules from a higher to lower concentration.
- The escape of gas molecules through a hole one after the other without collisions is called effusion.
- Force exerted by a gas on unit area of a container is called its pressure.
- The SI unit of pressure is pascal. One pascal is the force of one newton exerted on an area of one square meter.
- Compressibility of gases is due to the presence of large empty spaces between the gas molecules.
- Gases have tendency to expand and fill the entire available space.
- Boyle's law states that the volume of a fixed amount of a gas at a given temperature is inversely proportional to the applied pressure.
- Charles's law states that the given volume of a gas varies directly with absolute temperature at constant pressure.
- Conversion of a liquid to vapours is called vapourization or evaporation.
- The pressure exerted by vapours of a liquid in equilibrium with its liquid is called vapour pressure.
- A liquid boils when its vapour pressure is equal to the external pressure.
- Distillation is the process in which a liquid is heated to vapourize it and the vapour is cooled to condense it back to the liquid in a different container.
- Melting point is the temperature at which a solid turns in to a liquid.
- Sublimation is the process in which a solid changes into vapours without passing through the liquid state with or without heating.
- A crystalline solid is made up of orderly, repeating three dimensional arrangement of particles.
- The different forms of an element in the same physical state and phase are called allotropes.

REFERENCES FOR ADDITIONAL INFORMATION

- Collins, Advanced Molecular Sciences, Chemistry (AS) .
- Iain Brand and Richard Grime, Chemistry (11-14).
- Bob Berry, IGCSE, study guide for chemistry.



Review Questions

1. Encircle the correct answer:

(i) Which statement about the particles of a solid is not correct?

- (a) They move at great speeds
 (b) They are arranged in regular patterns.
 (c) There is a very little space between the particles.
 (d) The forces of attraction between the particles are strong.

(ii) The boiling points of some elements are given below:

- (a) Argon $\longrightarrow$ -186°C (b) Nitrogen $\longrightarrow$ -196°C
 (c) Oxygen $\longrightarrow$ -183°C (d) Xenon $\longrightarrow$ -108°C

$-210 + 20 = -190$
 A mixture of the above gases at -210°C was heated by 20°C . Which of the element will still be in the liquid state?

- (a) Oxygen only (b) Argon, oxygen and xenon
 (c) Argon, nitrogen and xenon (d) Xenon only (e) Oxygen and Xenon

(iii) The following table shows the melting and boiling points of four substances. Which substance is a solid at room temperature?

Substance	Melting Point ($^{\circ}\text{C}$)	Boiling Point ($^{\circ}\text{C}$)
A	-186	-183
B	-6	120
C	0	100
D	36	130

(iv) In which of the following processes speed of the particles decreases?

- (a) melting (b) boiling (c) sublimation
 (d) condensation (e) evaporation

(v) Consider the following table:

Sr. No	X	Y	X/Y
1	1.00	273	0.0037
2	1.37	373	0.0037
3	1.73	473	0.0037
4	2.10	573	0.0037

$x = y \times \text{constant}$
 $x = Y \times \text{constant}$
 Which of the following relationships does this table represent?

- (a) $x \propto y$ (b) $x = y$ (c) $x \propto \frac{1}{y}$ (d) $x = \frac{1}{y}$

(vi) A liquid boils when its vapour pressure becomes equal to

- (a) 760 cmHg (b) 1Pa (c) 101.325kPa (d) 0.1atm.



- (vii) Acetone has characteristic fragrant odour and is used to make finger nail polish. It melts at -94°C and boils at 56°C . What is the physical state of acetone at 25°C and 1 atm.?
- (a) gas (b) liquid (c) solid (d) cannot be predicated
- (viii) Water normally boils at 100°C , but it is possible to boil water at room temperature. What variable would you have to change to do this
- (a) increase external pressure (b) decrease external pressure
(c) decrease temperature (d) none of these.
- (ix) Bromine has a melting point of -7°C and a boiling point of 59°C . What is the physical state of bromine at 100°C .
- (a) gas (b) liquid (c) solid (d) cannot be predicted.
- (x) Which is not the property of crystalline solids?
- (a) have well defined shape (b) have orderly arrangements of particles
(c) have repeating three dimensional arrangement of particles
(d) are generally soft

2. Give short answers

i. Explain why volume of a gas decreases on increasing pressure on it at constant temperature? 82

ii. How does temperature effect vapour pressure of a liquid? 90

iii. Water boils at 120°C in a pressure cooker, why? $P_{\text{ext}} = 2 \text{ atm}$, $P_{\text{vp}} = P_{\text{ext}}$. B.P at 120°C

iv. Is evaporation a cooling process? When evaporation has occurred.

v. Can you make water boil at 70°C ?

vi. Express the pressure 400 mm Hg in kPa? $760 \text{ mmHg} = 101.325 \text{ kPa} \Rightarrow 400 \text{ mmHg} = \frac{101.325}{760} \times 400 = 53.32 \text{ kPa}$

3. Explain the effect on the volume of a gas by a change in the (a) pressure (b) temperature

4. Explain the following properties of gases (a) Diffusion (b) Effusion

5. Explain the terms (a) evaporation (b) vapour pressure (c) boiling point

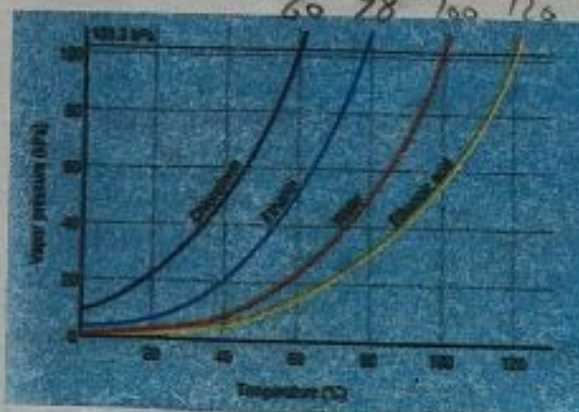
6. Explain the effect of external pressure on boiling point. 91, 92 *same conditions*

7. Differentiate between amorphous and crystalline solids. Def, e.g., M.P., forces

8. Explain the allotropic form of any two solids. 99

9. Plots of vapour pressure versus temperature of four liquids are given in the following figure:

- (a) Find the boiling point of each liquid when the atmospheric pressure is 1 atm.
- (b) At what temperature ethanol will boil when the atmospheric pressure is 51 kPa.
- (c) How can you make water to boil at 80°C ?
- (d) At what temperature chloroform will boil when the external pressure is 50 kPa.
- (e) Can you boil chloroform at 0°C .





- (f) Predict the boiling point of chloroform at 600 mm Hg.
- Q184 11. Why does evaporation lower the temperature of a liquid?
12. The air in a perfectly elastic balloon occupies 885 cm^3 , during the fall when the temperature is 20°C . During the winter, the temperature on a particular day is -10°C , the balloon occupies 794.39 cm^3 . If the pressure remains constant. Show that the given data proves the volume temperature relation according to the Charles's Law.
13. In the past, gas volume was used as a way to measure temperature using devices called gas thermometers. An experimenter obtains following data from a gas thermometer.

Volume (dm^3)	Temperature ($^\circ\text{C}$)
2.7	0°C
3.7	100°C
5.7	300°C

Show that gas thermometer obtained results according to Charles's law.

14. In automobile engine the gaseous fuel-air mixture enters the cylinder and is compressed by a moving piston before it is ignited. If the initial cylinder volume is 990 cm^3 . After the piston moves up, the volume is 90 cm^3 . The fuel-air mixture initially has a pressure of 1.0 atm, and final pressure 11.0 atm. Do you think this change occurs according to the Boyle's law.
- Q2 15. A sample of neon that is used in a neon sign has a volume of 1500 cm^3 at a pressure of 636 torr. The volume of the gas after it is pumped into the glass tube of the sign is 1213.74 cm^3 , when it shows a pressure of 786 torr. Show that this data obeys Boyle's law.
16. Instrumentation changes as science progresses, comments on it?



Think-Tank

1. Following table shows data collected from an experiment by a student:

Volume (dm^3)	Pressure (mm Hg)
400	353.5
320	442
240	589
200	707

Do you think that the student collected data carefully or carelessly? Defend. Interpret it.

2. The water level in an aquarium decreases slowly even though the tank does not leak. Interpret it.
3. Illustrate types of attractive forces you expect between the molecules of HF and HCl?
- Q16 4. Name two substances that are solids at 25°C . Name two substances that are liquid at 25°C .
- Q21 5. Identify the process occurring in each of the following:



- a) Mothballs slowly disappear.
b) A cold windshield becomes covered with ice when struck by raindrops.

6. An autoclave is used to sterilize surgical equipment. It is far more effective to produce steam by autoclave than steam produced from boiling water in the open atmosphere, because it generates steam at a pressure of two atmospheres. Argue why an autoclave is such an efficient sterilization device.

7. The following table shows the melting points and boiling points of four substances.

Substance	Melting point ($^{\circ}\text{C}$)	Boiling point ($^{\circ}\text{C}$)
A	-123	-79
B	-17	58
C	52	305
D	-6	120

- a) Write the physical state of each substance at room temperature and 1 atm.
b) Choose substance that exists as a liquid for the longest range of temperature.
c) Interpret what will happen to the substance B when it is heated from 0°C to 100°C .
d) Interpret what will happen to the substance D when it is cooled from 100°C to -10°C .



6

SOLUTIONS



This is a 18 days lesson

After completing this lesson, you will be able to:

- Explain the difference between saturated, unsaturated and supersaturated solutions.
- Explain the formation of solution (mixing gases into gases, gases into liquids, gases into solids) and give an example of each.
- Explain the formation of solutions (mixing liquids into gases, liquids into liquids, liquids into solids) and give an example of each.
- Explain the formation of solutions (mixing solids into gases, solids into liquids, solids into solids) and give an example of each.
- Explain what is meant by the concentration of a solution.
- Convert between the Molarity of a solution and its concentration in g/dm^3 .
- Define the terms: solution, aqueous solution, solute and solvent and give an example of each.
- Define Molarity.
- Define Percentage composition of a solution. Describe how to prepare a solution of given Molarity.
- Describe how to prepare dilute solution from concentrated solution of known Molarity. Solve problems involving the Molarity of a solution.
- Use the rule that "like dissolves like" to predict the solubility of one substance to another.



Pre- Reading

Why solutions are important for us? We encounter many substances in our daily life such as air, soft drinks, juices, shampoo, petrol, natural gas, diesel, kerosene, cough syrup and many others. Are these substances solutions? Most of the chemical reactions that take place in the bodies of living organisms occur in aqueous solutions.

Brass, steel, German silver are solid solutions. These solutions are widely used for making cooking utensils, surgical tools, cutlery, musical instruments and many other objects. How can we express their composition? This chapter will acquaint you of many things about solutions.



Reading

6.1 SOLUTION, AQUEOUS SOLUTION, SOLUTE AND SOLVENT

A solution is a homogeneous mixture of substances that has uniform composition throughout. In a solution the substance that is present in lesser amount is called solute. Whereas the substance that is present in larger amount is called solvent. For example, air is a solution containing oxygen, carbon dioxide, nitrogen and number of trace gases. In this



solution oxygen, CO_2 and trace gases are solutes, since they are present in lesser quantities whereas nitrogen is solvent, since it is present in larger amount. Consider another example, if you dissolve 10 g sugar in 100 g of water you will get sugar solution. Here sugar is solute and water is solvent. **A solution in which water is solvent is called an aqueous solution.** The word aqueous is derived from the Latin word **aqua** meaning water. Aqueous solutions are frequently used to dispense substances such as medicines.

6.2 SATURATED, UNSATURATED, SUPERSATURATED SOLUTIONS AND DILUTION OF SOLUTION



Activity 6.1

Take 100g water in a beaker. Add a tea spoon of sugar in it. Stir it, the sugar will dissolve. Repeat the process the added sugar will again dissolve in it. **A solution which can dissolve more of the solute at a given temperature is called an unsaturated solution.**

Go on adding sugar in the above solution till it starts settling down at the bottom of the beaker at a particular temperature. **The solution which cannot dissolve more solute at a particular temperature is called a saturated solution.**

Now heat the solution, stir it, add more sugar, it will dissolve. Go on adding more sugar and stir it. A stage will reach when no more sugar will dissolve and will start settling down at the bottom of the beaker. This solution is called supersaturated solution. **A solution that contains more of the solute than is contained in the saturated solution is called supersaturated solution.** How to know whether a solution is saturated or supersaturated? A supersaturated solution is not stable in the presence of crystals of solute. If you add a crystal of sodium thiosulphate to its saturated solution, it will simply drop to the bottom, without dissolving, but if you add a crystal of sodium thiosulphate to a supersaturated solution of sodium thiosulphate (see figure 6.1a), crystallization will start. When crystallization has finished, you will have a saturated solution in presence of sodium thiosulphate crystals. Figure 6.1(b) shows the crystallization from a supersaturated solution of sodium thiosulphate.

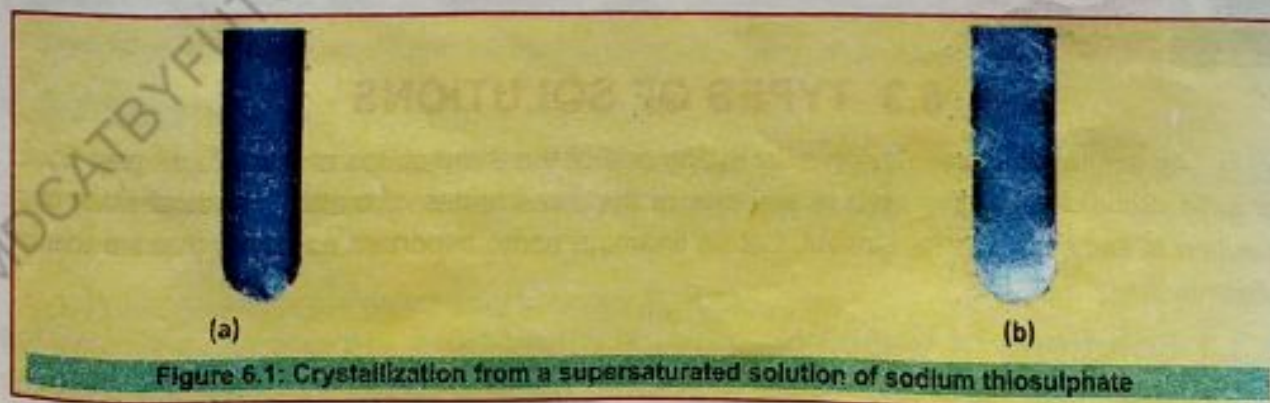


Figure 6.1: Crystallization from a supersaturated solution of sodium thiosulphate



Activity 6.2

Differentiating saturated, unsaturated and supersaturated solutions.

• **You will need:**

- Three test tubes.
- Sodium thiosulphate, $\text{Na}_2\text{S}_2\text{O}_3$
- Water

Carry out the following:

1. Take 20cm^3 water in each test tube and label them as A, B, C.
2. Add 4g $\text{Na}_2\text{S}_2\text{O}_3$ in A, 12g in B and 16g in C and shake.
3. Heat test tubes C till $\text{Na}_2\text{S}_2\text{O}_3$ is dissolved completely.
4. Cool test tube C at room temperature without disturbing.
5. Write down what happens in test tube C before heating and after heating.
6. What do you observe in test tube C after cooling?
7. Compare test tubes B and C and notice difference.
8. Identify test tubes containing unsaturated, saturated and supersaturated solution.



Self-Assessment Exercise 6.1

The maximum amount of sodium acetate that dissolves in 100g of water at 0°C is 119g and 170g at 100°C .

- (a) If you add 170g of sodium acetate in 100g of water at 0°C , how much will dissolve? *119g*
- (b) Is the solution saturated unsaturated or supersaturated?
- (c) If the solution is heated to 100°C , is the solution now saturated, unsaturated or supersaturated.
- (d) If the solution is cooled back to 0°C and no crystals appear. Is the solution saturated, unsaturated or supersaturated? *✓*



Reading

6.3 TYPES OF SOLUTIONS

As solute and solvents can exist in any one of the three states of matter i.e., gas, liquid or solid. Solution can also exist in any one of the three states of matter. Physical state of a solution is same as that for solvent. Let us examine some important solutions that we use or observe in our daily life.

6.3.1 Solutions of Gases *→ solute = gases*

Gaseous solutions are commonly used by chemical industries to prepare chemical substances. For example a gaseous mixture of Nitrogen (N_2) and hydrogen (H_2) is used to



prepare ammonia gas (NH_3). Ammonia is used to prepare fertilizers and nitric acid. For preparing fertilizer Urea, a gaseous mixture of NH_3 and CO_2 is used. For preparing nitric acid, mixture of NH_3 gas and oxygen gas is used. In all these cases a gaseous mixture or solution of gases is used. In these solutions, solute and solvent both are gases. The gas which is present in greater proportion is solvent and the other gas is solute. You can identify this solution as gas in gas.

Gases when dissolved in liquid also produce some important solutions. For example (water dissolves small amount of air giving a solution whose oxygen content (solute) is important for the survival of fish and other aquatic animals. Carbon dioxide gas dissolves readily in water, for this reason it is used in making carbonated drinks.)

In food industries, vegetable oil is converted into vegetable ghee by passing H_2 gas through the oil. Finely divided Ni is used as catalyst. Ni accelerates this process by adsorbing H_2 gas on its surface, producing a solution. What type of solution is this?

6.3.2 Solutions of Liquids $\rightarrow$ solute = liquid

Solution of liquid in gas, liquid or solid solvents are also very common. For instance, fog is a solution of water vapours in air. Other examples of similar solutions are clouds, mist etc. In these solutions solute is in the liquid state and solvent in the gaseous state.

Fermentation of cane sugar produces 95% v/v of ethyl alcohol called rectified spirit. This means every 100cm^3 of this solution contains 95cm^3 ethyl alcohol and 5cm^3 water. Vinegar is 5% (m/m) aqueous solution of acetic acid. In these solutions solute is a liquid dissolved in another liquid. We call such a solution as liquid in liquid solution.

The metal mercury is the only metal that is liquid at room temperature. It dissolves a number of metals to give a solution called amalgam. A solution of any metal in mercury is called an amalgam. Silver and tin amalgams are widely used to make dental filling. When silver or tin is dissolved in mercury, it forms a semi solid amalgam that can be shaped to fill a cavity. On standing it forms a hard solid and expands slightly. Therefore, it tightly fits within the cavity. Thus in amalgams solute is a liquid and solvent is solid, we call such a solution as liquid in solid.

6.3.3 Solutions of Solids $\rightarrow$ solute = solid

You must have observed smoke. It spreads in air forming solution that contains solid carbon particles. In this solution solid particles are solute and air is solvent. We call such a solution as solid in gas.

Water dissolves a wide range of substances. It dissolves many minerals in small amounts. For instance it dissolves limestone, which is a common component of well water. It contains dissolved salts such as NaCl . You also prepare many solutions by dissolving solid substances in water, such as common salt, sugar etc. In Laboratories mostly solutions are prepared by

Important information

Ozone is found in the upper atmosphere. It has an important biological function. It prevents most of sun's ultraviolet solar radiation from reaching the earth surface. Ultraviolet radiation causes changes in the structure of the genetic material, like DNA. Long exposure to this radiation can cause cancer.



dissolving solid substances in water. 0.85% m/v NaCl solution is used in intravenous solution that is given to persons suffering from dehydration. Such solutions that contain solid solute dissolved in a liquid are called as solids in liquids.

We often prepare a solution of solid in solid because they are used commercially. For example, pure gold is very soft, therefore cannot be used for making Jewelry. To make it harder, copper is added to the gold. This produces a solid solution of gold that melts at lower temperature than pure gold. Therefore, it is easier to cast. Most commercial metals are examples of solid solutions of various metals called alloys. For example, brass is an alloy of copper and zinc. Steel is an alloy of iron containing small amounts of carbon and silicon. Many naturally occurring gemstones are solid solutions. For example Ruby, Opal, in these solutions a solid solute dissolves in a solid solvent. We call these solutions as solids in solids. Table 6.1 shows nine types of solution, we have discussed in sections 6.3.1, 6.3.2 and 6.3.3.

Table 6.1: Common types of solutions

Solute	Solvent	State of resulting solution	Examples
Gas	Gas	Gas	Air
Gas	Liquid	Liquid	Soda water
Gas	Solid	Solid	H ₂ absorbed on Ni, , Pb, Pd
Liquid	Gas	Gas	Mist, fog, clouds
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Solid	Solid	Amalgams
Solid	Gas	Gas	Carbon particles in air
Solid	Liquid	Liquid	Sugar in water
Solid	Solid	Solid	Alloys



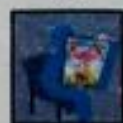
Self-Assessment Exercise 6.2

What are the physical states of solute and solvent in each of the following solutions. Also identify the type of solution.

- Deep sea divers use a breathing mixture of helium and oxygen. *Su = gas, Sv = gas*
- Brass contains 80% copper and 20% zinc. *Su = solid, Sv = solid, Type of sol = gas in gas*
- Dental filling. *Su = liquid, Sv = solid, Type of sol = gas in gas*
- Brine (salt in water). *Su = solid, Sv = liquid, Type of sol = liquid in solid*



- (e) Drinking water containing chlorine as disinfectant.
- (f) Gemstone, Ruby contains Cr_2O_3 and Al_2O_3 .
- (g) Conc. H_2SO_4 , we use in the laboratory is 98% H_2SO_4 and contains only 2% H_2O .

**Reading****6.4 CONCENTRATION UNITS**

When solid solutes are dissolved in water, their molecules or ions readily move about in solution. They come in contact with one another readily in solution, so they combine easily. **The quantity of a solute present in a given amount of solvent or solution is called concentration of solution.**

In grade VI, you have learned the difference between a dilute and a concentrated solution. A dilute solution is that whose concentration is relatively low while a concentrated solution is that solution whose concentration is relatively high. You can find dilute and concentrated solutions of common acids and bases in your chemistry laboratory. Chemists use many concentration units. Here we will discuss only percentage and molarity, other concentration unit you will learn in grade XI.

6.4.1 Percentage

Commercially available solutions that contain the maximum available concentration of solutes are referred as concentrated solutions. For example commercially available concentrated H_2SO_4 is 98% and conc HCl is 37%. This means concentrated H_2SO_4 contains 98g H_2SO_4 in 100g of solution and concentrated solution of HCl contains 37g HCl in 100g of solution. Thus percentage is the unit of concentration that specifies the quantity of solute in 100 parts of solution. Quantity of solute and solvent can be expressed by mass in grams or volume in cm^3 . Therefore, **by the percentage of a solution we mean the mass or volume of solute dissolved in 100g or 100cm^3 of solution.** For example, if you dissolve 10g NaCl in 90g water to make 100g of solution, the concentration of solution will be 10% m/m. But if you dissolve 10cm^3 of alcohol in sufficient water to make 100cm^3 of solution, the solution will be 10% v/v. If you dissolve 10g NaCl in sufficient water to make 100cm^3 solution, the resulting solution will be 10% m/v. If you dissolve 10cm^3 of alcohol in sufficient water to make 100g of solution, the solution will be 10% v/m. Therefore, there are four ways to express percentage of solutions.

**Self-Assessment Exercise 6.3**

- Write four ways to express percentage of solutions.
- A saline solution is administered intravenously to a person suffering from severe dehydration. This is labeled as 0.85% m/v of NaCl . What does it mean?

**Reading****6.4.2 Molarity (M)**

If you read label on the bottle of concentrated H_2SO_4 you will notice 98% H_2SO_4 by mass and also 18M H_2SO_4 . What does 18M stand for? This means there are 18 moles of H_2SO_4 in each dm^3 of solution. Similarly conc. HCl is 37% and 12.1 M HCl. This means there are 12.1 moles of HCl in each dm^3 of solution. We can express the concentration in terms of moles of solute in the given volume of solution.

Molarity is the concentration unit in which amount of solute is expressed in moles and quantity of solution in dm^3 .

"Molarity is defined as the number of moles of solute dissolved per dm^3 of solution".

Mathematically,

$$M = \frac{\text{moles of solute}}{\text{dm}^3 \text{ of solution}}$$

6.4.3 Problems Involving the Molarity of a Solution

Urea (NH_2CONH_2) is a white solid used as fertilizer and starting material for synthetic plastic. A solution contains 40g urea dissolved in 500cm^3 of solution. Calculate the molarity of this solution.

Problem Solving Strategy:

To calculate the molarity you need the number of moles of solute (which are not given) and the volume of solution in dm^3 . Convert volume of solution 500 cm^3 to dm^3 . Determine moles of solute from its mass using its molar mass.

Solution:

Mass of urea = 40g

Molar mass of urea (NH_2CONH_2) = $14 + 1 \times 2 + 12 + 16 + 14 + 1 \times 2$
= 60g/mole

Moles of urea = $\frac{40\text{g}}{60\text{g/mole}} = 0.667\text{ moles}$

Volume of solution = $500\text{cm}^3 = \frac{500}{1000} = 0.5\text{dm}^3$

Now

Molarity = $\frac{\text{moles of solute}}{\text{dm}^3 \text{ of solution}}$

Molarity = $\frac{0.667}{0.5}$

= 1.334 M

Example 6.1: Calculating molarity from moles of solute and volume of solution



Potassium permanganate (KMnO_4) is a dark blue-black compound. When it is dissolved in water, it forms a bright purple solution. It is used as disinfectant in water tanks. It is also known as pinky. A solution contains 0.05 moles of KMnO_4 in 600cm^3 of solution. Calculate molarity of this solution.

Problem Solving Strategy:

To calculate molarity, you need moles of solute which are given and volume of solution in dm^3 . But volume is given in cm^3 . So convert cm^3 to dm^3 by dividing with 1000. Use formula to calculate molarity.

Solution:

$$\begin{aligned}\text{Moles of } \text{KMnO}_4 &= 0.05 \\ \text{Volume of solution} &= 600 \text{ cm}^3 \\ &= \frac{600}{1000} \\ &= 0.6 \text{ dm}^3\end{aligned}$$

Now

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{dm}^3 \text{ of solution}}$$

$$\begin{aligned}\text{Molarity} &= \frac{0.05}{0.6} \\ &= 0.083\text{M}\end{aligned}$$



Self-Assessment Exercise 6.4

Potassium chlorate (KClO_3) is a white solid. It is used in making matches and dyes. Calculate the molarity of solution that contains. (a) 1.5 moles of this compound dissolved in 250cm^3 of solution (b) 75g of this compound dissolved to produce 1.25dm^3 of solution. (c) What is the molarity of a 50cm^3 sample of potassium chlorate solution that yields 0.25g residue after evaporation of the water.

Example 6.2: Converting the molarity of a solution into its concentration in g/dm^3

A flask contains 0.25M NaOH solution. What mass of NaOH is present per dm^3 of solution?

Problem Solving Strategy:

Molarity means number of moles per dm^3 of solution. So 0.25M NaOH means 0.25 moles of NaOH dissolved per dm^3 of solution. You need to convert moles of solute to mass using molar mass.

**Solution:**

$$\text{Molarity} = 0.25\text{M}$$

$$\text{Moles of NaOH} = 0.25$$

$$\text{Volume of solution} = 1\text{dm}^3$$

$$\begin{aligned}\text{Molar mass of NaOH} &= 23+16+1 \\ &= 40\text{g/mole}\end{aligned}$$

$$\text{Moles of solute} = \frac{\text{mass of solute}}{\text{molar mass of solute(g/mole)}}$$

$$0.25 \text{ moles} = \frac{\text{mass of NaOH}}{40\text{g/mole}}$$

$$\begin{aligned}\text{Mass of NaOH} &= 0.25 \text{ moles} \times 40\text{g/mole} \\ &= 10\text{g}\end{aligned}$$

Thus, solution contains 10g NaOH per dm^3 .

Example 6.3: Converting concentration in g/dm^3 into molarity

Potassium hydroxide (KOH) is used in the manufacture of shaving creams, paints and varnish. An analyst makes up a solution by dissolving 5.6g of KOH in one dm^3 of solution. Calculate the molarity of this solution.

Problem solving strategy:

To calculate the molarity, you need moles of solute per dm^3 of solution. Moles of solute are not given. But mass of solute per dm^3 of solution is given. Convert mass of solute into moles by using its molar mass.

Solution:

$$\text{Mass of KOH dissolved in one dm}^3 \text{ of solution} = 5.6\text{g}$$

$$\begin{aligned}\text{Molar mass of KOH} &= 39+16+1 \\ &= 56\text{g/mole}\end{aligned}$$

$$\begin{aligned}\text{Moles of KOH} &= \frac{\text{mass of KOH(g)}}{\text{molar mass of KOH(g)}} \\ &= \frac{5.6}{56} \\ &= 0.1 \text{ mole}\end{aligned}$$

Thus, solution contains 0.1 mole of KOH in one dm^3 of solution, so molarity of solution is 0.1M.

**Teacher's Point**

A teacher may give different concentration interconversion numericals to students as home work.

**Self-Assessment Exercise 6.5**

1. Sodium hydroxide solutions are used to neutralize acids and in the preparation of soaps and rayon. If you dissolve 25g of NaOH to make 1 dm³ of solution, what is the molarity of this solution?
2. A solution of NaOH has concentration 1.2M. Calculate the mass of NaOH in g/dm³ in this solution.
3. A solution is prepared by dissolving 10g of haemoglobin in enough water to make up 1dm³ in volume. Calculate molarity of this solution. Molar mass of haemoglobin is 6.51x10⁴g/mole.

Example 6.4: Preparing a solution of given molarity

Prepare 0.2M KMnO₄ solution. *of 100 cm³ volume*

Problem Solving Strategy:

For preparing solution you need a volumetric flask (figure 6.2). A volumetric flask has a long neck with an etched line. This line indicates the exact volume, when the flask is filled to this line. Volumetric flasks of capacity 1 dm³, 500cm³, 250cm³, 100cm³, 50cm³ can be used. Suppose you use a 100cm³ volumetric flask. First find the mass of KMnO₄ to give 100 cm³ of 0.2 M KMnO₄ solution.

Solution:

$$\text{Required volume of solution} = 100\text{cm}^3 = \frac{100}{1000} = 0.1\text{dm}^3$$

0.2 M KMnO₄ means:

1dm³ solution contains 0.2 moles of KMnO₄

$$\begin{aligned} \text{So, } 0.1\text{dm}^3 \text{ solution contains} &= 0.2 \times 0.1 \text{ moles of KMnO}_4 \\ &= 0.02 \text{ moles of KMnO}_4 \end{aligned}$$

Now convert moles of KMnO₄ into mass using molar mass of KMnO₄

$$\begin{aligned} \text{Molar mass of KMnO}_4 &= 39+55+16 \times 4 \\ &= 39+55+64 \\ &= 158\text{g/mole} \end{aligned}$$

$$\text{Moles of KMnO}_4 = \frac{\text{mass of KMnO}_4}{\text{molar mass of KMnO}_4}$$

$$0.02\text{mole} = \frac{\text{mass of KMnO}_4}{158\text{g/mole}}$$

$$\text{Mass of KMnO}_4 = 3.16\text{g}$$



To prepare 0.2M KMnO_4 in 100cm^3 volumetric flask, follow the following steps. Figure 6.2 shows these steps.



(a)



(b)



(c)

Figure 6.2: Preparation of 0.2M KMnO_4 solution

- Weigh out 3.16g of KMnO_4 (0.02 moles)
- Add this solid into a beaker, add some water to dissolve it.
- Transfer this solution to the 100 cm^3 volumetric flask and add more water.
- Keep adding water until the volume of solution rises to the etched line and mix the solution. This is desired solution.



Self-Assessment Exercise 6.6

- How can you prepare 500cm^3 of 0.2M KMnO_4 solution.
- How can you prepare 25cm^3 of 0.25M solution of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (Blue vitriol).

Example 6.5: Preparing a solution of given molarity by diluting a solution of known

Concentrated sulphuric acid is 18M H_2SO_4 . How many cm^3 of this acid is needed to produce 250cm^3 of 0.1M H_2SO_4 ?

Problem Solving Strategy:

Find the volume of given conc H_2SO_4 to be diluted to 250cm^3 to make 0.1 M H_2SO_4 . Use the molarity equation.

Solution:

Given H_2SO_4

M_1V_1

=

Desired H_2SO_4

M_2V_2

M_1 = molarity of given conc H_2SO_4 = 18M

V_1 = volume of conc H_2SO_4 needed to dilute = ?

M_2 = molarity of required H_2SO_4 solution = 0.1M



V_2 = volume of required H_2SO_4 = 250cm^3

$$\begin{aligned} M_1 V_1 &= M_2 V_2 \\ V_1 &= \frac{M_2 V_2}{M_1} \\ &= \frac{0.1 \times 250}{18} \\ &= 1.39\text{cm}^3 \end{aligned}$$

Transfer 1.39cm^3 of $18\text{M H}_2\text{SO}_4$ to a 250cm^3 volumetric flask and dilute it by adding water up to the mark and mix. Resulting solution is $0.1\text{M H}_2\text{SO}_4$.



Self-Assessment Exercise 6.7

1. A stock solution of hydrochloric acid is 12.1 M . How many cm^3 of this solution should you use to prepare 500cm^3 of 0.1 M HCl .
2. Potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$) is a red-orange compound. It is a strong oxidizing agent and is used in the estimation of iron content in ores. A stock solution is $2.5\text{M K}_2\text{Cr}_2\text{O}_7$. How many cm^3 of this solution you need to dilute to make 50 cm^3 of $0.05\text{ M K}_2\text{Cr}_2\text{O}_7$.
3. Commercial acetic acid is 17.8 Molar . How can you convert this into 0.1 M acetic acid.



Reading

6.5 SOLUBILITY

Recall that a solution that contains maximum amount of solute in a given amount of solvent at a particular temperature is called saturated solution.



Activity 6.3

Prepare saturated solution of sodium chloride in 100g of water. Take this solution in a pre-weighed china dish. Place china dish on the burner and heat it slowly till water evaporates completely. Cool china dish and weigh it. Calculate the mass of sodium chloride present in it.

Solubility of sodium chloride in 100g of water at room temperature is 35.7g . Similarly solubility of sodium thiosulphate is 50g per 100g of water. If you add 60g of sodium thiosulphate in 100g of water at 20°C , what will happen?

The amount of solute that dissolves in 100g of a solvent at a particular temperature is called its solubility.



Teacher's Point

A teacher may ask the students to check solubility of different substances practically.



Reading

6.5.1 Solubility and Solute-Solvent Interactions

The chemistry saying that "like dissolves like" is a guiding rule when considering the solubility of substances. Recall nature of intermolecular forces discussed in section 4.4.

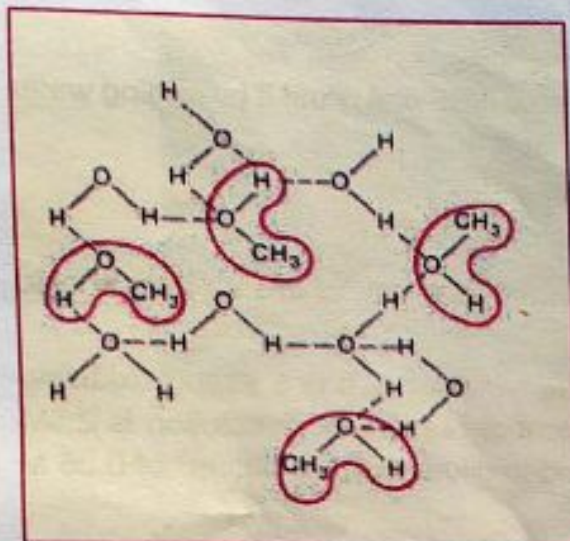


Figure 6.3: H-bonding between water and methanol molecules

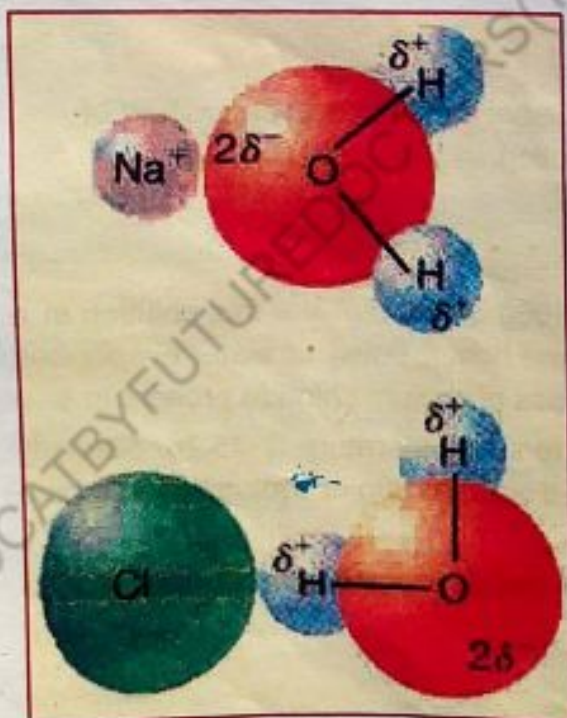


Figure 6.4. Attraction of Na^+ and Cl^- ions for water molecules

CH_3OH readily dissolves in water. To find the reason, look into the nature of intermolecular forces among their molecules. Water molecules are polar. Two H-atoms bonded to an O-atom are slightly positively charged and O-atom has a slight negative charge. Water molecules form hydrogen bonds with one another. (Methanol molecules are also polar and exhibit hydrogen-bonding. This means water and methanol molecules have similar structures and intermolecular forces. They can form hydrogen bonds with each other (see figure 6.3). Thus methanol and water are miscible.)

Similarly glucose, whose molecule has many $-\text{O}-\text{H}$ bonds, is very soluble in water. When you place a crystal of sodium chloride in water, it dissolves. Sodium chloride is an ionic compound. The negative end of water molecules is attracted to sodium ions and the positive end of water molecules is attracted to chloride ions. These attractive forces are strong enough to overcome the strong attractions that exist between ions in NaCl . Thus sodium chloride dissolves readily. Figure 6.4 shows attraction of Na^+ and Cl^- ions for water molecules.

On the other hand, gasoline and oils do not dissolve in water. Look into the nature of gasoline and oil molecules. (Gasoline and oil molecules are non-polar in nature, the attraction between a water molecule and an oil-gasoline molecule is very weak, so these liquids are insoluble in water.) But Gasoline and oil, which have similar non-polar molecules, are soluble in one another.

Thus we conclude substances that have similar structures and intermolecular forces tend to be soluble. Whereas substances that have dissimilar structures and intermolecular forces are insoluble.



Self-Assessment Exercise 6.9

1. Sodium Chloride and glucose both are soluble in water. But the Solubility of NaCl is greater than glucose. Explain why?
2. In which liquid of each of the following pairs you would expect KCl, an ionic solid, to be more soluble.
(a) H_2O or CCl_4 (b) CH_3OH or Benzene
3. Which of the following pairs of liquids are miscible?
(a) Water and benzene (b) Benzene and CCl_4 (c) An oil and benzene



Reading

6.5.2 Effect of Temperature on Solubility

If you add 34.7g of KCl to 100g of water at 20°C , it will dissolve. If you add more than 34.7g of KCl at 20°C , it will not dissolve. However, if you increase temperature it will readily dissolve. Keep on adding more KCl and increase temperature. You will observe that 56.7g of KCl dissolves in 100g of water at 100°C , when temperature of this solution is decreased to 20°C , 22.0g of KCl will crystallize out. This means temperature has a profound effect on the solubility of a substance. Figure 6.5 shows the variation of solubility with temperature.

Note that the solubility of an ionic compound generally increases with the increase in temperature. However, solubility of some solids decreases with temperature. For instance, solubility of Na_2SO_4 decreases with increase in temperature.

Heat water in a beaker, you will see small bubbles form at the side of the beaker before the water boils. These bubbles are composed of air. Since, air is less soluble in hot water than in cold water, air comes out of water in the form of bubbles. This means solubility of air in water decreases with increasing temperature. You might have observed in a home aquarium, that the fish show signs of stress on a hot day. This is because less oxygen from air dissolves in the warm water.

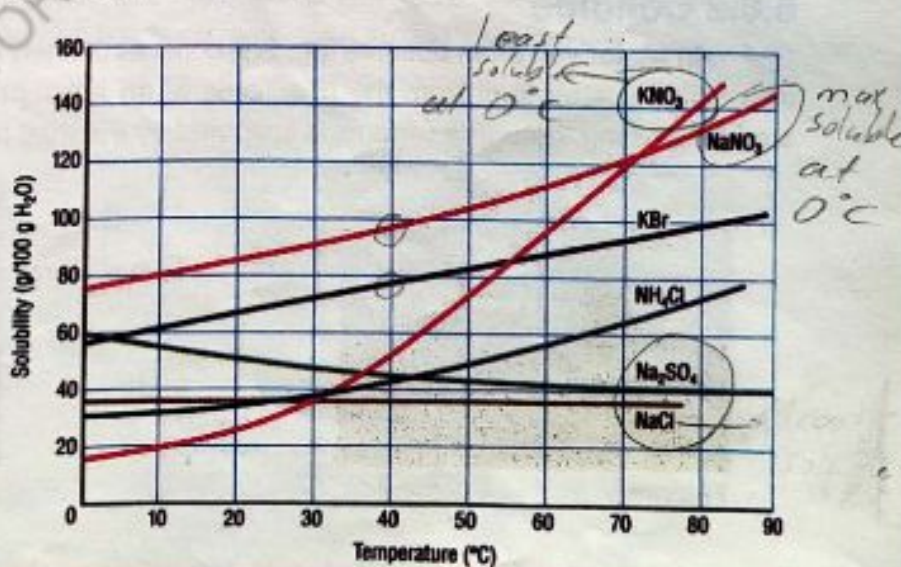


Figure 6.5: Variation of solubility with temperature



Self-Assessment Exercise 6.9

Use figure 6.5 and answer the following questions:

1. At what temperature the solubility of KNO_3 and KBr is same. 55°C
2. What is the solubility of KBr at 45°C . $80\text{ g}/100\text{ g H}_2\text{O}$
3. Which is greater at 40°C , the solubility of NaNO_3 or solubility of KBr .
4. Identify from the graph the compound whose solubility is little affected with increase in temperature. NaCl



Reading

6.6 COMPARISON OF SOLUTION, SUSPENSION AND COLLOIDS

In this section, you will learn the differences between solution, suspension and a colloid.

6.6.1 Solutions *e.g. = solutions of acids/bases in water OR aqueous sol. of acids/bases*

A solution is a homogeneous mixture in which the particles are individual molecules or ions distributed evenly throughout the surrounding fluid. Sizes of these particles are between $0.1\text{-}1\text{ nm}$. Therefore, these particles cannot be seen by the naked eye, ordinary microscope and electron microscope. These particles can pass through ordinary and ultra filter papers. A solution is unable to scatter light.

6.6.2 Colloids

In winter you often observe fog. Fog consists of finely divided droplets of water suspended in air. Fog scatter light from the headlights of an automobile. You might have observed that sunlight shining through a window is scattered by the dust particles. See figure 6.6.



Figure 6.6 (a): Fog is a colloid



(b): Dust particles in air is a colloid

Fog and dust particles suspended in air are colloids; Colloids differ from solution in scattering light (see figure 6.7). Solutions cannot scatter light.

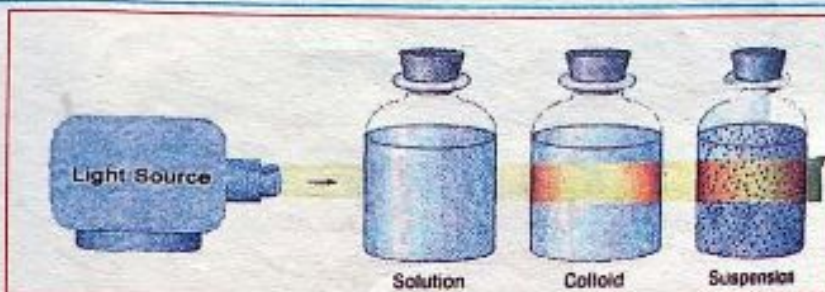


Figure 6.7: Scattering of light by colloid and suspension

A heterogeneous mixture of tiny particles of a substance dispersed through a medium is called a **colloid**. Apparently, a colloid looks homogenous like solutions. In a colloid dispersed phase is like solute in solutions and dispersion medium is like solvent in the solutions. Colloids are generally opaque but some are transparent as well. The size of colloidal particles varies from $1-10^3\text{nm}$. These particles can only be seen under electron microscope. Colloids can scatter light. Milk, butter, cheese, jam, jellies, halva, mayonnaise, are other examples of colloids. You will learn more about colloids in grade XI.

Science titbits

Many medicine bottles contain an insoluble solid in water. The bottle has to be shaken before use to produce a suspension, so that the solid is spread evenly throughout the bottle and the patient takes the correct amount of the medicine.

6.6.3 Suspensions

Mix a small quantity of sand in a glass of water. You will observe sand particles remain suspended at first, then gradually settle to the bottom. **Such a heterogeneous mixture containing particles large enough to be seen with the naked eye and clearly distinct from the surrounding fluid is called a suspension.** Particles of a suspension are bigger than 10^3nm and cannot pass through ordinary filter paper. Because of their large size, they can scatter light like colloids. Table 6.2 shows comparison of solutions, suspensions and colloids.

Table 6.2: Comparison of properties of solutions, suspension and colloids.

S.No.	Solutions	Suspensions	Colloids
1.	Homogeneous	Heterogeneous	Heterogeneous
2.	Particles size vary from 0.1 to 1nm	Particles size is greater than 10^3nm	Particles size vary from 1 to 10^3nm
3.	Particles are invisible by naked eye, ordinary microscope as well as electron microscope	Particles are visible by naked eye	Particles are invisible by naked eye and in ordinary microscope but visible under electron microscope
4.	Particles can pass through ordinary as well as ultra filter paper	Particles can not pass through ordinary as well as ultra filter paper	Particles can pass through ordinary filter paper but cannot pass through ultra filter paper
5.	Cannot scatter light	Scatter light	Scatter light



Society, Technology and Science

Most of the substances we need for our existence are solutions. The air we breathe is a gaseous solution containing N_2 , O_2 , CO_2 and rare gases. The water we use for drinking, cooking and washing is not pure. It contains dissolved gases and many minerals that are essential for our health. In fact natural water is a liquid solution. Beverages, vinegar, soft drinks etc are liquid solutions. Commercial products such as window cleaners, sanitary cleaner, shampoo, gasoline, kerosene, diesel, etc are also liquid solutions. Most medicines are dispensed in solution form. We also use many solid solutions in our daily lives. Gold is a solid solution of gold containing some copper. Brass and steel used for making utensils, musical instruments, buses, cars, trains etc are solid solution of metals. Parts of aeroplane are made of solid solution of metals such as Al and Mg. Dental fillings are liquid solutions of metals in mercury.



Key Points

- A solution is a homogeneous mixture of substances that has uniform composition throughout.
- A solution in which water is solvent is called an aqueous solution.
- A solution in which maximum amount of solute has been dissolved at a particular temperature is called saturated solution.
- A solution that contains more of a solute than is contained in the saturated solution is called supersaturated solution.
- Most commercial metals are examples of solid solutions of various metals and are called alloys, for example gold.
- The quantity of a solute present in a given amount of solvent or solution is called concentration.
- By the percentage of a solution we mean the mass or volume of solute dissolved in 100g or 100cm^3 of solution.
- Molarity is the number of moles of solute dissolved per dm^3 of solution.
- The amount of solute that dissolves in 100g of a solvent at a particular temperature is called solubility.
- A heterogeneous mixture of tiny particles of a substance dispersed through a medium is called a colloid.
- A heterogeneous mixture containing particles large enough to be seen with the naked eye and clearly distinct from the surrounding fluid is called a suspension.

REFERENCES FOR ADDITIONAL INFORMATION

- Raymond Chang, Essential Chemistry.
- David E. Goldberg, Fundamentals of Chemistry.



Review Questions

1. Encircle the correct answer:

- (i) The maximum amount of sodium acetate that dissolves in 100g of water at 0°C is 119g and 170g at 100°C. If you place 170g of sodium acetate in 100g of water at 0°C, the resulting solution would be
 (a) unsaturated (b) saturated (c) supersaturated (d) 1M
- (ii) How many moles of sodium atoms are present in 2.3g Na? $n = \frac{m}{M} = \frac{2.3}{23} = 0.1$
 (a) 1 (b) 1.5 (c) 0.1 (d) 0.15
- (iii) What is the mass of 5 moles of hydrogen gas?
 (a) 5g (b) 5.04g (c) 10.08g (d) 1.008g $1 \text{ mole} = 2.016g$
 $5 \text{ moles} = 2.016 \times 5 = 10.08$
- (iv) How many atoms are there in 28g of nitrogen gas?
 (a) 2 (b) 1 (c) 6.022×10^{23} (d) 12.044×10^{23} $N_2 = 28g/mol = 1 \text{ mole}$
- (v) How many atoms are there in 0.1 mole of carbon? $1 \text{ mole} = N_A$
 $0.1 = 0.1 \times 6.022 \times 10^{23} = N_A$
 (a) 6.022×10^{23} (b) 6.022×10^{22} (c) 6.022×10^{24} (d) 6.022×10^{21}
- (vi) A solution of NaOH has a concentration of $4g/dm^3$. What is the mass of NaOH contained in $250cm^3$ of this solution?
 $M = \frac{m}{V}$
 $4 = \frac{m}{250}$
 $m = 4 \times 250 = 1000$
 (a) 40g (b) 20g (c) 1g (d) 2g
- (vii) Which of the following solution is more dilute?
 (a) 1M (b) 2M (c) 0.1M (d) 0.009M $\frac{1 \text{ mol}}{1 dm^3} > \frac{2 \text{ mol}}{1 dm^3} > \frac{0.1 \text{ mol}}{1 dm^3} > \frac{0.009 \text{ mol}}{1 dm^3}$
- (viii) A solution of NaOH contains 20g of this compound in $2dm^3$ of solution. What is the molarity of this solution?
 $M = \frac{20}{4} = 5$
 (a) 2M (b) 1M (c) 0.25M (d) 0.5M $\frac{20}{4} = 5$
- (ix) Which quantity is the same for one mole of hydrogen gas and one mole of water?
 (a) mass (b) number of atoms
 (c) number of molecules (d) number of gram atomic mass
- (x) If one mole of Na contains x atoms of sodium, what is the number of moles contained in 46g of sodium? $1 \text{ mole Na} = x = 23g/mol$
 $molar = \frac{46}{23} = 2 \text{ moles} = 2x$
 (a) x (b) 2 (c) 2x (d) 1.5x

2. Give short answers

- Differentiate between saturated and unsaturated solution? Def, e.g., stages. Conversion
- Give example of a solid solution containing two solids.
- Can you call colloid a solution?
- Gasoline does not dissolve in water, why?
- Are gem stones solutions?

3. A tiny crystal of a solid substance is added to an aqueous solution of the same substance. What would happen if the original solution was:

- (a) supersaturated
 crystallization
 would occur
- (b) unsaturated
 dissolve into
 solution
- (c) saturated
 settles down
 to bottom



4. Explain why CH_3OH is soluble in water but C_6H_6 is not. *methanol* *Benzene*
5. How can you prepare 250cm^3 of 0.5M MgSO_4 from a stock solution of 2.5M MgSO_4 ? *50cm³*
6. Copy and complete the following table for aqueous solution of NaOH .

Mass of solute	Moles of solute	Volume of solution	Molarity
20g	0.5	500cm^3	1
10	0.25	1000	0.25
0.8	0.02	200cm^3	0.1

7. Give examples of the following solutions:

- (a) a liquid solution of a liquid solvent and gaseous solute *109, 110*
- (b) a solid solution of two solids *Q2 (ii)*

8. What is the molarity of a solution prepared by dissolving 1.25g of HCl gas into enough water to make 30cm^3 of solution. *1.19M*
9. Formalin is an aqueous solution of formaldehyde (HCHO), used as a preservative for biological specimens. A biologist wants to prepare 1dm^3 of 11.5M formalin. What mass of formaldehyde he requires? *345g*
10. A solution of $\text{Ca}(\text{OH})_2$ is prepared by dissolving 5.2mg of $\text{Ca}(\text{OH})_2$ to a total volume of 1000cm^3 . Calculate the molarity of this solution. *$7 \times 10^{-5}\text{M}$*
11. Calculate the number of moles of solute present in 1.25cm^3 of 0.5M H_3PO_4 solution. *6.3×10^{-4} moles*
12. Calculate the new molarity when 100cm^3 of water is added to 100cm^3 of 0.5M HCl . *0.25M*
13. How are solutions useful for society? Give three examples. *Pg# 122*



Think-Tank

- 1: A 10.0g of solid solute is placed in 100g of water at 20°C and all of it dissolves. Then another 4.0g of the solute is added at 20°C and all of it dissolves.
- (a) Is the first solution saturated, unsaturated or supersaturated? Interpret.
- (b) Is it possible to tell from this information that the final solution is unsaturated or saturated? Argue. *Yes*
- 2: Suggest a convenient way to change:
- (a) a saturated solution to an unsaturated solution. *add more solvent, ↑ T*
- (b) an unsaturated solution to a saturated solution. *add solute*
- 3: Knowing the molarity of a solution is more meaningful than knowing whether a solution is dilute or concentrated. Defend this statement.



- 4: Design an experiment to determine the solubility of table sugar in water at room temperature. *Pg # 107*
- 5: Design an experiment to prepare 10% mass by volume solution of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (Nelathota). $\rightarrow$ *Pg # 116*
- 6: Which solution is more dilute 50cm^3 of 0.2M NaOH or 100cm^3 of 0.1M NaOH ? Justify.
- 7: Which solution is more concentrated 100cm^3 of 0.1M HCl or 100cm^3 of 0.1M NaOH ? Justify.
- 8: Benzene is a common organic solvent. Its use is now restricted because this can cause cancer. The recommended limit of exposure to benzene is 0.32 mg per dm^3 of air. Calculate the molarity of this solution. *$4 \times 10^{-6}\text{M}$*
- 9: A patient in a hospital is often administered an intravenous (IV) drip containing an aqueous solution. (This aqueous solution contains 0.85% (mass by volume) of sodium chloride or 5% (mass by volume) of glucose.) Calculate the molarity of both these solutions. *$\text{NaCl} = 0.145\text{M}$, Glucose = 0.278M*
- 10: 100cm^3 of NaOH solution was heated to complete dryness, 1.5g residue left behind. What is the molarity of the solution. *0.375M*
- 11: Design an experiment to distinguish between a saturated and supersaturated solution.



7

ELECTROCHEMISTRY



This is a 22 days lesson

After completing this lesson, you will be able to:

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen.
- Define oxidation and reduction in terms of loss or gain of electrons.
- Define oxidizing and reducing agent in a redox reaction.
- Define oxidation state.
- Determine the oxidation number of an atom of any element in a compound.
- Describe the nature of electrochemical processes.
- Describe how a battery produces electrical energy.
- Identify the oxidizing and reducing agents in a redox reaction.
- Identify the direction of movement of cations and anions toward respective electrodes.
- List the possible uses of an electrolytic cell.
- State the common rules for assigning oxidation numbers to free elements, ions, simple and complex molecules, atoms.
- Sketch an electrolytic cell, label the cathode and the anode.
- Sketch a Daniell cell, labeling the cathode, the anode, and the direction of flow of the electrons.



Pre-Reading

What do rusting of iron objects, combustion of fuel in automobile engine, forest fire and metabolism of food in human and animal bodies have in common? All of these important processes involve oxidation-reduction reaction. When you turn on a flash light, mobile, calculator, electronic toy etc. a chemical reaction in a battery generates a current of electricity. House hold bleaching agents decolourize the colour bearing substances in stains. How? Electrochemical industries produce millions of tons of important metals such as aluminium, copper, sodium carbide.

magnesium and zinc. They also produce chemicals such as caustic soda, chlorine and silicon carbide.

Anodized aluminium sheets are widely used in buildings now- a-days. Anodized aluminium can absorb dyes. Dying of aluminium can produce metallic red, metallic blue or other metallic colours on the metal surface. This chapter will enable you to understand chemical processes involved in the formation of important metals and chemicals.



Reading

7.1 OXIDATION AND REDUCTION

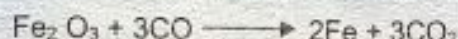
Oxidation and reduction can be defined in two ways:

- (a) Loss or gain of oxygen and hydrogen.
- (b) Loss or gain of electrons.



7.1.1 Oxidation and Reduction in Terms of Loss or Gain of Oxygen and Hydrogen

In steel mills iron ores, usually oxides of iron are converted to the pure metal commercially by the reaction with coke (carbon) in the blast furnace. The carbon first reacts with air to form carbon monoxide, which in turn reacts with iron oxide as follows.



Which substance in this reaction loses oxygen? Which substance gains oxygen?

Oxidation is defined as the gain of oxygen atoms or loss of hydrogen atoms by an element.

Which substance is oxidized, which substance is reduced in the above reaction?

Acetylene (C_2H_2) is commercially used for cutting and welding metals. When acetylene burns, it produces a very hot flame known as oxy-acetylene flame. Following reaction takes place when it burns.



Which substance loses hydrogen?

Which substance gains hydrogen?

Coal is burned in thermal power stations to produce electricity. When it burns, following reaction occurs. (Figure 7.1).



In this reaction carbon atoms undergo oxidation, since they gain oxygen atoms to form CO_2 .

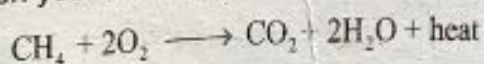
Another familiar example of an oxidation that does not involve burning is rusting. In the rusting of iron following reaction occurs (see figure 7.2)



Is iron oxidized in this reaction?

Example 7.1: Identifying the element undergoing oxidation

Following reaction occurs when you burn Sui gas.



Identify the element undergoing oxidation.



Figure 7.1: Burning coal

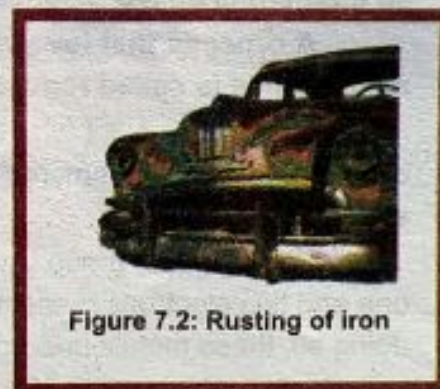


Figure 7.2: Rusting of iron

Problem solving strategy:

Identify the substance that gains O-atoms or loses H-atoms.

Solution:

Since C in CH_4 loses H-atoms and combines with oxygen atoms, thus C atoms undergo oxidation. At the same time O-atoms combine with H-atoms to form H_2O , thus O-atoms undergo reduction.

**Self-Assessment Exercise 7.1**

Identify elements undergoing oxidation and reduction in the following reactions:

1. $\text{N}_2 + 3\text{H}_2 \longrightarrow 2\text{NH}_3$
2. $2\text{H}_2 + \text{O}_2 \longrightarrow 2\text{H}_2\text{O}$
3. $\text{Fe}_2\text{O}_3 + 3\text{CO} \longrightarrow 2\text{Fe} + 3\text{CO}_2$
4. $4\text{Al} + 3\text{O}_2 \longrightarrow 2\text{Al}_2\text{O}_3$

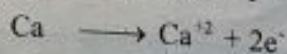
**Reading****7.1.2 Oxidation and Reduction in Terms of Loss or Gain of Electrons**

A process that involves the loss of electrons by an element is called oxidation.

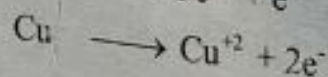
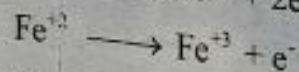
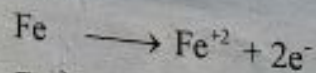
A process that involves the gain of electrons by a substance is called reduction.

In section 4.3.1 we discussed the formation of cations by metallic elements and anions by non-metallic elements.

For example, group IA and group IIA elements lose one and two electrons respectively to form cations. In doing so, these metals undergo oxidation.



Other examples of oxidation are,

**Society, Technology and Science****Redox in photography**

A photographic film is basically an emulsion of silver bromide, (AgBr) in gelatin. When the film is exposed to light, Silver bromide granules become activated. This activation depends on the intensity of the light falling upon them. When exposed film is placed in the developer solution that is actually a reducing agent. Hydroquinone which is a mild reducing agent is used as developer. In hydroquinone the activated granules of silver bromide are reduced to black metallic silver. Reduced silver atoms form image.



Inactivated silver bromide is removed from the film by using a solvent called a fixer. Sodium thiosulphate is used for this purpose. The areas of the film exposed to the light appear darkest because they have the highest concentration of metallic silver. Thus photography involves oxidation-reduction reaction.



Elements of group VIA and VIIA gain one and two electrons respectively to form anions. They undergo reduction.



Example 7.2: Identifying the element oxidized or reduced

In the following reaction identify which element is oxidized and which element is reduced

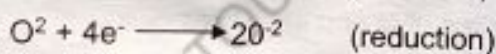
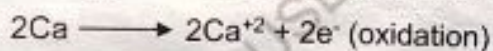


Problem solving strategy:

Ca being metal will form cation by losing electrons (oxidation) and oxygen being non-metal will gain electrons (reduction) to form anions.

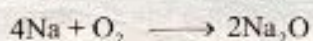
Solution:

We have learned that Group IIA metals form M^{+2} cations, and that Group VIA non-metals form X^{2-} anions. This means in this reaction each Ca atom loses two electrons to form Ca^{+2} , so it is oxidized. Each oxygen atom gains two electrons to form O^{2-} , so it is reduced.



Self-Assessment Exercise 7.2

In the following reactions, identify which element is oxidized and which element is reduced.



Reading

7.2 OXIDATION STATES AND RULES FOR ASSIGNING OXIDATION STATES

7.2.1 Oxidation States

Oxidation state or oxidation number is defined as the number of charges an atom will have in a molecule or a compound.



The elements that show an increase in oxidation number are oxidized. The elements that show a decrease in oxidation number are reduced. Do you think H in HCl is oxidized and Cl is reduced? Comparison of oxidation and reduction processes is given in table 7.1.

Table 7.1 Process leading to oxidation and reduction

Oxidation	Reduction
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electrons
Increase in oxidation number	Decrease in oxidation number

7.2.2 Rules for Assigning Oxidation States or Numbers

1. The oxidation state of any uncombined or free elements is always zero e.g., oxidation state of Zn, Na, H in H_2 , S in S_8 etc is zero.
2. In simple ions, oxidation state is same as their charge e.g., oxidation state of Na in Na^+ and Ca in Ca^{+2} is +1 and +2 respectively.
3. In a complex ion the total sum of oxidation states of atoms is equal to the charge on their ion. e.g., in CO_3^{2-} , the sum of oxidation states of C and 3O atoms is -2. Similarly, in NH_4^+ , the sum of oxidation states of N and 4H atoms is +1.
4. The oxidation number of each of the atoms in a molecule or compound is counted separately and their algebraic sum is zero e.g., In HCl, the sum of oxidation states of H and Cl atoms is zero. Similarly in CO_2 , the sum of oxidation states of one C and 2 oxygen atoms is zero.

Table 7.2 shows the oxidation states of some of the elements in binary compounds which rarely change.

Table 7.2 Oxidation states of some elements in binary compounds that rarely change

Elements	Oxidation State
Group-IA	+1
Group-IIA	+2
Group-IIIA	+3
H	+1 (except in metal hydrides where it is -1)
Group-VIIA	-1
O	-2(except peroxides and in OF_2)

In grade XI you will use oxidation numbers for balancing redox reactions.



7.2.3 Determining the Oxidation Number of an Atom in a Compound

Let's see how to use rules discussed in section 7.2.2 to determine the oxidation number of an atom of an element in a compound.

Example 7.3: Determining oxidation number

A device called Breath Alyzer is used by police to test a person's breath for alcohol. It contains an acidic solution of potassium dichromate $K_2Cr_2O_7$. It is a strong oxidizing agent. Determine oxidation state of Cr in it.

Problem Solving Strategy:

Use rules 1 to 4 and table 7.1 to get as many oxidation numbers as you can. Use rule 4 to get oxidation number that has not been assigned.

Solution:

1. The oxidation number of K is +1, since it belongs to Group-1A. There are 2 K atoms therefore, overall oxidation number for K is $2(+1) = +2$
2. There are 7 oxygen atoms, therefore overall oxidation state for O is $7(-2) = -14$
3. Suppose oxidation for Cr is x, since there are two Cr atoms, therefore, overall oxidation state for Cr is $2x$.
4. The sum of oxidation numbers must be zero.

$$+2 + 2x + (-14) = 0$$

$$2x - 12 = 0$$

$$2x = 12$$

$$x = +6$$

Thus oxidation state for Cr in $K_2Cr_2O_7$ is +6

Example 7.4: Determining oxidation state

Example 7.4:

Boric acid H_3BO_3 is used in eye wash. What is the oxidation state of B in this acid?

Problem solving strategy:

Use rules and table 7.2 to get the oxidation state of H and O- atoms. Use rule 4 to get the oxidation state of B.

Solution:

1. There are 3 H-atoms, therefore, overall oxidation state for H is

$$3(+1) = +3$$



Teacher's Point

A teacher may give home work of calculations of oxide, number of different substances.



2. There are 3 O-atoms, therefore, overall oxidation state for O is
 $3(-2) = -6$
3. Suppose the oxidation state for B is x.
4. The total oxidation states for all the atoms must be zero.

$$+3 + x + (-6) = 0$$

$$+3 + x - 6 = 0$$

$$x - 3 = 0$$

$$x = 3$$

Thus the oxidation state for B in H_3BO_3 is +3.



Self-Assessment Exercise 7.3

One major problem of air pollution is the formation of acid rain. Air pollutants such as SO_2 and NO_2 combine with oxygen and water vapours in the air to form H_2SO_4 and HNO_3 . These acids fall to the ground with the rain, making the rain acidic. Clouds can also absorb the acids and carry them hundreds of kilometer away from where the pollutants are released. Determine the oxidation number of N in NO_2 and HNO_3 , S in SO_2 and H_2SO_4 .



Reading

7.3 OXIDIZING AND REDUCING AGENTS

An oxidizing agent is the reactant containing the element that is reduced (gains electrons) in a chemical reaction. A reducing agent is the reactant containing the element that is oxidized (loses electrons) in the chemical reaction.

For example, in the reaction between sodium and chlorine to form sodium chloride.



Na is reducing agent as it is oxidised whereas Cl_2 is oxidizing agent as it is reduced.

7.3.1 Identifying Oxidizing and Reducing Agents



Activity

Prepare solutions of ferrous sulphate (FeSO_4) and potassium permanganate (KMnO_4) in separate beakers. Transfer about 10cm^3 of ferrous sulphate solution in a test tube. Add about 10cm^3 of dil. H_2SO_4 in it. Now add few drops of KMnO_4 solution in the test tube. What happens?



FeSO_4 reduces KMnO_4 , so its purple colour is discharged. KMnO_4 oxidizes FeSO_4 in this reaction. FeSO_4 is reducing agent whereas KMnO_4 is oxidizing agent.

How can you identify oxidizing and reducing agents in a chemical reaction?

Consider the following reaction that takes place in the manufacture of steel.



To identify the oxidizing and reducing agents, work out the oxidation states of all the elements involved in the reaction.



- Carbon is being oxidized because there is an increase in its oxidation state.
- Fe is being reduced because there is a decrease in its oxidation state.
- The reactant CO contains the C that is being oxidized, so CO is reducing agent.
- The reactant Fe_2O_3 contains the Fe that is being reduced. So Fe_2O_3 is oxidizing agent.

Oxidizing or reducing agent is the whole molecule or formula unit and not the atom that has undergone change in oxidation number.

Example 7.5: Identifying the oxidizing and reducing agents

Tungsten is used to make filaments for electric bulbs because it has the highest melting point and high electrical resistance. This metal is obtained from tungsten (VI) oxide, WO_3 by reducing it with hydrogen gas:



Identify the oxidizing and reducing agents in this reaction.

Problem solving strategy:

Step 1: Work out the oxidation states of all the elements involved in the reaction.

Step 2: Note the element that is undergoing an increase in its oxidation state. Since it is being oxidized, the reactant that contains this element is reducing agent.

Step 3: Note the element that is undergoing a decrease in its oxidation state. Since it is being reduced, the reactant that contains this element is oxidizing agent.

Solution:

First assign oxidation numbers to each atom.

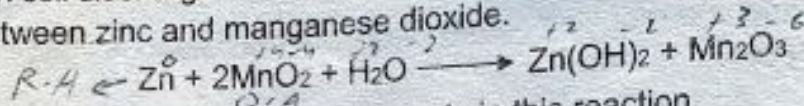


Because the oxidation number of W decreases, so WO_3 is an oxidizing agent. Similarly the oxidation number of H increases, therefore H_2 is reducing agent.



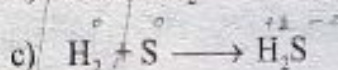
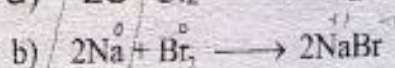
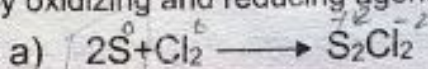
Self-Assessment Exercise 7.4

1. The torch cell discharges electricity because of an oxidation-reduction reaction that takes place between zinc and manganese dioxide.



Identify the oxidizing and reducing agents in this reaction.

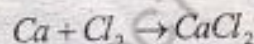
2. Identify oxidizing and reducing agents in the following reactions:



Reading

7.4 OXIDATION REDUCTION REACTIONS

In section 4.3.1 you have learned the formation of cations from metallic elements and formulation of anions from non-metallic elements. You have also learned that anions and cations combine to form ionic compounds. Consider the formation of CaCl_2 by the reaction of calcium and chlorine.



Calcium and chlorine contain neutral atoms. Since CaCl_2 is known to contain Ca^{+2} and Cl^- ions, this reaction must involve a transfer of electron from calcium atom to chlorine atom. Such a reaction which involve transfer of electrons is known as an oxidation-reduction reaction. These reactions are also called redox reactions. Such reactions are commercially very important. Most of the metals are recovered from their ores by redox reactions.

7.4.1 Method of Recovering Metals from their Ores

The process of producing a metal from its ores, always involves oxidation-reduction reaction. Most of the metals are found in nature as oxides or sulphide ores. After mining the ore, desired mineral is separated from the other

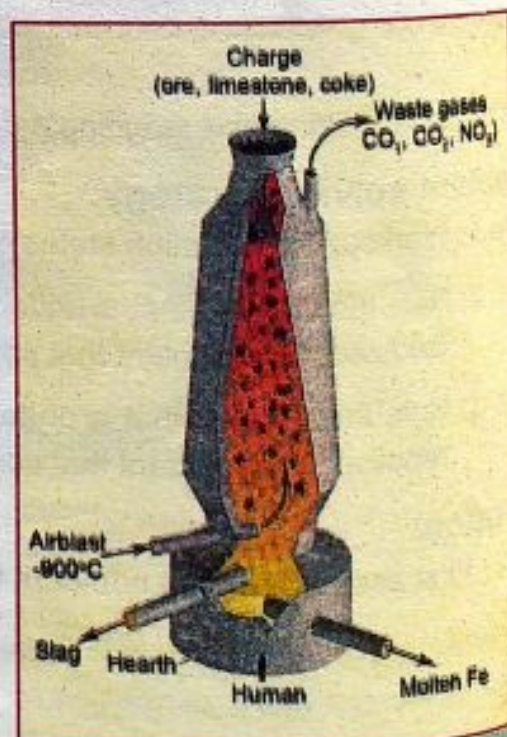


Figure 7.3 A blast furnace

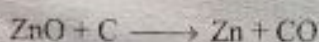


materials. Purified metal oxides are reduced to free metals by using a reducing agent. Aluminum, coke, carbon monoxide gas and hydrogen gas are generally used as reducing agents.

For example extraction of iron involves the chemical reduction of Hematite (Fe_2O_3) ore by coke (carbon) in a blast furnace (figure 7.3). Iron ore, lime stone (CaCO_3) and coke are introduced into the blast furnace from the top. A blast of hot air is forced up in the furnace from the bottom (hence the name blast furnace). The oxygen gas reacts with coke to form mostly carbon monoxide and some carbon dioxide. These reactions are highly exothermic. As hot CO rises, it reacts with iron oxide and reduces it to iron. Molten iron collects at the bottom of the furnace. Lime stone removes impurities from iron as slag.



Lead and Zinc metals occur naturally as sulphide ores. These ores are first converted into corresponding oxide by heating in oxygen. In this process sulphur is oxidized and oxygen is reduced. These oxides are then reduced by coke or carbon monoxide or hydrogen.



7.5 ELECTROCHEMICAL CELLS

Devices that convert chemical energy into electrical energy or vice versa are called electrochemical cells. Therefore, there are two types of electrochemical cells:

- (1) Electrolytic Cells
- (2) Galvanic Cells

An electrochemical cell that uses electrical energy to drive a chemical reaction is called an electrolytic cell. While an electrochemical cell that converts chemical energy into electricity is called a galvanic cell.

7.5.1 Nature of Electrochemical Process

Electrochemical processes are oxidation-reduction reactions in which chemical energy released by a spontaneous reaction is converted to electricity or in which electrical energy is used to drive a non-spontaneous reaction. Whether an electrochemical process releases or requires energy, it always involves the transfer of electrons from one substance to another. This means that this process always involves an oxidation-reduction or a redox reaction.

Science Tidbit

A physical or chemical change that occurs by itself is called a spontaneous process. Spontaneous processes do not require a source of energy to make them happen. For instance water flows from higher level to lower level. Iron placed in moist air, rusts. The flow of electrons through a conductor can be obtained from a spontaneous oxidation-reduction reaction. This is the basis for how batteries work.

On the other hand a physical or chemical change that requires a source of energy to make them happen is called non-spontaneous process. For example water can be made to flow from lower level to higher level by using a pump.



7.5.2. Concept of Electrolytes

When an ionic compound is dissolved in water, it splits up into its positive and negative ions, these ions are capable of conducting electricity. Such compounds are referred as electrolytes.

A substance that conducts electrically when it is dissolved in water or in the molten state is called electrolyte. Whereas, a substance that cannot conduct electricity when dissolved in water or in the molten state is called non-electrolyte.

Examples of electrolytes are NaCl, KCl, HCl, NaOH etc.

Example of non-electrolytes are urea, glucose, sucrose, benzene etc.



Activity 7.1

Passing electricity through aqueous solution

You will need:

- Beaker
- Graphite rod
- A battery and Conducting wires
- Aqueous solutions of NaCl, NaOH, urea and glucose

Carry out the following:

1. Set up the circuit as shown in the figure 7.4.
2. Fill beaker with one of the given solutions.

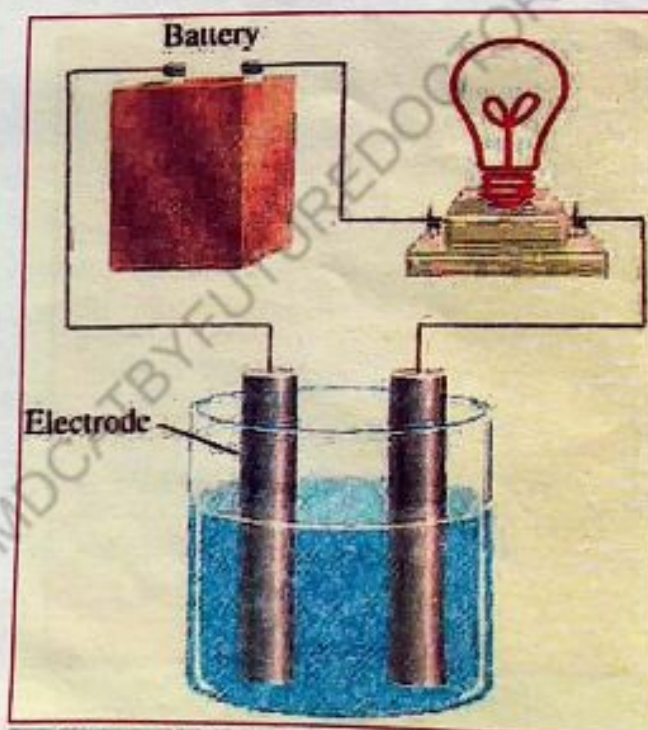


Figure 7.4: Assembly for distinguishing electrolytes and non-electrolytes

Do you know?

An electrode is a conductor in a circuit that carries electrons to or from a substance other than a metal for example graphite rods are used as electrodes.

3. Dip graphite rods in the solution.
4. Note if the bulb is lighted or not.
5. Repeat the experiment with other solution one by one.
6. Identify electrolytes.

Figure 7.4 shows an easy method of distinguishing between electrolytes and non-electrolytes. When graphite electrodes are dipped in pure water, no current flows. But if you add a small quantity of solution of KCl, the bulb glows. KCl is an ionic compound. It breaks up into K^+ and Cl^- ions, as it dissolves in water. These ions move towards the opposite electrodes. Thus current flows through the solution.

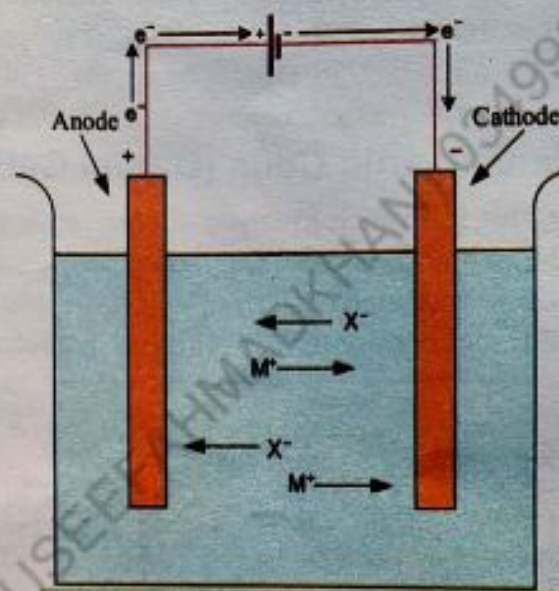
**Reading****7.5.3 Electrolytic Cells**

An electrochemical cell in which electrical energy is used to drive a chemical reaction is called an electrolytic cell. Figure 7.5 shows the sketch of an electrolytic cell.

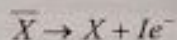
An electrolytic cell consists of

- A vessel containing an electrolyte (MX)
- Two inert electrodes
- A battery

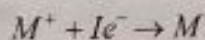
The figure shows that electrons move from anode to cathode in the outer circuit, in the solution the cations move towards cathode and anions towards anode. At anode anions oxidize by losing electrons. At cathode cations reduce by gaining electrons. This means oxidation occurs at anode and reduction at cathode.



At Anode



At cathode

**Self-Assessment Exercise 7.5**

- Sketch an electrolytic cell for the electrolysis of fused KCl
- Electrolytic cell shown in figure 7.6 is used for the electrolysis of fused sodium chloride. Indicate the direction of flow of electrons. Identify anode and cathode.

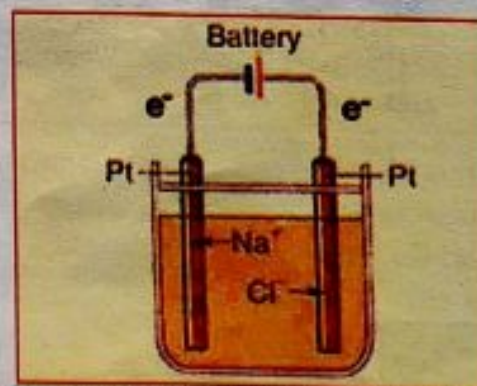


Figure 7.6: Electrolytic cell for the electrolysis of fused NaCl

**Reading****7.5.4 Uses of Electrolytic Cells**

Possible uses of electrolytic cells are as follows:

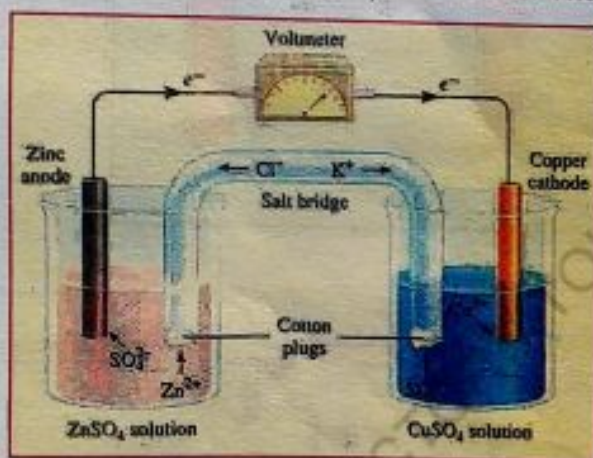
- Down's Cell is used for the commercial preparation of sodium metal. It produces chlorine gas as by product.



- (b) Nelson's Cell is used for the commercial preparation of sodium hydroxide. It also produces chlorine and hydrogen gas as by product.
- (c) Electrolytic cells are used for the commercial preparation of calcium and magnesium metals.
- (d) It is used to produce aluminum metal commercially.
- (e) It is used for the purification of copper.
- (f) Electrolytic cells are used to electroplate metals such as tin, silver, nickel etc on steel.
- (g) Electrolytic cells are used to prepare anodized aluminum. Anodized aluminum can absorb dyes. Dyeing of anodized aluminum can produce metallic red, metallic blue or other metallic colours on the metal surface.

7.5.5. Galvanic Cells (Daniel Cells) (Voltaic Cells) $\rightarrow$ G.D.V Cells

The cell which involves spontaneous redox reaction to generate electricity is called a galvanic or voltaic cell. The name Voltaic is given to this cell because Alessandro Volta discovered first such cell. The English chemist Fredrick Daniel constructed first voltaic cell using zinc (Zn) and copper (Cu) electrodes. Therefore this cell is named as Daniel Cell. A galvanic or



Daniel cell is shown in figure 7.7.

A galvanic cell consists of the following parts:

- (1) A zinc bar dipped into a 1M ZnSO_4 solution.
- (2) A copper bar dipped into a 1M CuSO_4 solution.
- (3) A salt bridge which is inverted U tube containing an inert electrolyte such as KCl. Its ions do not react with the electrodes or with the ions in solution. It makes the electrical contacts between the solutions through which ions can move.
- (4) A voltmeter to measure current.

Thus a Daniel cell consists of two half-cell joined in series. When circuit is complete, electrons begin to flow from Zn rod through the external wires to Cu rod. Thus Zn half-cell acts as anode and Cu half-cell as cathode. Note that a half cell consists of a metal rod dipped in the solution of its salt.



Self-Assessment Exercise 7.5

Sketch a voltaic cell labeling the cathode, the anode and the direction of flow of the electrons. Use the following chemicals:

Silver, Zinc, Silver Nitrate (AgNO_3) and zinc sulphate (ZnSO_4).

(Hint: Zn is more active than Ag)



Teacher's Point

A teacher may ask the students to make a model of Galvanic cell.

*Ag reduces (even more ^{pot.} than Cu)
Zn oxidizes*

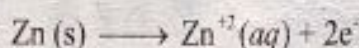


Reading

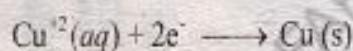
7.5.6 Reactions in a Daniel Cell

In Daniel cell, electrons flow from Zn rod, through the external wire to Cu rod. This is because Zn has more tendency to undergo oxidation than Cu. Zn atoms from the rod go into the solution as Zn^{+2} ions leaving electrons on the rod. These electrons flow in the external circuit. Thus oxidation half reaction occurs at anode compartment. Cu^{+2} ions in copper sulphate solution capture electrons from Cu electrode and are reduced. Reduction half reaction occurs at the cathode compartment. Such oxidation and reduction reactions are called half cell reactions.

At Anode (Oxidation half reaction):



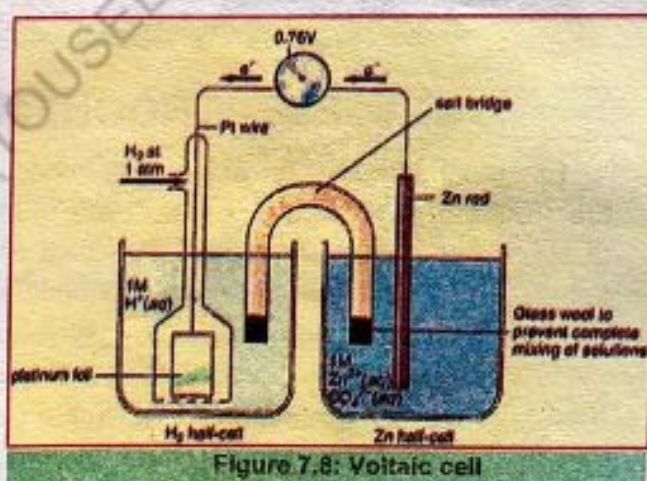
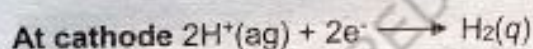
At Cathode (Reduction half reaction):



Self-Assessment Exercise 7.6

Identify the half-cell in which oxidation occurs and the half cell in which reduction occurs in the following voltaic cell (figure 7.8).

Following reactions occur in the cell.



Reading

7.5.7 Battery-a Source of Electrical Energy

A battery is a galvanic cell or a group of galvanic cells joined in series. It generates electric current by a redox reaction. When connected in a circuit its anode oxidizes by releasing electrons. These electrons through the external circuit begin to flow towards the cathode. At cathode these electrons reduce oxidizing agent present in the electrolyte. Examples of batteries are dry cell, storage cell, mercury battery etc.

DRY CELL

The dry cell batteries are used to power many flashlights, toys and small appliances. The anode is the zinc metal of the container and the cathode is an inert graphite rod at the center of



the container in contact with a mixture of MnO_2 and carbon (charcoal) see fig 7.9. The electrolyte is a mixture of moist NH_4Cl and ZnCl_2 . Following reactions take place in it,

At Anode



At Cathode



This cell produces a potential of 1.5V.

$$\begin{aligned} x + y &= +1 \\ x &= +1 - y \\ &= -3 \end{aligned}$$

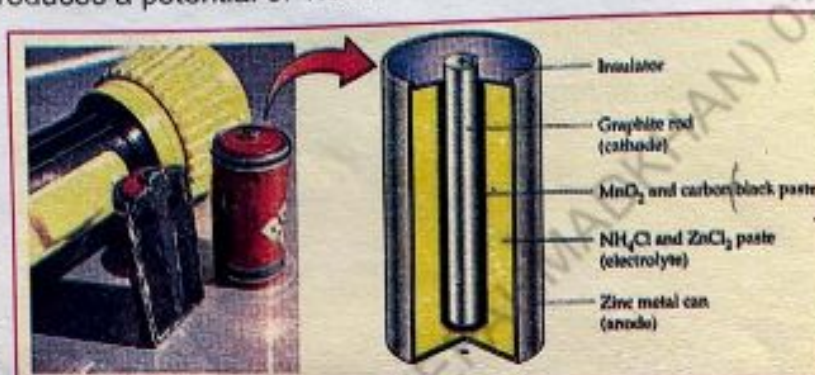


Figure 7.9: A dry cell

7.6 ELECTROCHEMICAL INDUSTRIES

Electrical energy is extensively consumed by the chemical process industries. Electrochemical industries use electricity to bring chemical change and produce wide variety of substances. Such industries are called electrolytic industries. These industries produce substances that cannot be made economically in another way. For example, aluminum and calcium carbide, sodium hydroxide, magnesium and hydrogen. These industries also include electrolytic refining and electroplating industries. Electricity is also used to produce heat that produces high temperature required in the electrothermal chemical industries.

We will discuss the electrochemical preparation of the following:

1. Manufacture of sodium metal from fused sodium chloride.
2. Manufacture of NaOH from brine.
3. Electrolytic refining of copper.
4. Electroplating of Zinc, tin and Chromium on steel.

7.6.1 Manufacture of Sodium Metal from Fused Sodium Chloride

On the large scale sodium metal is produced by the electrolysis of fused NaCl . An electrolytic cell called Down's Cell is used for this purpose. (Figure 7.10)

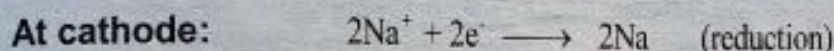
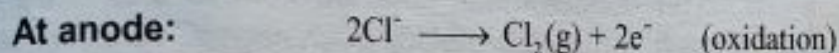


Teacher's Point

A teacher may show a cut piece of Dry cell to students.



The electrodes are iron cathode and graphite anode. Chlorine is obtained as by-product. In molten sodium chloride (NaCl), Na^+ and Cl^- ions are free to move about. Under the influence of electric current, Na^+ ions move towards the cathode and Cl^- ions towards the anode. At the electrodes following reactions occur:



Molten sodium is collected into a sodium collecting ring, from where it is periodically drained. Whereas, chlorine gas is collected into the funnel at the top of the cell.

Other alkali metals can also be obtained by the electrolysis of their fused salts. For this purpose ores are first purified than electrolyzed.

7.6.2 Manufacture of Sodium Hydroxide from Brine

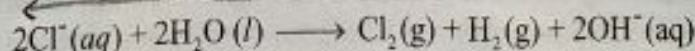
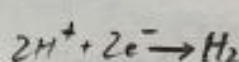
Electrolysis of brine, a concentrated aqueous solution of sodium chloride is used for the industrial production of sodium hydroxide. Electrolysis of brine produces simultaneously three important industrial chemicals, chlorine gas, hydrogen gas and sodium hydroxide. The electrolytic cell called Nelson's cell is shown in figure 7.11.

During electrolysis chloride ions move towards anode. At anode chloride ions are oxidized to produce chlorine gas. Sodium ions move towards cathode. However sodium ions do not reduce to sodium metal in this process. This is because water molecules are more easily reduced than sodium ions. The reduction of water molecules produces hydroxide ions and hydrogen gas. Thus the electrolyte in solution becomes sodium hydroxide (NaOH). Following reactions occur in the electrolytic cell.

At anode:



At cathode:



Overall Reaction:

A

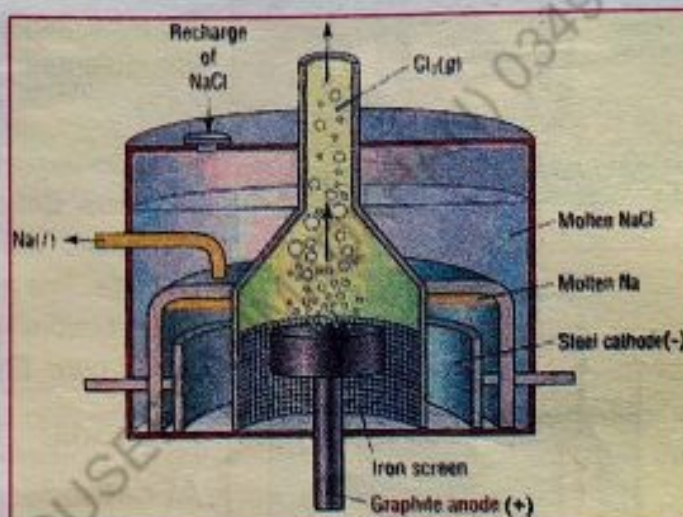
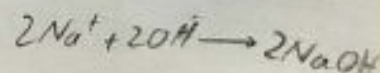


Figure 7.10: Down's Cell

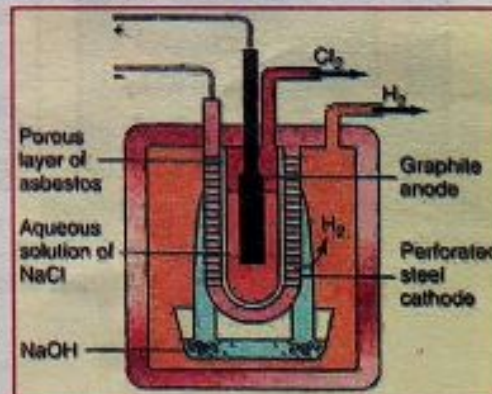


Figure 7.11: Nelson's Cell

The solution contains Na^+ and OH^- ions. Evaporation of water from this solution produces relatively pure solid sodium hydroxide.

7.6.3 Electrolytic Refining of Copper

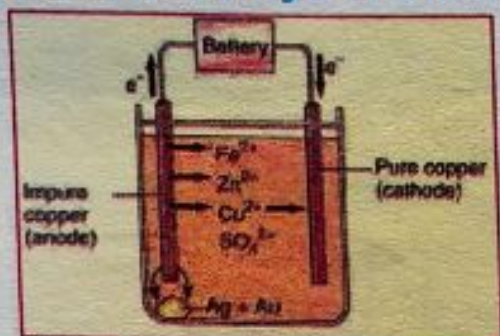


Figure 7.12: Electrolytic refining of copper

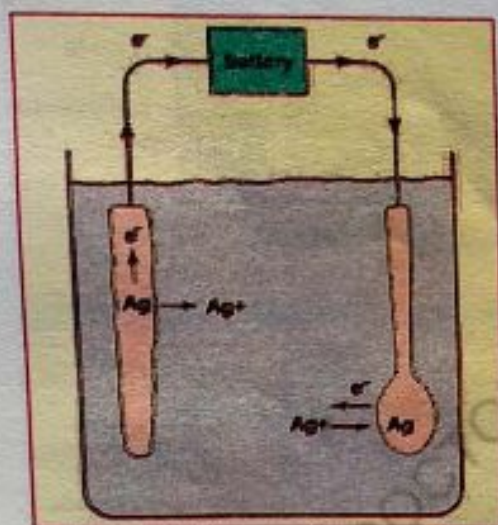


Figure 7.13: Electroplating of silver on spoon

Conditions for good electroplating are as follows:

- High current density.
- Low temperature.
- High concentration of metal in its electrolyte.

7.6.4 Electroplating on Steel

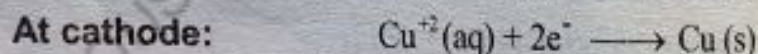
Steel objects are often protected from corrosion by electroplating with zinc, tin and chromium. *anode* *cathode*

Zinc Plating

Zinc plating on steel is done by using zinc metal as anode. A solution of potassium zinc cyanide $\text{K}_2[\text{Zn}(\text{CN})_4]$ containing little sodium cyanide is used as electrolyte. The steel object is

The copper metal obtained from its ores is usually impure. It contains impurities such as zinc, iron, silver and gold. These impurities are removed by the process of electrolysis. In this process impure copper bars act as anode and pure copper bars as cathode. CuSO_4 solution containing little sulphuric acid is used as the electrolyte. (See figure 7.12)

On passing electricity, copper anode dissolves forming Cu^{2+} ions. Cations move towards the cathode at which only Cu^{2+} ions are reduced. Thus pure copper deposits at cathode. The less electropositive metals, silver and gold fall to the bottom of the cell. Copper obtained in this process is 99.5% pure. Following reaction occur in this process.



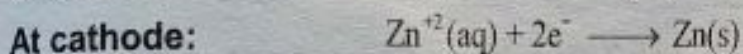
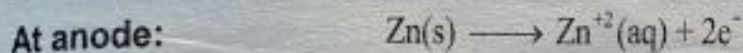
Electrolytic process used to deposit one metal on another metal is called **electroplating**. An object can be electroplated by making it cathode in an electrolytic tank containing ions of the plating metal. The plating metal is made anode. On passing electricity through the electrolytic tanks, a thin layer of anode metal is deposited on the surface of the object. Figure 7.13 shows silver plating of a teaspoon.



made cathode. During the electrolysis zinc at the anode enters the solution as Zn^{+2} ions, which are deposited at the cathode. The electrolyte ionizes as follows.



Following reactions occur at the electrodes:



Sodium cyanide prevents the hydrolysis of electrolyte.

(Awareness against pollution)

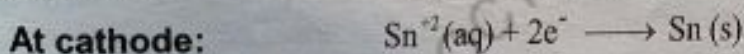
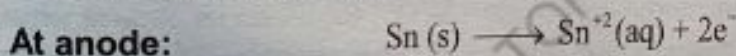
Cyanide ions are extremely toxic. Therefore, solution containing cyanide ions must not be dropped into rivers and streams. This is responsible for killing fish and other animals.

Tin Plating

Food cans are generally tin plated. Tin plating on steel is done by using tin as anode and a solution of stannous sulphate, (SnSO_4) as electrolyte. Few drops of dil H_2SO_4 are added in the electrolyte to prevent its hydrolysis. The electrolyte ionizes as follows.



During the electrolysis following reactions occur:

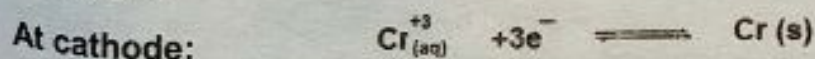
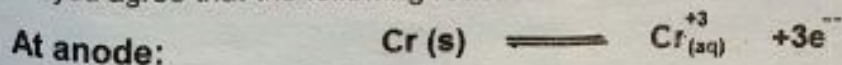


Chromium Plating

Since chromium metal does not adhere strongly to the steel therefore steel is first plated with copper or nickel and then with chromium. For electroplating chromium, chromium metal is used as anode and chromium sulphate, $\text{Cr}_2(\text{SO}_4)_3$ as an electrolyte. Few drops of dil H_2SO_4 are added in the electrolyte to prevent its hydrolysis. The electrolyte ionizes as follows:



Do you agree that the following reactions will occur at the electrodes?



Chromium plated steel is used to make automobile parts (see figure 7.14).

Society, Technology and Science

Tin plated steel is used to make cans. Food and beverages industries use tinned steel cans. This is because the components of food, beverages and the preservatives contain organic acids or their salts. They may form toxic substances by reacting with iron. These acids and salts are corrosive. Tin plating is non-poisonous and prevents corrosion.

steel
→ corrode
iron



Figure 7.14: Motor bike

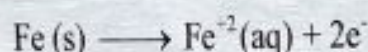


7.7 CORROSION AND ITS PREVENTION

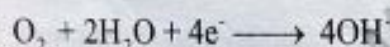
Corrosion is the process in which a metal reacts with oxygen and moisture in the atmosphere. It is a natural process that converts refined metals to the more stable metal oxides.

7.7.1 Rusting of Iron

Most familiar example of corrosion is the formation of rust on iron. Oxygen and water are necessary for iron to rust. A region of metal surface that has relatively less moisture, acts as anode. Will Fe oxidize in this region?



Another region on the surface of metal that has relatively more moisture acts as cathode. The electrons released in the oxidation process reduce atmospheric oxygen to hydroxyl ions.



The Fe^{+2} ions formed at the anodic regions flow to the cathodic regions through the moisture on the surface. Here Fe^{+2} ions further react with oxygen to form rust, $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$

7.7.2 Corrosion of Aluminum

Corrosion is not limited to iron. Aluminium is extensively used in the construction of aircraft, ships, cars, cooking utensil, window frames, soda canes etc. Aluminium has much higher tendency to oxidize than iron. Therefore, we might expect to see aircraft, ships, cars, cooking utensils, soda canes transformed into piles of corroded aluminium. However, this does not occur. This is because a tough layer of insoluble aluminium oxide (Al_2O_3) forms on its surface when metal is exposed to air. This layer firmly adheres to the metal and serves to protect the underlying aluminium layers from further corrosion. On the other hand, the insoluble layer of rust, $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ that forms on the surface of iron is too porous to protect the underlying metal. This layer flakes away and exposes metal for further corrosion.

7.7.3 Prevention of Corrosion

Prevention of corrosion is an important way of conserving our natural resources. Following methods have been devised to protect metals from corrosion:

1. Coating with paint:

Corrosion can be prevented by painting the metal, so that it does not come in contact with oxygen and moisture and other harmful agents. Paint is cheap and easily applied. Paint is used to protect many everyday steel objects such as cars, trucks, trains, bikes, bridges etc. Paint also provides visual appeal.

2. Alloying:

The tendency of iron to oxidize can be greatly reduced by alloying it with other metals. For example, stainless steel is an alloy of iron chromium and nickel. It is protected from corrosion by an outer layer of Cr_2O_3 . Fe Cr Ni



3. Coating with a thin layer of another metal:

Metals that readily corrode can be protected by coating with a thin layer of another metal that resists corrosion. This can be done by:

- (a) Tinning
- (b) Galvanizing
- (c) Electroplating

(a) Tinning:

In the process of tin plating, clean iron sheet is dipped in a bath of molten tin. It is then passed through hot pair of rollers. Tin protects iron effectively, since, it is very stable.

(b) Galvanizing (Coating with Zinc):

The process of galvanizing consists of dipping a clean iron sheet in a hot zinc chloride bath and heating. After this sheet is rolled into zinc bath and cooled.

(c) Electroplating:

In electroplating an electrolytic process is used to deposit one metal on another metal. For detail see section 7.6.3.

Society, Technology and Science

Silver is very soft metal. Silver atoms have weak interactions and are loosely packed together. Silver tarnishes in air when it comes in contact with trace quantities of H_2S or SO_2 in the atmosphere or food such as eggs, that are rich in sulphur compounds. Silver tarnish is silver sulphide that gives silver blackish appearance. Due to this reason decorative and practical objects made of solid silver gradually turn black and lose shining appearance. Decorative and practical objects are plated with a thin layer of silver. Atoms in thin layers firmly adhere to the metal atoms of the object and form a durable layer. An article thickly plated with silver contains many layers of silver atoms. Such layers form soft covering. These layers gradually turn black.

4. Cathodic Protection:

Cathodic protection is the process in which the metal that is to be protected from corrosion is made cathode and is connected to metals such as magnesium or aluminum. These metals are more active than iron, so they act as anode and iron as cathode. The more active metals themselves oxidize and save iron from corrosion. Cathodic protection is employed to prevent iron and steel structures such as pipes, tanks, oil rigs etc in the moist underground and marine environment (see figure 7.16).

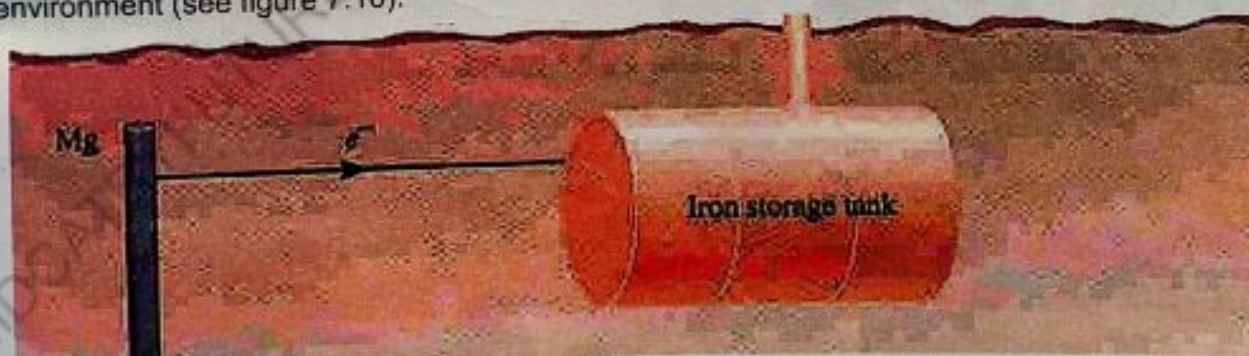


Figure 7.16: Cathodic Protection



Teacher's Point

A teacher may show different articles of Ni and Cr plating to students.

**Key Points**

- Oxidation is the gain of oxygen atom or loss of hydrogen atom or loss of electrons by a substance.
- Reduction is the loss of oxygen atom or gain of hydrogen atom or gain of electrons by a substance.
- Oxidation state or oxidation number is defined as the number of apparent charges that an atom will have in a molecule.
- The sum of oxidation state of all the atoms in a molecule of compound is zero.
- An oxidizing agent is the reactant containing the element that is reduced in a reaction.
- A reducing agent is the reactant containing the element that is oxidized in a reaction.
- An electrochemical cell in which electrical energy is used to drive a chemical reaction is called an electrolytic cell.
- Oxidation always occurs at the anode.
- Reduction always occurs at cathode.
- An electrochemical cell in which a spontaneous oxidation-reduction reaction generates electricity is called a galvanic or voltaic cell.
- A galvanic cell consists of two half cells which are joined in series.
- In a galvanic cell oxidation half reaction occurs in anode compartment and reduction half reaction occurs in the cathode compartment.
- A battery is a galvanic cell or a group of galvanic cells joined in series.
- On the large scale, sodium metal is produced by the electrolysis of fused sodium chloride.
- Electrolysis of brine is used for the industrial production of sodium hydroxide.
- Corrosion is the process in which a metal reacts with oxygen and moisture in the atmosphere.
- Electrolytic process used to deposit one metal on another metal is called electroplating.
- Cathodic protection is the process in which metal that is to be protected from corrosion is made cathode and is connected to metals such as magnesium or aluminum.

REFERENCES FOR ADDITIONAL INFORMATION

- B.Earl and LDR Wilford, Further Advanced Chemistry.
- B.Earl and LDR Wilford, Introduction to Advanced Chemistry.
- David E. Goldberg, Fundamental of Chemistry.
- Addison Wesley, Chemistry.



Review Questions

1. Encircle the correct answer:

- (i) In which of the following changes, the nitrogen atom is reduced.
(a) N_2 to NO (b) N_2 to NO_2 (c) N_2 to NH_3 (d) N_2 to HNO_3
- (ii) Which of the following changes reaction is an example of oxidation
(a) Chlorine molecule to chloride ion (b) Silver atoms to silver (I) ion
(c) Oxygen molecule to oxide ion (d) Iron (III) ion to iron(II) ion
- (iii) Which of the following elements in the given reaction is reduced?
$$\text{ZnO} + \text{H}_2 \longrightarrow \text{Zn} + \text{H}_2\text{O}$$

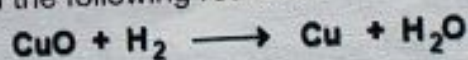
(a) H_2 (b) ZnO (c) Zn (d) O
- (iv) Consider the following reaction:
$$\text{H}_2\text{S} + \text{Cl}_2 \longrightarrow 2\text{HCl} + \text{S}$$

In this reaction H_2S behave as
(a) Reducing agent (b) Oxidizing agent
(c) Catalyst (d) Electrolyte
- (v) The oxidation state of Cr in $\text{K}_2\text{Cr}_2\text{O}_7$ is
(a) 12 (b) 6 (c) 3 (d) Zero
- (vi) Which of the following statement is not correct about the galvanic cell.
(a) Cations are reduced at cathode (b) Anions are oxidized at anode
(c) Electrons flow from cathode to anode (c) Oxidation occurs at anode
- (vii) Which of the following is not true about the Daniel cell
(a) Half-cell of an active metal acts as cathode.
(b) Half-cell contains an element in contact with its ions in aqueous solution.
(c) A salt bridge connects the two half cells.
(d) A spontaneous oxidation-reduction reaction generates electricity.
- (viii) Which of the following do not involve electrolytic process?
(a) Refining of copper
(b) Manufacture of sodium from NaCl
(c) Electroplating of steel
(d) Reduction of metal oxide by a reducing agent.
- (ix) Galvanizing is
(a) Coating with Sn (b) Coating with Zn
(c) Coating with Cr (d) Coating with Cu
- (x) Which of the following is true for the Nelson cell.
(a) Sodium metal is produced at anode
(b) Chlorine gas is produced at anode
(c) Hydrogen gas is produced at anode
(d) Sodium ions are not reduced at cathode



2. Give short answers

- What is oxidation state?
- What is the oxidation number of Cr in chromic acid (H_2CrO_4)?
- Identify reducing agent in the following reaction



- Write chemical reactions that occur in Nelson's cell.
- Why tin plated steel is used to make food cans?
- Explain one example from daily life which involves oxidation-reduction reaction?

- Define oxidation and reduction in terms of loss or gain of oxygen or hydrogen.
- Define oxidation and reduction in terms of loss or gain of electrons.
- List the possible uses of electrolytic cell.
- Sketch a Daniel Cell, labeling the cathode, anode, and the direction of flow of the electrons.
- Describe how a battery produces electrical energy.
- Describe the method of recovering metal from its ores.
- Explain electrolytic refining of copper.
- Compare the effects of Al_2O_3 and Fe_2O_3 formation on their parent metals and cite examples from daily life.
- Explain how food and beverage industries deal with corrosion.
- Explain how chemistry interacts with photography.
- Electrolysis has a major role in electrochemical industries.
 - Sketch an electrolytic cell, label the anode and cathode and indicate the direction of electron transfer.
 - Describe the nature of electrochemical process.
 - Distinguish between electrolytic and voltaic cell.
- State the substances which are oxidized or reduced. Give reason for your answer.
 - $\text{N}_2 + 3\text{H}_2 \longrightarrow \text{NH}_3$
 - $\text{CO}_2 + 2\text{Mg} \longrightarrow 2\text{MgO} + \text{C}$
 - $\text{Mg} + \text{H}_2\text{O} \longrightarrow \text{MgO} + \text{H}_2$
 - $\text{H}_2\text{S} + \text{Cl}_2 \longrightarrow 2\text{HCl} + \text{S}$
 - $2\text{NH}_3 + 3\text{CuO} \longrightarrow 3\text{Cu} + \text{N}_2 + 3\text{H}_2\text{O}$
- Define oxidation number or oxidation state.
 - Find the oxidation state of nitrogen in the following compounds.

(i) NO_2	(ii) N_2O	(iii) N_2O_3	(iv) HNO_3
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- Find the oxidation state of S in the following compounds.

(a) H_2S	(b) H_2SO_3	(c) $\text{Na}_2\text{S}_2\text{O}_3$
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- Define oxidizing and reducing agents.
 - Identify the oxidizing agents and reducing agents in the following reactions:



- (i) $\text{H}_2\text{S} + \text{Cl}_2 \longrightarrow 2\text{HCl} + \text{S}$
- (ii) $2\text{FeCl}_2 + \text{Cl}_2 \longrightarrow 2\text{FeCl}_3$
- (iii) $2\text{KI} + \text{Cl}_2 \longrightarrow 2\text{KCl} + \text{I}_2$
- (iv) $\text{Mg} + 2\text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$

18. Hydrogen peroxide reacts with silver oxide and lead(II) sulphide according to the following equations.

- (i) $\text{H}_2\text{O}_2 + \text{Ag}_2\text{O} \longrightarrow 2\text{Ag} + \text{H}_2\text{O} + \text{O}_2$
- (ii) $4\text{H}_2\text{O}_2 + \text{PbS} \longrightarrow \text{PbSO}_4 + 4\text{H}_2\text{O}$

Is hydrogen peroxide an oxidizing or reducing agent in these reactions. Argue.

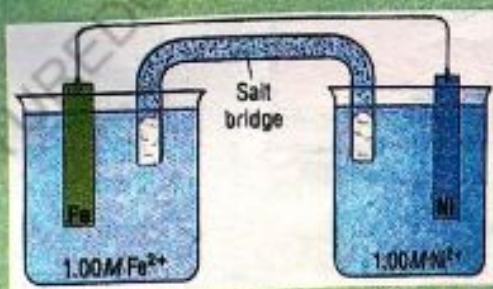
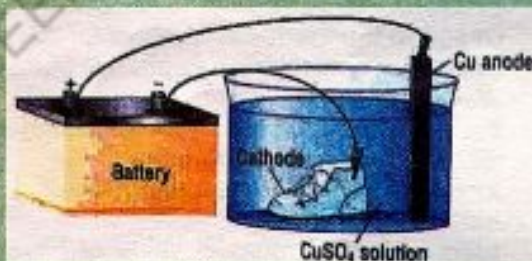


Think-Tank

- What materials do you need to electro plate copper onto an iron nail. Construct a diagram showing how these materials should be arranged.
- Evaluate the process that is occurring in the following illustration. Shoe has steel strips.
- Following redox reaction occurs in the voltaic cell illustrated below:



Select the anode, cathode and indicate the direction of flow of electrons.



- Design an experiment to demonstrate cathodic protection from corrosion.
- Would you expect corrosion to occur more rapidly in a desert or in a rainforest? Justify your stand.



8

CHEMICAL REACTIVITY



This is a 10 days lesson

After completing this lesson, you will be able to:

- Compile some important reactions of halogens.
- Describe the position of sodium in Periodic Table, its simple properties and uses.
- Describe the position of calcium and magnesium in Periodic Table, their simple properties and uses.
- Differentiate between soft and hard metals (Iron and Sodium).
- Describe the inertness of noble metals.
- Explain why alkali metals are not found in the Free State in nature.
- Explain the differences in ionization energies of alkali and alkaline earth metals.
- Identify elements as an alkali metal or an alkaline earth metal.
- Identify the commercial value of Silver, Gold and Platinum.
- Name some elements, which are found in uncombined state in nature.
- Show how cations and anions are related to the terms metals and non-metals.



Pre- Reading

Chemical reactions are heart and soul of chemistry. When an iron object is placed in moist air, it rusts. But when gold or platinum ornaments are placed in the moist air, they do not rust why? Copper is one of the important metals. It is widely used for making electrical cables, ornaments, coins and alloys such as brass, bronze etc. Commercially low grade copper ore is treated with dil. H_2SO_4 to convert ore into copper sulphate solution. From this solution copper metal is recovered by dipping iron rods in it. Iron displaces copper from copper sulphate solution. Can copper metal be used to displace iron or other metals? Iodine is an essential element for thyroid gland

and an important disinfectant. It is used for making tincture iodine. Commercially it is prepared by the oxidation of iodide salt by Cl_2 . Can iodine displace Cl_2 or Br_2 from their salts?

How is the reactivity of elements related to the outermost electrons in atoms? What does periodic table tell you about reactivity of elements? This chapter will make foundation to deal with such inquiries.



Reading

8.1 METALS

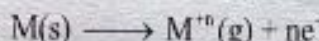
You are familiar with metals like aluminum, iron and copper. They can be transformed into a variety of shapes by melting, molding, casting, hammering and rolling. Metals occupy left



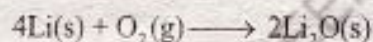
and lower three-quarters of the periodic table. They are typically shiny solids, have moderate to high melting points and are good conductor of heat and electricity. They tend to lose electrons in chemical reactions.

8.1.1 Electropositive Character

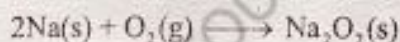
Chemically the property of an element that makes it a metal is its ability to lose valence electrons to form metal cation. **The tendency of a metal to lose electrons is called electropositivity.**



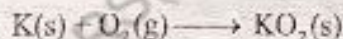
where M stands for any metal. Metals have large atomic size and low ionization energies. As ionization energy decreases down the group, the electropositivity increases. Thus sodium is more electropositive than lithium. Similarly, magnesium is more electropositive than beryllium. On the other hand, the tendency to lose electron decreases as we move from left to right in a period. For instance in the second period Li and Be are metals, B is metalloid, whereas C, N, O, F and Ne are non-metals. Also Be is less metallic than Li. These trends in electropositivity are reflected in chemical reactivity of metals. In a group the reactivity of metals with oxygen increases. For instance, alkali metals on exposure to air show an increase in affinity for oxygen as we move down the group. Lithium forms normal oxide.



Sodium forms per oxide.



K, Rb and Cs form super oxide.



In periods as we move from left to right, reactivity of metals with oxygen decreases. For instance, in the third period sodium readily reacts with oxygen to form sodium peroxide, at room temperature. Mg reacts with oxygen only on ignition. Al reacts with oxygen on heating. It is superficially oxidized at room temperature to form Al_2O_3 layer which serves as a protective covering that prevents further reaction.

You can also understand trends in electropositivity from the nature of the oxides. Metal oxides are generally basic in character. This is because they produce bases in water. Non-metallic oxides are generally acidic since they produce acids in water. The basic character of oxides decreases as we move along a period. This indicates decrease in metallic character. For example, elements of the third period form oxides such as Na_2O , MgO , Al_2O_3 , SiO_2 , P_4O_{10} , SO_2 , Cl_2O_7 . These oxides turn from strongly basic through weakly basic, amphoteric, weakly acidic to strongly acidic. See figure 8.1 to understand this trend.

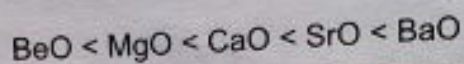
Table 8.1: Nature of oxides of elements in the period No. 3

Group	IA	IIA	IIIA	IVA	VA	VIA	VIIA
3 rd period	Na	Mg	Al	Si	P	S	Cl



Oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₂	Cl ₂ O ₇
Nature of oxide	Strongly basic	Basic	Amphoteric (both acidic and basic)	Weakly acidic	Acidic	Strongly Acidic	Very strongly acidic

On the other hand basic character of metal oxides of main group elements increases down the group. For example, in group IIA, the basic character of oxides increases in the following order.



Example 8.1: Identifying the more metallic element in a pair of elements.

In each of the following pairs of elements, identify which element is more metallic.

- (a) Na, K (b) Be, Na

Problem Solving Strategy:

3. Apply the trends of increasing electropositive character down a group and decreasing electropositive character across a period. Check the periodic table to see which elements is more metallic.

Solution:

- (a) K is more metallic (b) Na is more metallic



Self-Assessment Exercise 8.1

3. In each of the following pairs identify, which element is more metallic?
 (a) Be, B (b) Si, Al (c) K, Li
4. Using the periodic table, rank each set of elements in order decreasing metallic character
 (a) Na, Li, K (b) Al, Na, Mg
5. Which is more basic Li₂O or Na₂O? Argue.



Reading

8.1.2 Comparison of Reactivity of Alkali and Alkaline Earth Metals

To make a comparison of reactivity of alkali and alkaline earth metals. We need to know their position in the periodic table and their relative tendencies to lose electrons.



Position of Alkali Metals and Alkaline Earth Metals in the Period Table

The group IA elements except hydrogen are known as alkali metals. These include lithium, sodium, potassium, rubidium, cesium and francium. These elements have general electron configuration ns^1 in their valence shell. The Group IIA elements are known as alkaline earth metals. These include beryllium, magnesium, calcium, strontium, barium and radium. They have general electron configuration ns^2 in their valence shell.



Self-Assessment Exercise 8.1

Identify the position of Mg and Ca in the periodic table.

Ionization Energies of Alkali and Alkaline Earth Metals

Table 8.2 shows ionization energies of alkali and alkaline earth metals (in KJ / mole)

Because alkali metals have low ionization energies, therefore have a great tendency to lose the single valence electron to form cations. They are so reactive that they are never found in the Free State. The alkali metals show increased reactivity down the group. This is because of decrease in ionization energies down the group. This means Na is more reactive than Li, but less reactive than K. They are all so reactive that they are kept under a liquid such as kerosene oil.

Alkaline earth metals also have low ionization energies, so they also have great tendency to lose both the valence electrons to form divalent cations. They are less reactive than alkali metals. Think Why? These metals also show increased reactivity down the group. Thus Mg is more reactive than Be but less reactive than Ca. Like alkali metals, they are also kept under a liquid such as kerosene oil to prevent contact with moist air.

Table 8.2	
Li	Be
520	899
Na	Mg
495	738
K	Ca
419	590
Rb	Sr
403	549
Cs	Ba
376	503

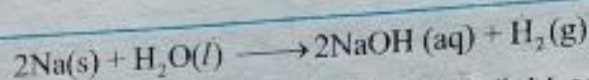


Reading

Some Important Alkali and Alkaline Earth Metals

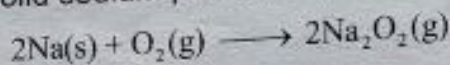
The most important alkali metal is sodium. We know that iron is the most commonly used metal. It is hard and strong. So when the term metal comes in our mind, we think it to be as hard as iron. But sodium is a soft metal can be cut with a knife. It melts at 97.8°C and boils at 881.4°C . The reason is that in iron, metal atoms are tightly packed due to small size. They are held through strong metallic bonds due to many electrons in the valence shell. Whereas sodium consists of bigger atoms. Because of large size and only one electron, sodium atoms have weak metallic bonds. Thus sodium is a soft metal.

Sodium is a silvery white soft metal. It is an extremely reactive metal. It reacts with water violently to form sodium hydroxide and liberates hydrogen gas.



This reaction is highly exothermic and proceeds with light explosion. For this reason, it should not be picked with fingers.

In a limited supply of oxygen sodium burns to form sodium oxide (Na_2O). But in excess of oxygen it forms pale yellow solid sodium peroxide (Na_2O_2).



Sodium mixed with potassium forms a liquid alloy that is used as a coolant in nuclear reactors. Sodium is also used in sodium lamps to illuminate highways. The petroleum industry utilizes sodium in the production of antiknock compounds for gasoline.

Do You Know?

Strontium-90, a radioactive isotope, is a major product of an atomic bomb explosion. Constant exposure of the body to high energy radiations emitted by the Sr-90 can lead to anaemia, leukemia and other chronic illnesses.

Society, Technology and Science

Copper, silver and gold have been used for thousand of years to make jewelry and coins. These metals have beautiful high luster and malleability. Pure gold is too soft to be used alone for making jewelry. It is alloyed with silver or copper to make it harder. A karat system is used to express purity of gold. Pure gold is 24 Karat. Coinage

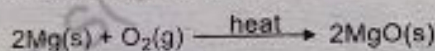
gold is 22 Karat or $\frac{22}{24} \times 100 = 92\%$

gold. Copper is a common coinage metal though modern pennies are only copper plated. Coins are made from Cu-Ni alloy.

Magnesium and calcium are the most common alkaline earth metals. Metallic bonding in these metals involve two electrons. Therefore these metals are relatively soft but are much harder than alkali metals. They are relatively reactive but less reactive than alkali metals. Magnesium reacts with steam but calcium reacts with cold water.



Magnesium reacts with oxygen at high temperature whereas calcium reacts with oxygen at room temperature.

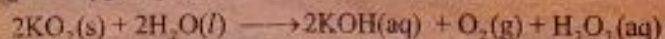


Hence Ca is more reactive than Mg.

Magnesium has relatively low density and it is used in making light weight alloy with aluminum that are used for making frames of automobiles, air craft and spaceship, cameras etc. Since magnesium burns brilliantly, it is used in photo flash guns. Mg(OH)_2 called milk of magnesia is commonly used as antacid.

Science Tit Bits

Potassium in excess air forms Potassium Super Oxide KO_2 . Potassium super oxide reacts with water to give oxygen gas.



This reaction is utilized in breathing equipments. Exhaled air contains both moisture and carbon dioxide. The moisture reacts with KO_2 in the apparatus to produce O_2 gas. KO_2 also reacts with exhaled CO_2 which produces more oxygen.

Calcium is an important component of teeth and bones. Calcium ions are found in all living cells. They are involved in clotting of blood. A proper Ca^{+2} and K^+ ions balance is required for normal heart function.

8.1.3 Inertness of Noble Metals

The chemistry of metals is characterized by their ability to lose electrons to form



cations. Some metals such as copper, silver, gold and platinum are relatively difficult to oxidize. Therefore, these metals are often called noble metals. Gold and platinum exist mostly as free elements in nature. Copper and silver exist in both free and combined states. All active metals react with HCl but noble metals do not react with HCl. Copper and silver react with strong oxidizing agents such as conc. HNO_3 and HClO_4 . Gold and platinum react only with aqua regia. Aqua regia is a mixture of 3 parts by volume of conc HCl and one part by volume of conc HNO_3 .

8.1.4 Importance of Silver, Gold and Platinum

Silver and gold have been admired for thousands of years and are still used in jewelry and fine flatware and ornaments. These metals have beautiful luster and high malleability, therefore can be transformed easily into different shapes. However both are too soft to be used for these purposes. Both are alloyed with copper to make them harder. In former times gold and silver were used in coins. But now-a-days copper-nickel alloys are universally used for making coins. Silver is an excellent conductor of both electricity and heat. It is used for making utensils. You might have had the painful experience of using a silver spoon to take soup. The light sensitive silver halides are used in photographic processes. Gold alloys are extensively used in dentistry. Some compounds of gold have useful therapeutic applications. For example salts of gold are used to treat rheumatoid arthritis.

Do You Know?

Rolled gold is a thin layer of gold alloy that is bonded onto brass or nickel-silver alloy.

Platinum is widely used as catalyst for many types of industrial processes. For example, 100% pure sulphuric acid is prepared by the contact process. In this process platinum is used as catalyst. Platinum is used as electrode as a part of hydrogen electrode and in fuel cells. Automobile exhaust is a major source of air pollution. Therefore, most new cars are equipped with catalytic converters. These converters contain platinum. These converters catalyze the complete combustion of CO and hydrocarbons. A platinum compound called cis-platin is useful as an anti cancer agent.

8.2. NON-METALS

Elements that conduct heat and electricity poorly are called non-metals. Non-metals appear in the upper right portion of the periodic table except hydrogen.

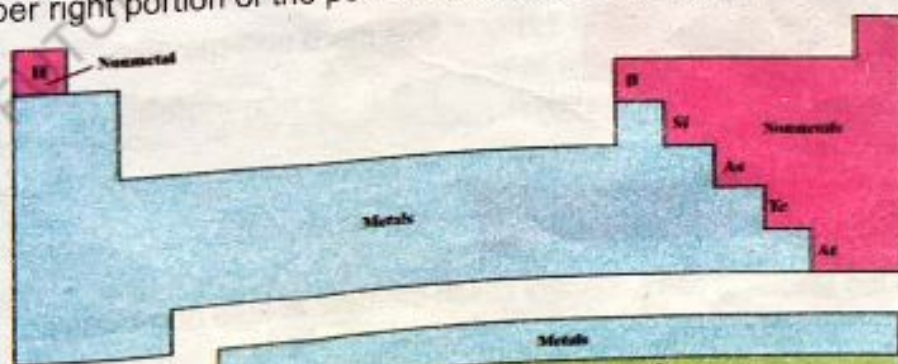


Figure 8.3: Metals and non-metals



Teacher's Point

A teacher may give examples of metals, nonmetals and metalloid.



Non-metals are generally gases or dull, brittle solids at room temperature except bromine which is liquid. Several elements near the stepped line have mixed properties of metals and non-metals. These elements are called metalloids. Some of these play important role in modern electronic materials. For example, silicon and germanium are metalloids. These elements are used in the manufacture of computer chips and solar cells.

8.2.1 Electronegative Character

Electronegative character is the reverse of electropositive character. It is the tendency to gain electrons to form anions. As you move from left to right across a period, the tendency to gain electrons increases.



This is due to increase in nuclear charge and decrease in atomic size, which binds the extra electron more tightly to the nucleus. But in a group going from top to bottom, electronegative character decreases, due to increase in atomic size. Halogens are the most electronegative elements in their respective periods. Non-metal oxides are generally acidic since they yield acids in water. Acidic strength of these oxides increases from left to right in a period.

Example 8.1: Identifying the more non-metallic element in a pair of elements.

Do You Know?

Computers exist because of an amazing invention called the integrated circuit known as the silicon chip or just chip. Chips are microprocessors which can perform all the function of a computer. These are so small that one will fit on the tip of your finger. A chip is made from a very thin flake of an element called silicon. Specific areas of the chip are cut or etched, in order to make a pattern called circuit. This is a kind of pattern along which electric current flows. Chips contain millions of working parts that are so small that they can be seen only with a microscope.

In each of the following pairs of elements, identify which element is more non-metallic.

- (a) N or O (b) F or Cl

Problem Solving Strategy:

Apply the trends of increasing electronegative character in a period and decreasing electronegative character in a group. Check the periodic table to see which element is more electronegative or non-metallic.

Solution:

- (a) O is more non-metallic
(b) F is more non-metallic



Self-Assessment Exercise 8.3

- In each of the following pairs, identify which element is less non-metallic.
(a) B or C (b) C or Si (c) Cl or Br.
- Using periodic table rank each set of elements in order of increasing non-metallic character.
(a) N, F, O (b) Cl, Br, I (c) Si, S, P

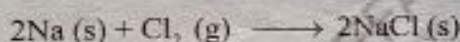
Reading

8.2.2 Comparison of Reactivity of the Halogens

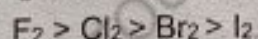
The elements in group VII-A are called halogens. The name halogen is derived from the Greek words "halous" meaning salt and "gen" meaning former. Halogens include fluorine, chlorine, bromine, iodine and astatine (astatine is a radio active element. Little is known about its properties.) All halogens are reactive non-metals. They all exist as diatomic molecular substances.

Fluorine (F ₂)	→	Pale yellow gas
Chlorine (Cl ₂)	→	Greenish yellow gas
Bromine (Br ₂)	→	Reddish-brown liquid
Iodine (I ₂)	→	Bluish-black solid

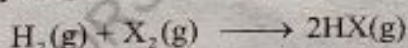
All the halogens react with metals to form salts called halides. In these reactions metals are oxidized. So, halogens act as oxidizing agents. Fluorine is the most reactive element known. Chlorine is less reactive, but combines vigorously with many metals. For instance sodium metal burns in chlorine gas to form sodium chloride.



Bromine and iodine react with metals less vigorously. The reactivity of halogens decrease down the group. Thus, the order of decreasing oxidizing power is

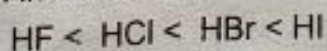


The halogens react with hydrogen to form hydrogen halides.



Where X = F, Cl, Br, I

Reaction of fluorine with hydrogen is explosive. The hydrogen halides dissolve in water to form hydrohalic acid e.g., hydrochloric acid HCl, hydrofluoric acid HF etc. Except HF, other hydrohalic are strong acids. The acidic strength increases in the following order:



Oxidizing power of F₂ is the highest and that of I₂ is lowest. Due to the relative strength as oxidizing agent, it is possible for a free halogen to oxidize the ion of halogen next to it in the group. This means F₂ can oxidize all the halide ions to free halogen. For example,



Similarly Cl₂ can oxidize Br⁻ and I⁻ ions. But I₂ can not oxidize any halide ion.

Teacher's Point

A teacher may show Bromine or Iodine to students in class.

**Self-Assessment Exercise 8.4**

Which of the following displacement reactions will occur?

1. $\text{Cl}_2(\text{g}) + 2\text{NaF}(\text{aq}) \longrightarrow 2\text{NaCl}(\text{aq}) + \text{F}_2(\text{g})$
2. $\text{Br}_2(\text{g}) + 2\text{KI}(\text{aq}) \longrightarrow 2\text{KBr}(\text{aq}) + \text{I}_2(\text{g})$
3. $\text{I}_2(\text{g}) + 2\text{KBr}(\text{aq}) \longrightarrow 2\text{KI}(\text{aq}) + \text{Br}_2(\text{l})$
4. $\text{Cl}_2(\text{g}) + 2\text{KBr}(\text{aq}) \longrightarrow 2\text{KCl}(\text{aq}) + \text{Br}_2(\text{l})$
5. $\text{Cl}_2(\text{g}) + 2\text{NaI}(\text{aq}) \longrightarrow 2\text{NaCl}(\text{aq}) + \text{I}_2(\text{s})$

**Activity 8.1**

Identifying cations such as NH_4^+ , Ca^{+2} , Mg^{+2} , Ba^{+2} and Zn^{+2}

You will need:

- Test tubes.
- Four Glass watches.
- Glass rods.
- Conc HCl, $\text{Co}(\text{NO}_2)_2$, NaOH.
- A platinum wire.
- Salts containing NH_4^+ , Ca^{+2} , Mg^{+2} , Ba^{+2} and Zn^{+2} ion.

Carry out the following:

1. Make a thick paste of each salt with conc HCl. Take a little of the paste on a clean platinum wire and place it into the non-luminous flame of burner. Note the colour of the flame. Apple green flame indicates Ba^{+2} ion. Brick red flame indicates Ca^{+2} ion. Other cation NH_4^+ , Mg^{+2} , Zn^{+2} do not give coloured flame.
2. Shake each salt with a few drop of cobalt nitrate solution in separate test tubes. Dip a piece of filter paper in each test tube, dry it and burn. Note the colour of the ash. This test is called as filter ash test. Pink colour of ash indicates Mg^{+2} salt while green colour of ash indicates Zn^{+2} salts.
3. Take each salt in separate test tubes and add NaOH solution to it and boil. Note the test tube in which pungent odour gas evolves. Test this gas with a glass rod dipped in conc HCl. This gas will give dense white fumes of NH_4Cl . This test will confirm that the salt contains NH_4^+ ion.

**Activity 8.2**

Identifying anions such as CO_3^{+2} , SO_4^{+2} , Cl^- , I^- and NO_2^-

You will need:

- Test tubes, glass rods.



8 Chemical Reactivity

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- Chemical such as dil H_2SO_4 , conc H_2SO_4 , FeSO_4 , NH_4OH , starch solution, lime water.
- Salts containing CO_3^{2-} , SO_4^{2-} , Cl^- , I^- and NO_3^- .

Carry out the following:

Take small quantity of each salt in separate dry test tubes. Add few drops of dil H_2SO_4 solution. Note the test tube in which a gas evolves. Test each gas as follows:

- (a) Pass colourless odourless gas through lime water. It will turn milky. It indicates CO_3^{2-} .
- (b) If reddish brown gas evolves, test it with a filter paper dipped in FeSO_4 solution. You will notice, paper will turn black. It indicates $\rightarrow \text{NO}_2^-$.
- (c) Take small quantities of remaining salts in dry test tubes and this time add few drops of conc H_2SO_4 heat, and analyze as follows:

- i. Note the salt that gives colourless gas having pungent odour. Test this gas with a glass rod dipped in liquid NH_3 or NH_4OH solution. It will give dense white fumes. It must be HCl . This indicates $\rightarrow \text{Cl}^-$ ion.
- ii. Note the salt that gives violet fumes having pungent odour. Test these vapours with starch paper, it will turn blue. This indicates $\rightarrow \text{I}^-$ ion.
- iii. Now, after these tests only one salt is left that contains SO_4^{2-} ion. To test this ion, dissolve small amount of this salt in water in a test tube and add BaCl_2 solution. You will observe white precipitate in the test tube. Add conc HNO_3 to the ppt and shake. You will observe ppt remains undissolved. This indicates $\rightarrow \text{SO}_4^{2-}$ ion.

8.1.2 Comparison of Reactivity of Alkali Metals



Activity 8.3

Metal displacement reaction in aqueous medium

You will need:

- Test tube.
- Spatula.
- Zinc plate.
- Copper (II) Sulphate.

Carry out the following:

1. Mix one spatula full of copper (II) sulphate in 20cm^3 of water in the test tube.
2. Place zinc plate in the copper (II) sulphate solution and leave for few hours (as shown in figure 8.4).



Figure 8.4: Zinc dipped in Copper (II) sulphate



3. Look for the signs of red-brown particles on the surface of zinc plate.
4. Also notice the change in the colour of copper (II) sulphate solution.
5. Write conclusion:
6. Is zinc more active than copper?



Key Points

- The tendency of a metal to lose electrons is called electropositivity.
- As ionization energy decreases down the group, the electropositivity increases.
- Metal oxides are generally basic in character since they produce bases in water.
- Non-metal oxides are generally acidic since they produce acids in water.
- Basic character of metal oxides of main group elements increases as we move down the group.
- The ionization energies of Group IIA elements are greater than those for the elements of Group IA. So alkaline earth metals are less reactive than alkali metals.
- Sodium mixed with potassium forms an alloy that is used as a coolant in nuclear reactors.
- Metals such as copper, silver, gold and platinum are relatively difficult to oxidize. Therefore, these metals are often called noble metals.
- Elements that conduct heat and electricity poorly are called non-metals. (Except graphite)
- All the halogens react with metals to form salts called halides.

REFERENCES FOR ADDITIONAL INFORMATION

- B. Earl and LDR Wilford, Further Advanced Chemistry.
- B. Earl and LDR Wilford, Introduction to Advanced Chemistry.
- David E. Goldberg, Fundamental of Chemistry.
- Addison Wesley, Chemistry.



Review Questions

1. Encircle the correct answer:

- (i) Which of the following oxides is the most basic oxide?
(a) MgO (b) BaO (c) CaO (d) SrO
- (ii) Which of the following oxides is amphoteric oxide?
(a) Na₂O (b) Al₂O₃ (c) MgO (d) SO₂
- (iii) Which of the following elements is most metallic?
(a) Al (b) Na (c) Mg (d) Li
- (iv) Valence shell electronic configuration of alkaline earth metals is:
(a) ns¹ (b) ns²np¹ (c) ns² (d) ns²np⁶
- (v) Which of the following metals is not noble metal.
(a) Au (b) Cu (c) Hg (d) Ag



- (vi) Which of the following elements is the least non-metallic?
(a) Si (b) S (c) P (d) Cl
- (vii) Oxidizing power of _____ is the highest.
(a) I_2 (b) Cl_2 (c) Br_2 (d) F_2
- (viii) Chlorine cannot oxidize _____ ion.
(a) Iodide (b) Fluoride (c) Bromide (d) Astatine
- (ix) Na is more reactive than Li, but less reactive than
(a) Mg (b) Al (c) Rb (d) All of these
- (x) Which of the following is the strongest acid?
(a) HI (b) HCl (c) HF (d) HBr

2. Give short answers

- In a group the reactivity of metals with oxygen increases? Give example.
- Which element is more metallic Mg or Al? Explain.
- What is the importance of Mg?
- Arrange the following in order of increasing acidic strength
HF, HI, HBr, HCl
- Can F_2 oxidize all the halides ions to free halogen?

**3. Arrange the following oxides in order of decreasing basic character.
BeO, CaO, MgO, SrO****4. Rank the each set of elements in order of increasing metallic character.**

- (a) Al, Na, Mg
(b) Na, Li, K

5. Which of the following displacement reaction will not occur? Give reason.

- (a) $Cl_2 + 2NaF \longrightarrow 2NaCl + F_2$ (b) $Br_2 + 2KI \longrightarrow 2KBr + I_2$
(c) $I_2 + 2KBr \longrightarrow 2KI + Br_2$

6. Give some important applications of platinum.**7. What is a half reaction?****8. What is aqua regia?****Think-Tank**

- It is advisable, not to pick sodium metal with fingers. Defend this statement.
- How do electrons participate in half-reactions?
- Design an experiment to show that iron is more reactive than copper.
- How are voltaic cells and electrolytic cells alike and how are they different?
- An iron bar is to be silver plated which metal would be the anode and which the cathode?
- Compare corrosion of Al and Fe.
- Compare and contrast alkali and alkaline earth metals.
- Choose three metals that would not be suitable for coinage. Elaborate your choice.



Glossary

Absolute temperature is the temperature on Kelvin scale on which zero is marked as -273°C .

Absolute zero is the lowest possible temperature and is equal to -273°C .

Allotropes are different physical forms of a pure element.

Alloy is a solid solution of metals, it may contain small amount of non-metals.

Alpha particle is doubly positively charged helium nuclei.

amu is the abbreviation for atomic mass unit. It is exactly $1/12^{\text{th}}$ mass of one atom of carbon-12.

Anion is an atom or group of atoms having a negative charge.

Anode is an electrode at which oxidation occurs.

Atmosphere is the unit of pressure of air at the earth's surface.

Atom is the smallest particle of an element to retain the properties of the element.

Atomic mass is the exact mass of one atom of an element.

Atomic number is the number of protons in the nucleus of an atom.

Boiling point is the temperature at which vapour pressure of a liquid equals the atmospheric or external pressure over the liquid.

Bonding is the term used to describe ways in which atoms are held together in elements and compound.

Cathode is the electrode at which reduction occurs.

Cation is an atom or group of atoms having a positive charge.

Charles's Law describes relationship between volume occupied by a gas and temperature at constant pressure.

Chemical bond is the force that holds two or more atoms together in a substance.

Chemistry is the study of matter, its composition, properties and the interaction of atoms as they react to form new substances.

Colloid is a heterogeneous mixture of a substance dispersed through a medium.

Compound is a substance made up of elements in fixed ratio.

Concentration is the measure of the amount of solute dissolved in a solution.

Condensation is the change of a gas into a liquid.

Conductor is a substance that can conduct electric current.

Covalent bond is a bond formed by the mutual sharing of electrons between two atoms.

Distillation is the process for purifying a liquid by evaporating it, condensing the vapours and collecting the liquid in a clean container.

Double bond is a covalent bond formed by two pairs of shared electrons.



Glossary

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Electrolyte is a substance which in fused form or in solution form allows current to flow through it.

Electron is the sub-atomic negatively charged particle in an atom.

Electron configuration is the arrangement of electrons in orbitals.

Electron shells are sets of energy levels corresponding to a single value of n .

Electronegativity is the power of an atom to attract electrons to itself.

Element is the pure substance in which all the atoms have same atomic number.

Energy levels are the discrete energies of electron in atom.

Evaporation is the escape of molecules of a liquid through the surface.

Formula mass is the sum of atomic masses of all the atoms present in the compound. It can also be defined as the mass of one formula unit or molecule in amu or mass of one mole of the compound in grams.

Formula unit is the simplest formula of an ionic compound.

Freezing point is the temperature at which a liquid changes to a solid.

Gas is one of the physical states of matter that has neither any definite shape nor fixed volume.

Gold is an elemental metal, its symbol is Au.

Gram formula mass is the mass of 1 mole of a compound or element in grams.

Group refers to the vertical column on the periodic table.

Halogens are the elements present in VIIA group on the periodic table.

Hydrochloric acid is an aqueous solution of hydrogen chloride.

Hydrogen bond is an intermolecular force due to the interaction of lone pair present on N, O or F atom and the partially positively charged hydrogen of another molecule attached with N, O or F.

Inorganic refers to a compound made up primarily of atoms other than carbon.

Intermolecular forces are the attractive forces between the molecules.

Ion is an atom or a molecule possessing an electric charge.

Ionic bond is the bond formed by the attraction of oppositely charged ions.

Isotopes are atoms of an element that differ in the number of neutrons.

Kelvin is a temperature scale, the zero of which corresponds to -273°C .

Liquid is one of the states of matter that has definite volume but no definite shape.

Lone pair is a pair of non-bonding valence electron or a pair of electron that are not shared in covalent bonding.

Main group elements are elements belonging to sub-group A on the periodic table.

Mass number is the sum of the number of protons and neutrons in an atom.

Matter is anything that occupies space and has mass

Melting point is the temperature at which a substance melts.



Metals are elements that are good conductor of electricity and heat and tends to form positive ions.

Molar mass is the mass of one mole of a compound or element in grams.

Molarity is the number of moles of solute per dm^3 of solution.

Mole is Avogadro's number of atoms, molecules or formula units. It can also be defined as the atomic mass, formula mass or molecular mass of a substance expressed in grams.

Molecular mass is the sum of atomic masses of all the atoms present in a molecule.

Molecule is a group of atoms held by a chemical bond and behaves as an independent unit.

Neutron is a sub-atomic neutral particle in the nucleus of an atom.

Non-metal is an element that is a bad conductor of electricity and tends to form negative ion.

Nucleus is the positively charged center of an atom.

Octet is eight electron configurations in the valence shell.

Orbitals are regions of space around the nucleus where electron is likely to be present.

Organic chemistry is the study of carbon compound or the study of hydrocarbons and their derivatives.

Organic compounds are compounds mainly composed of carbon atoms.

Oxidation is the loss of electron by an element.

Oxidation state is the actual charge on an ion or apparent charge on an atom in a molecule.

Oxidizing agent is the reactant containing the element that is reduced.

Period refers to the horizontal row in the periodic table.

Periodic table is a systematic arrangement of elements based on electron configuration and periodic behaviour.

Polar covalent bond is a bond in which electrons are shared unequally between two atoms. Thus one atom bears partial positive charge and the other a partial negative charge.

Proton is the positively charged sub-atomic particle in the nucleus of an atom.

Reactant is a substance consumed in a chemical reaction.

Redox reaction is an oxidation and reduction reaction.

Reduction is the gain of electron by an atom.

Reducing agent is the reactant containing the element that is oxidized.

Solid is one of the three states of matter, it has a definite shape and definite volume.

Solubility is the amount of solute that dissolved in 100g of water at a particular temperature.

Solute is the component of solution that is present in relatively small amount.

Solvent is the component of solution that is present in relatively larger amount.

S-orbital is a spherical shape orbital.

Sublimation is the conversion of a solid directly into a gas.

Technology is the application of scientific knowledge to serve some practical purpose.



Glossary

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Torr is the unit of pressure equal to 1mm of Hg.

Triple bond is a covalent bond involving the sharing of three pairs of electrons.

Valence electrons are the electron in the outer most energy level of the atom.

Valence shell is the outermost energy level of the atom.

Vapour pressure is the pressure exerted by the vapours of a liquid in equilibrium with its liquid at a particular temperature.

Zero group refers to VIIIA group on the periodic table.

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Answers

Chapter 1 (Review Questions)

(a) 0.6 (b) 0.021 (c) 0.256 (d) 1.25 (e) 37.5

Q. 7

(a) 46.8g (b) 150g (c) 4.5g (d) 261.96g (e) 14.7g

Q. 8

(a) 1.505×10^{24} (b) 2.05×10^{24} (c) 5.56×10^{23} (d) 6.02×10^{21}

Q. 14

(a) 2.05×10^{24} (b) 6.02×10^{23} (c) 3.01×10^{24}

Q. 15

(a) 3.01×10^{-4} (b) 9.3×10^{-4} (c) 2.99×10^2 (d) 1.346×10^{-16}

Chapter 6

Q. 8 1.14M

Q. 9 245g

Q. 10 $7 \times 10^{-5}M$

Q. 11 $6.3 \times 10^{-4}M$

Q. 12 0.25M

Chapter 1. (Think Tank)

(1) 23g (2) 48g (3) 1.66×10^{-24} (4) 1.806×10^{24}

(5) 1.204×10^{24}

Chapter 6 (Think Tank)

Q. 8 $4 \times 10^{-6}M$

Q. 9 0.278M

Q. 10 0.3