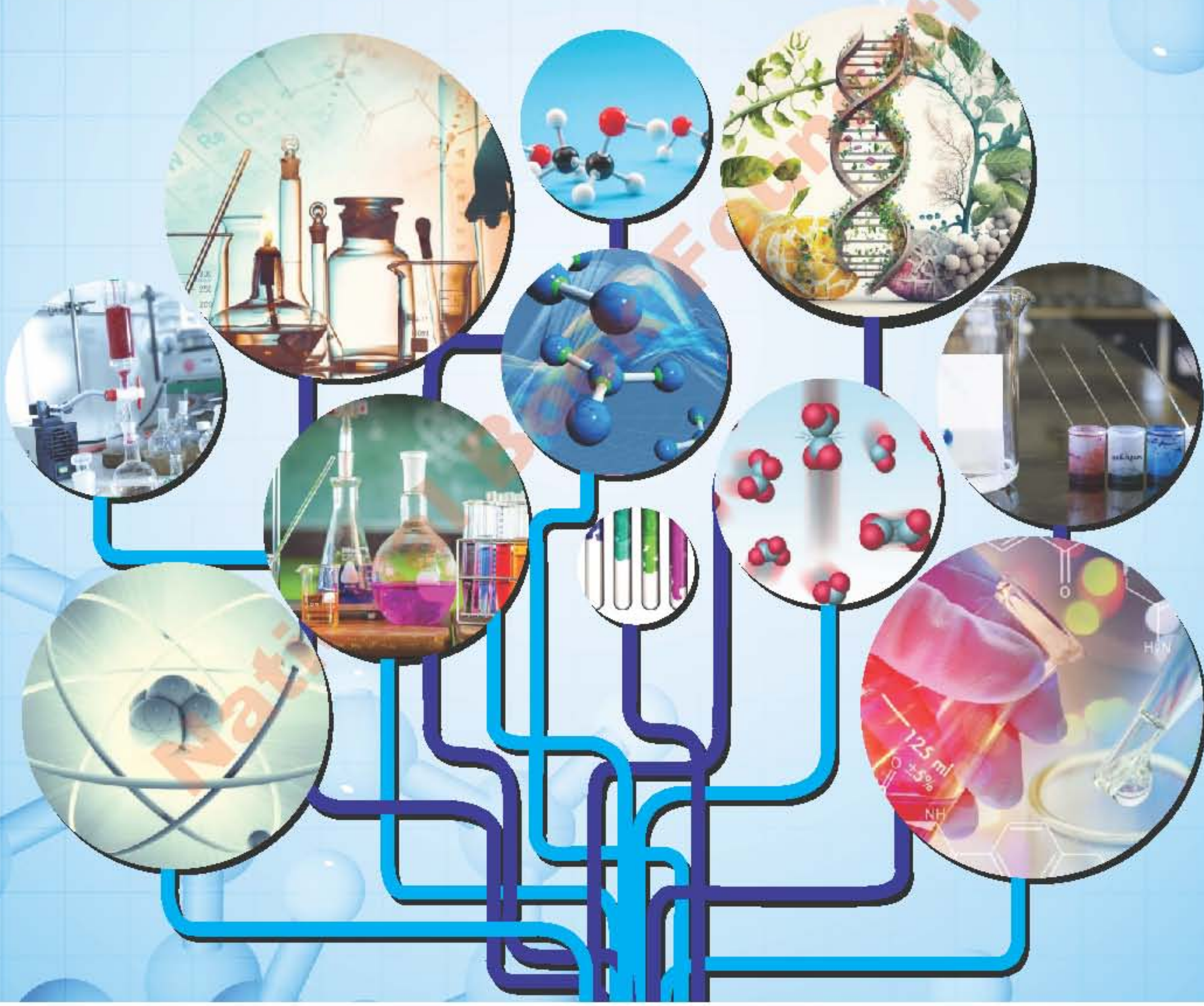
Model Textbook of

CHEMISTRY

Grade 9

Based on National Curriculum of Pakistan 2022-23



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Model Textbook of **Chemistry**
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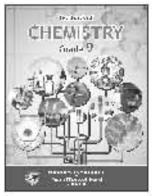

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PREFACE

In a historic footstep, the national curriculum of Pakistan 2022-2023 has introduced a new era for schooling in the country, This is the first-ever core curriculum in the 75-year history of Pakistan. It is in line with the protected right to school education by Article 25-A.

Chemistry might be a difficult subject for someone, but it holds significance for those who embrace a systematic approach to understanding its concepts.

This new Textbook has been developed as a model Textbook for Pakistan. The book consolidates critical thinking methodologies, guiding scientific reasoning, and thinking abilities. The book incorporates problem-solving strategies, which will guide students toward analytical thinking and skills. These skills would be invaluable for both academic as well as practical life.

The book also inspires concept assessment exercises in every unit, which have been designed to evaluate acquired knowledge and promote critical thinking and analyzing data..

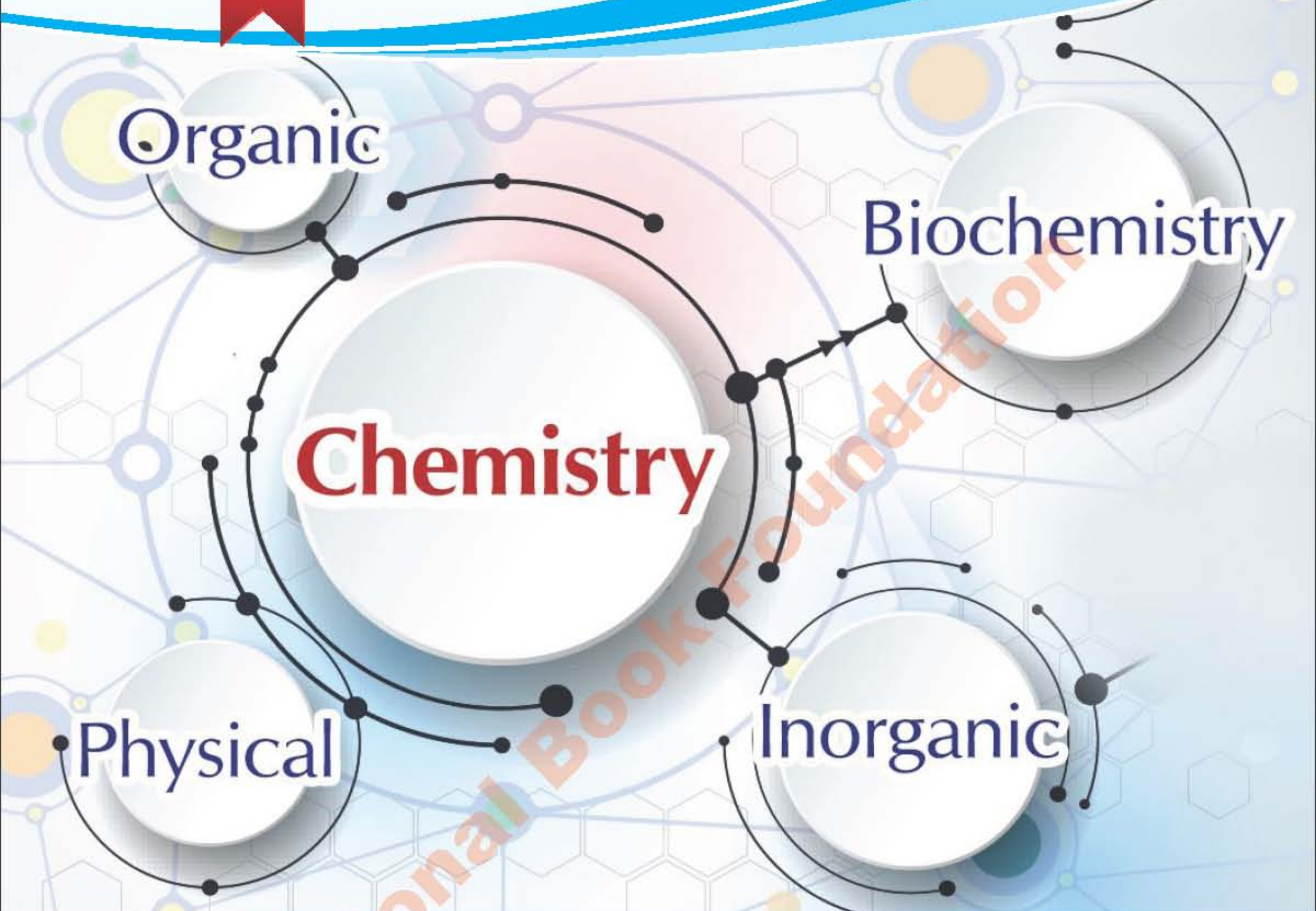
One of the book`s distinctive features is the key points at the end of each unit, which serve as a quick reference to reinforce the salient features of each unit.

Dr. Raja Mazhar Hameed

Managing Director

CONTENTS

UNIT 01



NATURE OF SCIENCE IN CHEMISTRY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define chemistry as the study of matter, its properties, composition, and its interactions with other matter and energy.
- Explain with examples that chemistry has many sub-fields and interdisciplinary fields.
- Formulate examples of essential questions that are important for the branch of chemistry.
- Differentiate between 'science' 'technology' and 'engineering' by referring to examples from the physical sciences.

1.1 DEFINITION OF CHEMISTRY AND ITS INTERACTION WITH OTHER MATTER AND ENERGY:

Chemistry is defined as the science that investigates the materials of the universe and the changes that these materials undergo. Chemistry deals with the composition, structure, properties, behavior, and changes of matter and energy. Understanding the fundamental concepts of chemistry help to explain natural phenomena and apply them to the formation of new substances, drugs, and technologies.

DO YOU KNOW?

- How green chemistry is helpful in understanding and reducing pollution?
- Green Chemistry is the model of chemical products and processes that reduce the use of hazardous substances.

1.2: BRANCHES OF CHEMISTRY

Chemistry is a diverse field of study, surrounding numerous sub-fields and interdisciplinary areas.

1. Organic Chemistry

Organic chemistry is a branch of chemistry that deals with substances containing carbon (except carbonates, bicarbonates, oxides, and carbides).

2. Inorganic Chemistry

Inorganic chemistry is a branch of chemistry that deals with elements and their compounds except organic compounds.

3. Physical Chemistry

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

4. Analytical Chemistry

Analytical chemistry is a branch of chemistry that deals with the methods and instruments for determining the composition and properties of matter.

5. Biochemistry

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

6. Environmental Chemistry

Environmental chemistry is the branch of chemistry that deals with the study of chemical and toxic substances that pollute the environment and their adverse effects on human beings.

7. Industrial Chemistry

Industrial chemistry is the branch of chemistry that deals with the large-scale production of chemical substances.

8. Medicinal Chemistry

The branch of chemistry deals with the study of the interaction between drugs and biological targets, as well as the development of new medicinal agents.

9. Polymer Chemistry

The branch of chemistry that focuses on the study of polymers, their types, properties, uses, importance, and types of polymerizations is called polymer chemistry. Examples of synthetic polymers include nylon bearings, plastic bags, polyethylene cups, polyester, Teflon coated cook ware, and epoxy glue etc.

10. Geochemistry

Geochemistry is the branch of chemistry that deals with the study of chemical composition, distribution, and transformation of elements and compounds in the Earth's crust, such as rocks, minerals, soils, water, and the atmosphere.

11. Nuclear Chemistry

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

12. Astrochemistry

Astrochemistry is a branch of chemistry that deals with the study of chemical processes and reactions that occur in astronomical environments, such as stars, planets, comets, and interstellar space.

1.3 EXAMPLES OF ESSENTIAL QUESTIONS THAT ARE IMPORTANT FOR THE BRANCHES OF CHEMISTRY

Some essential questions for various branches of chemistry that can help enhance understanding are as follows:

Physical Chemistry

1. What is the structure of an atom, and how does it influence chemical behavior?
2. How do different types of chemical bonds (ionic, covalent, metallic) form and function?

Organic Chemistry

1. Why carbon is considered the backbone of organic compounds?
2. What are the major functional groups in organic molecules, and how do they affect chemical properties?

Inorganic Chemistry

1. What distinguishes inorganic compounds from organic compounds?
2. How does Periodic table helps to organise elements?

Analytical Chemistry

1. How are analytical methods used to identify and quantify chemical substances?

Biochemistry

1. How do biomolecules such as carbohydrates, proteins, nucleic acids, and lipids contribute to the structure and function of living organisms?

Environmental Chemistry

1. How do human activities contribute to air pollution, and what are the consequences for the environment?
2. What role do greenhouse gases play in climate change, and how can we mitigate their effects?

Medicinal Chemistry

1. How are drugs designed and developed for specific therapeutic purposes?

Polymer Chemistry

1. What are polymers, and how do their structures affect their properties?

Geochemistry

1. How do geological processes influence the distribution of elements in the Earth's crust?

Nuclear Chemistry

1. How do nuclear reactions differ from chemical reactions, and what are their applications?
2. What is the role of radioisotopes in medicine and industry?

Astronomy

1. What types of reactions occur in astronomical environments?

These questions can serve as a foundation for exploring the key concepts within each branch of chemistry.

1.4: DAILY LIFE APPLICATIONS OF CHEMISTRY

Organic Chemistry

To treat diseases, organic chemists synthesize new medicines that interact with specific targets like proteins or enzymes.

Inorganic Chemistry

Lithium-ion (Li-ion) batteries are used as rechargeable batteries for electronics, toys, wireless headphones, handheld power tools, small and large appliances, electrical storage devices, and electric vehicles.

Analytical Chemistry

Forensic chemistry is the application of analytical chemistry. It involves the examination of physical traces, such as body fluids, bones, fibers and drugs. It can be used to identify an unknown compound. For example drugs are often found in various colored powders and are analyzed to determine their content.

Physical Chemistry

Physical chemistry is a part of our everyday life. The batteries in our vehicles are built on the principle of electrochemistry.

Environmental Chemistry

Environmental chemistry is used to protect water that has been poisoned by soil, and dust by using different methods e.g., sedimentation, filtration, and disinfection.

1.5: 'SCIENCE' 'TECHNOLOGY' AND 'ENGINEERING'

Science

Science is the systematic process of constructing and organizing knowledge about the universe. Thus, science seeks to understand the natural world. For example, chemists seek to understand the behaviour and properties of materials, chemical reactions, and the fundamental principles that control the behaviour of matter.

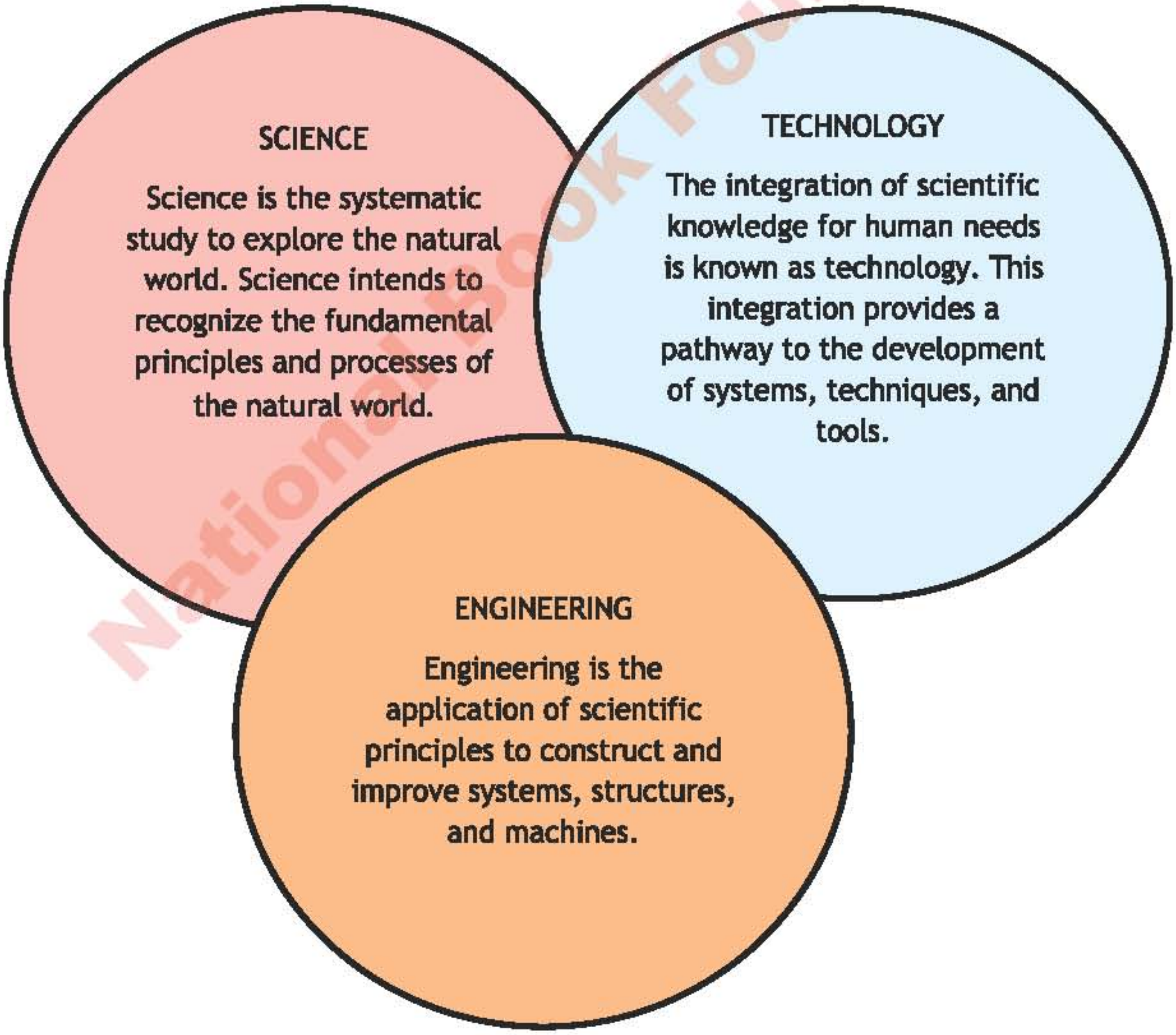
Technology

Technology is the process of applying scientific knowledge to practical applications, resulting in the creation of tools, machines, and systems that enhance our lives.

Science and technology play a major role in the field of chemistry by providing tools, machines, techniques and methods which can help in discovery and ,development of new materials .These also help in improving quality of products .Technology has revolutionized the field of chemistry, making research and applications more efficient. It has enabled chemists to more effectively analyze and identify substances. Their work is beneficial for chemists working in pharmaceutical and other chemical industries.

Engineering

Engineering is the use of science and mathematics to design and construct systems, structures, and tools for various processes. Chemical engineers develop and design manufacturing processes for the production of chemicals , fuels, food, medicines, polymers , detergents, paper etc. They often work to maximize productivity and product quality.



1.6: APPLICATIONS OF SCIENCE AND TECHNOLOGY AND ENGINEERING

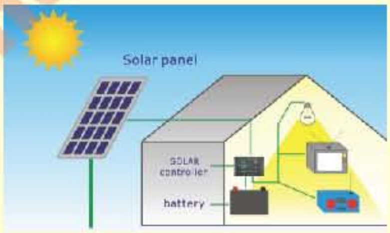
Let's take a look at how science, technology and engineering work together to solve problems in real-world situations. For example:

Example 1.1: Investigating rusting of iron.

Imagine trying to figure out why a bike or car will rust over time. Scientists could investigate the chemical reactions that occur between iron, water and oxygen that cause rust to form. Experiments could be conducted to understand the factors that influence this process and help develop strategies to prevent rust.

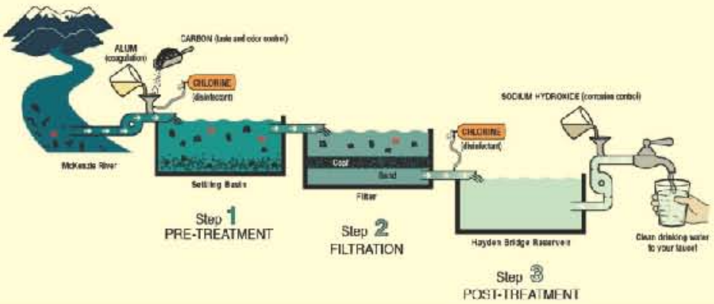
Example 1.2: Harnessing Solar Energy

Scientists may study the principles of photovoltaic cells to understand how sunlight can be converted into electricity. Technologists can develop solar panels based on the scientific principles discovered. Engineers play their role in designing and implementing large-scale renewable energy systems. For instance, an electrical engineer might design the wiring and connections of a solar power plant, a civil engineer could be involved in designing the infrastructure. In this example, science helps us understand the underlying principles of converting sunlight into electricity. Technology transforms this knowledge into practical applications, such as solar panels and energy storage systems. Engineering takes these technologies and implements them on a larger scale. Together, science, technology, and engineering contribute to the development and utilization of sustainable energy sources.



Example 1.3: Designing a Water Filtration System

let's look at how engineers design water filtration systems. Chemical engineers can help develop processes to remove contaminants from water while mechanical engineers design physical components. Together, they create a solution to clean drinking water for a community. From understanding chemical reactions to using technological devices to solving practical problems through engeneering.



Example 1.4: Organic Chemistry in Action

How do you make french fries. The oil used to fry potatoes contains carbohydrates, which are organic molecules. Scientists study carbohydrates to learn more about how they work, so food technologists extract oil from seeds. Chemical engineers design oil production equipment and processes so that oil is produced efficiently and safely for cooking.

Example 1.5: Plastic Bags

Think about the science behind plastic bags. Scientists study the small building blocks known as monomers. When monomers combine, they form long chains known as polymers. One of those long chains is polyethylene, which is one of the many polymers found in plastic bags! Engineers and technicians use these discoveries to create bags that are durable, flexible, and easy to make.

These examples demonstrate how science, technology, and engineering work together in various aspects of our daily lives. Whether it's understanding chemical reactions, using technological devices, or solving practical problems through engineering solutions, these concepts are interconnected and contribute to advancements that impact the world around us.

KEY POINTS

- Chemistry is the study of matter around us.
- The branch of chemistry deals with carbon compounds (except bicarbonates, carbonate oxides, and carbides.
- The branch of chemistry that deals with the elements and their compounds except organic compounds is called inorganic chemistry.
- Industrial chemistry is concerned with the large-scale production of chemical substances.
- The branch of chemistry that deals with the laws and theories to understand the structure and changes of matter is called physical chemistry.
- Science is defined as the study of nature.
- Technology is the application of science.

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Which branch of chemistry is the study of elements and their compounds except for organic compounds?

(a) Physical Chemistry (b) Organic Chemistry
(c) Inorganic Chemistry (d) Geochemistry Chemistry

(ii) Which branch of chemistry helps to protect water that has been poisoned by soil?

(a) Environmental Chemistry (b) Organic Chemistry
(c) Inorganic Chemistry (d) Geochemistry Chemistry

(iii) Which area of Chemistry improves to gauge the behavior of pollutants and develop techniques for pollution control?

(a) Analytical Chemistry (b) Organic Chemistry
(c) Environmental (d) Geochemistry

(iv) The branch of chemistry that helps to treat diseases, organic and to synthesize new medicines.

(a) Physical (b) Organic
(c) Inorganic (d) Environmental

(v) The branch of science helps to understand chemical products and processes that reduce the use of hazardous substances:

(a) Analytical Chemistry (b) Physical chemistry
(c) Green Chemistry (d) astrochemistry

(vi) To identify the concentration of a particular solution through titration is and application of

(a) Astrochemistry (b) Analytical Chemistry
(c) Geochemistry (d) Organic chemistry

(vii) The batteries in our vehicles are built on the principle of electrochemistry. It is the application of:

(a) Astrochemistry (b) Analytical Chemistry
(c) Organic chemistry (d) Physical chemistry

(viii) The branch of chemistry that is concerned with the large-scale production of chemical substances is:

(a) Industrial chemistry (b) Physical chemistry
(c) Inorganic chemistry (d) Environmental Chemistry

(ix) The branch of chemistry that focuses on the study of polymers, their types, properties, uses is called:

(a) Industrial Chemistry (b) Polymer chemistry
(c) Organic Chemistry (d) astrochemistry

(x) The study of the interaction between drugs and biological targets, as well as the development of new medicinal agents.

(a) Organic chemistry (b) Medicinal chemistry
(c) Inorganic chemistry (d) Environmental Chemistry

2. **Give short answer.**

(i) How does chemistry help a doctor to know about the chemical nature of medicine?

(ii) In what ways does technological innovation help to understand the development of new materials?

(iii) Differentiate between geochemistry and astrochemistry.

(iv) With the help of an example correlated the use of science, technology, and engineering.

(v) With the help of the Venn diagram compare and contrast organic and inorganic chemistry.

(vi) What are the uses of nuclear chemistry?

3. Define chemistry and its interactions with other matter and energy.
4. Describe the applications of inorganic chemistry and its importance in our daily lives?
5. With the help of few examples highlight the relation between science, technology and engineering.
6. Evaluate the role of chemistry in environmental science.
7. How does geochemistry help us to solve the problems such as pollution and climate change?
8. How is organic chemistry applied in medicines, biochemistry and industrial science?

PROJECT

1. Draw figure of a tree showing different branches of chemistry.
2. Composting is a great way to recycle materials that might be thrown into landfill. It takes years to decompose them. Make an indoor composter and determine how readily different materials decompose.

UNIT 02

MATTER

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define matter as a substance having mass and occupying space.
- State the distinguishing macroscopic properties of commonly observed states of solids, liquids, and gases in particular density, compressibility and fluidity.
- Identify that state is a distinct form of matter (examples could include familiarity with plasma, intermediate states and exotic states e.g. BEC or liquid crystals).
- Explain the allotropic forms of solids (some examples may include diamond, graphite, and fullerenes).
- Explain the differences between elements, compounds, and mixtures.
- Identify solutions, colloids and suspensions as mixtures and give an example of each.
- Explain the effect of temperature on solubility and formation of unsaturated and saturated solutions

INTRODUCTION

The study of chemistry revolves around the study of matter which is all around us; not only is the entire world made up of matter but so are we, so are the objects that we use. From this we can derive the definition of matter:

Anything that has mass and occupies space is called matter. This makes air, water, rocks, and even people are examples of matter. Different types of matter can be described by their mass. Matter is itself composed of the atom. The atom is the building block of all matter and it is the various combinations of these atoms that make up all the matter that we see around us. You may ask yourself how the book you are reading and the water you are drinking are both matter. They neither look nor feel nothing alike. So how can they both fall into the definition of matter? From there we reach the conclusion that there are states of matter which differ from each other in the way that the atoms that make them up are arranged:

2.1 STATE OF MATTER

There are four states of matter

1. **Gas**
2. **Liquid**
3. **Solid**
4. **Plasma**

Each state is a distinct form of matter.

1. States of matter are the different forms in which matter can exist. These are solids, liquids, gases, and plasmas. These states are determined by the arrangement and movement of particles and the strength of intermolecular and atomic forces.
2. Energy can change matter into different states. For example, solids become liquids or gases when heated. At very high temperatures or when subjected to a strong electric field, the gas transforms into plasma. Under normal conduction, most substances remain in one distinct state: solid, liquid, or gas. Temperatures and energy levels on the Earth are not sufficient to ionize atoms and create plasma.
3. When heated, some crystalline solids turn into cloudy liquids that completely dissolve. This cloudy state is called liquid crystal. Liquid crystal states have many properties of liquids and some properties of solids. This form exists within a certain temperature range. When heated further, the state of the liquid crystal changes to a transparent liquid.
4. Furthermore, there are other states such as Bose-Einstien Condensates(BEC) which is defined as the state of matter in which separate atoms cooled to temperatures very close to absolute zero. BEC is observable under extreme conditions of cold temperature. Superfliud and superconductors are the two main materials which contain BEC.

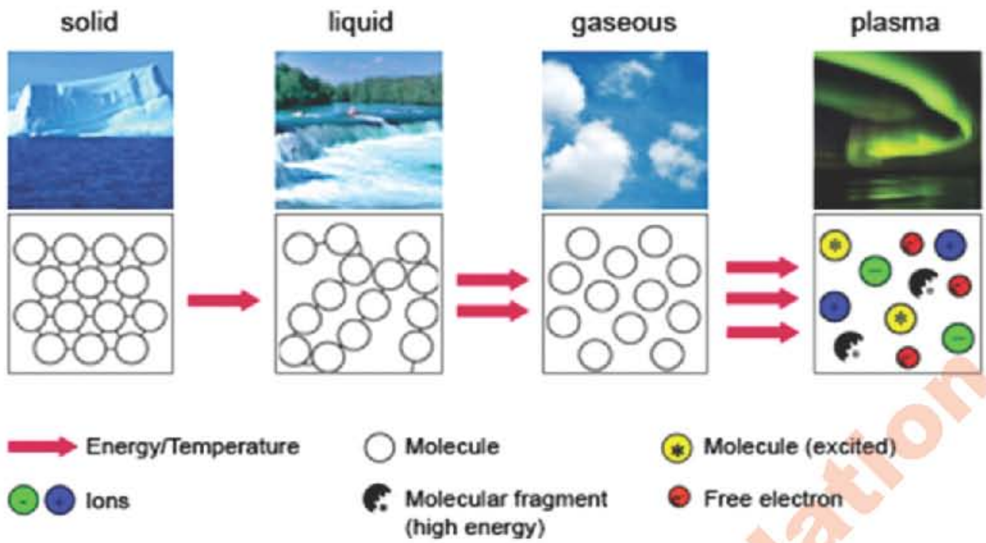


Figure 2.1: State of MAtter

Macroscopic properties that can be visualized by the naked eye and we can take measurements easily. Some common examples of macroscopic properties of matter include density, fluidity, compressibility.

Table 2.1: Properties of different states of matter

State of matter	Gas	Liquid	Solid
Density	Low density at normal condition due large spaces between molecules	High density at normal condition	High density at normal condition
Compressibility	Very compressible because of large empty spaces	moderately compressible	not compressible
Fluidity	Can flow	Can flow	Can not flow

Have you ever boiled water on a stove? What do you observe when the water heats up? Bubbles form and the water turns into the gas. This tells us a very important fact about the states of matter. Though the states of matter are distinct and are easily distinguishable from the other, through physical techniques we can convert one state of matter into the other. Physical techniques are techniques where we manipulate the physical aspects of matter such as the temperature or pressure. However, the chemical composition of matter stays the same.

2.2 ELEMENTS, COMPOUNDS AND MIXTURES

Matter can be described with both physical properties and chemical properties. Matter can be classified as

1 **Pure substance**
   a) Element
   b) Compound
2 **Mixture**
   a) Homogeneous
   b) Heterogeneous mixtures are
      i) Colloid
      ii) Suspension

Earlier, we talked about the atom and how atoms make up all of matter. Same types of atoms are called *elements*. An element consists of atoms that have the same atomic number also known as the proton number. This is the simplest form of matter which cannot be broken down through chemical means. While a physical change alters the physical properties of a substance, a chemical change forms a new substance completely.

***Element*:** the simplest form of matter made up the same type of atoms

So we have learnt that matter is made up of atoms and the atoms that have the same proton number are called elements. The combination of these different elements makes up the diversity of objects we see around us. When two or more elements chemically combine, meaning undergo a chemical reaction to form a new substance, this is called a compound. As this is a completely new substance, it is completely different from the elements that were used to make it.

Compound: A substance formed when two or more different atoms chemically combine.

Mixtures are the physical combinations of substances. A mixture that does not contain the same types of particles. If you were to examine the chemical composition of the particles in a mixture, the particles would be chemically different from each other. Tea is an example of a mixture. Tea is made up of milk, water, tea leaves and sugar all of which have different chemical compositions.

Mixture: It is a substance formed when two or more substances physically combine.

2.3 ALLOTROPES

The property of an element to exist in different physical forms is called allotropy. These different forms in the same physical state are called allotropes. Atoms of the same element arranged in different manners in the same physical state in allotropes. They are different structural forms of the same element. For example,

Diamond, graphite and buckyballs are three important allotropes of carbon.

Graphite:

Graphite is composed of flat two dimensional layers of hexagonally arranged carbon atoms. In a layer, each C-atom is covalently bonded to three other Carbon atoms. Weak intermolecular bonds exist between each layer which allows the layers to slide over one another without breaking the bonds. This arrangement makes graphite soft and slippery, making it ideal to be used as a lubricant. Graphite is a good conductor of electricity.

Diamond:

Diamond is the hardest and the purest crystalline allotrope of carbon. In its structure, each C-atom is covalently bonded to four other carbon atoms forming a rigid network of tetrahedral shape. This tetrahedral, three-dimensional arrangement makes it the hardest substance with a very high melting point. Since all the Carbon atoms are bonded with other carbon atoms, no free electrons are present resulting in the structure being non-conductive. Diamond is non-conductor of electricity.

Buckyballs (C-60):

Buckyballs, also known as fullerenes, have a football like fused hollow ring structure made up of twenty hexagons and twelve pentagons. Each of its 60 carbon atoms are bonded to 3 carbon atoms.

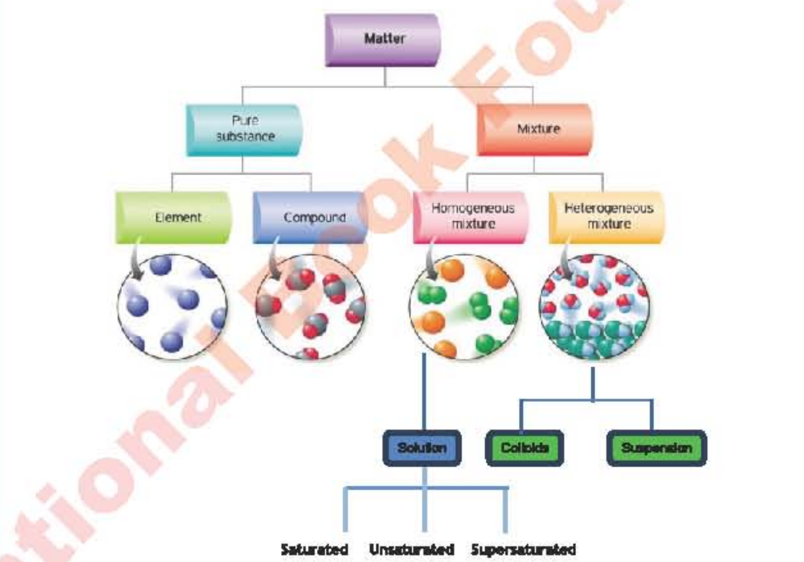


Fig 2.2: Classification of matter

2.4 Solution

A solution is a homogeneous mixture of two or more substances in which one substance is dissolved in the other. Homogeneous means that no particles or parts of different substances can be seen. When one substance dissolves, the solution looks exactly the same. **A substance that is dissolved is called a solute and a substance in which it is dissolved is called a solvent.** In solution, the particles are microscopic, less than 1 nm in diameter. A solution is a very stable mixture and the solute does not separate from the solvent itself.

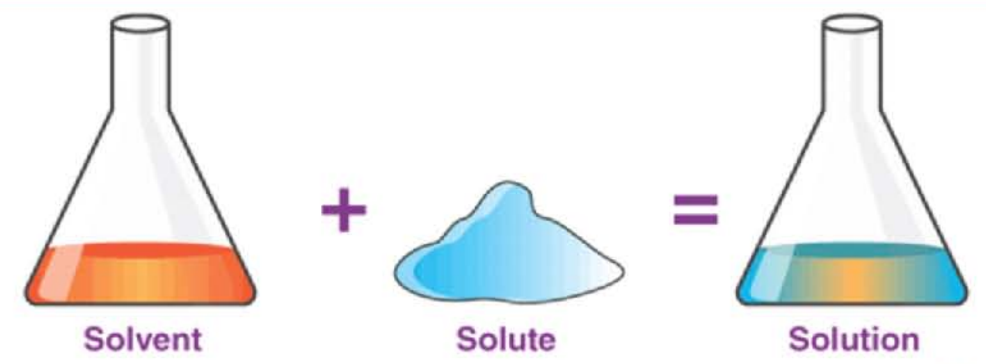


Fig 2.3: Formation of Solution

In salt solution, salt is the solute and water is solvent. More than one solute may be present in a solution. For example, in soft drinks, water is a solvent while other substances like sugar, salts and CO_2 are solutes. Consider the example of air where Nitrogen gas is solvent and Oxygen, carbon dioxide and trace gases are solute. On the basis of physical states of solvent and solute can be categorized as solid, liquid and gaseous solutions. Generally, solutions are found in three physical states depending upon the physical state of the solvent, e.g. air is a gaseous, sea water is a liquid solution and alloy is a solid solution in real life.

Gaseous Solutions

In Gaseous Solutions solvent is a gas and solute can be a gas or liquid or solid. For example a mixture of nitrogen and hydrogen used in Haber's process (ammonia formation) and other is mixture of ammonia and carbon dioxide used for urea preparation. Fog, clouds and mist are examples of solutions where liquid water (solute) is dissolved in air (solvent). Smoke is a solution of carbon particle in gaseous air in our daily life.

Liquid Solutions

Carbonated drinks are solutions where solvent is liquid water and solute is gaseous carbon dioxide. Rectified spirit produced by fermentation of sugar cane, Vinegar (acetic acid in water), are examples of solutions where liquid dissolved in liquid. Brine and sugar syrup are solutions of solid salt and sugar in water.

Solid Solutions

Hydrogen gas on the nickel metal surface is used in ghee industry where hydrogen gas is solute and nickel catalyst is solvent. Solution of any metal (solid) in liquid mercury is called amalgam. Alloy industry is very common these days. Alloys are formed by mixing different metal (Brass, Bronze, steel).

2.4.1 Aqueous Solutions

Aqueous solution is formed by dissolving a substance in water. The dissolved substances in an aqueous solution may be solids, gases, or other liquids. In order to be a true solution, a mixture must be stable. For example, sugar in water and table salt in water. Water is called a universal solvent because it dissolves majority of compounds present in earth's crust. Aqueous solutions are mostly used in the laboratories.

Depending on amount of solute solution can be classified as

Solution

Saturated Solution | **Unsaturated Solution** | **Supersaturated solution**

2.4.2 Saturated Solution

A solution containing maximum amount of solute at a given temperature is called saturated solution.

When a small amount of solute at given temperature is added in a solvent, solute dissolves very easily in the solvent. If the addition of solute is kept on, a stage is reached when solvent cannot dissolve any more solute. At this stage, further added solute remains undissolved and it settles down at the bottom of the container. On the particle level, a saturated solution is the one, in which undissolved solute is in equilibrium with dissolved solute. At this stage, dynamic equilibrium is established. Although dissolution and crystallization continue at a given temperature, but the net amount of dissolved solute remains constant.

2.4.3 Unsaturated Solution

A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution. Such solutions have the capacity to dissolve more solute to become a saturated solution.

2.4.5 Supersaturated Solution

When saturated solutions are heated, they develop further capacity to dissolve more solute. Such solutions contain greater amount of solute than is required to form a saturated solution and they become more concentrated. The solution that is more concentrated than a saturated solution is known as supersaturated solution. Supersaturated solutions are not stable. Therefore, an easy way to get a supersaturated solution is to prepare a saturated solution at high temperature. It is then cooled to a temperature where excess solute crystallizes out and leaves behind a saturated solution.

Activity 2.1

Take 100g water in a beaker. Add a tea spoon of sugar in it. Stirr it. The sugar will dissolve.Repeat the process and the and 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 and 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 2.4), crystallization will start. When crystallization has finished, you will have a saturated solution in the presence of sodium thiosulphate crystals.

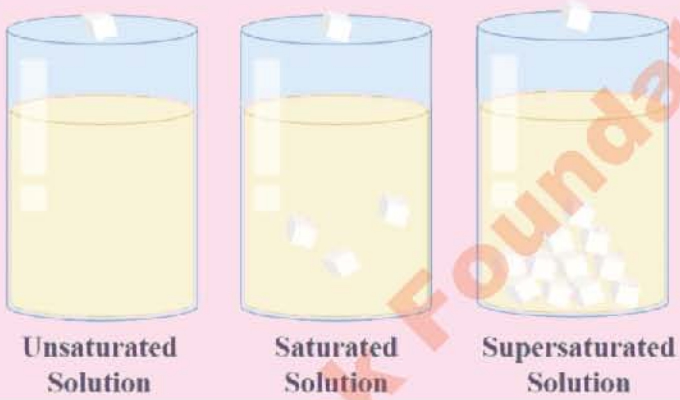


Fig 2.4: Different types of solutions

2.4.6 Concentrated and Dilute Solution

The solutions are classified as dilute or concentrated on the basis of relative amount of solute present in them. Dilute solutions are those which contain relatively small amount of dissolved solute in the solution.

Concentrated solutions are those which contain relatively large amount of dissolved solute in the solution. For example, brine is a concentrated solution of common salt in water. These terms describe the concentration of the solution. Addition of more solvent will dilute the solution and its concentration decreases.

2.4.7 Solubility

Solubility is the maximum amount of solute which dissolves in a specified amount of solvent at a specific temperature. The solubility of a substance depends on the solvent used, as well as temperature and pressure. See Table 2.2.

2.4.8 Effect of Temperature on Solubility

The solubility of solutes depends on temperature. Depending on the nature of solute there is either:

a) **Increase in solubility with temperature e.g., KCl, NH_4Cl**

b) **Decrease in solubility with temperature e.g., Na_2SO_4 $Ca(OH)_2$**

Table 2.2: Solubility of some salts g/100g of solvent at Different Temperatures		
(Solute)	Solubility (Amount of solute in 100g of solvent at 20°C	Solubility (Amount of solute in 100g of solvent at 100°C
NaCl	36.5g/100g H_2O	39.2g/100g H_2O
KCl	34.7g/100g H_2O	56g/100g H_2O
NH_4Cl	37.5g/100g H_2O	77g/100g H_2O
$Ca(OH)_2$	0.173g/100g H_2O	0.066g/100g H_2O

Example:

An example of a solute Whose decreases in solubility with increasing temperature is calcium hydroxide, which can be used to treat chemical burns and as an antacid.

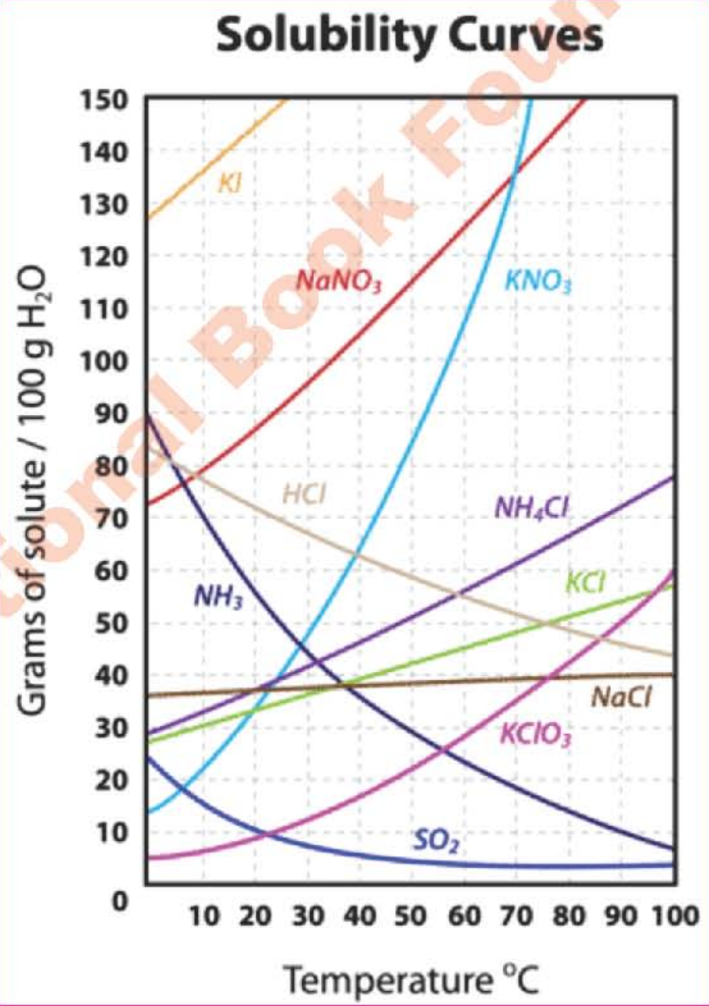


Fig.2.4 Effect of temperature on solubility

2.5 COLLOIDS & SUSPENSIONS

Colloid

These are heterogeneous mixtures in which the solute particles are larger than those present in the true solutions but not large enough to be seen by naked eye. A colloid is a mixture that has particles ranging between 1 and 1000 nanometers in diameter, yet are still able to remain evenly distributed throughout the solution. These are also known as colloidal dispersions because the substances remain dispersed and do not settle to the bottom of the container. The particles in such system dissolve and do not settle down for a long time. But particles of colloids are big enough to scatter the beam of light. It is called Tyndall effect. We can see the path of scattered light beam inside the colloidal solution. Tyndall effect is the main characteristic which distinguishes colloids from solutions. Hence, these solutions are called false solutions or colloidal solutions. Examples are starch, albumin, soap solutions, blood, milk, ink, jelly and toothpaste, etc.

Suspension

A suspension is defined as a heterogeneous mixture in which the solid particles are spread throughout the liquid without dissolving in it. It is mixture of undissolved particles in a given medium. Particles are big enough (greater than 1000nm) to be seen with naked eyes. Examples are chalk in water (milky suspension), paints and milk of magnesia (suspension of magnesium oxide in water). For better understanding of true solutions, false solution and suspension, a comparison of their characteristics is given in table 2.3

Table 2.3

S.No	Solution	Colloids	Suspension
1	A homogeneous mixture of two or more components	A heterogeneous mixture of two or more components	A heterogeneous mixture of two or more components
2	Particle size vary from 0.1-1nm. Not visible by naked eye	Particle size vary from 1-10^3 nm. Visible by naked eye by naked eye	Particle size greater than 10^3 nm. Visible by naked eye by naked eye
3	Particles can pass through normal as well as ultra-filter paper	Particles can pass through normal filter paper but not through ultra-filter paper	Particles cannot pass through normal as well as ultra-filter paper
4	Cannot Scatter the light (due to small size)	Can Scatter the light (Tyndal effect)	Can Scatter the light (Tyndal effect)
5	Does not separate	Does not separate	Separate or settles down when stationary
Examples	Sea water	Milk	Muddy water

KEY POINTS

- Anything that has mass and occupies space is called matter.
- Plasma is an electrically charged gas, which is affected by electrical and magnetic fields.
- The property of an element to exist in different physical forms is called allotropy.
- Element: the simplest form of matter made up the same type of atoms
- Compound: A substance formed when two or more different atoms chemically combine.
- A homogeneous mixture of two or more components is called solution.
- Aqueous solution is formed by dissolving a substance in water.
- A solution containing maximum amount of solute at a given temperature is called saturated solution.
- A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution
- A colloid is a mixture that has particles ranging between 1 and 1000 nanometers in diameter
- A suspension is defined as a heterogeneous mixture in which the solid particles are spread throughout the liquid without dissolving in it
- **References for additional information**
- Matter and its properties: Joseph Midthun, Paul Kobasa
- Cambridge IGSE™ Chemistry 5th Edition
- Cambridge International AS & A Level Chemistry (9701)

REVIEW QUESTIONS

1. **Encircle the correct answer.**

(i) Anything that has mass and occupies space is called.

(a) Liquid	(b) Gas
(c) solid	(d) Matter

(ii) Following are states of matter

(a) Gas	(b) Liquid
(c) Solid	(d) All of these

ii) Macroscopic properties are properties that can be visualized by

(a) the naked eye	(b) microscope
(c) electron microscope	(d) telescope

(iv) Matter can be described by both its

(a) physical properties and chemical properties.

(b) physical properties (c) chemical properties.

(v) A substance formed when two or more different combine chemically.

(a) atom (b) compound

(c) element (d) solution

2. **Give short answer.**

(i) Can you write the formula of the carbon dioxide gas that we exhale?

(ii) Define the element, Compound, Mixture

(iii) Differentiate between compound and mixture

(iv) Differentiate between concentrated and dilute solution

3. Define the term Allotropes Explain the allotropes of Carbon
4. What is difference between Homogeneous and heterogeneous solution?
5. Differentiate between the Colloids, Suspension
6. How can you identify solvent and solute?
7. If there are 18 protons in the Argon atom, then what is the atomic number of Argon?
8. Describe State of matter with example
9. Differentiate between the following.
   a. Colloids and Suspensions
   b. Elements and Compounds
   c. Concentrated and Dilute solutions
10. Examine the concept of solubility.

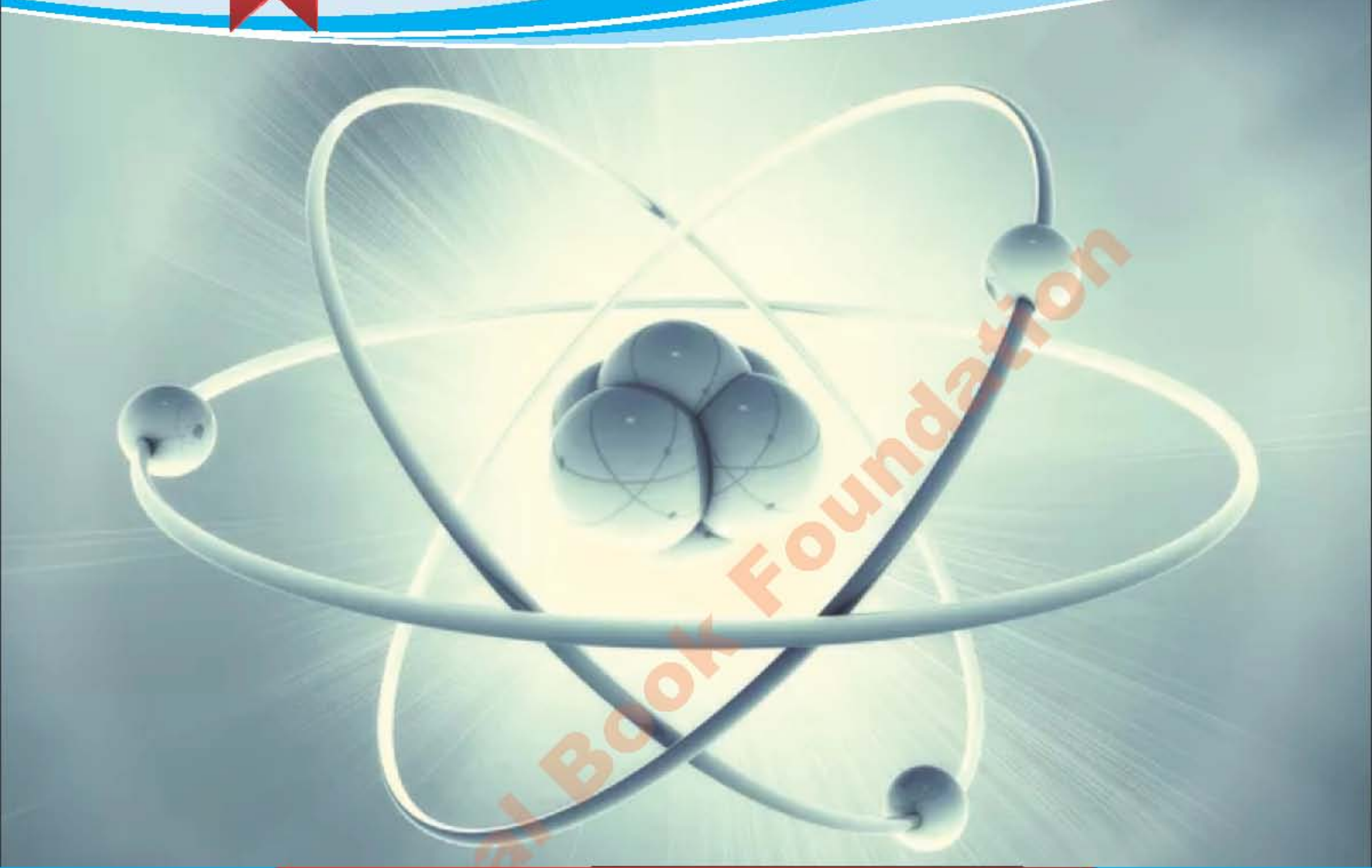
THINK TANK

11. **Why is a solution considered mixture?**
12. **How will you test weather given solution is a colloid or a solution?**

PROJECT

Create a poster that illustrates the various form of matter in the students everyday environment.

UNIT 03



ATOMIC STRUCTURE

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Explain the structure of the atom as the central nucleus containing neutrons and protons surrounded by electrons in shells.
- State that, orbits(shells) are energy levels of electrons and a larger shell implies higher energy and greater average distance from nucleus.
- State the electrons are quantum particles with probabilistic paths whose exact paths and location cannot be mapped (with reference to uncertainty principle)
- Explain that nucleus is made up of protons and neutrons held together by strong nuclear force.
- Explain that an atomic model is an aid to understand the structure of an atom.
- State the relative charge and relative masses of a subatomic particles (an electron, proton, and neutron).
- Interpret the relationship between a subatomic particle, their mass, and charge.

- Illustrate the path that positively and negatively charged particles would take under the influence of a uniform electric field.
- Define proton number/atomic number as the number of protons in the nucleus of an atom.
- Explain that the proton number is unique to each element and use to arrange elements in periodic table.
- State that radioactivity can change the proton number and alter an atom`s identity.
- Define nucleon number/atomic mass as sum of protons and neutrons in the nucleus of an atom.
- Define isotopes as different atoms of the same element that have same number of protons but different neutrons.
- State that isotopes can affect molecular mass but not chemical properties of an atom.
- Determine the number of protons and neutrons of different isotopes.
- Define relative atomic mass as the average mass of isotopes of an element compared to1/12th of the mass of carbon-12
- State that isotopes can exhibit radioactivity.
- Discuss the importance of isotopes using carbon dating and medical imaging as examples.
- Describe the formation of positive(cation) and negative(anion) ions from atoms.
- Interpret and use the symbols for atoms and ions.
- Calculate the relative atomic mass from relative masses and abundance of isotopes.
- Calculate the relative mass of an isotope given relative atomic mass and abundance of all stable isotopes.

INTRODUCTION

This chapter presents the historical development of atomic theory to the modern atomic model. One of the basic concepts of atomic structure is atomic number and mass number, which define an element and its isotopes. Understanding the structure of atoms is essential to understanding many scientific phenomena.

3.1 ATOMIC MODELS

The concept of the atomic model evolved over time as our understanding of atomic structure deepened through experimental observations and theoretical advances. Several important models of the atom had been proposed throughout history, each contributing to the understanding of atomic behaviour and properties. The most important atomic models are:

Dalton's model

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 laws of chemical combinations. His brilliant work became the main stimulus for the rapid progress of the chemistry during nineteenth century. However, series of experiments 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.

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

Rutherford's Experiment

Rutherford bombarded a very thin gold foil about 0.00004cm thickness with α-particles. (figure 3.1). 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.

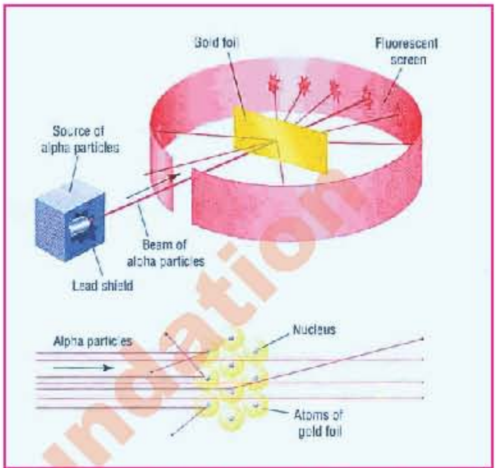


Figure 3.1 Rutherford's Experiment

Rutherford made 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 centrifugal force due to the revolution of electrons balances the electrostatic force of attraction between the nucleus and the electrons.

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.

Bohr's Atomic Theory

In1913 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 Plank'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.

$$\Delta E = E_2 - E_1$$

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

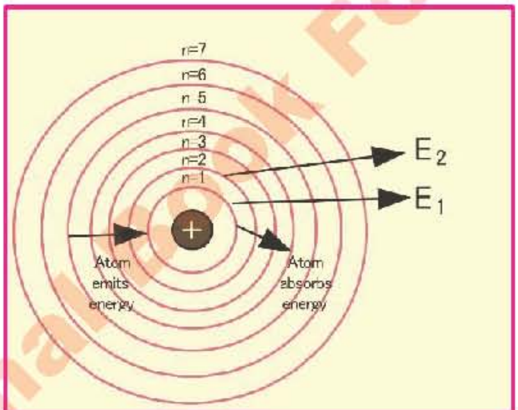


Figure 3.2 Bohr's model of the atom

Bohr model does not depict the three dimensional aspect of an atom.

Quantum Mechanical Model:

This is the current model used by modern science to describe the structure of the atom. It incorporates the principles of quantum mechanics and treats electrons as wave-particle entities. Instead of exact orbits, it defines probability regions, called orbitals, where electrons are likely to be found.

The Heisenberg Uncertainty Principle :

Heisenberg uncertainty principle is one of the fundamental concepts of quantum mechanics and is named after the German physicist Werner Heisenberg, who formulated it in 1927.

This principle states that it is impossible to simultaneously determine the exact location and future trajectory of an electron. As a result, plotting the electron orbit around the nucleus becomes an irresistible challenge.

Imagine that you have a single hydrogen atom and you decide to observe the position of that single electron at a given moment. Shortly after you repeated this process, the electron moved to another position. This means that from the original location to the next one is completely unknown to you. Continuous repetition of this process allows the gradual construction of a three-dimensional map representing the likely locations where the electron is expected to exist. You cannot know for sure where an electron is and where it goes next. This makes it impossible to draw the orbit of the electron around the nucleus.

In hydrogen, the electron has the potential to exist anywhere in the spherical region surrounding the nucleus. 95% (or whatever you want) of the time, the electron will be in a relatively simple region of space close to the nucleus, called an orbital. An orbital is the region of space where the electron lives.

Louis de Broglie, a French physicist, in 1924 proposed duel nature of electrons. He suggested that sub-atomic particles like electrons, can exhibit both particle-like and wave-like behaviour. His idea opened the door for new possibilities in understanding behaviour of sub-atomic particles. This concept made a significant contribution to the development of quantum mechanics.

In 1927, Davisson and Germer, experimently confirmed the de Broglie hypothesis that electron has wave like behaviour. This discovery laid the foundation for the Modern Quantum Mechanics.

Understanding Atomic Models

An atomic model is a tool for understanding the structure and behavior of atoms and their interactions in chemical reactions. Any atomic model helps us understand the structure of an atom. An atomic model is not a physical model, but represents a conceptual imagination. This helps to explain experimental observations of atomic behavior. The atomic model gives us a simplified representation of complex reality. As research and technology progress, scientists continue to improve our knowledge and atomic models.

A simple view of the structure of an atom

The nucleus of an atom is in the center. It contains protons and neutrons. Protons and neutrons are collectively called nucleons. The nucleus is surrounded by electrons in shells. Protons and neutrons are massive particles. The mass of electrons is so small. So, in practice, the mass of an atom is concentrated in the nucleus.

Nuclear Force

The nucleus contains protons and neutrons. Protons are positively charged and neutrons are neutral. The nucleus has no negative charge. The positively charged protons must cancel each other out and the nucleus must break apart. But atoms are stable and have existed for billions of years. Therefore, there must be some kind of attraction that connects them. No electrostatic or magnetic forces occur within the core. This is because these forces involve both attraction and repulsion. Therefore, the force that binds protons and neutrons together is a strong force. This force is called strong nuclear force. This is defined as the strong attractive force that binds protons and neutrons together. This force is stronger than electrostatic or magnetic forces. This force exists between neutrons and neutrons, protons and protons, and neutrons and protons.

3.2 SUBATOMIC PARTICLES

Subatomic particles are the fundamental particles that make up atoms. The three main subatomic particles are:

Proton:

- Relative charge: +1
- Relative mass: Approximately 1 atomic mass unit (amu) or 1.6726×10^{-27} kg

Neutron:

- Relative charge: 0 (neutral)
- Relative mass: Approximately 1 atomic mass unit (amu) or 1.6749×10^{-27} kg

Electron:

- Relative charge: -1
- Relative mass: Approximately 1/1836 amu or 9.11×10^{-31} kg

Protons and neutrons are found in the nucleus of an atom, whereas electrons orbit around the nucleus in energy levels or shells. They play crucial roles in determining the properties and behaviour of atoms and molecules. Neutrons and protons are held together in the nucleus by a strong nuclear force. This force exists between neutron-neutron, proton-proton, and neutron-proton.

Relationships between subatomic particles:

Protons and neutrons have roughly the same mass, around 1 amu. This mass contributes significantly to the total mass of the atom. Electrons have much less mass, so their contribution to the total mass of an atom is usually negligible.

The interaction between the negatively charged electrons and positively charged protons in the nucleus is what holds the atoms together.

The behavior of protons, neutrons and electrons in an electric field

What happens when a beam of these particles passes between two electrically charged plates?

figure

- Protons are positively charged and are deflected on a curved path toward the negative plate.
- Electrons are negatively charged and are deflected on a curved path toward the positive plate.
- Neutrons have no charge, go straight ahead.
- If the electrons and protons are traveling at the same speed, the electrons being lighter are deflected far more strongly than the heavier protons.

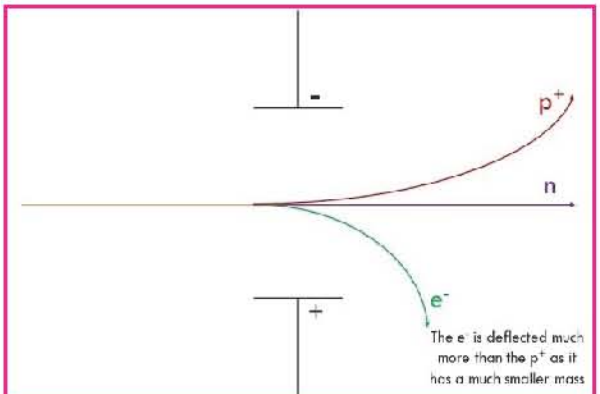


Figure 3.3: Path of positively and negatively charged particles through the uniform electric field.

Charge Neutrality

Atoms are electrically neutral because the number of protons (positively charged) in the nucleus is equal to the number of electrons (negatively charged) in the electron cloud. The charges balance each other so there is no net charge on the atoms.

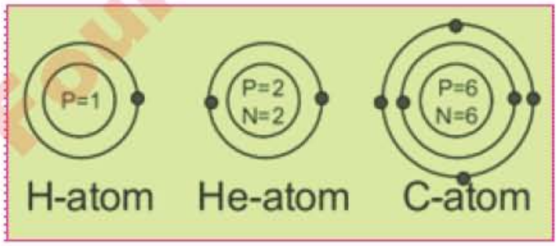
Radioisotopes

Different isotopes of the same element have the same number of protons in their atomic nuclei but differing numbers of neutrons. Some isotopes of an element are unstable and show radioactive decay. Radioactive isotopes of an element can be defined as atoms that contain an unstable combination of neutrons and protons, or excess energy in their nucleus. For example, hydrogen-3(protium), carbon-14, uranium-238 etc.

3.3 PROTON OR ATOMIC NUMBER

What determines the identity of an element?

Proton number refers to the number of protons in the nucleus of an atom. It is also known as the atomic number and is indicated by the symbol "Z".



Protons have a positive electrical charge. In neutral atoms, the number of protons is equal to the number of electrons. This balances the positive charge of the protons. This means that the proton number also indicates the number of electrons in the atom. For example, there is only one proton in the nucleus of a 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 proton number of C-atom?

Uniqueness of proton number

Each element has a unique proton number that distinguishes it from other elements. It determines the various properties of an element and its position in the periodic table. In the periodic table, elements are arranged based on their atomic or proton number. Therefore, the number of protons is related to the position of the element in the periodic table. Thus, the number of protons determines each particular element. This will tell you what element you are talking about.

For example, if an atom has a proton number of 6, it must be carbon. If an atom has 11 protons, it must be sodium. Similarly, each nitrogen atom has 7 protons, each oxygen atom has 8 protons, etc. You can identify each atom by the number of protons.

Nucleon number or Atomic mass

The total number of protons and neutrons in an atom is known as its mass number or nucleon number.

Some atoms of an element have different number of neutrons, such atoms are called isotopes.

No. of neutrons = mass number – atomic number

Example 3.1: 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

Radioactivity

The proton number determines the identity of the element. In stable elements, the nuclear force is balanced. In some elements, the nuclear forces are not naturally balanced. The nucleus of these atoms decays and becomes another atom. This process is called radioactive decay and this phenomenon is called radioactivity. This process continues until the forces in the nuclear core are balanced. In radioactive decay, when an atom emits a neutron, it changes to another isotope of that atom. But when it emits a proton, it becomes another atom. This means that radioactivity can change the number of protons in an atom and thus change the identity of the atom. For example;

1. Carbon-14 is a radioactive isotope of carbon. It is naturally present in the atmosphere. When any living organism takes in carbon dioxide from the air, it incorporates both C-14 and C-12 atoms into its tissues. The radioactive C-14 undergoes radioactive decay, transforms into nitrogen-14.
2. Uranium-238 is a radioactive isotope of uranium. It decays over time and finally transforms into stable lead-206 atom.

3.4 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.

$$\text{Mass of one C-12 atom} = 12 \text{ amu}$$

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

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

$$= \frac{8.40}{100} \times 12 \text{ amu}$$
$$= 1.008 \text{ amu}$$

Similarly, relative atomic masses of O, Na, Al are 15.9994 amu, 22.9898 amu, 26.9815 amu respectively. Table 3.1 shows the relative atomic masses of some elements.

Table 3.1: 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

3.5 ISOTOPES

Figure 3.1 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.

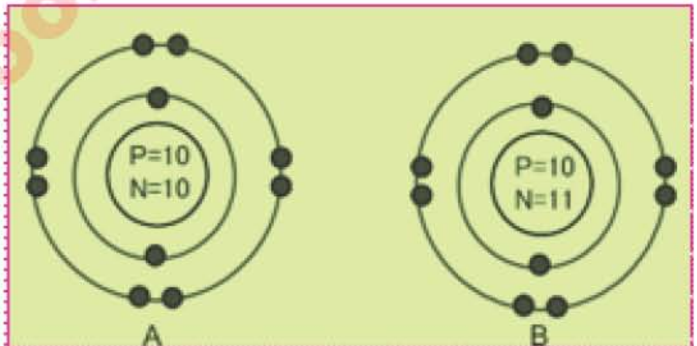


Figure 3.1

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?

3.5.1 Isotopes of Hydrogen

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

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 3.2 Shows some physical properties of ordinary water and heavy water.

Table 3.2 - Comparision of ordinary water and heavy water.

Property	Ordinary water	Heavy water
Melting Point	$0.00^{0}C$	$3.81^{0}C$
Bioling point	$100^{0}C$	$101.2^{0}C$
Density at $25^{0}C$	$0.99701\ g/cm^3$	$1.1044\ g/cm^3$

Isotopes affect molecular mass of a substance, can change physical properties but do not change chemical properties.

At what temperature would a sample of heavy water freeze?

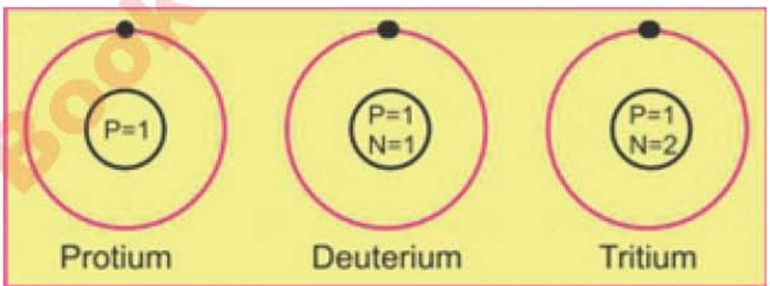


Figure 3.2 Isotopes of hydrogen

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 highly unstable.

3.5.2 Isotopes of Carbon

Carbon has three isotopes. Carbon-12, carbon-13 and carbon -14. Almost all the carbon is carbon-12. Its symbol is. $^{12}_{6}C$ It has six neutrons and six protons. Carbon-13 has symbol. $^{13}_{6}C$ It has seven neutrons and six protons. Carbon-14 has eight neutrons and six protons. Its symbol is $^{14}_{6}C$ 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 3.1

Carbon has three isotopes $^{12}_{6}C$, $^{13}_{6}C$, $^{14}_{6}C$ Figure 3.3 shows incomplete structure of isotopes of carbon. Can you complete it?

Natural abundance of isotopes of carbon is as follows

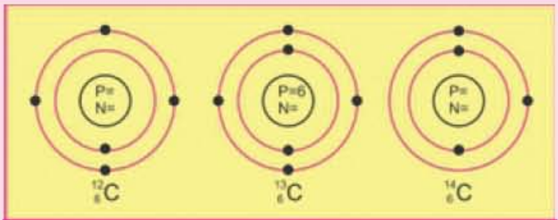


Figure 3.3 Isotopes of Carbon

$$^{12}_{6}C = 98.8\%, \quad ^{13}_{6}C = 1.1\%, \quad ^{14}_{6}C = 0.009\%$$

3.5.3 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 3.2

Chlorine has two isotopes. Figure 3.4 shows the structure of isotopes of chlorine. Can you write isotope symbol for each?

Isotope symbols: ________ ________

Natural abundance 75.77% 24.23%

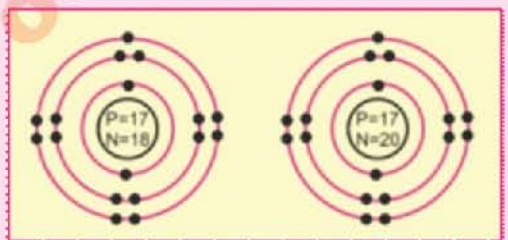


Figure 3.4 Isotopes of chlorine

3.5.4 Isotopes of Uranium

Activity 3.3

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

$$^{234}_{92}U, \quad ^{235}_{92}U, \quad ^{238}_{92}U$$

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

$$^{234}_{92}U = 0.006\%, \ ^{235}_{92}U = 0.72\%, \ ^{238}_{92}U = 99.27\%$$

Fill in the blanks?

$^{234}_{92}U$has ___ protons, ___ electrons and ___ neutrons

$^{235}_{92}U$has ___ protons, ___ electrons and ___ neutrons

$^{238}_{92}U$ 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.

$$^{238}_{92}U \longrightarrow ^{234}_{90}Th + ^{4}_{2}He$$

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.

3.5.5 Determination of Relative Atomic Mass

The relative atomic mass of an element can be calculated from the relative masses of its isotopes and their relative abundance.

Natural abundance of isotopes of carbon is as follows

$^{12}_{6}C = 98.8\%$, $^{13}_{6}C = 1.1\%$, $^{14}_{6}C = 0.009\%$

Calculate relative atomic mass of carbon.

Solution:

The relative atomic mass is a weighed average of the all the naturally occurring isotopes of an element, taking into consideration of their natural abundance. Use general formula

$$\text{Relative atomic mass of C} = \frac{\text{RA of C-12 x at. mass of C-12} + \text{RA of C-13 x at. mass of C-13} + \text{RA of C-14 x at. mass of C-14}}{100}$$

$$\text{Relative atomic mass of C} = \frac{98.8 \times 12 + 1.1 \times 13 + 0.009 \times 14}{100}$$

$$\text{Relative atomic mass of C} = \frac{1185.6 + 14.3 + 0.126}{100}$$

$$\text{Relative atomic mass of C} = 12.00026\ \text{amu}$$

CONCEPT ASSESSMENT EXERCISE 3.1

An element has two isotopes A and B.

The relative atomic mass of element is 35.5 amu. Relative abundance of isotope A is 75.77 % and its isotopic mass is 35. Find the isotopic mass of B if its relative abundance is 24.23 %.

3.5.6 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.

(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. 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.

(vi) Radioactive isotopes are also used to study the mechanism of chemical reactions.

(vii) Radioactive isotopes are used to date rocks, soils, archaeological objects, and mummies.

3.5.7 Carbon Dating

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.

3.6 CATIONS AND ANIONS:

Cations:

Cations are positively charged ions that form when an atom loses one or more electrons. Cations are usually formed from metal atoms that tend to lose electrons to achieve a stable electron configuration similar to a noble gas. When an atom loses one or more electorns, it forms a cation. The resulting cation has the electronic configuration of a noble gas. Neutral atoms have equal number of protons and electrons. When an atom loses one or more electrons, the number of protons becomes greater than electrons, as a result atom acquires positive charge.

Example 3.1: Describing the formation of cations

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

Problem Solving Strategy:

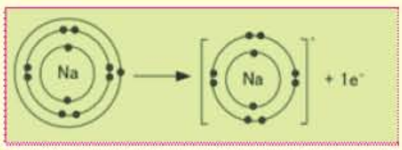
1. Sodium belongs to Group IA on the periodic table. It has only one electron in the valence shell. Sodium atom loses its valence electron and is left with an octet. Represent this by drawing the complete electronic configuration or using an electron dot structure.
2. Magnesium belongs to Group IIA in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

$$Na\ 1s^2 2s^2 2p^6 3s^1 \xrightarrow{-e^-} Na^+ 1s^2 2s^2 2p^6$$

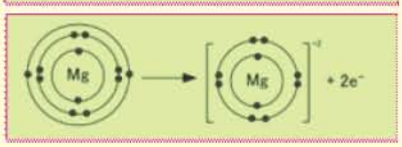
You can also represent this by following electron dot structure,



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

$$Mg\ 1s^2 2s^2 2p^6 3s^2 \xrightarrow{-2e^-} Mg^{+2}\quad 1s^2 2s^2 2p^6$$

You can also represent this by electron dot structure,



CONCEPT ASSESSMENT EXERCISE 3.2

Describe the formation of cations for the following metal atoms:

(a) Li (atomic no 3)

(b) Al (atomic no. 13)

Anions

Anions are negatively charged ions that form when an atom gains one or more electrons. This process usually occurs when an atom has a relatively high electron affinity, meaning that it can easily attract and capture more electrons to achieve a stable electron configuration similar to a noble gas. When an atom gains one or more electrons, the number of electrons becomes greater than protons, so it acquires negative charge.

Example 3.1: Describing the formation of anions.

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

(a) Oxygen(atomic no.8) (b) Fluorine (atomic no. 9)

Problem Solving Strategy:

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

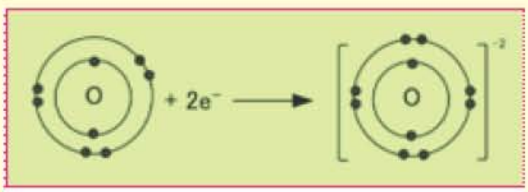
Solution:

(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.

$$O\ 1s^2 2s^2 2p^4 + 2e^- \longrightarrow O^{-2}\ \ 1s^2 \underbrace{2s^2 2p^6}_{\text{octet}}$$

You can also represent this by electron dot structure,

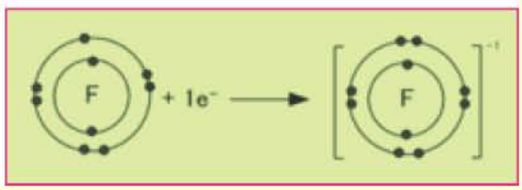


(b) Formation of anion by fluorine atom

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.

$$F\ 1s^2 2s^2 2p^5 + e^- \longrightarrow F^-\ \ 1s^2 \underbrace{2s^2 2p^6}_{\text{octet}}$$

You can also represent this by electron dot structure,



CONCEPT ASSESSMENT EXERCISE 3.3

Describe the formation of anions by the following non-metals.

(a) Sulphur (atomic No. 16)

(b) Chlorine(atomic No. 17)

3.7 ELECTRONIC CONFIGURATION

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

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.

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 as 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.

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

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.

Symbols for atoms and ions

The symbol for an atom represent the element . It consists of one or two-letters, the mass number as a left superscript, the atomic number as a left subscript, and the charge as a right superscript. For example;

$^{24}_{12}Mg^{2+}$

This number is often omitted. This diagram shows symbol for magnesium "Mg" which stands for magnesium. The number to the upper left of the symbol is the mass number, which is 24. The number to the upper right of the symbol is the charge which is positive 2. The number to the lower left of the symbol is the atomic number which is 12.

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. Isotopes are atoms of an element that differ in the number of neutrons.
- $^{235}_{92}U$ 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.
- There are four types sub-shell s, p, d, and f.
- 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, Introducion to Advanced Chemistry.
- Iain Brand and Richard Grime, Chemistry (11-14).

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Chlorine has two isotopes, both of which have

(a) same mass number. (b) same number of neutrons.

(c) different number of protons. (d) same number of electrons.

(ii) Number of neutrons in ${}^{27}_{13}M$ are

(a) 13 (b) 14

(c) 27 (d) 15

(iii) Which isotope is commonly used to irradiate cancer cells?

(a) Iodine-123 (b) Carbon-14

(c) Cobalt-60 (d) Iodine-131

(iv) M shell has sub-shells:

(a) 1s, 2s (b) 2s, 2p

(c) 3s, 3p, 3d (d) 1s, 2s, 3s

(v) A sub-shell that can accommodate 6 electrons is

(a) s (b) d

(c) p (d) f

(vi) ${}_{11}Na$ has electronic configuration:

(a) $1s^2,2s^2,3s^1$ (b) $1s^2,2s^2,2p^7$

© $1s^2,2s^2,2p^5,3s^2$ (d) $1s^2,2s^2,2p^6,3s^1$

(vii) 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

(viii) Which isotope is used in nuclear reactors?

(a) U-234 (b) U-238

(c) U-235 (d) All of these

2. Give short answer.

(i) Distinguish between shell and sub-shell

(ii) Why an atom is electrically neutral?

(iii) How many sub-shells are there in N shell.

(iv) Give notation for sub-shells of M shell.

(v) List the sub-shells of M Shell in order of increasing energy

(vi) 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.

4. $x = 1s^2, 2s^2, 2p^4, 3p^2$, $y = 1s^2, 2s^1, 2p^1$, $z = 1s^2, 2s^2, 2p^5, 3s^1$

5. Which orbital in each of the following pairs is lower in energy?

6. (a) 2s, 2p (b) 3p, 2p (c) 3s, 4s

7. Draw Bohr's Model for the following atoms indicating the location for electron, protons and neutrons:

8. (a) Potassium (Atomic No 19, Mass No. 39)

   (b) Silicon (Atomic No. 14 Mass No. 28)

   (c) Argon (Atomic No. 18 Mass No. 39)

9. Write electronic configuration for the following elements:

10. (a) $^{28}_{14}Si$ (b) $^{24}_{12}Mg$ © $^{27}_{13}Al$ (d) $^{40}_{18}Ar$

11. State the importance and uses of isotopes in various fields of life.

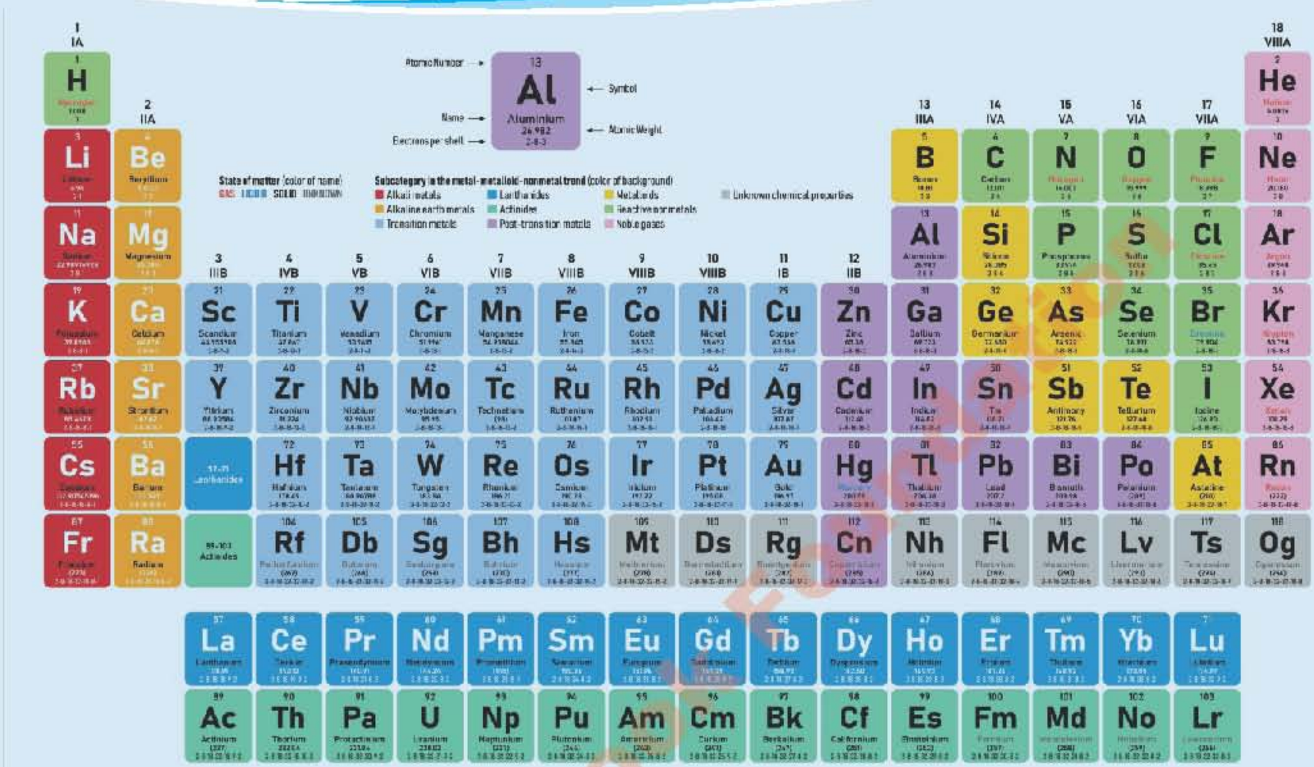
12. The atomic number of an element is 23 and its mass number is 56.

    a. How many protons and electrons does an atom of this element have?

    b. How many neutrons does this atom have?

13. The atomic symbol of aluminium is written as $^{27}_{13}Al$. What information do you get from it?

UNIT 04



PERIODIC TABLE AND PERIODICITY OF PROPERTIES

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define the periodic table as an arrangement element in periods and group, in order of increasing proton number/atomic number.
- Identify the group or period or block of an element using its electronic configuration (only the idea of subshells related to the blocks can be introduced).
- Explain the relationship between group number and the charge of ions formed from elements in the group in terms of their outermost shells.
- Explain similarities in the chemical properties of elements in the same group in terms of their electronic configuration.
- Identify trends in groups and periods, given information about the elements, including trends for atomic radius, electron affinity, electronegativity, ionization energy, metallic character, reactivity, and density.
- Use terms like alkali metals, alkaline earth metals, halogens, noble gases, transition metals, lanthanides, and actinides in reference to the periodic table.

- Predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity.
- Deduce the nature, possible position in the Periodic Table and the identity of unknown elements from given information about their physical and chemical properties.
- Define Group 1 Alkali metals as relatively soft metals with general trends down the group limited to decreasing melting point, increasing density and increasing reactivity.
- Predict properties of other elements in Group 1, given information about the elements.
- Predict properties of elements in Group 1 in order of reactivity given relevant information.
- Define Group VII halogens as diatomic non-metals with general trends limited to increasing density and decreasing reactivity.
- Identify the appearance of halogens at rtp as fluorine as pale yellow gas, chlorine as yellow-green gas, bromine as red-brown liquid, iodine as grey-black solid.
- Explain the displacement reactions of halogens with other halide ions and also as reducing agents.
- Predict the properties of elements in group VII, given information about the elements.
- Analyze the relative thermal stabilities of the hydrogen halides and explain these in terms of bond strengths.
- Describe the transition elements as metals that: have high densities, high melting points, variable oxidation numbers, form coloured compounds and act as catalysts for industrial purposes. (some example include catalysts being used are the Haber process, catalytic converters, Contact process and manufacturing of margarine).
- Define the Group 18 noble gases as un-reactive mono-atomic gases.
- Explain this interms of electronic configuration.
- Compare the general physical properties of metals and non-metals. (specifically in terms of
  - a. Thermal conductivity
  - b. Electrical conductivity
  - c. Malleability and ductility
  - d. Melting points and boiling points

INTRODUCTION

Welcome to the exciting world of chemistry, where the elements come to life thanks to the remarkable periodic table. From its humble beginnings, where only 23 elements were known until the end of the 18th century, to its development of 118 elements today. It is very difficult and impossible to remember information about the reactions, properties, and atomic masses of elements. So we obviously need a way to organize our information 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 example, if you know the physical and chemical properties of one element in a group, you can predict the physical and chemical properties of any other element in the same group. The periodic table allows you to relate the reactivity tendencies of elements to their atomic structure. You can also predict which elements can form ionic or covalent bonds.

4.1 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.

4.1.1 Periods and Groups of Elements.

The most commonly used form of the periodic table is shown in figure 4.1. 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.

International Union of Pure and Applied Chemistry (IUPAC) has recently renamed newly discovered elements and placed them on the periodic table.

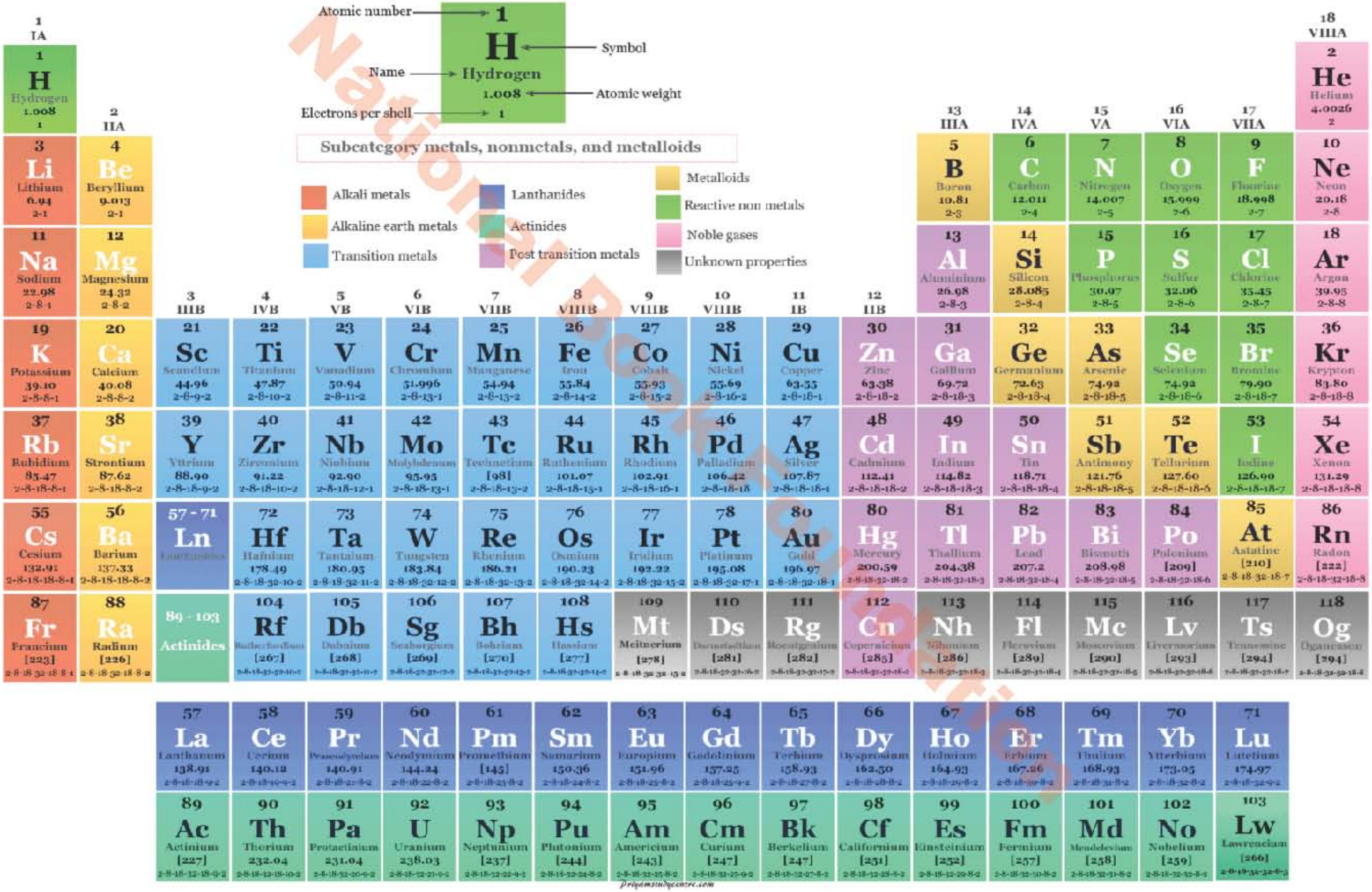
Activity 4.1

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

Table Number of elements in the periods of the periodic table

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

Figure 4.1: Periodic Table of Elements



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.**

Two numbering systems are often used to designate groups. You should know both. In the traditional system and the old IUPAC, the letters A and B are used. The first two groups are IA and IIA, while the last six groups are IIIA to VIIIA and the middle groups are in group B. The International Union of Pure and Applied Chemistry (IUPAC) decided that the groups would be 1-18 from left to right.

The elements in the same group have same number of valence electrons. Group number indicates the number of valence electrons in an element. For example, Group1 and Group 2 elements have 1 and 2 valence electrons respectively. In Groups 13 elements have 3, Group 14 have 4, Group15 have 5 valence electrons and so on. It is important to note that in Groups 13 to 18 (p block elements), the number of valence electrons is equal to group number minus 10.

Group A elements are called normal or representative elements. They are also called main group elements. Group B elements are called transition elements.

Names of Some Groups in the Periodic Table

Some groups of elements in the periodic table have been given group names. For example metallic elements in Group 1 are called alkali metals. Group 2 elements are called alkaline earth metals. The elements in Group 17 or VIIA are called halogens. The Group 18 or VIIIA elements are called noble gases because they do not readily undergo chemical reactions.

Recall that all elements have a unique identificcation nunber known as the atomic number or proton number. The atomic number of an element represents the number of electrons or protons present in the atom of the element. Aufbau's Principle helps in determining the order in which the electron orbitals get filled.

Electronic Configuration

According to Aufbau's principle, the order in which the orbitals fill up is as follows:

1s,2s,2p,3s,3p,4s,3d,4p,5s,4d,5p,6s,4f,5d,6p,7s,5f,6d,7p and so on.

Each orbital has a fixed capacity for the maximum number of electrons accommodated, s-orbitals have the capacity of 2 electrons, while p orbitals have the capacity for 6 electrons, d orbitals have the capacity for 10 electrons and f orbitals have the capacity for 14 electrons.

Using these concepts, we can determine the electronic configuration of the given element.

Block of an element: When you have filled all the electrons, the orbital in which the last electron is in, represents the block in which the element is placed.

Period of an element: Now, to determine the period in which the element is placed, you need to look at the principal quantum number(n) of the valence electron. This number repersents period number of element

Group of an element: To determine the group, we need to understand some rules:

(a) If the element is in s block, then the group number is equal to the number of valence electrons.

(b) If the element is in the p block, then the number of the group can be determined by the formula: (number of valence electrons + 10).

For example, the atomic number of sodium is 11.

Hence its electronic configuration is: $1s^2,2s^2,2p^6,3s^1$

Since the valence electron is in the 3s subshell, sodium belongs to belongs to the s block.

The principal quantum number of the valence electron of Na is 3. Hence, it belongs to the 3rd period.

Since Na belongs to the s block, its group number is equal to a number of electrons in valence subshell s. This is equal to 1

Hence, sodium belongs to the Group 1.

Note:

we can start filling the orbitals in the order mentioned by the Aufbau principle.

1s,2s,2p,3s,3p,4s

Example 4.1: Identifying the group and period of an element

Identify the group, period, and block of following elements on the basis of electronic configuration.

1. Al (atomic number= 13)
2. K(atomic number = 19)

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 if element belongs to s block. If it belongs to p block, then group number is equal to the total number of valence + 10.

Solution:

1. Electromnic configuration of Al (atomic no. 13) = $1s^2,2s^2,2p^6,3s^2,3p^1$

   Valence sub- shells is 3p, so Al belongs to p block

   As n = 3, Al is present in the 3^{rd} period.

   Total number of electrons in the valence shell = 2+1= 3

   Group number of Al= total number of electrons in the valence sub-shells + 10

   = 3+ 10

   = 13

   Hence Al belongs to Group 13

2. Electronic configuration of K (atomic no. 19) = $1s^2,2s^2,2p^6,3s^2,3p^6,4s^1$

   Valence shells is 4s. hence K belongs to s block

   As n = 4, K is present in the 4^{th} period.

   Total number of electrons in the valence shell = 1

   Group number of K= total number of electrons in the valence sub-shells

   = 1

   Hence K belongs to Group 1

CONCEPT ASSESSMENT EXERCISE 4.2

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

(a) $^{28}_{14}Si$, (b) $^{32}_{16}S$, (c) $^{19}_{9}F$, (d) $^{40}_{18}Ar$

Example 4.2: 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:

	II A	V A	VII A	
2	A $2s^2$	B $2s^2 2p^3$	C $2s^2 2p^5$	Period
3	D $3s^2$	F $3s^2 3p^3$	E $3s^2 3p^5$	Period

CONCEPT ASSESSMENT EXERCISE 3.1

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

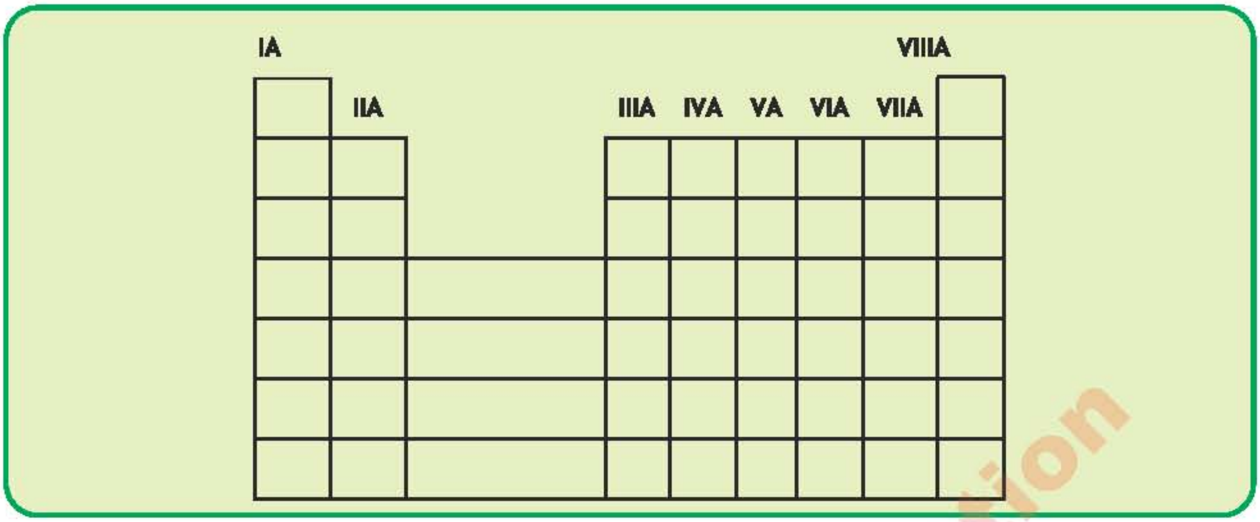
$P = 1s^2, 2s^2, 2p^2$ $\quad$ $Q = 1s^2, 2s^2, 3p^1$

$R = 1s^2$ $\quad$ $S = 1s^2, 2s^2$

$T = 1s^2, 2s^1$ $\quad$ $W = 1s^2, 2s^2, 2p^6$

$X = 1s^2, 2s^2, 2p^6, 3s^2, 3p^2$ $\quad$ $Y = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

$Z = 1s^2, 2s^2, 2p^1$



4.1.2 s and p Blocks in the Periodic Table

Group 1 and Group 2 elements contain their valence electrons in the s sub-shell. Therefore, these elements are called s-block elements. Elements in groups 13 to 18 (except He) are known as p-block elements because their valence electrons are located in the p sub-shell. Lanthanides and actinides are known as f-block elements since their valence electrons lie in f sub-shell. Figure 4.2 shows the blocks of the periodic table.

Li = 1^{s2},$2s^1$, as valence electron is in s sub-shell, Li belongs to s-block.

C=$1s^2$,$2s^2$,$2p^2$, as valence electron is in sub-shell p, C belongs to p- block.

4.2 GROUP NUMBER AND CHARGE ON AN ION

The group number of an element in the periodic table can provide information about the charge of an ion formed by an element. Valence electrons are involved in the formation of ions. The relationship between group number and ions formed by elements is based on the number of valence electrons in the element.

The group number of an s-block element in the periodic table corresponds to its number of valence electrons.

Whereas in the case of p-block elements, the number of valence electrons is equal to Group number minus 10.

Some elements tend to lose electrons. Why? Elements tend to achieve a stable electron configuration such as the noble gases. Remember that the 2 or 8 electron configuration is the most stable configuration. Elements with 1-3 electrons in their valence shell tend to lose those electrons and form +1, +2, +3 ions respectively. Elements with 5-7 electrons in their valence shell tend to gain 3, 2, 1 electrons respectively and form negatively charged ions with -3, -2, -1 charges respectively. Elements with 4 valence electrons can lose 4 electrons to form +4 ions. They can also gain 4 electrons and form -4 ions.

Group 1 (alkali metals): Group 1 elements such as lithium (Li), sodium (Na), and potassium (K) have one valence electron and belong to s block. S block elements lose electrons equal to their group number.They tend to lose this electron to form a +1 ion, also known as a mono-valent cation. For example: Lithium (Li) loses one valence electron to form Li^+ Sodium (Na) loses one valence electron to form Na^+. Potassium (K) loses one valence electron to form K^+. These elements after losing an electron acquire 8 electron configuration of a noble gas.

$Na = 1s^2,2s^2,2p^6,3s^1$

$Na^+ = 1s^2,2s^2,2p^6$

Group 2 (alkaline earth metals): Group 2 **elements such as beryllium** (Be), magnesium (Mg), and calcium **(Ca)** have two valence electrons and are s block element. They tend to lose these two electrons to form + 2 ions, also **called divalent cations**. For example: Beryllium (Be) loses two valence electrons to form **Be^{2+}** Magnesium (Mg) loses two valence electrons to form **Mg^{2+}** Calcium (Ca) loses two valence electrons to form **Ca^{2+}**

$Mg = 1s^2,2s^2,2p^6,3s^2$

How many electrons Mg can lose to achieve stable electron configuration?

Magnesium will lose 2 electrons to achieve stable configuration and this no. is same as it group number i.e., 2

$Mg^{2+} = 1s^2,2s^2,2p^6$

Some elements tend to gain electrons to achieve noble gas configuration. For example, Group 17 (Halogens): Group **17 elements such as** fluorine (F), chlorine (Cl), and bromine **(Br)** have seven valence electrons. They tend to gain one electron to **reach** a stable octet and form **- 1 ion**, also **called a** monovalent **anion**. For example: Fluorine (F) gains one electron to form F^-. Chlorine (Cl) gains one electron to form Cl^-. Bromine (Br) gains one electron to form Br^-.

$F = 1s^2,2s^2,2p^5$

$F^{1-} = 1s^2,2s^2,2p^6$

Similarly, $Cl = 1s^2,2s^2,2p^6,3s^2,3p^5$

$Cl^{1-} = 1s^2,2s^2,2p^6,3s^2,3p^6$

Group 16 (chalcogens): Group 16 **elements such as** oxygen (O), sulfur (S), and selenium **(Se)** have six valence electrons. They tend to gain two electrons **to reach** a stable octet and form **a -2 ion,** also **called a** divalent **anion**. For example: Oxygen (O) gains two electrons to form O^{2-} Sulfur (S) gains two electrons to form S^{2-}.

Group 18 (precious gases): Group 18 **elements such as** helium (He), neon (Ne), and argon **(Ar)** have full valence electron shells (except helium, which has only two valence electrons). They are chemically stable and do not form ions under normal conditions. Noble gases are known for their low reactivity due to their stable electronic configuration .

Example 4.3: Obtaining the position of element in the periodic table from the electronic configuration

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

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

Problem Solving Strategy:

Write electronic configuration of the element. Identify the valence shell configuration, co-efficient 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

Number of valence electrons = 2 + 3 = 5

N belongs to p-block

So, Group number = 5 + 10 = 15

$\therefore$ Nitrogen is present in the 2nd period of Group 15

b) Electronic configuration of oxygen = $1s^2, 2s^2, 2p^4$

Valence shell has configuration = $2s^2, 2p^4$

So, Period number = 2

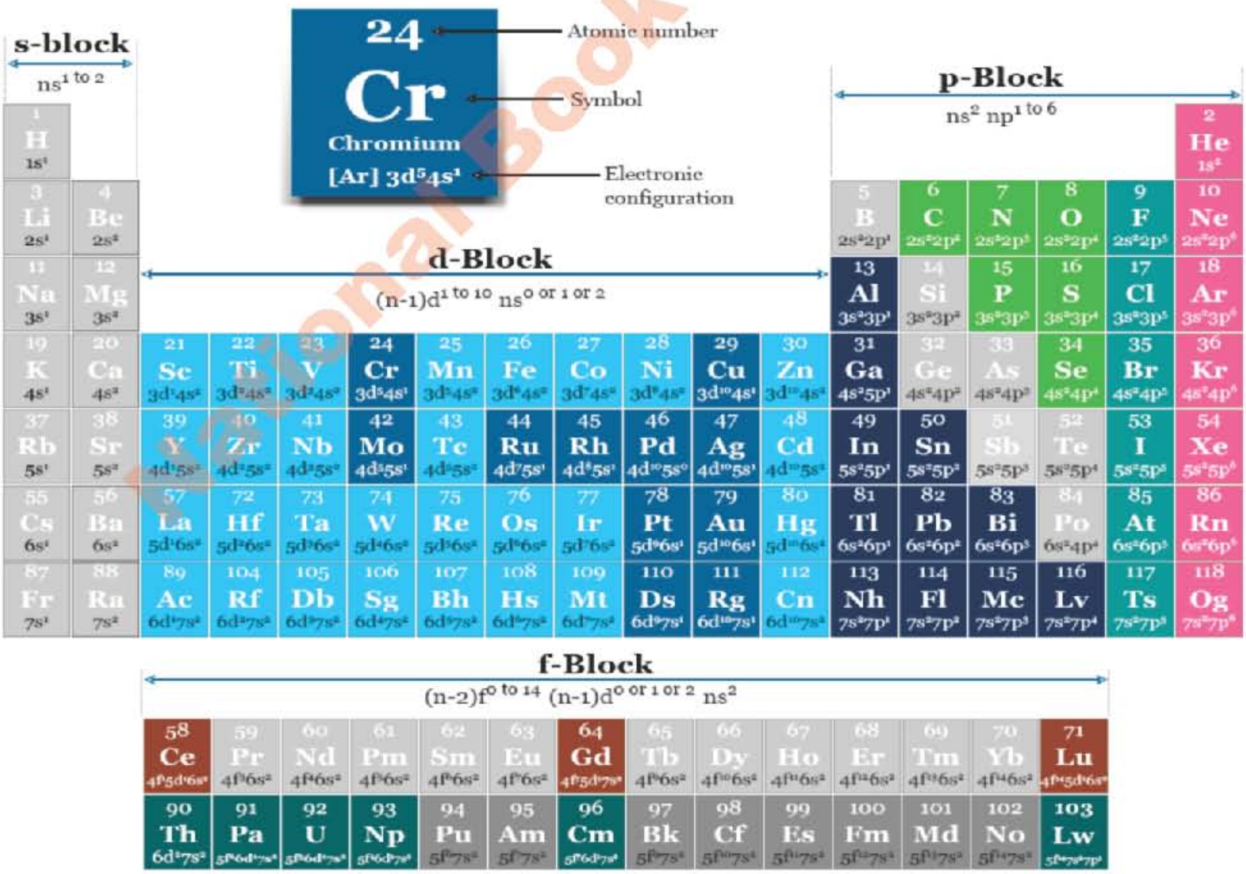
Number of valence electrons = 2 + 4 = 6

O belongs to p-block

So, Group number = 6+10 = 16

Oxygen is present in the 2nd period of Group 16

Figure 4.2: Periodic Table of Elements



CONCEPT ASSESSMENT EXERCISE 4.4

Obtain the valence shell configuration of Al and S from their position in the periodic table

4.3 PERIODICITY OF PROPERTIES

There is a periodic fluctuation in the electronic configuration of the elements as the atomic number increases. Therefore, the physical and chemical properties of the elements very in a periodic manner. Elements with a similar valence shell electronic configuration are placed in the same group, one below the other. Chemical properties depend on the electronic configuration of the valence shell. Because all the elements in a given group have similar valence shell electronic configurations, they have similar chemical properties. Physical properties depend on the size of atoms. Because the sizes of atoms change gradually from top to bottom in a group. Therefore, the elements show a gradation of physical properties within the same group. In the period of the periodic table, the number of electrons in the valence shell increases gradually from left to right. Their chemical and physical properties also differ in the same way. In this section, you will learn about the variation of physical properties of certain elements within a group and across a period.

4.3.1 Shielding Effect

Figure 4.3 shows electronic configuration of Li, Be and Mg.

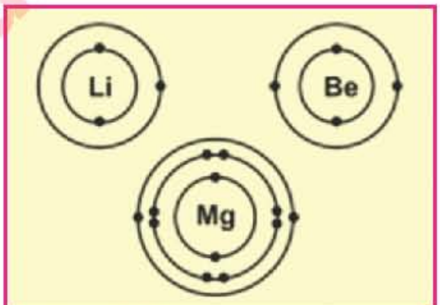


Fig 4.3 Electronic structure of Li, Be and Mg

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.

Example 4.4: 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

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.

CONCEPT ASSESSMENT EXERCISE 4.5

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 elements depend on the electronic configuration of their atoms. We now consider some properties of atoms that are affected by electronic configuration: atomic size, ionization energy, electron affinity and electronegativity. They usually increase and decrease repeatedly throughout the periodic table. That is, they show consistent changes or trends within a group or a period. These tendencies are correlated with behavior.

4.3.2 Atomic Size

The size of an atom depends on its electronic configuration. **Atomic size is the average distance between the atomic nucleus and the electronic outer shell.** Figure 4.4 shows the atomic radii of the main group elements. Figure 4.4 shows the variation of atomic radii within a period and within a group. You can see two general trends in atomic radii.

1) The atomic radius decreases in each period as you move across the period. This is because as you move from one element in the sequence to the next, to the right of it. Another electron is added to the same valence shell. At the same time, the positive charge of the core also increases by one. The attraction of the nucleus to the electron in the valence shell increases. Therefore, the size of the shell

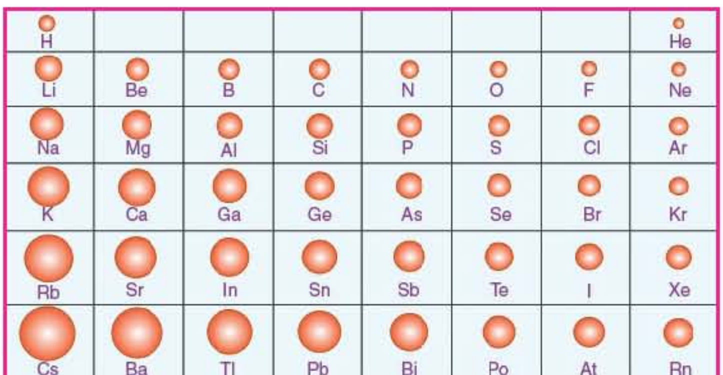


Figure 4.4: Atomic sizes of the main group elements

and the radius of the atom decreases. For example, going from lithium to beryllium, the atomic size decreases. You can understand this from the electronic configuration of the valence shell of Li ($2s^1$) and B ($2s^2$). Moving from Li to Be, the number of shells does not change, but the atomic number increases from 3 to 4. Therefore, the strength of the nucleus on the valence shell electron increases. Therefore, the atomic radius decreases.

2) Atomic radius increases in each main group as you move down the element group. This is because the size of an atom is determined by the size of its valence shell. As you move down the group to the next lower element, the atom has an additional shell of electrons. This increases the radius of the atom. For example, going from Li to Na, the atomic radius increases. Consider the electron configuration of Li ($1s^2\ 2s^1$) and Na ($1s^2$, $2s^2$, $2p^6$, $3s^1$). A new electron shell is added, increasing the size of the atom.

Example 4.5: 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:

(a) Period lies further to the left in the periodic table.
(b) Group lies closer to the bottom in the periodic table.
(c) Check the periodic table and choose the element.

Solution:

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

CONCEPT ASSESSMENT EXERCISE 4.6

Using the periodic table but without looking at the figure 4.4, 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

4.3.3 Ionization Energy

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".**

$$M(g) + \text{ionization energy} \longrightarrow M^{+}(g) + \acute{e}$$

Period	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 4.5 Ionization energies of the main group elements

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 4.5 shows the ionization energies of the main group elements. Values are given in units of kJ/mole^{-1}or kJ/mole.

Trends in ionization energy values.

The value of the ionization energy decreases from top to bottom in the group. This is because the shielding effect of the atoms increases down the group. Greater shielding effects result in a weaker attraction of the valence electrons to the nucleus. So they are easier to remove. This leads to a decrease in ionization energy from top to bottom in the group. Which atom has a greater shielding effect, Li or Na? As you move from left to right in the period, the shielding effect remains unchanged. But little by little the nuclear charge increases. The stronger attraction between the nucleus and the valence electron increases. As a result, the ionization energy increases from left to right in a period. Which atom has the higher ionization energy, Li or Be?

Example 4.6: 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.

CONCEPT ASSESSMENT EXERCISE 4.7

Which atom has the smaller ionization energy?

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

4.3.4 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 uni negative gaseous ion.**

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

Figure 4.6 shows electron affinities of main group elements.

Factors affecting electron affinity are nuclear charge, atomic radius and shielding effect.

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

Electron affinity decreases from top to bottom in a group. This is due to an increase in the shielding effect. Due to the increased shielding effect and increase in atomic radius, the 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 4.6 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 higher grade.

4.3.5 Electronegativity

Electronegativity is the ability of an atom to attract electrons toward itself in a chemical bond. Figure 4.7 shows as a scale of electronegativities of the elements devised by Linus Pauling. The American chemist Linus Pauling devised a method for calculating the 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 4.7 the electronegativities of elements.

Activity 4.2

Determining the general trends in the electronegativities

You will need:

- Figure 4.7

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 electronegativites across a period and write reason.
4. Move from top to bottom in Groups IA and IIA and note down the variation in electronegativites value.

5. Move from top to bottom in Groups VIA and VIIA and note down the variation in electronegativities value.

Make generalization about the trend in electronegativity values in a group. Give reason.

4.4 CHARACTERISTIC PROPERTIES

Characteristic properties of an element in a given group are based on periodicity and chemical reactivity. For example, in Group 1 (alkali metals) such as lithium, sodium, potassium, are highly reactive metals. They have general electron configuration ns^1. Their reactivity trend increases as you move down the group. Lithium, being at the top of the group is the least reactive metal among alkali metals. As you move down the group, the atomic size increases and the outer most electron is further from the nucleus, leading to be lost easily. This leads to increased reactivity. So, sodium is more reactive than lithium. Which is more reactive sodium or potassium ?

Similar trend is observed in Group 2 (alkaline earth metals). Which is more reactive Mg or Ca ?

4.4.1 Metallic Character

Metallic nature refers to a property of elements in the periodic table that determines how easily they can lose electrons and form positive ions (cations). Elements with metallic character have a strong tendency to lose electrons and easily form cations. The metallic character of an element is affected by its position in the periodic table.

Metallic character increases as you move down a group in the periodic table. This is primarily due to addition of new electronic shells. The outermost electrons are farther from the nucleus and experiences weaker attractive forces, making it easier for them to be lost. This promotes metallic character.

Metallic character decreases as you move across a period from left to right in the periodic table. This is because effective nuclear charge increases across a period, while the number of shells remains the same. The stronger attractive forces make it more difficult for valence electrons to be lost.

Example 4.6: Identifying the element that has higher metallic character.

Choose the element you expect to have higher metallic character in each of the following pairs.

(a) Na or K

(b) Na or Mg

Problem Solving Strategy

Remember that metallic character :

(a) Increases down the group. The element that has higher metallic character will be closer to the bottom.

(b) Decreases across a period. The element that has higher metallic character is further to the left.

(c) Check the periodic table to choose the element.

Solution:

(a) K

(b) Mg

CONCEPT ASSESSMENT EXERCISE 4.8

Which element has lower metallic character?

(a) Li or K

(b) Mg or Ca

(c) Compare and contrast ionization energy and electron affinity

4.4.2 Reactivity

The capability of an element to react with other elements to form new compounds is called it reactivity. Reactivity of elements generally increases as you move down a group. This is due to the increase in atomic size. The outermost electrons are farther from the nucleus and experience weaker attractive forces, making it easier for them to participate in chemical reaction.

Reactivity tends to very across a period. Elements on the left side of a period (Group 1 and 2) are highly reactive due to their strong tendency to lose electrons and form positive ions. Elements on the right side of a period (Group 16 and 17) are highly reactive as well but tend to gain electrons to form negative ions.

4.4.3 Density

Density of elements generally increases as you move down a group. This is due to the increasing atomic mass and the larger size of atoms. As the number of protons and neutrons in the nucleus increases, the atomic mass increases. This results in higher density.

Density can wary across the period. In general, density tends to increase from left to right until it reaches a maximum around the middle of the period, and then it starts to decrease.

4.4.4 Characteristic Properties of Alkali Metals

Some characteristic properties of alkali metals are as follows:

(a) Highly reactive metals:

Alkali metals are highly reactive metals in the periodic table. They readily lose valence electron to form a +1 cations. This trend increases down the group.

(b) Softness and low density:

Alkali metals have low densities and are relatively soft, which allows them to be easily cut with a knife. This trend increases down the group the group.

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

(d) They have low melting points.

(e) They are highly reactive and monovalent elements.

(f) They react with H_2O to give H_2 and alkali metal hydroxides.

Which is more soft Na or K?

4.4.5 Prediction of properties of other elements in Group 1

In Group 1 lithium, sodium and potassium are a collection of relatively soft metals showing a trend in melting point and reaction with water.

The metals in group I are called alkali metals.

- They are very soft.
- Their melting and boiling points decrease down the group.
- When alkali metals react with water, they produce a metal salt and hydrogen .

  metal + water ——–> salt + hydrogen
- The alkali metals become more reactive down the group.

Activity 4.5

Predict the properties of other elements in Group I, from the data given above.

The element after Potassium is Rubidium and you can predict that its reaction with water will be much more violent. We can also predict that Rubidium will have a lower melting and boiling point than the three elements above it. And the elements below Rubidium will be even more reactive and have very low melting and boiling points.It will also react with water to for salt and hydrogen.

4.4.6 Position of Unknown Element in the Periodic Table

You can place an unknown element accurately at appropriate position in the periodic table, and can predict about its properties.

The electronic configuration of an element strongly influences its chemical behaviour. Elements within the group have similar electronic configuration, and therefore similar properties. By examining the electronic configuration of unknown element and comparing it to the known elements in the periodic table, it`s likely position in the periodic table can be determined.

The periodic table allows for the identification of trends and patterns across periods and groups. These trends include variations in atomic size, ionization energy, electron affinity, electronegativity, reactivity, and other properties. By analyzing these trends, it becomes possible to estimate the general properties of unknown element and make guess about its position in the periodic table.

Example 4.7: Identifying position of an unknown element in the periodic table

Suppose you have an unknown element having atomic number 19, and

you want to determine its position in the periodic table.

Problem Solving Strategy

1. Write its electronic configuration.
2. Use valence electronic configuration to locate its position i.e., find its group and period.

Solution:

Electronic configuration : $1s^2,2s^2,2p^6,3s^2,3p^6,4s^1$

Valence shell electronic configuration is $4s^1$, which shows it is an alkali metal, because, Group 1 elements have one electron in valence sub- shell s. As n value of valence sub-shell is 4, this unknown element must lie in the 4^{th} period in the periodic table.

From its position in the periodic table, you can predict its properties. For example will it possess higher or lower melting point, density, reactivity, etc. than the element above or below it.

4.5 TRANSITION ELEMENTS

Elements located in d-block (Group 3 to 12) in the periodic table are called transition elements. These elements exhibit several characteristic properties, which set them apart from other elements. Some of their properties are as follows.

1. High Density

   Transition elements generally possess high densities due to their higher atomic masses and closely packed structures. For example, iron(Fe) has a density of 7.87 g/cm^3, tunguston(W) has a density of 19.3g/cm^3.

2. High Melting Points

   Transition elements have high melting points. This is because their metallic bonding is stronger, which in tern is due to the presence of partially filled d-sub shells. For example tunguston has a melting point of 3422 °C, platinum(Pt) has a melting point of 1768 °C.

3. Variable Oxidation States

   Transition elements exhibit multiple oxidation states. This is because of d-sub shell can also participate in bonding along with s-sub shell. For example, iron(Fe) has oxidation states +2 and +3, copper(Cu) has oxidation states +1 and +2.

4. Coloured Compounds

   Transition elements often exhibit vibrant colours. For example, copper compounds appear blue or green, chromium compounds are often red or green.

5. Catalysts for Industrial Processes

   Transition metals and their compounds are widely used as catalyst in various industrial processes. For example,

   (a) Iron is used in the Haber Process for the synthesis of ammonia.

   (b) Platinum and palladium are used as catalyst in catalytic converters to reduce harmful emissions in automobiles and industrial units.

   (c) Nickel is used as catalyst in the manufacture of margarine.

   (d) Platinum is used as catalyst in the contact process for the manufacture of sulphuric acid.

4.6 LANTHANIDES & ACTINIDES

Lanthanides also known as "rare earth elements" are series of 14 elements located at the bottom of the periodic table. They include elements with atomic number 57 to 71.

Actinides are another series of 14 elements located just below lanthanides. They include elements with atomic number 89 to 103.

4.7 HALOGENS

The elements in group 17 (or Gruop 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, astatine, and tenessine (astatine and tenessine are radio- active elements. Little is known about their properties). All halogens are reactive non-metals and exist as diatomic molecules.

4.7.1 Appearance of halogens

They all exist as diatomic coloured molecular substances. The colour of halogen become darker as you go down the group. At room temperature and pressure(RTP) fluorine(F_2) exist as pale yellow gas, chlorine(Cl_2) as yellow-green gas, bromine(Br_2) as red-brown liquid and iodine(I_2) as grey-black solid. Iodine easily turn into a dark purple vapours on warming.

Electronic Configuration

Halogens possess 7 electrons in their valence shell. They have general electronic configuration ns^2np^5. They need only one electron to complete their valence shell. Consequently, they tend to gain one electron to form univalent negative ions, F^-, Cl^-, Br^-, I^-.

Density of halogens

As you move down the group the number of electrons and protons increases, the size of the atom increases and the volume increases. However, the increase in mass exceeds the increase in volume, so the density, which is mass per unit volume, increases in general. Also fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid . So, the forces of attraction between molecules increase down the group. Solid iodine has molecules that are highly attracted and tightly packed together than bromine. Therefore, as you go down the group of halogens, the forces of attraction increase and the density of the halogen increases.

Densities of halogens

Halogen	Density (g/cm^3 at 25 °C)
Fluorine	0.0017
Chlorine	0.0032
Bromine	3.1028
Iodine	4.933

Reactivity of halogens

The reactivity of halogens is directly related to their ability to gain an electron and form a halide ion (fluoride ion F^-, chloride ion Cl^-, bromide ion Br^-, iodide ion I^-) when they react with other elements. Fluorine has the greatest tendency to gain electrons and form a halide ion, making it the most reactive halogen. As you move down the group, the electronegativity of the halogens decreases. This leads to a decrease in reactivity. Which halogen is the least reactive? Bromine or iodine. Because halogens have a strong tendency to gain electrons, they have a strong oxidizing power, and this power decreases down the group. Thus, the order of decreasing oxidizing power is.

$$F_2 > Cl_2 > Br_2 > I_2$$

Displacement reactions of halogens

Oxidizing power of F_2 is the highest and that of I_2 is lowest. Due to the relative strength as oxidizing agent, it is possible for a free halogen to oxidize or displace the ion of halogen next to it in the group from their aqueous solutions. This means F_2 can oxidize and displace all the halide ions to free halogen. For example,

$$F_{2(g)} + 2KCl_{(aq)} \rightarrow 2KF_{(aq)} + Cl_{2(g)}$$
$$F_{2(g)} + 2KBr_{(aq)} \rightarrow 2KF_{(aq)} + Br_{2(g)}$$
$$F_{2(g)} + 2NaI_{(aq)} \rightarrow 2NaF_{(aq)} + I_{2(g)}$$

Similarly Cl_2 can oxidize Br^- and I^- ions. But I_2 can not oxidize any halide ion.

Hydrogen halides and their thermal stabilities

Halogens react with hydrogen to form hydrogen halides.

$H_2 + X_2 \rightarrow HX$ Where X =F_2, Cl_2, Br_2, I_2

The strength of the hydrogen-halogen bond is related to the electronegativity difference between the hydrogen and halogen atoms. A larger electronegativity difference results in a stronger bond. As we move from HF to HI, the electronegativity difference between the hydrogen and halogen atoms decreases, resulting in weaker bonds in HCl, HBr, and HI. So, the relative thermal stability of hydrogen halides gradually decreases from HF to HI.

Consequently, the energy needed to break H-X decreases in the following orders

HF> HCl> HBr > HI

Prediction of properties of elements in Group VIIA or Group 17

The elements present in Group 17 or VIIA are called halogens. They are poisonous non-metals that have low melting and boiling points that increase down the group. As a result of this increasing boiling and melting points, the state of the halogens at room temperature, changes from gas to liquid to solid down the group (fluorine and chlorine, the 1st and 2nd halogens, are a gas; bromine, the 3rd halogen is a liquid; and iodine, the 4th halogen, is a solid). The colours of halogens also get darker from top to bottom.

Activity 4.6

Predict the properties of other elements in Group VII, from the given data given above.

From this data you can predict how the halogens will behave up and down the group. Astatine, the fifth halogen, will have high melting and boiling points so will be solid at room temperature, and will have a very dark colour.

CONCEPT ASSESSMENT EXERCISE 4.9

Which of the following displacement reactions will occur?

1. $Cl_{2(g)} + 2NaF_{(aq)} \longrightarrow 2NaCl_{(aq)} + F_{2(g)}$
2. $Br_{2(g)} + 2KI_{(aq)} \longrightarrow 2KBr_{(aq)} + I_{2(g)}$
3. $I_{2(g)} + 2KBr_{(aq)} \longrightarrow 2KI_{(aq)} + Br_{2(l)}$
4. $Cl_{2(g)} + 2KBr_{(aq)} \longrightarrow 2KCl_{(aq)} + BR_{2(l)}$
5. $Cl_{2(g)} + 2NaI_{(aq)} \longrightarrow 2NaCl_{(aq)} + I_{2(s)}$

4.8 NOBLE GASES

Noble gases, also known as inert gases, are a group of chemical elements found in Group18 (or Group VIII-A) of the periodic table. They have general electron configuration ns^2,np^6 except He, which has $1s^2$. They are characterized by unique properties. They are odorless, colorless mono-atomic gases and possess very low reactivity with other elements. This low reactivity is due to the presence of a complete valence shell, which makes them stable and unlikely to form chemical bonds with other elements under normal conditions. Noble gases include elements: Helium (He) neon (Ne) argon (Ar) crypto (kr) xenon (Xe) radon (Rn), and oganesson (Og).

DO you know?

Due to their non-reactive nature, noble gases are used in many ways, such as in lighting (e.g. neon signs), refrigeration systems and welding. They are also used in special applications, including filling gas exhaust lines and as a shielding gas in certain industrial processes.

Table 4.1: Electronic Configuration of Nobal Gasses

Element	Atomic Number	Electronic Configuration
Helium	2	$1s^2$
Neon	10	$1s^22s^22p^6$
Argon	18	$1s^22s^22p^63s^23p^6$
Krypton	36	$1s^22s^22p^63s^23p^63d^{10}4s^24p^6$
Xenon	54	$1s^22s^22p^63s^23p^63d^{10}4s^2$ $4p^64d^{10}5s^25p^6$

4.9 COMPARISON OF GENERAL PHYSICAL PROPERTIES OF METALS AND NON-METALS:

Thermal Conductivity:

Metals generally have high thermal conductivity, which means they can conduct heat easily. On the other hand, non-metals tend to have poor conductivity, making them less efficient at conducting heat.

Electrical Conductivity:

Metals are good conductor of electricity, because they have free electrons that can move freely in the metal lattice. Non-metals, with few exception such as graphite, are poor conductor of electricity because they lack free electrons.

Adaptability:

The metals are malleable and ductile. So, they can be hammered, drawn into wires or transformed into thin sheets without breaking. This property is due to metallic bonds which allow atoms exchange easily under pressure. Non-metals are neither malleable nor ductile rather they are brittle.

Melting Points and Boiling Points

Metals generally have high melting points and boiling points due to strong metallic bonds that require a lot of energy to break. On the other hand non-metals often have lower melting points and boiling points because their atoms and molecules are held by weaker bonds such as covalent bonds, van der Waals bonds, or hydrogen bonds that require less energy to break.

CONCEPT ASSESSMENT EXERCISE 4.10

Compare the general properties of metals and non-metals

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 the same group possess similar chemical properties.
- 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 uni-negative gaseous ion.

References for additional information

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

REVIEW QUESTIONS

1. Encircle the correct answer.

(i) Number of periods in the periodic table are:
(a) 8 (b) 7
(c) 16 (d) 5

(ii) Which of the following groups contain alkaline earth metals?
(a) 1A (b) IIA
(c) VIIA (d) VIIIA

(iii) Which of the following elements belongs to VIIIA?
(a) Na (b) Mg
(c) Br (d) Xe

(iv) Main group elements are arranged in _________ groups.
(a) 6 (b) 7
(c) 8 (d) 10

(v) Period number of $^{27}_{13}Al$ is:
(a) 1 (b) 2
(c) 3 (d) 4

(vi) 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) $3s^2,3p^2$

(vii) Which of the following elements you expect to have greater shielding effect?
(a) Li (b) Na
(c) K (d) Rb

(viii) 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 (c) shielding effect

(ix) 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 answer.**

   (i) Write the valence shell electronic configuration of an element present in the 3rd period and Group IIIA.

   (ii) Define halogens.

   (iii) Which atom has higher shielding effect, Li or Na?

   (iv) Explain why, Na has higher ionization energy than K?

   (v) 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.

   A = $1s^2,2s^2$ B = $1s^2,2s^2,2p^6$ C = $1s^2,2s^2,2p^3$

   D = $1s^2$ E = $1s^2,2s^2,2p^6,3s^2,3p^3$ F = $1s^2,2s^1$

   G = $1s^2,2s^2,2p^6,3s^1$ H = $1s^2,2s^2,2p^6,3s^2$

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

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.

$$^{27}_{13}Al,\ ^{32}_{16}S,\ ^{39}_{19}K,\ ^{8}_{6}O$$

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

a. Alkali metals b. Alkaline earth metals

c. Halogens d. Noble gases

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

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

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

(a) An element present in period 3 of Group VA

(b) An element present in period 2 of Group VIA

4. 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 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) $1s^2$

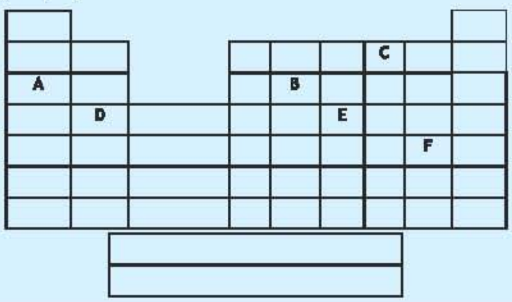
THINK TANK

14. What types of elements have the highest ionization energies and what types of elements have the lowest ionization energies. Argue.

15. i. 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.

ii. 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.

iii. Letter A, B, C, D , E, F indicates elements in the following figure:



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 dipositve 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?

16. 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

17. Argue in what region of the periodic table you will find elements with relatively

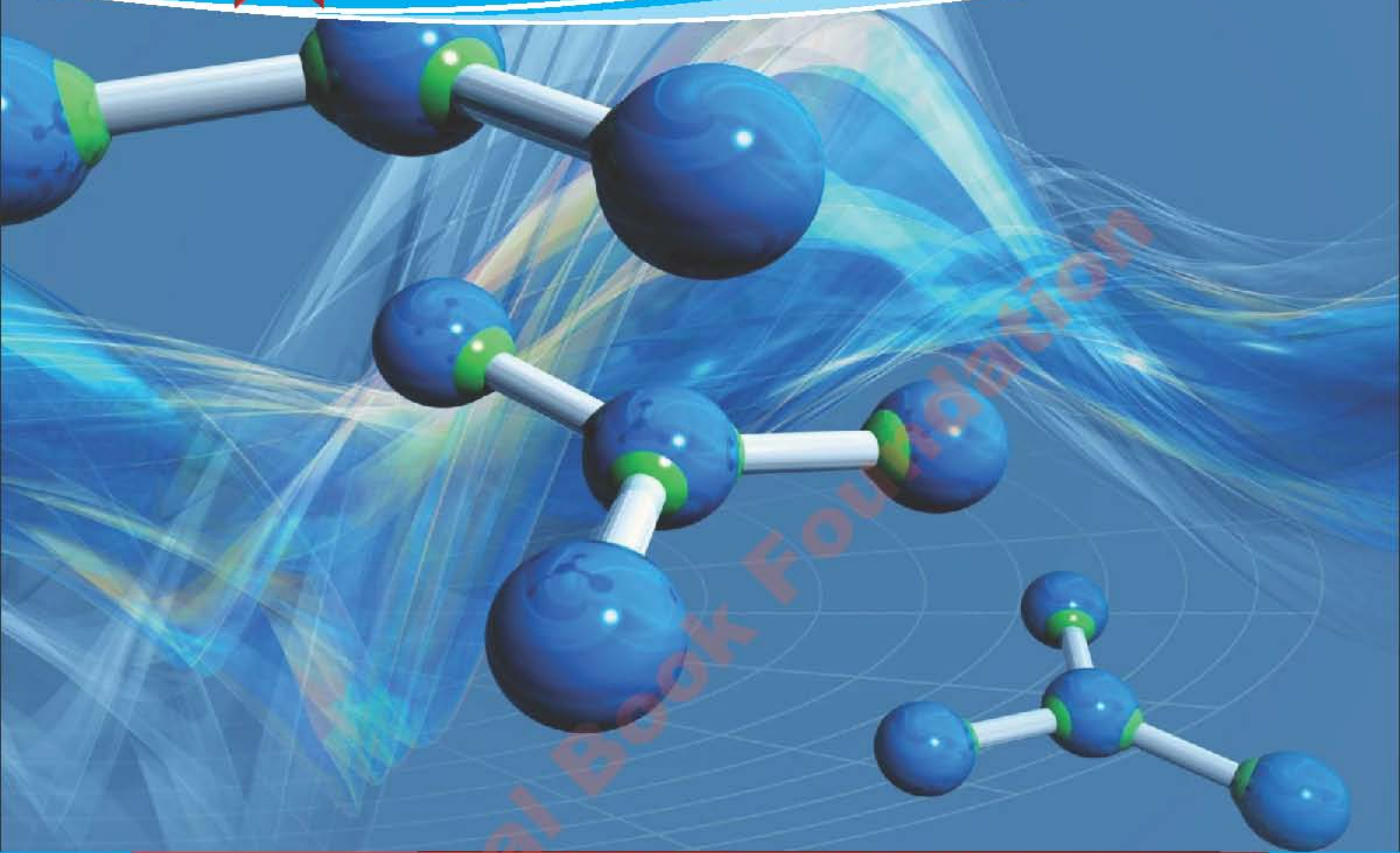
a) high ionization energies

b) low ionization energies

PROJECT

Prepare 3D model of the periodic table (Group Activity)

UNIT 05



CHEMICAL BONDING

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Describe noble gas electronic configuration, octet and duplit rules help predict chemical properties of main group elements.
- Compare between the formation of cations and anions.
- Account for the electropositive and electronegative nature of metals.
- Define ionic, covalent, coordinate covalent and metallic bonds.
- Differentiate between ionic compounds and covalent compounds. (the following points to be included in the respective definitions: a. Ionic bond as strong electrostatic attraction between oppositely charged ions. b. Covalent bond as strong electrostatic attraction between shared electrons and two nuclei. C. Metallic bonds as strong electrostatic attraction between cloud/sea of delocalized electrons and positively charged cations.

- Explain the properties of compounds in terms of bonding and structure.
- Compare properties and use of materials such as strength and conductivity as determined by the type of chemical bond present between their atoms.
- Interpret the strength of forces of attraction and their impact on melting and boiling points of ionic and covalent compounds.
- Justify the availability of free charged particles (electrons or ions) for conduction of electricity in ionic compounds (solid and molten) covalent compounds and metallic bonds.
- Recognize some substances can ionize when dissolved in water (e.g. acids dissolve in water and conduct electricity).
- Justify the suitability of usage of graphite, diamond and metals for industrial purposes (some example may include; a. graphite as lubricant or an electrode. b. diamond in cutting tools. c. metals for wires, and sheets).
- Draw the structure of ionic and covalent compounds along with their formation. (some examples may include: a. ionic bonds in binary compounds such as NaBr, NaF, $CaCl_2$, using dot-and-cross diagrams and Lewis-dot structures. b. simple molecules including H_2, Cl_2, O_2, N_2, H_2O, CH_4, NH_3, HCl, CH_3OH, C_2H_4, CO_2, HCN, and similar molecules using dot-and-cross diagrams and Lewis -structures.

INTRODUCTION

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.

5.1 WHY DO ATOMS REACT?

There are eight groups of normal elements (IA, IIA, IIIA, IVA, VA, VIA VIIA VIIIA) in the periodic table. Group VIIIA consist of the noble gases or zero group elements because they are all very stable and chemically inert under ordinary condition. They exist in atomic form in the atmosphere. They have general electronic configuration = ns^2, np^6 (8 electrons in valence shell) except He ($1s^2$). These noble gases have completely filled valence shells (s and p subshells). Their octet is complete, so they do not participate in ordinary chemical reactions and are called inert gases. They have eight electrons in their valence shell, except He, which has two electrons in its valence shell.

$_{2}He = 1s^2$

$_{10}Ne = 1s^2, 2s^2, 2p^6$

$_{18}Ar = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

In 1916 a chemist G. N. Lewis used the concept of octet (eight electrons) and duplet (2 electrons) electronic rule to explain the reactivity and stability of molecules.

Octet Rule

The octet rule states that an atom is most stable when its valence shell contains eight electrons. This principle is derived from the observation that atoms of the major group elements tend to participate in chemical bonding in the form of eight electrons per atom in the resulting molecule. This rule only applies to the major group element. The chemical behaviour of the main group elements can be predicted with the help of the octet rule. This is because the rule only involves s and p electrons. Molecules such as oxygen, nitrogen, and carbon follow the octet principle. Hydrogen, helium, and lithium follow the duplet rule because their electrons lie in s orbital.

$_{11}Na = 1s^2, 2s^2$, sp6, $3s^1$ (unstable, reactive, incomplete octet) Loss of one electron

$Na^+ = 1s^2, 2s^2, 2p^6$ which is same as that of Ne

$_{17}Cl = 1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ (unstable, reactive, incomplete octet) $_{18}Cl^{-1} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ which is same as that of $_{18}Ar$

Duplet rule

The tendency of atoms to acquire two electronic configuration in their outermost shell during bond formation is called duplet rule. They attain electronic configuration like Helium.

For Example

$_3Li = 1s^2, 2s^1$ lose 1 electron to form Li^+ $(1s^2)$

$_4Be = 1s^2, 2s^2$ loses two electrons to form Be^{+2} $(1s^2)$

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 electronic configuration in their valence shell. For example, hydrogen, lithium and beryllium etc. tend to achieve two electron configuration in the valence shell.

5.2 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 loosing or gaining electrons. When atoms gain or lose electron they acquire the configuration of next noble gas element. The tendency of metal atoms to lose electrons is called electropositivity. Where as the tendency of non-metal atoms to gain electrons is called electronegativity. So, metals are electropositive and non-metals are electronegative elements.

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

Electropositive and Electronegative Elements

Metals are electropositive in nature because all metal atoms lose electrons from their outermost shell in order to become stable and become positively charged. They have low ionization energy and low electronegativity allowing them to easily lose electrons. Therefore, they can form positive ions by losing electrons.

Example: $Na \rightarrow Na^{+} + e^{-}$

$Mg \rightarrow Mg^{2+} + 2e^{-}$

Non-metals are electronegative in nature because all non-metals gain electrons in order to become stable and hence become negatively charged. They have high electronegativity and have high electron affinity. So they can easily form negative ions by gaining electrons. For example:

$F + e^{-} \rightarrow F^{-}$

$O + 2e^{-} \rightarrow O^{2-}$

5.3 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

5.3.1 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.

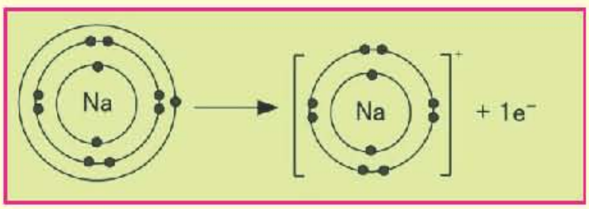
Example 5.1: Describing the formation of cations

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. The 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. A 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.

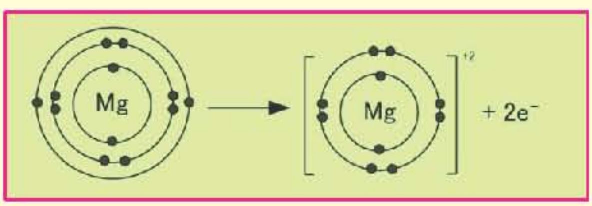


Solution:

(a) Formation of Na^{+} ion

$${}_{11}Na(1s^2 2s^2 2p^6 3s^1) \xrightarrow{-e^-} Na^{+} \ (1s^2 2s^2 2p^6)$$

You can also represent this by following electron dot structure,



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

$$_{12}Mg\ (1s^2 2s^2 2p^6 3s^2) \xrightarrow{-2e^-} Mg^{2+}\ (1s^2 2s^2 2p^6)$$

You can also represent this by electron dot structure,

CONCEPT ASSESSMENT EXERCISE 5.1

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

Example 5.2: Describing the formation of anions.

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

(a) Oxygen(atomic no.8) (b) Fluorine (atomic no. 9)

Problem Solving Strategy:

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

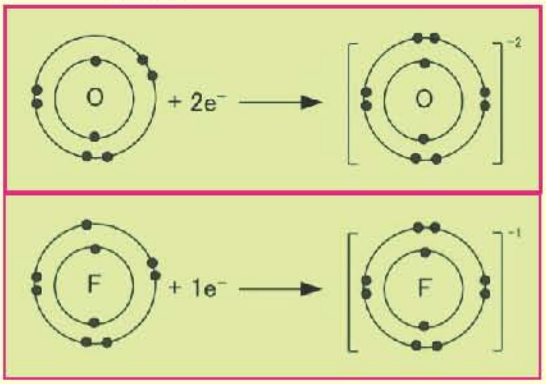
Solution:

(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.

$$_8O(1s^2 2s^2 2p^4) + 2e^- \longrightarrow O^{2-}\ (1s^2\ \underbrace{2s^2 2p^6}_{\text{octet}})$$

You can also represent this by electron dot structure,



(a) 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.

$$_9F(1s^2 2s^2 2p^5) + e^- \longrightarrow F^-\ (1s^2\ \underbrace{2s^2 2p^6}_{\text{octet}})$$

You can also represent this by electron dot structure,

CONCEPT ASSESSMENT EXERCISE 5.2

1. Describe the formation of anions by the following non-metals.
   (a) Sulphur (atomic No. 16) (b) Chlorine(atomic No. 17)
2. Represent the formation of anions by the following non-metals using electron dot structures.
   (a) N (b) P (c) Br (d) H
3. Compare differences between the formation of cations and anions.

Anions and cations have opposite charges. They attract one another by strong electrostatic forces. **"An ionic bond is a strong electrostatic attraction between positively charged metal ions and negatively charged non-metal ions".** 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 are electrically neutral as a whole.

Example 5.3: 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.

(a) Na and Cl (b) Mg and F

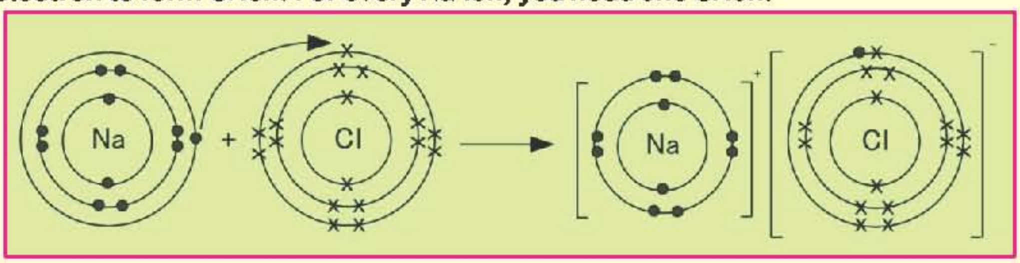
Problem Solving Strategy:

1. The metal atoms form cations and non-metal atoms form anions.
2. 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 and electron cross structures.

Solution:

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

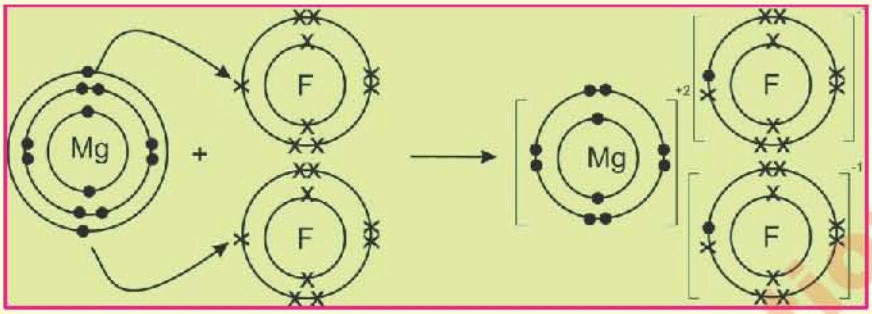
Metal atom tends to lose electrons and non - metal atoms tends to gain electrons to acquire electronic configuration of nearest noble gas. Since a Na atom has one electron in the outer most shell. It losses one electron to form Na^+ ion. Since a Cl atom has seven electrons in outermost shell, it needs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



$$2Na + Cl_2 \longrightarrow 2NaCl \text{ (Sodium Chloride)}$$

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

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



$Mg + F_2 \longrightarrow MgF_2$ (Magnesium Fluoride)

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

CONCEPT ASSESSMENT EXERCISE 5.3

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 5.4: 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. A Mg atom has two electrons in outermost shell. So it loses two electrons to form Mg^{+2} ion. Since an 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. A Na atom has one electron in outmost shell. So it loses one electron to form Na^+ ion. Since a 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.

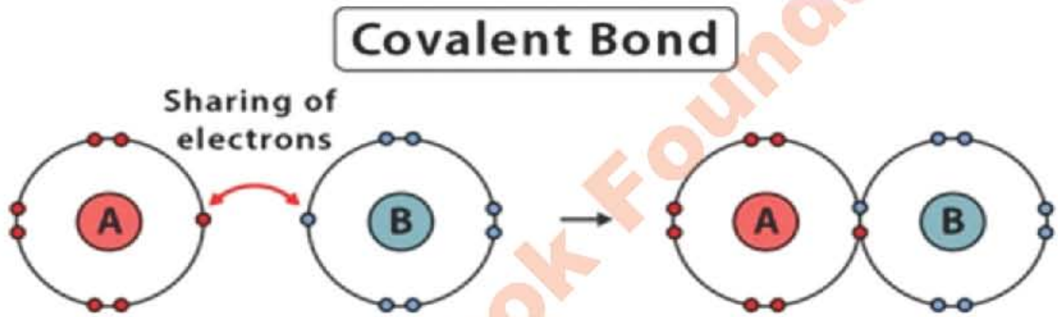
CONCEPT ASSESSMENT EXERCISE 5.4

Recognize the following compounds as having ionic bonds:

(a) KCl (b) $AlCl_3$ (c) MgF_2 (d) NaF (e) NaBr

5.3.2 Covalent Bonds

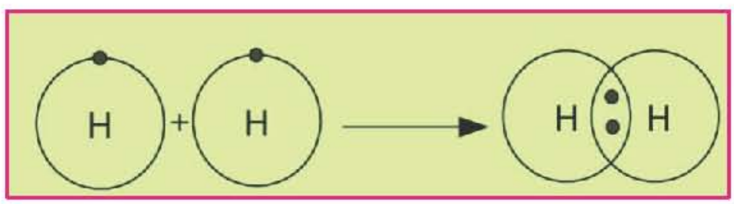
Nonmetal atoms tend to share electrons with each other or with other nonmetal atoms, forming a chemical bond called a covalent bond. A chemical bond formed by mutual sharing of electrons between two atoms is called a covalent bond. General representation of a covalent bond is given below.



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

$$H^{\cdot} + {}^{\cdot}H \longrightarrow H{:}H$$

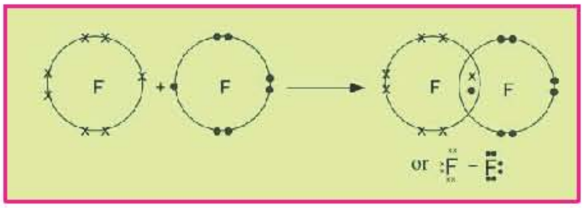
shared pair of electrons



In the formation of this molecule, each hydrogen atom reaches the electronic configuration of the noble gas helium with two valence electrons. An electron pair in the region between two atoms attracts both hydrogen nuclei. This creates a strong electrostatic attraction between the shared electrons and the two nuclei. This means that the situation is more stable than in individual atoms. Because of this stability, the two atoms form a covalent bond.

In a covalent bond, a strong electrostatic forve of attraction between the bonding electrons and two atomic nuclei binds them together.

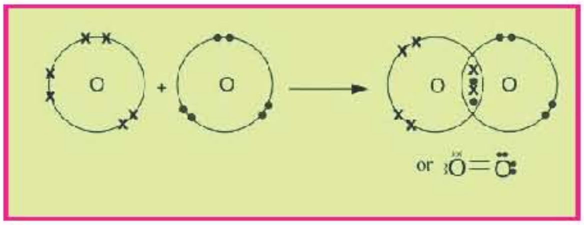
A covalent bond between two atoms can be represented by using electron-dot and electron-cross symbols for the atoms and the resulting molecule. As already discussed valence electrons are represented by dots. Just to understand sharing, we represent valence electrons in one atom by dots and in the other 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.

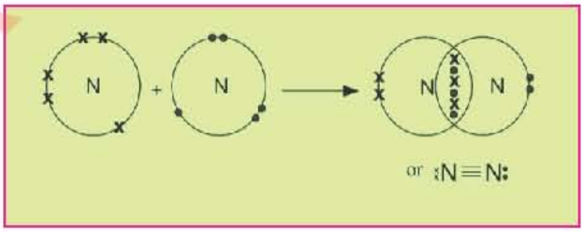
Pairs of valence electrons that are not shared between atoms are called **lone** pairs or lone pairs. A covalent bond formed by sharing **one pair of electrons** is called a single covalent bond. So both H_2 and F_2 molecules contain **single** covalent bond.

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 an 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.**

Consider the formation of O_2 molecules. 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 molecules?

Example 5.5: 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.

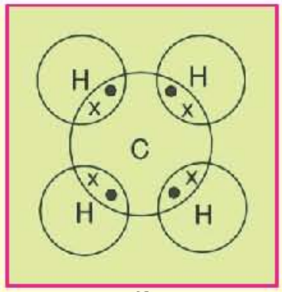
Problem Solving Strategy:

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

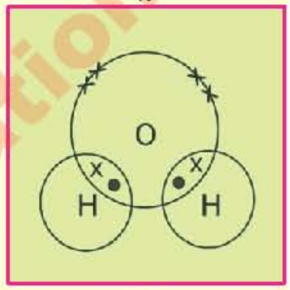
(i) C has four electrons in the valence shell and needs four electrons to complete its octet. H has only one valence electron and needs one electron to complete the duplet. So C can form four single bonds with four H-atoms. C is the central element.



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

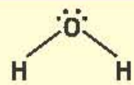
(b) H_2O

(i) O has six valence electrons $\ddot{\text{O}}$ and each hydrogen atom has one valence electron. $\dot{\text{H}}$ So O-atom needs two electrons to complete the 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)



CONCEPT ASSESSMENT EXERCISE 5.5

Draw electron cross and dot structures for the following molecules:

(a) NH_3

(b) HCl

(c) CH_3OH

Example 5.6: Drawing electron cross and dot structures for molecules containing multiple bonds

Draw electron cross and dot structures for the following molecules:

(a) CO_2, a component of air and is responsible for greenhouse effect.

(b) HCN, used as insecticide.

Problem Solving Strategy:

1. 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.
2. 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.
3. Connect central atom with the other atoms by electron pair or pairs to satisfy the octet rule.

Solution:

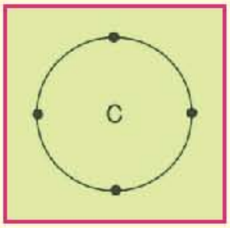
(a) CO_2

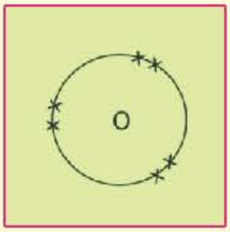
(i) C has four electrons in the valence shell. It needs four electrons to complete octet.

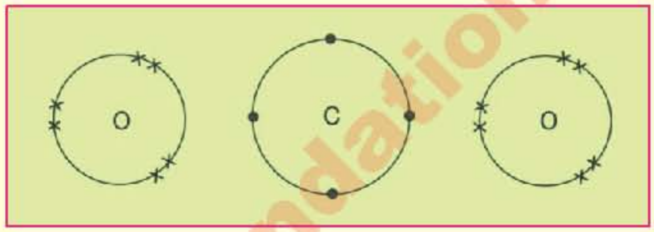
(ii) Each oxygen atom has six valence electrons and needs two electrons to have an octet.

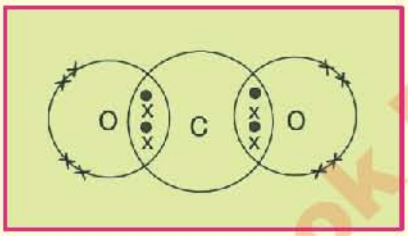
(iii) C is central atom, arrange O-atoms around it.

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







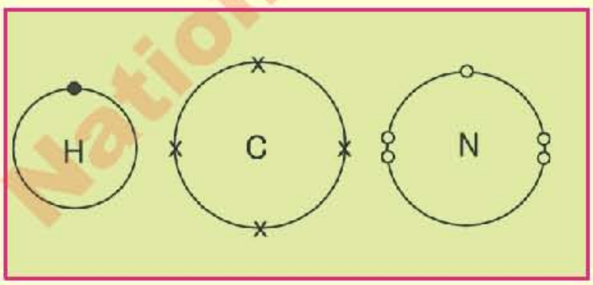


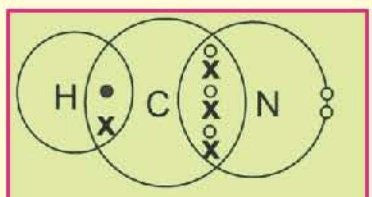
or $:\ddot{O}=C=\ddot{O}:$

(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.





or $H-C\equiv N:$

CONCEPT ASSESSMENT EXERCISE 5.6

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.

(c) C_2H_4, ethane, a component of natural gas.

5.3.3 Types of covalent bond on the basis of polarity:

Non-Polar Covalent bond:

A covalent bond can **form** between two similar atoms such as in H_2, N_2, O_2, Cl_2 etc. It can also occur between two **different** atoms, as in, HCl, H_2O, NH_3, HCN, CO_2etc. When two identical atoms share electron pairs, both atoms exert the same force on the shared electron pairs. Such a covalent bond is called a **nonpolar** covalent bond. For example, bonds H-H, O = O, etc. are non-polar covalent bonds.

Polar Covalent bond:

On the other hand, when two different atoms share an electron pair, both atoms exert different forces on the shared electron pair. **The more** electronegative atom pulls **the** shared electron pairs towards **itself** with a greater force than the **other atom. Thus, the** more electronegative atom **attracts some of the** electron density towards itself. This makes it partially negatively charged and the other atoms partially positively charged. Such a covalent bond is called a polar covalent bond. The forces of attraction between molecules are called intermolecular forces. For example, $\overset{\delta+}{H}-\overset{\delta-}{Cl}$:

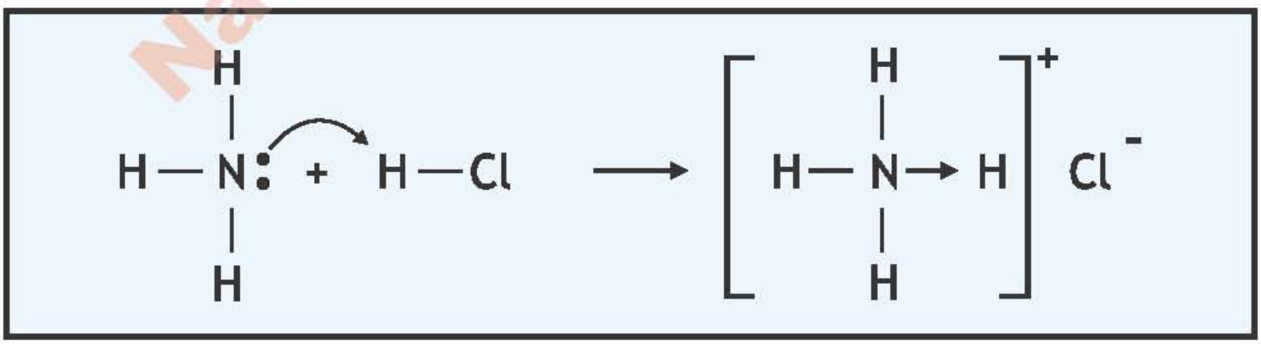
5.3.4 Coordinate Covalent Bond

A coordinate covalent bond is a type of covalent bond where the shared electron pair comes from a single atom (called donor). Atoms are held together because both nuclei attract a pair of electrons. Once a covalent bond is formed, it is impossible to distinguish the origin of the electrons. Such bonding is usually observed when metal ions bind to ligands. However, nonmetals can also participate in this bond. The reaction between a Lewis acid and a base is a covalent coordinate bond.

Examples of coordinate covalent bonds:

1. **Ammonium (NH_4^+) ion**

   The ammonium ion is formed from the reaction of ammonia (NH_3) gas with hydrogen chloride (HCl) gas. In NH_4+, the fourth hydrogen is attached by a coordinate covalent bond because only the hydrogen's nucleus is transferred from the chlorine to the nitrogen. The



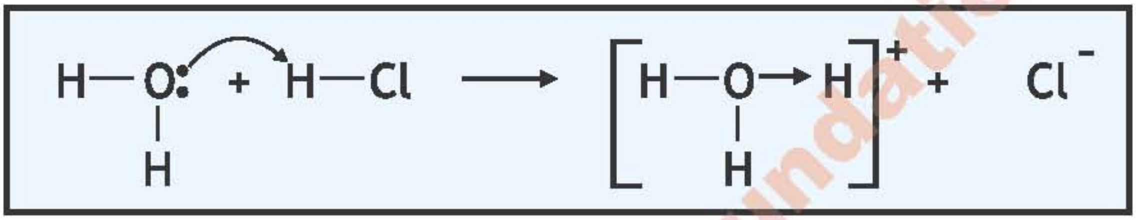
hydrogen's electron is left behind on the chlorine to form a negative chloride (Cl^-) ion.

2. **Hydronium ion (H_3O^+)**

   When hydrogen chloride (HCl) gas dissolves in water to make hydrochloric acid (HCl aq.), a coordinate covalent bond is formed in the hydronium ion. The hydrogen (H) nucleus is transferred to the water (H_2O) molecule, which has a lone pair of electrons to form hydronium. So, H does not contribute any electrons to the bond.

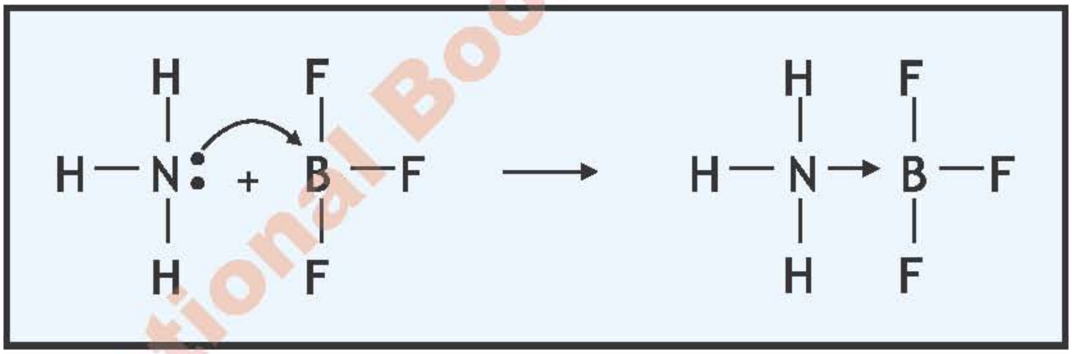
3. **Ammonia Boron Trifluoride (NH_3-BF_3)**

   Boron trifluoride (BF_3) is a compound that does not have a noble gas structure around the boron (B) atom. The boron only has three pairs of electrons in its valence shell and requires a pair to complete the orbital. Hence, BF_3 is electron deficient. The lone pair on the nitrogen (N) of the ammonia (NH_3) molecule is used to overcome that deficiency, and a



CONCEPT ASSESSMENT EXERCISE 5.7

1. Differentiate between polar and non-polar covalent bonds.
2. How is coordinate covalent bond different from normal covalent bond?



5.4 INTERMOLECULAR FORCES

An intermolecular force is the attractive force that exist between the molecules.

Dipole-dipole forces

Dipole-dipole interactions occur between polar molecules. Figure 5.1 shows these interactions.

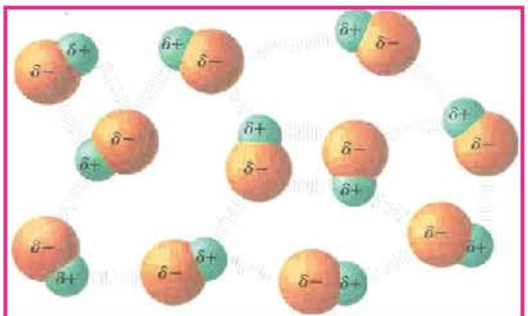


Figure 5.1: Dipole-Dipole interactions

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?

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.

Hydronging bonding

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 5.2 shows hydrogen bonding in water molecules.

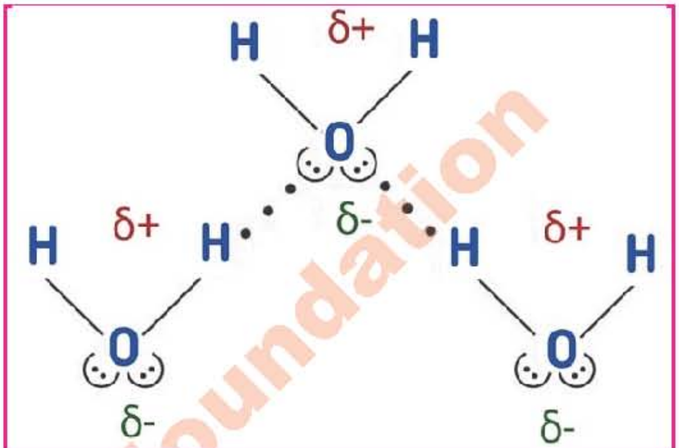


Figure 5.2: Hydrogen bonding in water

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.

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

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.

5.5 NATURE OF BONDING, STRUCTURE AND PROPERTIES

Three main factors are important when determining the properties of a substance:

1. **Type of Particles**

   The types of elementary particles contained. The substance can contain atoms, ions or molecules. For example, if it contains ions (such as sodium chloride), it will conduct electricity when melted or dissolved in water. In order to be soluble in water, the substance must contain ions or polar molecules.

2. **The way elementary particles are connected to each other.**

Particles may have ionic, covalent, metallic, or weak intermolecular forces. The stronger the bond, the higher the melting/boiling point and hardness of the substance.

For example, silicon dioxide (SiO_2) has strong covalent bonds, connecting each atom to several other atoms to form a giant covalent structure. The atoms in silica are difficult to separate, making it very hard and difficult to melt.

On the other hand carbon dioxide has strong covalent bonds between the C and O atoms. But these molecules have weak intermolecular forces between them. The molecules are therefore easily separated and so CO_2 has a low melting/boiling point.

3. **The arrangement of particles**

Particles may be arranged in planes (for example, polymers), in layers (for example, clays, graphite) or in a variety of three-dimensional networks. In graphite atoms are arranged in 2-dimensional layers. This allows the layers of graphite to move over one another .(for example, graphite pencil writing). Diamonds have a large three-dimensional network of carbon atoms, which make it the hardest substance on earth. Metals also have giant structures. metallic bonding is stong , most metals have very high melting and boiling points and are thermally stable.

Conduction of electricity in ionic compounds

Electrical conductivity is achieved by the movement of charged particles. Ionic compounds cannot conduct electricity in the solid state because their ions remain in a fixed position and cannot move. When an ionic compound is melted or dissolved in water. It is ionized, its ions move freely in molten or aqueous solution. Therefore electricity can pass through a molten ionic compound or its aqueous solution.

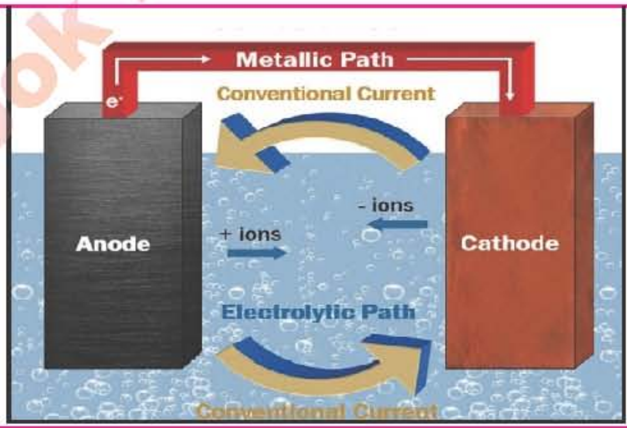


Figure 5.4: conduction of electricity through molten NaCl

Conduction of electricity through acids

Covalent compounds have no free charged particles, so they do not conduct electricity. However, some covalent compounds conduct electricity when dissolved in water. For instance, acids like HCl, H_2SO_4, HNO_3 etc. When these acids are dissolved in water, they ionize and form high concentrations of H+ ions and negatively charged ions. These ions can move freely in aqueous solution. Therefore, aqueous solutions of acids conduct electricity.

Metals are good conductor of electricity because they have free electrons. These electrons are not associated with a single atom. These electrons begin to flow under the influence of electricity. Therefore metals allow electricity to pass through.

Compounds that consists of covalent molecules are called covalent compounds. The intermolecular forces between their molecules are much weaker than the covalent bonds. Therefore, covalent compounds have low melting and boiling points. Since their molecules do not contain any free electrons or ions, they are poor conductors of electricity.

Intermolecular Forces and Their Influence on the Melting and Boiling Points

Tables shows melting and boiling points of some common covalent and ionic compounds.

Table 5.1: Melting point and boiling points of some covalent compounds

Compound	Melting Point (^0C)	Boiling Point (^0C)
Water (H_2O)	0	100
Mehtane (CH_4)	-183	-162
Ethanol (CH_3CH_2OH)	-117	78

Table 5.2: Melting point and boiling points of some ionic compounds

Compound	Melting Point (^0C)	Boiling Point (^0C)
Sodium Chloride (NaCl)	801	1465
Sodium Fluoride (NaF)	996	1695
Magnesium Chloride ($MgCl_2$)	714	1412

Covalent compounds usually have much lower melting points than ionic compounds. For example, a common covalent compound of water has a melting point of 0°C and a boiling point of 100°C. The melting points and boiling points of the common ionic compound sodium chloride are 801°C and 1465°C. This is because ionic compounds involve breaking the ionic bond. Breaking the electrostatic forces between ions requires large amounts of energy. Thus, ionic compounds have high melting points and boiling points. Melting of covalent solids involves the breaking of intermolecular forces, which are much weaker than electrostatic forces. Thus, less energy is required to break the intermolecular forces between covalent molecule

5.5.1 Graphite

Graphite's name is derived from the Greek word "graphein," meaning "to write." It is commonly called black lead. Graphite is an allotrope of carbon. Graphite is formed when carbon is subjected to the intense heat and pressure of the earth's crust and upper mantle.

Structure of Graphite

In graphite, each carbon atom is linked with 3 other carbon atoms by a single covalent bond resulting in the hexagonal ring arranged in a layer. It has a 2-dimensional layers structure. The 4th valence of the carbon atom is satisfied by weak Vander walls forces between 2 layers.

Uses of graphite

1. Graphite is a unique material since it has both metal and non-metal qualities. Moreover, it is a soft mineral with black colour, slippery surface and lustre. These properties are due to layered structure of graphite. Its major uses include:

2. Due to its stability in high temperatures and chemical inertness, graphite is used in many refractory items such as carbon refractory bricks.
3. The electrodes of graphite are used in electrical metallurgical furnaces. It is used as an anode in electrolytic processes.
4. Graphite is used in making moderator rods and reflector components in a nuclear reactor. It is used in the manufacturing of carbon brushes andelectric motors.
5. Graphite material is used in engineering sectors in the making of thrust and journal bearing, piston rings, and vanes.
6. Other applications of graphite include metallurgy, as lubricants, and in the production of paints and pencils.

All these uses are a testament to the unique properties of graphite. The patterned bonding and layered structure make it suitable for such diverse applications.

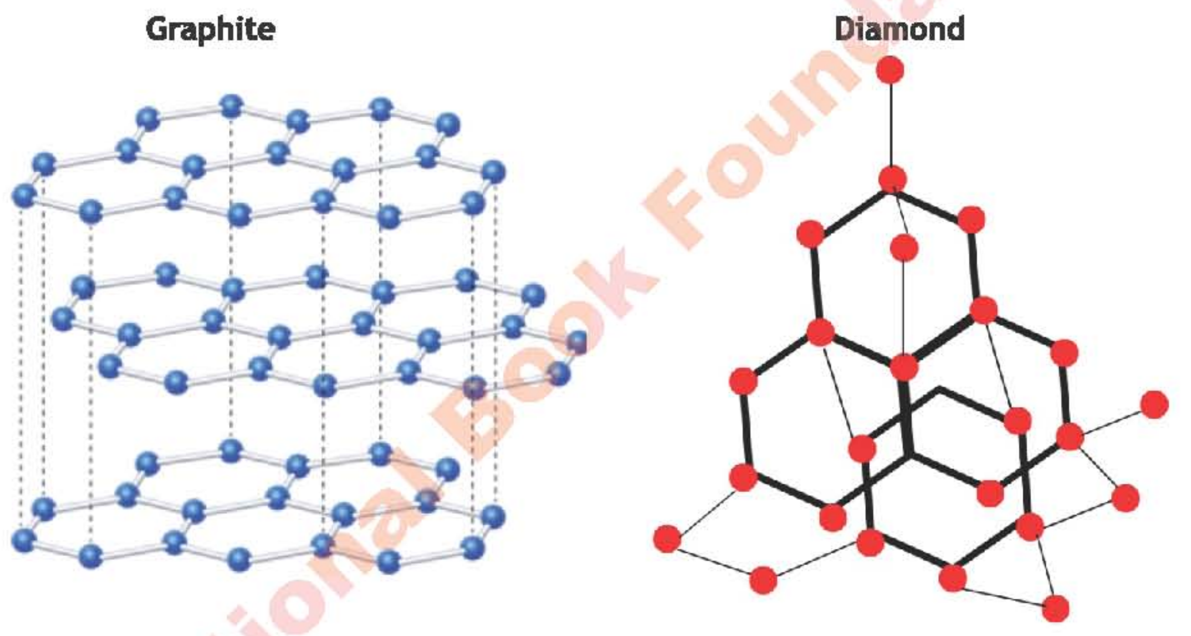


Figure 5.5: Structure of graphite and dimond

5.5.2 Diamond

Diamond is an allotrope of carbon in which the carbon atoms are arranged in a diamond cubic crystal lattice. Thanks to the presence of **strong covalent bonds** and a rigid tetrahedral structure, Diamond is the **hardest** material ever discovered.

Structure of Diamond

In a diamond, the carbon atoms are arranged tetrahedrally. Each carbon atom is attached to four other carbon atoms 1.544×10^{-10} meter away with a C-C-C bond angle of 109.5°. It is a strong, rigid three-dimensional structure that results in an infinite network of atoms. This accounts for diamond's hardness, extraordinary strength and durability and gives diamond a higher density than graphite (3.514 grams per cubic centimeter).

Properties and uses of Diamond

The giant structure and extensive covalent bonding in diamond renders it extraordinary hardness, elasticity, high yield strength, less conductivity, and chemical inertness. Owing to these properties diamond has variety of applications like:

1. Diamonds are most commonly used in ornaments like rings, necklace, earrings, etc. In the gem industry, the value of diamonds is very high. They are used in making jewellery because of their durability and lustre property.
2. Its property of hardness is useful to drill, grind or cut materials. Hence, some blades used for cutting and drills in the industry used diamonds. They are present on the edges and tips in small sizes.
3. Diamonds are used in making medicines and beauty products. They are also used in making medical tools, like tools used in cataract surgery. Nano-diamonds have potential health benefits.
4. Diamonds produce high-quality sound because they are hard and vibrate easily at high speed. It is also used in DJ equipment and high-quality recorders.

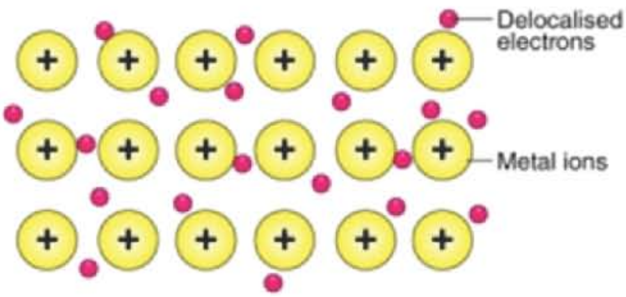
5.5.3 Contrasting ionic and covalent compounds and their uses

The type of chemical bonds significantly influences the properties and uses of materials.

- Ionic compounds are strong in compression, but they are brittle, i.e. they can break easily. In the solid state ionic compounds are poor conductors of electricity. But when they melt or dissolve in water, they conduct electricity due to the free movement of ions. Therefore, batteries and fuel cells use ionic compounds as electrolytes.
- Covalent compounds with giant structures, such as diamond, quartz, silica, etc. are usually very strong and hard. Because of its hardness, diamond is used in cutting and drilling tools. Quartz and silicon dioxide are used in the production of abrasives. Graphite, quartz and silica, because they are stable at high temperatures, are used to make ceramics, glass and refractories. Most covalent compounds are poor conductors of electricity

5.6 METALLIC BONDS

A special type of bonding occurs in metals. In metals, the valence electrons are not confined to individual atoms. These electrons are called free electrons. Metal atoms lose these electrons and form positive ions. The free electrons can move throughout the entire metal structure. This leads to the forming a sea of delocalized electrons called the electron sea. The metal cations are held together by the strong electrostatic attractive forces between the metal cations and negatively charged electron sea. This force gives metals their unique properties. This type of bonding is called metallic bonding.



The properties of metals that are a consequence of metallic bonding include:

Malleability

Ductility

High melting and boiling point

High electrical and thermal conductivity Metallic lustre

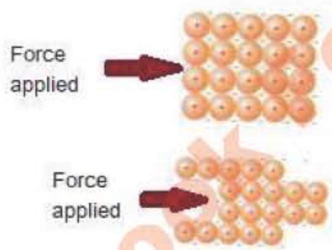
5.6.1 Structure and Properties of Metals Which make it Suitable for Industrial Purposes

1. Metals have giant structures. Metallic bond is strong due to which metals have very high melting and boiling points. This makes them thermally stable.
2. The layers are able to slide over each other, which makes the metals to bent and shaped. This makes them malleable and ductile. They can be drawn into wires and sheets.
3. Metals are good conductors of electricity because the delocalised electrons can move freely. The delocalised electrons can also transfer energy from one place to another and conduct thermal energy.

High melting point

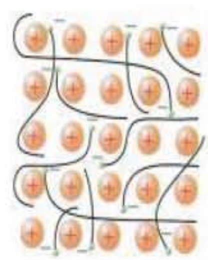
Strong forces of attraction between cations and electrons

Malleable



Rows of ions slide over each other

Electrons moving between the cations



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?

	Atomic Number	Mass Number		Atomic Number	Mass Number
(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.

	Atomic Number		Atomic Number
(a)	20	(b)	18
(c)	17	(d)	15

(iv) Which of the following atoms obey duplet rule?

(a) O_2 (b) F_2

(c) F_2 (d) N_2

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

(a) 2 (b) 3

(c) 4 (d) 6

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

(a) 2 (b) 3

(c) 4 (d) 5

(vii) In the formation of AlF_3, aluminum atom loses _____ electrons.

(a) 1 (b) 2

(c) 3 (d) 4

(viii) Which of the following is not true about the formation of Na_2S:

(a) Each sodium atom loses one electron

(b) Sodium forms cation

(c) Sulphur forms anion

(d) Each sulphur atom gains one electron

(ix) Identify the covalent compound

(a) NaCl (b) MgO

(c) H_2O (d) KF

2. **Give short answer.**

(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 from 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. A sample of sulphur from a volcano was analysed to give the following composition of isotops (At no of S = 16)

Isotope	Abundance (%)
S - 32	95.0
S - 33	0.76
S - 34	4.22

(a) Define the term isotope

(b) Define the term relative atomic mass

(c) Calculate the relative atomic mass of sulphur

(d) Complete the following table.

	Protons	Neutrons	Electrons
S - 32			
S - 34			

(e) Where will you place S in the periodic table.

(f) How many electrons S will loss or gain to acquire stable configuration.

(g) How many atoms of S are there in 0.3 mole of sulphur.

11. 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 by this atom.

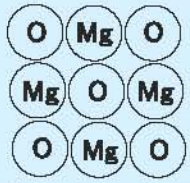
(d) Write electronic configuration of this element.

(e) Point out its group in the periodic table.

(f) Point out its period in the periodic table.

THINK TANK

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

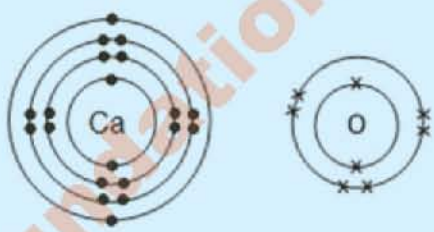


(a) What is the charge on these ions.

(b) How these ions get these charges.

(c) Show with electron cross-dot diagrams the formation of these ions.

13. 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.

14. 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

(a) Which substance is a metal?

(b) Which substance is an ionic compound?

(c) Which substance is a covalent compound?

(d) Which substance is a non-metal?

PROJECT

Prepare a chart displaying different types of bonds with example.

Approved by National Curriculum Council, Secretariat
Ministry of Federal Education & Professional Training
vide letter No. Dated:

قومی ترانہ

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

پاک سرزمین کا نِظام قُوتِ اُخوتِ عوام
قوم ، مُلک ، سلطنت پائندہ تابندہ باد!
شاد باد منزلِ مُراد!

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