A TEXTBOOK

OF

CHEMISTRY

for

Class IX

By

Dr. N. K. Sharma

With learners for life

macmillan
education

MACMILLAN

First published 2015
Revised edition 2017

MACMILLAN PUBLISHERS INDIA PRIVATE LIMITED
Delhi Bengaluru Chennai Kolkata Mumbai
Ahmedabad Bhopal Chandigarh Coimbatore
Cuttack Guwahati Hyderabad Jaipur Lucknow Madurai
Nagpur Patna Pune Thiruvananthapuram Visakhapatnam

ISBN: 978-9350374086

Typeset at: Kalyani Computer Services, New Delhi

Published by Macmillan Publishers India Private Ltd,
21, Patullos Road, Chennai 600002, India

Printed at

Preface

A Textbook of Chemistry for **Class IX** has been written in strict conformity with the **Science Curriculum Framework** prepared by **Department of Curriculum Research and Development, Ministry of Education, Royal Government of Bhutan**.

A sincere attempt has been made for developing this book to adhere the spirit of 'learning without burden' and presenting science as a live and growing body of knowledge with many open areas of investigation, rather than a finished product.

The book also follows a pedagogy that is hands-on and inquiry based. The examples have been taken from things that are directly related to a child's experience.

Students in Bhutan have been learning science in the form of Physics, Chemistry and Biology from class IX onwards. Some of the challenges in learning science are rote learning without much experimentation and lack of opportunities to test the concepts learnt. Learning of scientific knowledge, skills and values from experience and verification through right experiments becomes rational, meaningful and worthwhile.

Science curriculum in Bhutan has undergone a change in recent years, and students learn science till class VIII as an integrated subject. Therefore, a chapter on 'Introduction to Chemistry' has been incorporated which act as a bridge to this specialized science. Other special features of this textbook include competency based questions, topic end questions, relevant solved examples after topics and a model question paper at the end of the book. Utmost care has been given to introduce the new concepts by relating to the existing knowledge of the learners and befitting activities are designed to either induce or consolidate their learning. Children will have the opportunity to learn all the 21st century skills for learning through active learning strategies and at the same time assess their learning.

Every effort has been made to keep this book error free. All suggestions and constructive feedbacks are welcome and we shall try our best to accommodate them in subsequent editions.

– Author

Acknowledgement

We would like to thank all individuals for making contributions in the form of suggestions, feedbacks and comments towards the writing of this textbook. Special mention and gratitude goes to Mr. Wangpo Tenzin, Curriculum Specialist, and Mr. Surjay Lepcha, Curriculum Developer, REC, Paro for helping us and guiding us on every aspect on the development of textbook.

Our gratitude and appreciation also goes to the following teachers for their time and space to attend the review works at Gelephu during the winter vacation of 2014. Their feedbacks and comments were very useful in bringing the book to the current shape.

Mr. Bhoj Raj Rai
Chief/Unit Head,
STEM Division,
Royal Education Council

Mr. Wangpo Tenzin
Curriculum Specialist (Science),
Curriculum Development Centre, REC

Mr. Surjay Lepcha
(Coordinator)
Curriculum Developer (Science),
Curriculum Development Centre, REC

Mr. Basant Pradhan
Lecturer, College of Science & Technology

Mrs. Sonam Zangmo
Chemistry Teacher, Phuentsholing HSS

Mrs. Hari Maya Gurung
Lecturer, Paro College of Education

Mr. Bhim Kumar Sharma
Chemistry Teacher, Damphu HSS

Mr. Cheten Tshering
Chemistry Teacher, Tendu HSS

Mr. Khaganath Gajmer
Chemistry Teacher, Shari HSS

Mr. Rinchen Dorji
Chemistry Teacher, Punakha HSS

Syllabus

Strand: Material and their Properties

1. Classifying Materials

(i) Bonding

- Describe the structure and bonding in metal elements.
- Explain duplet and octet rules.
- Explain the formation of ions when atoms gain or lose electrons to form ionic compounds.
- Explain that the giant ionic lattices are held together by the attraction between oppositely charged ions.
- Explain formation of covalent bonds when atoms share electrons.
- Explain that substances which have covalent bonding can be elements (e.g. H_2, O_2, N_2, C) or compounds (e.g. CH_4, CO_2, SiO_2).
- Describe that substances with covalent bonds can form simple molecular structures or giant structures.
- Outline the physical properties of substances with giant structures (metallic, ionic or covalent) and those with simple molecular structures.
- Explain that the physical properties of substances with giant structures differ (metallic, ionic or covalent) from those with simple molecular structures.

2. Materials and Change

(i) Organic Chemistry

(a) Hydrocarbons – alkanes, alkenes and alkynes.

- Explain that alkanes are saturated hydrocarbons, alkenes and alkynes are unsaturated hydrocarbons.
- Apply the general formula for alkanes (C_nH_{2n+2}), alkenes (C_nH_{2n}) and alkynes (C_nH_{2n-2}).
- Apply the IUPAC rules to the nomenclature of alkanes, alkenes and alkynes.
- Draw and name structural isomers for simple alkanes, alkenes and alkynes, e.g. hexane.
- State the products of burning hydrocarbons and write balanced chemical equations for these reactions.
- Identify unknown hydrocarbon, as being saturated or unsaturated, by tests such as burning, adding bromine water or acidified potassium manganate (VII).
- Describe that petroleum is a mixture consisting mainly of alkane hydrocarbons.
- Describe the process of fractional distillation of crude oil.
- State the uses of the different fractions of crude oil.
- State that cracking involves the breaking of C–C bonds in alkanes and cracking can be used to obtain more useful alkanes and alkenes.

- Describe the formation of addition polymers from the products of crude oil by cracking and polymerization.
- State some uses of addition polymers, e.g. polyethene and polyvinylchloride (PVC).

(ii) Metallurgy

- Describe using balanced chemical equations, the reactions of common metals (e.g. Ca, Mg, Zn, Fe, Pb, Cu) with acids, oxygen and water.
- Construct a reactivity series for metals (e.g. Ca, Mg, Zn, Fe, Pb, Cu) by using the knowledge of the reactions of these metals with acids, water and oxygen.
- Apply the reactivity series to predict the reactions of other metals.

(iii) Green Chemistry

(a) The nitrogen cycle

- Describe the different parts of the nitrogen cycle.
- Explain the importance of converting nitrogen to ammonia for agriculture.
- Explain the manufacturing of nitrogenous fertilizers and their effect on plant growth.
- Explain the environmental consequences of the over-use of chemical fertilizers.

(b) The carbon cycle and global climate change

- Describe that the Earth's atmosphere and oceans have changed over time.
- Describe the role of carbon cycle in maintaining the atmospheric composition.
- Explain that the burning of fossil fuels can upset the balance of the carbon cycle resulting to global climate change.

(c) The problems of plastics

- State the environmental hazards caused by polymers.
- State a number of measures to prevent environmental hazards caused by polymers e.g. how polymers can be processed, the development of biodegradable plastics and the removal of toxic products during the disposal of halogenated plastics (e.g. PVC).

3. Patterns in Chemistry

(i) Patterns in the Periodic Table

(a) The organisation of the Periodic Table

- Explain that the Periodic Table shows all the elements, arranged in order of ascending atomic number.
- Identify that each element has a specific number of protons in the nucleus.
- Explain the connection between the arrangement of outer electrons and the position of an element in the Periodic Table and predict the group of the given elements.
- Explain that elements in the same group of the Periodic Table have similar properties and justify with reasons.
- Explain periodic properties and their variations across the period and down the group.

(b) Group 1 – the Alkali Metals Li, Na and K

- State the physical properties for example, melting points and boiling points of the alkali metals and describe the changes in these properties as the order in group descends.
- Describe the reactions of the alkali metals, Li, Na and K with water, oxygen and chlorine and write balanced chemical equations for each reaction.
- Investigate the reactions of the alkali metals with water, oxygen and chlorine to describe the trends in reactivity as the order in group descends and predict the reactions of Cs and Fr with water, oxygen and chlorine.

(c) Group 18 – The Noble Gases

- State the physical properties for example, melting points and boiling points of the Noble Gases.
- Explain the changes in physical properties of the Noble Gases as the order in group descends.
- State uses of the Noble Gases based on their properties.

(ii) Chemical reactions, Conservation of Mass and Stoichiometry

- Represent chemical reactions by balanced symbol equations to justify that mass is conserved in all chemical reactions.

(iii) Rates of Reaction

- Give examples of different chemical reactions where there is a great variation in the rates at which these reactions take place.
- Describe alteration of the rates of reactions by varying temperature or concentration, or by changing the surface area of a solid reactant, or by adding a catalyst.

(iv) Energy Transfer in Reactions

- Classify reactions as exothermic reaction or endothermic reaction depending on the temperature change that takes place during the course of the reaction.

Assessment

Assessment in science involves detailed process of measuring students' achievement in terms of knowledge, skills, and attitude. The progress of learning is inferred through analysis of information collected. The accuracy and objectivity of assessment determines its validity. The modality and components of assessment should be clearly conveyed to the students. The teacher's expectations should be made clear to students and appropriate learning outcomes should be set. The teachers can play an important role in the students' achievement by effectively monitoring their learning, and giving them constructive feedback on how they can improve, and provide the necessary scaffolding for the needy learners as identified through reliable assessment techniques and tools.

PURPOSE OF ASSESSMENT

Assessment is used to:

- **Inform and guide teaching and learning:** A good assessment plan helps to gather evidences of students' learning that inform teachers' instructional decisions. It provides teachers with information about the performance of students. In addition to helping teachers formulate the next teaching steps, a good classroom assessment plan provides a road map for students. Therefore, students should have access to the assessment so they can use it to inform and guide their learning.
- **Help students set learning goals:** Students need frequent opportunities to reflect on what they have learnt and how their learning can be improved. Accordingly, students can set their goals. Generally, when students are actively involved in assessing their own next learning steps and creating goals to accomplish them, they make major advances in directing their learning.
- **Assign report card grades:** Grades provide parents, employers, other schools, governments, post-secondary institutions and others with summary information about students' learning and performances.
- **Motivate students:** Students are motivated and confident learners when they experience progress and achievement. The evidences gathered can usher poor performers to perform better through remedial measures.

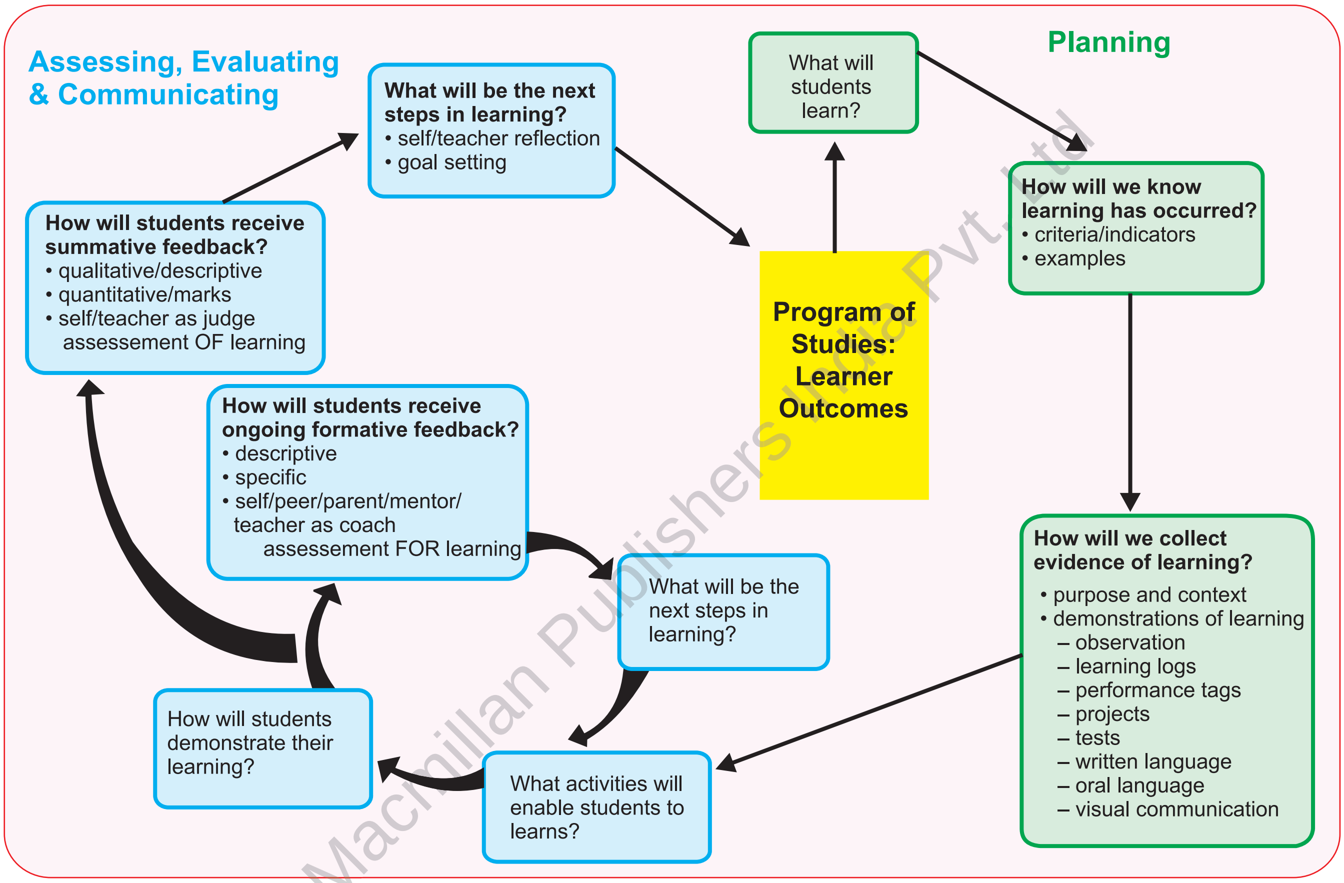
The achievements and performances of the learners in chemistry are assessed on the following three domains:

- **Scientific knowledge:** Basic knowledge and understanding of formation of substances with varying properties; hydrocarbons and their uses; reactivity series of metals that influence their properties; patterns in the periodic table with focus on group 1 and group 0 elements; chemical reactions, their kinetics and energy transfer; Biogeochemical cycles, and the impact of toxic chemicals on the environment and the community.
- **Working scientifically:** Basic understanding of the nature of science, and how science works. Demonstration of logical and abstract thinking and comprehension of complex situations. Explore how technological advances are related to the scientific ideas underpinning them. Compare, contrast, synthesize, question and critique the different sources of information, and communicate their ideas clearly and precisely in a variety of ways, including the use of ICT.
- **Scientific values and attitudes:** Consider the power and limitations of science in addressing social, industrial, ethical and environmental issues, and how different groups in the community and beyond may have different views about the role of science. They make informed judgments on statements and debates that have a scientific basis, and use their learning in science for planning positive action for the welfare of themselves, others in their community and the environment.

The Assessment Process

Effective classroom assessment in Science:

- assesses specific outcomes in the program of studies.
- the intended outcomes and assessment criteria are shared with students prior to the assessment activity.
- assesses before, during and after instruction.
- employs a variety of assessment strategies to provide evidence of students' learning.
- provides frequent and descriptive feedback to students.
- ensures students can describe their progress and achievement, and articulate what comes next in their learning.
- informs teachers and provides insight that can be used to modify instruction.



Scheme of Assessment in Science

The following schemes of assessment are used to assess students' performance:

1. ***Continuous Formative Assessment (CFA)***

   Formative assessment is used to provide feedback to teachers and students, so that teaching and learning can be improved through the provision of regular feedback and remedial learning opportunities. It also enables teachers to understand what teaching methods and materials work best.

   CFA facilitates teachers to diagnose the learning needs of learners and recognize the individual differences in learning. Through the constructive feedback, students are able to understand their strengths and weaknesses. It also empowers them to be self-reflective learners, who monitor and evaluate their own progress.

   CFA should happen daily throughout the teaching-learning processes of the academic year. It is NOT graded, as it is only to give continuous feedbacks to the students.

2. ***Continuous Summative Assessment (CSA)***

Continuous Summative Assessment is another form of continuous assessment (CA). It helps in determining the student's performance and the effectiveness of instructional decisions of teachers. The evidences from this assessment help students to improve learning, and mandate teachers to incorporate varied teaching strategies and resources to ensure quality teaching and learning in the science classes. This assessment also empowers students to be self-reflective learners, who monitor and evaluate their own progress.

In CSA, the students' performances and achievements are graded. This ensures active participations of learners in the teaching and learning processes.

3. ***Summative Assessment (SA)***

Summative assessment (SA) is conducted at the end of the first term and at the end of the year to determine the level of learning outcomes achieved by students. The information gathered is used by teachers to grade students for promotion, and to report to parents and other stakeholders.

The identified techniques for SA are term examinations - first term and annual examinations. The questions for the term examinations should cover all the three domains of science learning objectives, using the principles of Bloom's taxonomy.

Assessment Matrix

Types of assessment	CFA			CSA			SA	
Definition	It is a continuous process of assessing student's problems and learning needs and to identify the remedial measures to improve student's learning. It also enables teachers to understand what teaching methods and materials work best.			It is a continuous process of grading student's performances and achievements. Teachers provide feedbacks for improvement. It also enables teachers to understand what teaching methods and materials work best.			Assesses student's cumulative performances and achievements at the end of each term.	
Domains	Scientific knowledge (SK)	Working scientifically (WS)	Scientific values and attitudes (SV)	Scientific knowledge (SK)	Working scientifically (WS)	Scientific values and attitudes (SV)	SK, WS & SV	SK, WS & SV
Techniques	Quiz and debate, class presentation, homework, class work, immediate interaction with students.	Immediate interaction with students, class work, home work, experiments, exhibition, case studies	Observation of student's conduct, in group work, field trip, excursion, etc.	Home work and chapter end test.	Practical work	Project Work	Term exam.	Term exam.
Assessment Tools	Q&A, checklist and anecdotal records.	Checklist and anecdotal records.	Checklist and anecdotal records.	Rubrics (HW) and paper pencil test (Chapter end test).	Rubrics (Practical work)	Rubrics (Project work)	Paper pencil test	Paper pencil test
Frequency interval (when and how)	Checklists and anecdotal records must be maintained for each topic throughout the academic year.			HW-for every chapter, Chapter end test – for every chapter.	Practical work once in each term	Project Work – Once for the whole year but assessed two times (half yearly)	Once in a term.	Once in a year.
Format in Progress Report				SK	WS	SV	Mid-Term	Annual Exam
Weightings				T1 = 2.5 T2 = 2.5	T1 = 5 T2 = 5	T1 = 2.5 T2 = 2.5	T1 = 30	T2 = 50

Assessment Techniques and Tools

The following techniques and tools are used in assessing students' performances with objectivity.

(a) Observation Checklist

Observing students as they solve problems, model skills to others, think aloud during a sequence of activities, or interact with peers in different learning situations provides insight into student's learning and growth. The teacher finds out under what conditions success is most likely, what individual students do when they encounter difficulty, how interaction with others affects their learning and concentration, and what students need to learn next. Observations may be informal or highly structured, and incidental or scheduled over different a period in different learning contexts.

Observation checklists are tools that allow teachers to record information quickly about how students perform in relation to specific outcomes from the program of studies. Observation checklists, written in a yes/no format can be used to assist in observing student performance relative to specific criteria. They may be directed toward observations of an individual or group. These tools can also include spaces for brief comments, which provide additional information not captured in the checklist.

Tips for using Observation Checklists

(i) Determine specific outcomes to observe and assess.

(ii) Decide what to look for. Write down criteria or evidence that indicates the student is demonstrating the outcome.

(iii) Ensure students know and understand what the criteria are.

(iv) Target your observation by selecting four to five students per lesson and one or two specific outcomes to observe. Date all observations.

(v) Collect observations over a number of lessons during a reporting period and look for patterns of performance.

(vi) Share observations with students, both individually and in a group. Make the observations specific and describe how this demonstrates or promotes thinking and learning.

(vii) Use the information gathered from observation to enhance or modify future instruction.

Sample Checklist

Name	Topic : Chemical bonding									Teacher's comments
	Scientific knowledge			Working scientifically			Scientific values			
	Explains the covalent bond	explains the structures of various molecules based on the concept of chemical bonds	Draw electron dot structure to represent simple ionic compounds	Follows correct experimental procedures	Handles equipment, apparatus, chemicals safely	Demonstrates ability to set up experiments	Evaluate the effects of chemicals on environment views	Shows curiosity to learn science	Demonstrates concern for oneself and others	
Ugyen										
Rinchen										
Pramod										
Choden										

(b) Anecdotal Notes

Anecdotal notes are used to record specific observations of individual student **behaviours, skills, and attitudes** in relation to the outcomes of the science teaching and learning process. Such notes provide cumulative information on students' learning and direction for further instruction. Anecdotal notes are often written as ongoing observations during the lessons, but may also be written in response to a product or performance of the students. They are generally brief, objective, and focused on specific outcomes. The notes taken during or immediately following an activity are generally the most accurate. Anecdotal notes for a particular student can be periodically shared with the student, or be shared at the student's request.

The purpose of anecdotal notes is to:

- provide information regarding a student's development over a period of time.
- provide ongoing records about individual instructional needs.
- capture observations of significant behaviours that might otherwise be lost.

Tips for maintaining Anecdotal Notes

(i) Keep a notebook or binder with a separate page for each student. Write the date and the student's name on each page of the notebook.

(ii) Following the observations, notes are recorded on the page reserved for that student in the notebook.

(iii) The pages may be divided into three columns: Date, Observation and Action Plan.

(iv) Keep notes brief and focused (usually no more than a few sentences or phrases).

(v) Note the context and any comments or questions for follow-up.

(vi) Keep comments objective. Make specific comments about student strengths, especially after several observations have been recorded and a pattern has been observed.

(c) Project Work

Project work is one of the best ways to practice the application of scientific conceptual ideas and skills. The very purpose of including project work is to provide opportunity to explore and extend their scientific knowledge and skills beyond the classroom. Students learn to organize, plan and piece together many separate ideas and information into a coherent whole. Through project work, students learn various scientific techniques and skills, including data collection, analysis, experimentation, interpretation, evaluation and drawing conclusion; and it fosters positive attitude towards science and environment.

The new science curriculum mandates students to carryout project work to help them to:

(i) develop scientific skills of planning, designing and making scientific artefacts, carrying out investigations, observation, analysis, synthesis, interpretation, organization and recording of information.

(ii) enhance deeper understanding of social and natural environment.

(iii) develop student's ability to work in group and independently.

(iv) provide opportunity to explore beyond the classroom in enhancing their scientific knowledge and skills, which will contribute towards the development of positive attitudes and values towards science and environment.

(v) understand how science works and the nature of scientific knowledge.

(vi) develop oral and written communication skills.

Teachers can facilitate students to carry out the project work by considering the following suggested guidelines.

- Allow students to select their own project ideas and topics.
- Encourage students to be scientifically creative and productive.
- Provide a clear set of guidelines for developing and completing projects.
- Help students to locate sources of information, including workers in science-related fields who might advise them about their projects.
- Allow students the option of presenting their finished projects to the class.
- Inform students about the general areas on which assessment may be made. For example, scientific content or concepts, originality of ideas, procedures, and the presentation.
- Advice students to contact their teacher for further assistance or consultations, for, students must be closely guided by the teacher starting from the selection of the topic, doing investigations, data collection, and analysis to writing report in a formal style.

At the end of the project work, every student must prepare a project work report, about 2000 to 2500 words, in the formal format, suggested in the following section. The product of the project work must be inclusive of write ups, illustrations, models, or collection of real objects.

Each student is assigned a Project Work for the academic year. The project work is assessed out of 28 marks, which should be converted to 5 marks for the whole year. Students can share their project work findings, either in the form of class presentation or display.

Following are some of the useful steps that students may follow.

1. ***Select a topic for the science project***

   The first step in doing science project is selecting a topic or subject of your interest. Teachers guide students in identification and selection of the topic. The concerned teacher has to approve the topic prior to the commencement of the project work.

2. ***Gather background information***

   Gather information about your topic from books, magazine, Internet, people and companies. As you gather information, keep notes from where you got the information as reference list.

3. ***Write your hypothesis***

   Based on your gathered information, design a hypothesis, which is an educated guess in the form of a statement, about what types of things affect the system you are working with. Identifying variables is necessary before one can make a hypothesis. For example, the rate of a chemical reaction is affected by the size of particles of reactants. Develop a research question supported by a few questions to test your hypothesis. For example, how does the particle size affect the rate of a chemical reaction? Sub-questions may include, what is the reaction rate when granulated forms of reactants are used? What happens to the reaction rates when powdered reactants are used?

4. ***Identify variables***

   The hypothesis and the research questions should guide you to identify the variables. When you think you know what variables may be involved, think about ways to change one at a time. If you change more than one at a time, you will not know what variable is causing your observation. Sometimes, variables are linked and work together to cause something. At first, try to choose variables that you think act independently of each other.

5. ***Design an experiment or observation method***

   Having made the hypothesis, design an experiment to test the hypothesis and devise the method of observation. Make a systematic list of what you will do or observe to answer each question. This list is known as experimental or observational procedure. For observations or an experiment to give answers, one must have a 'control'. A control is a neutral 'reference point' for comparison that allows you to see what changing or dependent variable does by comparing it to not changing anything. Without a control, you cannot be sure what variable causes your observations.

6. ***Write a list of material***

Make a list of materials useful to carry out your experiment or observations.

7. ***Write experiment results***

Experiments are often done in series. A series of experiments can be done by changing one variable at a time. A series of experiments are made up of separate experimental "runs". During each run, you make a measurement of how much the variable affected the system under the study. For each run, a different amount of change in the variable is used. This produces a different degree or amount of responses in the system. You measure these responses and record data in a table form. The data from the experiments and observations are considered as a "raw data" since it has not been processed or interpreted yet. When raw data is processed mathematically, for example, it becomes result.

8. ***Write a summary of the results***

Summarize what happened. This can be in the form of a table of processed numerical data, or graphs. It could also be a written statement of what occurred during experiments. It is from calculations using recorded data that tables and graphs are made. Studying tables and graphs, one can see trends or patterns that tell you how different variables cause to change the observations. Based on these trends, you can draw conclusions about the system under the study. These conclusions help to confirm or deny your original hypothesis. Often, mathematical equations can be made from graphs. These equations can help you to predict how a change will affect the system without the need to do additional experiments. Advanced levels of experimental science rely heavily on graphical and mathematical analysis of data. At this level, science becomes even more interesting and powerful.

9. ***Draw conclusions***

Using the trends in your experimental data and your experimental observations, try to answer your original questions. Is your hypothesis correct? Now is the time to pull together what happened in the form of conclusion, and assess the experiments you did. Describe, how variables have affected the observations, and synthesize a general statement. For example, the rate of chemical reaction increases with increase in surface area of reactants.

10. ***Write a report on the project***

Having completed all the steps of experiment and investigation with appropriate results and conclusion drawn, the last thing is to write a report. The report should start with an introduction on the topic related to your hypothesis, purpose of the study, literature review, methods used, findings, and

The **Format for Project Work** write-up (report) should include the following aspects:

- The title of the project work.
- Acknowledgement: Show courtesy to thank the people and organizations for the help received.
- Table of content.
- Introduction: What is the topic about, and why was the topic chosen? hypothesis, research question.
- Background information: Scientific concepts, principles, laws and information on the topic.
- Methodology: Methods of data collection – sampling, tools used, etc; data sorting.
- Data analysis: Data tabulation, data processing, findings, etc. presented in a logical order with illustrations, photographs, and drawings where appropriate and necessary to support the findings.
- Conclusion: Reflection of the findings, learner's experiences and opinions regarding the project.
- Bibliography: List of the sources of the information.

conclude with conclusions. Do not forget to acknowledge the support provided by all individuals and organizations. Write a bibliography to show your references in any form. Such information includes the form of document, name of writer, publisher, and the year of publication.

The teacher uses the 'Rubric for the Project Work' given below to assess the student's project work. Random viva voce is necessary to guide and support students' work during the course of project work.

Criteria for the Project Work

Name	Criteria						Bibliography (4)	Total scores (28)
	Problem and hypothesis (4)	Background research on the hypothesis (4)	Experimental design / materials / procedure (4)	Investigation (4)	Analysis (4)	Format and editing (4)		
Nima								
Dawa								

Rubrics for the Project Work

Criteria	Scoring				Total Score (28)
	4	3	2	1	
Problem and Hypothesis	• Problem is new, meaningful and well researched. • Hypothesis is clearly stated in the "IF...THEN" format.	• Problem is not new but meaningful. • Hypothesis is clearly stated.	• Problem is stated but neither new nor meaningful. • Hypothesis is not clearly stated.	• Problem is not stated and • Hypothesis is unclear.	
Background research on the hypothesis	• Research is thorough and specific. • All the ideas are clearly explained.	• Research is thorough but not specific. • Most ideas are explained.	• Research is not thorough and not specific. • Few ideas are explained.	• Research not thorough and • Ideas are not explained.	
Experimental design / materials / procedure	• Procedure is detailed and sequential. • All materials are listed. • Safety issues have been addressed.	• Procedure is detailed but not sequential. • Most materials are listed. • Safety issues have been addressed.	• Procedure is not detailed and not sequential. • Few materials are listed. • Few safety issues have been addressed.	• A few steps of procedure are listed. • Materials list is absent. • Safety issues are not addressed.	
Investigation	• Variables have been identified, controls are appropriate and explained. • Sample size is appropriate and explained. • Data collected from at least 4 sources.	• Variables have been identified and controls are appropriate but not explained. • Sample size is appropriate. • Data collected from at least 3 sources	• Variables have somewhat been identified, controls are somewhat known. • Sample size is not appropriate. • Data collected from at least 2 sources.	• Missing two or more of the variables or the controls. • Sample size is not considered. • Data collected from only 1 source.	

Criteria	Scoring				Total Score (28)
	4	3	2	1	
Analysis & conclusion	• Appropriate tool used for analysis. • Explanation is made for how or why the hypothesis was supported or rejected. • Conclusion is supported by the data. • Reflection is stated clearly.	• Appropriate tool used for analysis. • Conclusions are supported by the data. • Not enough explanation is made for how or why the hypothesis was supported or rejected. • Reflection is stated.	• No appropriate tool used for analysis. • Not enough explanation is made for how or why the hypothesis was supported or rejected. • Conclusion is not appropriate. • Reflection is not clear.	• No appropriate tool used for analysis. • Not enough explanation is made for acceptance and rejection of hypothesis. • Conclusion is absent. • Reflection is not stated.	
Format and editing	• Correct format followed throughout. • Report is free of errors in grammar, spelling or punctuation.	• Only one aspect of format is incorrectly done. • Report contains a few errors in grammar, spelling, and punctuation.	• Only two aspects of format are incorrectly done. • Report contains some errors in grammar, spelling, punctuation	• Three or more aspects of format are missing. • Report contains many errors in grammar, spelling, and punctuation.	
Bibliography	• Five or more references are cited in APA format and referenced throughout the paper and presentation.	• Three or four references are cited and referenced throughout the paper and presentation.	• One or two references are cited and referenced throughout the paper and presentation.	• No references made.	
				TOTAL SCORE	

(d) Practical Work

Learning by doing is fundamental to science education. Practical work is one of the means that helps students to develop their understanding of science, appreciate that science is evidence driven and acquire hands-on skills that are essential to science learning and in their future lives. The practical work as defined by SCORE (2009a) is *'a "hands-on" learning experience which prompts thinking about the world inwhich we live'*.Therefore, the purposes of doing practical in science classes are to:

(i) help students to gain or reinforce the understanding of scientific knowledge.

(ii) develop students' understanding of the methods by which the scientific knowledge has been constructed.

(iii) increase a student's competence to engage in scientific processes such as in manipulating and/or observing real objects and materials with due consideration for safety, reliability, etc.

(iv) develop technical and scientific skills that improve science learning through understanding and application.

(v) develop manipulative skills, knowledge of standard techniques, and the understanding of data handling.

(vi) Inculcate excitement of discovery, consolidation of theory, and the general understanding of how science works.

Practical work is integral to the aspects of thinking and working scientifically in science, and must be built in as a full learning experience for students. Students are engaged in a range of practical activities to enable them to develop their understanding through interacting with apparatus, objects and observations.

The assessment of students' scientific skills and their understanding about the scientific processes through practical work is crucial in the process of science learning. To ensure the validity, assessment needs to sample a range of activities in different contexts; and reliability is ensured through the appropriate moderation procedures so that fairness in assessment is maintained.

The new science curriculum envisages that students are given the opportunity to undertake work in which they make their own decisions. They should be assessed on their ability to plan, observe, record, analyze, communicate and evaluate their works.

To ensure that the assessment in the practical is evidence-based and objective, rubrics is used. The rubrics are scored out of 16, which must be reduced to 5% each for the two terms.

Criteria for the Practical Work

Name	Criteria				Total scores (16)
	Scientific operation & report format (4)	Results & data representation (4)	Analysis & discussion (4)	Conclusions (4)	
Sonam					
Wangmo					

Rubrics for the Practical Work

Criteria	Scoring				Total Score (16)
	4 (Very good)	3 (Good)	2 (Fair)	1 (Poor)	
Scientific operation	• Purpose is clear purposeful. • All the procedures are followed systematically. • Full attention is given to relevant safety for oneself and others.	• Purpose is clear purposeful. • All the procedures are followed but not done systematically. • Work is carried out with some attention to relevant safety procedures.	• Purpose is inaccurate, general or extraneous. • A few procedures are skipped. • Safety procedures were frequently ignored	• Purpose is vague or inaccurate. • Procedures are not followed • Safety procedures are ignored completely.	
Results & data representation	• Representation of the data/results in tables and graphs with correct units of measurement. • Transformations in the results/data are evident. • Graphs and tables are scaled correctly, with appropriate titles and labels.	• Representation of the data/results in tables and graphs with some error in units of measurement. • Transformations in some of the results/data are evident. • Graphs and tables are scaled correctly with appropriate titles but no labels.	• Representation of the data/results in tables and graphs numerous error in units of measurement. • Transformations in most of the results/data are not evident. • Graphs and tables are scaled correctly, but without appropriate titles and labels.	• Representation of the data/results in tables and graphs are not relevant. • Transformations in the results/data are not evident. • Some attempts are evident to produce graphs from the data/results.	

Analysis & discussion	• All the tools used for analysis are appropriate. • A comprehensive discussion, containing a comparative analysis is evident. • The experimental findings are significant to the purpose of the experiment.	• Most of the tools used for analysis are appropriate. • A comprehensive discussion, containing some comparative analysis is evident. • The experimental findings do not have strong significance to the purpose of the experiment.	• Only a few tools are used for analysis. • A comprehensive discussion, containing a few comparative analysis is evident. • The experimental findings have weak significance to the purpose of the experiment.	• No appropriate tools are used for analysis. • Comprehensive discussion is absent. • The experimental findings have no significance to the purpose of the experiment.	
Conclusions	• Conclusions are drawn from the findings and are significant to objectives of the experiment. • Limitations of experiment are identified, and ways to improve are evident.	• Conclusions are drawn from the findings but less significant to objectives of the experiment. • Limitations of experiment are identified.	• Conclusions are not drawn from the findings and have no significance to objectives of the experiment. • Some limitations of experiment are identified.	• No valid conclusions drawn from the findings. • Limitations of experiment are not identified.	
				TOTAL SCORE	

Chapter-wise Weighting and Time allocation

Chapters	Chapter title	Maximum time required (mins)	Weighting (%)
Chapter 1	Periodic Table	729	15%
Chapter 2	Chemical Bonding	729	15%
Chapter 3	Metal Reactivity	486	10%
Chapter 4	Chemical Reactions,Conservation of Mass and Stoichiometry	486	10%
Chapter 5	Rate of Reaction and Energy Transfer	583	12%
Chapter 6	Green Chemistry	729	15%
Chapter 7	Organic Chemistry	1118	23%
	Total	4860	100%

The total time required to complete the topics is 4860 minutes or 96 periods of 45 minutes in a period.

Contents

INTRODUCTION TO CHEMISTRY

INTRODUCTION

The modern human experience places a large emphasis upon the material world. From the day of our birth to the day we die, we are frequently preoccupied with the world around us. Whether struggling to feed ourselves, occupying ourselves with modern inventions, interacting with other people or animals or simply mediating on the air we breathe, our attention is focused on different aspects of the material world. Everything else is somehow related to chemistry, the scientific discipline which studies the properties, composition and transformation of matter.

To put it simply, chemistry is a science that deals with how materials form, their composition, various attributes, how they transform and the energy they consume or release through the transformations.

Chemistry relates to each and everything you can sense, right down from the elemental level to massive complex structures. It deals with how things are constructed and formulated to how they break apart and the energy that is either absorbed or released in each of these processes.

This is a complete course in chemical stoichiometry, which is a set of tools, chemists use to count molecules and determine the amount of substances consumed and produced by reactions.

The course is set in a scenario that show how stoichiometry calculations are used in real world situations. Chemistry deals with composition, structure and properties of matter. These aspects can be described in terms of basic constituents of matter: atoms and molecules. That's why chemistry is called the science of atoms and molecules.

SCOPE IN CHEMISTRY

A qualification in chemistry opens doors to a wide range of careers. Chemistry is involved in our everyday lives and there is a vast range of jobs and careers open to those who have studied chemistry at any level. Great career opportunities exist both inside and outside the lab. Nobody knows what the jobs of the future will look like, but many of them will be involved in chemistry to solve global challenges such as human health, energy and the environment.

As well as practical knowledge of the subject, chemistry students develop many other skills prized by employers such as problem solving, numeracy, communication, creativity and data analysis. Gaining these skills means that you can have a future in all sorts of careers from finance to public relations.

The career options in chemistry are practically endless! However, your employment options depend on how far you have taken your education. A 3-year degree in chemistry won't get you very far. You could work in some labs washing glassware or assist at a school with lab preparation, but you wouldn't have much advancement potential and you could expect a high level of supervision.

BRANCHES OF CHEMISTRY

The five major branches of chemistry are organic, inorganic, analytical, physical and biochemistry. These divide into many sub-branches.

Organic Chemistry

Organic chemistry involves the study of the structures, properties and preparation of chemical compounds that consist primarily of carbon and hydrogen.

Organic chemistry overlaps with many areas including

1. Medicinal chemistry—the design, development and synthesis of pharmaceutical drugs. It overlaps with pharmacology and bioanalytical chemistry.
2. Organometallic chemistry— the study of chemical compounds containing bonds between carbon and a metal.
3. Polymer chemistry— the study of the chemistry of polymers.
4. Physical organic chemistry— the study of the interrelationship between structure and reactivity in organic molecules.
5. Stereochemistry— the study of the spatial arrangements of atoms in molecules and their effects on the chemical and physical properties of substances.

Inorganic Chemistry

Inorganic chemistry is the study of the properties and behaviour of inorganic compounds. It covers all chemical compounds except organic compounds.

Inorganic chemists study things such as crystal structures, minerals, metals, catalysts and most elements in the periodic table.

Branch of inorganic chemistry includes:

1. Bioinorganic chemistry, 2. Geochemistry, 3. Nuclear chemistry, 4. Organometallic chemistry

5. Solid-state chemistry

Analytical Chemistry

Analytical chemistry involves the qualitative and quantitative determination of the chemical compounds of a substance.

Examples of areas using analytical chemistry include:

1. Forensic chemistry, 2. Environmental chemistry, 3. Drug testing

Physical Chemistry

Physical chemistry is the study of the chemical structure on the physical properties of substance.

Physical chemistry is typically a study of the rate of a chemical reaction, the interaction of molecules with radiation, and the calculation of structures and properties.

Sub-branches of physical chemistry include:

1. Photochemistry, 2. Surface chemistry, 3. Chemical kinetics, 4. Quantum chemistry, 5. Spectroscopy

Biochemistry

Biochemistry is the study of biological processes in an attempt to explain them in chemical terms.

Biochemical research includes cancer and stem cell biology, infectious disease and cell membrane and structural biology.

It spans molecular biology, genetics, pharmacology, clinical biochemistry and agricultural biochemistry.

Thus, although there are five main branches of chemistry, there are many sub-branches.

METHODOLOGY OF CHEMISTRY

Each step of a synthesis involves a chemical reaction. Reagents and conditions for each of these reactions must be designed to give an adequate yield of pure product, with as little work as possible. A method may already exist in the literature for making one of the early synthetic intermediates and this method will usually be used rather than an effort to reinvent the wheel. However, most intermediates are compounds that have never been made before, and these will normally be made using general methods developed by methodology researches. To be useful, these methods need to give high yields and to be reliable for a broad range of substrates. For practical applications, additional hurdles include industrial standards of safety and purity. Methodology research usually involves three main stages–discovery, optimisation, and studies of scopes and limitations. The discovery requires extensive knowledge of and experience with chemical reactivities of appropriate reagents. Optimisation is a process in which one or two starting compounds are tested in the reaction under a wide variety of conditions of temperature, solvent, reaction time, etc., until the optimum conditions for product yield and purity are found. Finally, the researcher tries to extend the method to a broad range of different starting materials, to find the scope and limitations. Total synthesis are sometimes used to showcase the new methodology and demonstrate its value in a real-world application. Such applications involve major industries focused especially on polymers and pharmaceuticals.

IMPORTANCE OF CHEMISTRY

Chemistry plays a vital role in science and it is often intertwined with other branches of science. Some of the important applications of chemistry are given below:

1. It has provided chemical fertilizers, pesticides, fungicides and weedicides.
2. It has helped to protect the crops by use of certain insecticides, fungicides and pesticides.
3. It has helped to preserve food by use of preservatives.
4. It has given method to test the presence of adulterants.
5. Many life saving drugs such as **cisplatin** and **taxol** are effective in cancer therapy.
6. **AZT** (azidothymidine) used for helping AIDS victims, have been isolated from plants and animals or prepared by synthetic methods.
7. Discovery of anaesthetics has made surgical operation more and more successful.
8. Synthetic vitamins and tonics have contributed towards good health.
9. It has lead to the production of superconducting ceramics, conducting polymers and optical fibres.

10. Refrigerant like Chlorofluorocarbons (CFCs) which destroys the ozone layer have been replaced by environment friendly chemicals.
11. Chemistry has helped in the production of synthetic fibres, building materials, domestic articles (detergents, oils, fats, sugar, paints, varnish, cosmetics, perfumes, cooking gas, etc.)
12. Trinitrotoluene (TNT), nitroglycerine, dynamite, mustard gas, phosgene, etc. are used in wars.
13. Water (H_2O) is composed of hydrogen and oxygen atoms and it is considered as one of the essentials of life.

Web Links:

For more information on the following topic, visit the website listed below.
https://www.thoughtco.com/what-is-the-importance-of-chemistry-604143

CHAPTER 1

PERIODIC TABLE

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Understand the organisation of the Periodic Table
  - Explain that the Periodic Table shows all the elements, arranged in order of ascending atomic number.
  - Identify that each element has a specific number of protons in the nucleus.
  - Explain the connection between the arrangement of outer electrons and the position of an element in the Periodic Table and predict the group of the given elements.
  - Explain that elements in the same group of the Periodic Table have similar properties.
  - Explain periodic properties and their variations across the period and down the group.
- Know Group 1 – the Alkali Metals Li, Na and K
  - State the physical properties for example, melting points and boiling points of the alkali metals and describe the changes in these properties as the order in group descends.
  - Describe the reactions of the alkali metals, Li, Na and K with water, oxygen and chlorine and write balanced chemical equations for each reaction.
  - Investigate the reactions of the alkali metals with water, oxygen and chlorine to describe the trends in reactivity down the group.
- Know Group 18 – the Noble Gases
  - State the physical properties for example, boiling points and melting points of the Noble Gases.
  - Explain the changes in physical properties of the Noble Gases.
  - State uses of the Noble Gases based on their properties.

1.1 INTRODUCTION

You must have visited a library. There are thousands of books in a large library. In spite of this, if you ask for a particular book, the library staff can locate it easily. How is it possible? In library the books are classified into various categories and sub-categories. Therefore, locating the exact position of books becomes easy.

Same is the story with the chemical elements. A large number of elements and compounds are known today. But a systematic classification of these elements has made their study possible and easy.

The well organised and tabulated classification of elements as we know today is called the periodic table. It not only helps to locate, identify and characterise the element and its properties but also point out the directions in which new investigation is made.

1.1.1 Mendeleev's Periodic Table

During the same time as of Lothar Meyer's, a Russian scientist Dmitri Mendeleev put forward the periodic law which forms the basis of modern classification of elements in the form of a table, known as **periodic table**. In 1872, he published a periodic law which stated that "*the physical and chemical properties of the elements are periodic functions of their atomic weights*".

Mendeleev's Periodic Table consists of six horizontal rows called **periods**, and eight vertical columns known as **groups**.

TABLE 1.1 Mendeleev's Periodic Table

Group ⇒ Periods	I A	I B	II A	II B	III A	III B	IV A	IV B	V A	V B	VI A	VI B	VII A	VII B	VIII Transition series
1	H 1.008														
2	Li 6.919		Be 9.012		B 10.81		C 12.011		N 14.007		O 15.999		F 18.998		
3	Na 22.99		Mg 24.31		Al 26.98		Si 28.09		P 30.974		S 32.06		Cl 35.453		
4.	K 39.102	Cu 63.54	Ca 40.08	Zn 65.37	Ga 69.72	Sc 44.96	Ge 72.59	Ti 47.90	As 74.92	V 50.94	Se 78.96	Cr 52.00	Br 79.909	Mn 54.94	Fe 55.85 Co 58.93 Ni 58.71
5.	Rb 85.47	Ag 107.87	Sr 87.62	Cd 112.40	In 114.82	Y 85.91	Sn 118.69	Zr 91.22	Sb 121.75	Nb 92.91	Te 127.60	Mo 95.94	I 126.90	Te 99	Ru 101.07 Rh 102.91 Pd 106.4
6	Cs 132.90	Au 196.97	Ba 137.34	Hg 200.59	Tl 204.37	La 138.91	Pb 207.19	Hf 178.49	Bi 208.98	Ta 180.95	Po (210)	W 183.85	At 210	Re 186.2	Os 190.2 Ir 192.2 Pt 195.09

1.2 MODERN PERIODIC TABLE OR LONG FORM OF THE PERIODIC TABLE

Mendeleev's Periodic Table is based on the atomic weights. As the properties of an element depends on the number of electrons and their arrangement in various energy shells of the atom, the atomic number can thus be considered as the fundamental property of an element. The Modern periodic table based on atomic number was given by Henry Moseley in 1913.

The Modern periodic table thus classified elements based on their electronic configurations.

1.2.1 Characteristics of Periods

1. There are seven periods in Modern periodic table.
2. Each period starts with an alkali metal and finishes at an inert gas, except first period.
3. Except first period, the number of electrons increases from 1 to 8 in the outermost shell of elements in a period from left to right.
4. Elements of third period (group 1 to 2 and 13 to 17) are called typical elements because these elements have typical properties of their respective family.
5. The valency of the elements with respect to hydrogen and chlorine increases from 1 to 4 and then decreases to zero.

1.2.2 Characteristics of Groups

1. Elements of group 1, 2 and 13 to 17 are called normal or typical or representative elements.
2. Elements of group 3 to 12 are called transition elements.
3. Elements of a sub-group show similar properties.
4. All the elements of a sub-group have identical electronic configurations.
5. Elements of group 18 are called noble gases.
6. In representative elements, the valence is generally equal to either number of valence electrons or equal to eight minus the number of valence electrons.

In the modern periodic table, all the elements are arranged according to the increasing order of their atomic numbers. The horizontal rows are called **'periods'** and vertical columns are called **'groups'**. A modern periodic law may be stated as follows — *the physical and chemical properties of the elements are the periodic functions of their atomic numbers.* There are many forms of the periodic table — the long form is the most convenient and the most widely used one. In the long form of periodic table, there are seven periods. The first period contains 2 elements. The subsequent periods consist of 8, 8, 18, 18 and 32 elements. The seventh period is incomplete.

According to the latest recommendations of the International Union of Pure and Applied Chemistry (IUPAC), the groups are numbered from 1 to 18 and there is no sub-division of a group till date. The Periodic Table based on the electronic configuration of the elements is called the long form of the periodic table. This is also called Bohr's periodic table as it follows Bohr's scheme for the arrangement of various electrons around the nucleus.

The present form of the periodic table is called the **extended form of the periodic table** or the **long form of the periodic table**. In this table, there are 7 periods and 18 groups (Table 1.2).

For elements with no stable isotopes, the mass number of the isotope with the longest half-life is in parentheses.

Period	No. of elements
I	2
II	8
III	8
IV	18
V	18
VI	32
VII	26

1.2.3 Advantages of the Modern Periodic Table or Long Form of the Periodic Table

1. This table is based on the atomic numbers of elements. The atomic number of an element is a fundamental property.
2. It correlates the position of the elements with their electronic configurations more clearly.
3. There is a complete segregation of metals and non-metals.
4. There is a gradual change in the properties of elements with increase in their atomic numbers, i.e. the periodicity can be easily visualised.
5. The properties of new elements can be easily predicted even before their actual discoveries.
6. The arrangement of elements is easy to remember and reproduce.

TABLE 1.2 Modern Periodic Table

	IA 1	IIA 2	IIIB 3	IVB 4	VB 5	VIB 6	VIIB 7	VIII 8	VIII 9	VIII 10	IB 11	IIB 12	IIIA 13	IVA 14	VA 15	VIA 16	VIIA 17	Zero 18
1.	1 **H** Hydrogen 1.008																	2 **He** Helium 4.003
2.	3 **Li** Lithium 6.941	4 **Be** Beryllium 9.012											5 **B** Boron 10.81	6 **C** Carbon 12.01	7 **N** Nitrogen 14.01	8 **O** Oxygen 16.00	9 **F** Fluorine 19.00	10 **Ne** Neon 20.18
3.	11 **Na** Sodium 22.99	12 **Mg** Magnesium 24.31	TRANSITION ELEMENTS										13 **Al** Aluminium 26.98	14 **Si** Silicon 28.09	15 **P** Phosphorous 30.97	16 **S** Sulphur 32.07	17 **Cl** Chlorine 35.45	18 **Ar** Argon 39.95
4.	19 **K** Potassium 39.10	20 **Ca** Calcium 40.08	21 **Sc** Scandium 44.96	22 **Ti** Titanium 47.90	23 **V** Vanadium 50.94	24 **Cr** Chromium 52.00	25 **Mn** Manganese 54.94	26 **Fe** Iron 55.85	27 **Co** Cobalt 58.93	28 **Ni** Nickel 58.70	29 **Cu** Copper 63.55	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.72	32 **Ge** Germanium 72.59	33 **As** Arsenic 74.92	34 **Se** Selenium 78.96	35 **Br** Bromine 79.90	36 **Kr** Krypton 83.80
5.	37 **Rb** Rubidium 85.47	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.91	40 **Zr** Zirconium 91.22	41 **Nb** Niobium 92.91	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium 98.91	44 **Ru** Ruthenium 101.1	45 **Rh** Rhodium 102.9	46 **Pd** Palladium 106.4	47 **Ag** Silver 107.9	48 **Cd** Cadmium 112.4	49 **In** Indium 114.8	50 **Sn** Tin 118.7	51 **Sb** Antimony 121.8	52 **Te** Tellurium 127.6	53 **I** Iodine 126.9	54 **Xe** Xenon 131.3
6.	55 **Cs** Caesium 132.9	56 **Ba** Barium 137.3	57 ***La** Lanthanum 138.9	72 **Hf** Hafnium 178.5	73 **Ta** Tantalum 180.9	74 **W** Tungsten 183.9	75 **Re** Rhenium 186.2	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.2	78 **Pt** Platinum 195.1	79 **Au** Gold 197.0	80 **Hg** Mercury 200.6	81 **Tl** Thallium 204.4	82 **Pb** Lead 207.2	83 **Bi** Bismuth 209.0	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7.	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **†Ac** Actinium (227)	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (263)	107 **Bh** Bohrium (262)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadtium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Uut** Ununtrium (284)	114 **Fl** Flerovium (289)	115 **Uup** Ununpentium (288)	116 **Lv** Livermorium (293)	117 **Uus** Ununseptium (294)	118 **Uuo** Ununoctium (294)

6.	*** Lanthanoid Series**	58 **Ce** Cerium 140.1	59 **Pr** Praseodymium 140.9	60 **Nd** Neodymium 144.2	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.4	63 **Eu** Europium 152.0	64 **Gd** Gadolinium 157.3	65 **Tb** Terbium 158.9	66 **Dy** Dysprosium 162.5	67 **Ho** Holmium 164.9	68 **Er** Erbium 167.3	69 **Tm** Thulium 168.9	70 **Yb** Ytterbium 173.0	71 **Lu** Lutetium 175.0
7.	**† Actinoid Series**	90 **Th** Thorium 232.0	91 **Pa** Protactinium (231)	92 **U** Uranium 238.0	93 **Np** Neptunium 237	94 **Pu** Plutonium (242)	95 **Am** Americium (244)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

1.2.4 Disadvantages of Modern Periodic Table or Long Form of the Periodic Table

1. The position of hydrogen is not fixed in modern periodic table.
2. Lanthanoids and actinoids have not been accommodated in the main body of this modern periodic table.
3. Position of helium is not properly placed in the periodic table. Helium should have been placed in *s*-block element, but it is in *p*-block.

1.2.5 Short Description of the Modern Periodic Table

1. **Normal Elements:** The elements constituting groups 1, 2 and 13 to 18 are called **normal** or **representative elements**. They are so called because the electrons in their valence shells are filled in accordance with the normal rules for filling of electrons in the orbitals.
2. **Typical Elements:** The elements of second and third periods are called **typical elements** as they summarise the properties of their respective groups, e.g. Li, Na in group 1 and Be and Mg in group 2.
3. **Transition Elements:** The elements of groups 3 to 12 are known as **transition elements**. These are so named because there is a transition from highly electropositive elements of groups 1 and 2 to highly electronegative elements of group 17.
4. **Lanthanoids or Lanthanides and Actinoids or Actinides:** These are also called **inner transition elements** as they lie inside the transition elements actually.

Some important families of the periodic table are:

(*i*) Alkali metals (elements of group 1)

(*ii*) Alkaline earth metals (elements of group 2)

(*iii*) Coinage metals (elements of group 11)

(*iv*) Halogen (elements of group 17)

(*vii*) Noble gases (elements of group 18 or zero group).

1.2.6 Characteristics of Periods and Groups

The long form of the periodic table consists of eighteen vertical columns and seven horizontal rows. They have been obtained by arranging the elements in increasing order of atomic number in such a way that the elements with similar electronic configuration are placed, under each other in the same vertical column. These vertical columns are known as groups and the horizontal rows are called periods.

A **1.1**

Divide the class into the groups of the periodic table and discuss the various properties and characteristics and the variation down the group. Repeat the same by dividing the class into the periods of a periodic table.

SELF EVALUATION

I. After having a close look at the Mendeleev's periodic table, answer the following questions.

(*i*) What is the basis of the classification of elements?

(*ii*) How many groups and periods are present in the periodic table?

(*iii*) Name the elements present in group 1B.

II. According to modern periodic law the physical and chemical properties of elements depend on the atomic numbers. Modern periodic law is the basis of the classification of the elements into groups and periods. There are 7 periods and 18 groups in the long form of periodic table.

(*i*) What is the general name of the element in group 17?

(*ii*) What are the significance of modern periodic table?

(*iii*) Write the importance of the classification of the elements.

III. Matching

Match the items of Column I with the corresponding items of Column II:

Column I	Column II
1. Typical element	(*a*) coinage elements
2. Alkaline earth metals	(*b*) Li, Na in group 1 and Be, Mg in group 2.
3. Position of hydrogen	(*c*) group 2
4. Noble gases	(*d*) zero group
5. Elements of group 11	(*e*) is not fixed in modern periodic table

1.3 TRENDS IN THE MODERN PERIODIC TABLE

Periodicity: The repetition of the similar properties of the elements after a regular interval of time in the periodic table, is known as periodicity.

Cause of Periodicity: Repetition of the same valence shell electronic configuration is the cause of periodicity.

Periodic Properties: Important properties which exhibit periodicity are melting point, boiling point, density, atomic radius, valency, oxidation states, electronegativity, etc. Properties such as specific heat, colour, refractive index are not periodic in nature.

1.3.1 Valence Electron

The electrons present in the valence shell or outermost shell are called **valence electrons**. These electrons determine the valency of the atom.

Valency: The combining capacity of an element to complete its octet (8 electrons in the outermost shell) is called its valency. Usually, it is found with reference to oxygen, hydrogen or chlorine.

Variation in a Period: The number of valence electrons increases in a period. In normal elements, the valency increases from 1 to 8 across a period from left to right. The first element in each period has 1 valence electron and the last element has 8 valence electrons i.e., it reaches 8 in group 18 elements which practically show no chemical activity under normal conditions and their valency is taken as zero. The valency of the elements with respect to hydrogen and chlorine increases from 1 to 4 and then decreases from 4 to 1. Valency of normal elements with respect to oxygen increases from 1 to 7 as shown below for elements of third period, the valency is equal to the number of valence electrons or group number for group 1 and 2 or groups 13 to 17.

The group number or the number of electrons in the last orbit gives the valency of the element. Valency with respect to hydrogen increases from 1 to 4 and then decreases to 0 in a period, but with respect to oxygen, it increases from 1 to 7.

Variation in a Group: On moving down a group, the number of valence electrons remains same and, therefore, all the elements in a group exhibit same valency. For example, all the elements of group 1 (alkali metals) have valency equal to one and those of group 2 (alkaline earth metals) have valency equal to two.

Group	1	2	13	14	15	16	17
Elements	Na	Mg	Al	Si	P	S	Cl
Valency	1	2	3	4	5	6	7
Valency w.r.t. O	1	2	3	4	5	6	7

Group	1	2	13	14	15	16	17
Elements	Li	Be	B	C	N	O	F
Valency	1	2	3	4	5	6	7
Valency w.r.t. H	1	2	3	4	3	2	1

1.3.2 Atomic Size

The **atomic size** may be defined as the distance between the centre of the nucleus and the outermost shell of an isolated atom.

Variation in a Period: In a given period, atomic radius generally decreases from left to right because nuclear charge increases. For example, in second period the atomic radii are in the order:

$$\text{Li} > \text{Be} > \text{B} > \text{C} > \text{N} = \text{O} > \text{F} < \text{Ne}$$

Variation in a Group: The atomic radius increases from top to bottom in a group because the number of shells increases. Thus shielding or screening effect of inner electrons also increases, and hence effective nuclear charge decreases. For example, in the group of alkali metals, the atomic size increases from Li to Cs, i.e.

$$r\text{Li} < r\text{Na} < r\text{K} < r\text{Rb} < r\text{Cs}$$

(where r is the atomic radius)

Screening Effect: The fact that the valence electrons are attracted less by the nucleus due to the intervention of electrons in the inner shells. This effect produced by the electrons in the inner shell is called screening effect or shielding effect.

1.3.3 Metallic Character or Electropositive Character

The tendency of atoms of an element to lose electrons and form positive ion is called **metallic** or **electropositive character**.

Variation in a Period: When moving across the period in the periodic table from left to right, the electropositive or metallic character decreases. This is because, due to decrease in the size of atom, electron releasing tendency decreases. For example, in the second period, lithium and beryllium are metals. Boron is a semi-metal whereas carbon, nitrogen, oxygen and fluorine are non-metals.

Variation in a Group: When moving down the group the electropositive character increases. This is due to the increase in atomic size while going down a group, the electron releasing tendency increases.

1.3.4 Atomic Volume

The **atomic volume** is defined as the volume occupied by one gram atom of an element. Mathematically,

$$\text{Atomic volume} = \frac{\text{Gram atomic weight}}{\text{Density in solid state}}$$

Units of atomic volume is c.c./mole.

Lower atomic volume generally leads to higher density, increases hardness and brittleness, high melting and boiling points, less malleability and ductility.

Variation in a Period: We do not observe a regular trend in the value of atomic volume in a period. The value of atomic volume depends upon the following factors:

(*i*) *Nuclear Charge:* Atomic volume decreases on moving from left to right in a period. This is because nuclear charge increases from left to right, whereas the number of shells remains the same. Thus, the outer electrons are attracted with greater force on moving across the period from left to right. This results in the decrease of atomic size in the period from left to right. This results in the decrease of atomic volume.

(*ii*) *Number of Valence Electrons:* The orbitals are completely filled towards the end of the period. A large number of electrons in these orbitals cause repulsion between them leading to increase in atomic volume.

The net result of these factors is that atomic volume first decreases, becomes minimum in the middle of the period and then increases.

Variation in a Group: On moving down a group, addition of new shells takes place. Hence, the atomic volume increases down the group. Greater the number of shells, greater is the atomic volume.

1.3.5 Ionization Enthalpy or Ionization Energy or Ionization Potential

The minimum energy required to remove the outermost electron from neutral gaseous atom to form ion is called the **ionization enthalpy** (or **ionization energy**). This process is written as:

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

where X (g) and $X^+(g)$ represent the gaseous atom and the gaseous cation respectively.

We have taken gas phase because in the gas phase, the atom and the ion are isolated from all external influences.

Ionization enthalpy is expressed in kJ mol^{-1} or kcal mol^{-1}. When expressed in eV per atom it is called ionization potential. Ionisation energy is a measure of force of attraction between the nucleus and the outermost electron. Stronger the attraction, greater is the value of ionisation energy.

If only one electron is removed, the ionisation energy is known as first ionization energy. If second electron is removed the ionization energy is known as second ionization energy and so on.

Variation in a Period: On moving across the period from left to right, atomic size decreases due to increase in force of attraction between the nucleus and valence electron. Thus, ionization energy increases across the period.

Variation in a Group: As we go down a group, atomic size goes on increasing and hence the force of attraction between the nucleus and the valence electrons goes on decreasing. As a consequence, the ionization energy goes on decreasing down the group.

1.3.6 Electron Affinity or Electron Gain Enthalpy

Electron gain enthalpy *or* **electron affinity** of an element may be defined as the energy released when a neutral isolated gaseous atom accepts an extra electron to form the gaseous negative ion, i.e. anion.

This process may be represented as:

$$\underset{\text{Neutral gaseous atom}}{X(g) + e^-} \longrightarrow \underset{\text{Anion}}{X^-(g)} + \text{Energy}$$

Evidently, greater the amount of energy released in the above process, higher is the electron affinity of the element.

Electron affinity values are influenced by (*a*) size of the atom (*b*) atomic number (*c*) screening effect (*d*) electronic configuration.

Thus, electron affinity is a measure of tightness of binding of an extra electron to an atom. Electron affinity depends upon the size and effective nuclear charge.

Variation in a Period: As we move across a period from left to right, the atomic size decreases and the nuclear charge increases. Both these factors tend to increase the attraction by the nucleus for the incoming electron and hence electron affinity, in general increases in a period from left to right.

Variation in a Group: As we move down a group, both the atomic size and the nuclear charge increase. But the effect of the increase in atomic size is much more pronounced than the nuclear charge. Thus, with increase in atomic size, the attraction of the nucleus for the incoming electron decreases and, hence, the electron affinity decreases. For example, the electron affinity of halogens decreases in going from chlorine to iodine.

The electron affinities of some of the elements of second period, e.g., N, O and F are, however, lower than those of the corresponding elements (i.e. P, S and Cl) of the third period.

This anomalous behaviour can be explained by saying that the elements of second period have smallest atomic size amongst the elements in their groups. As a result, there are considerable electron-electron repulsions within the atom itself and, hence, the additional electron is not accepted easily. For example, electron affinity of N is essentially zero while that of P is 74.3 kJ mol^{-1}. Similarly, the electron affinity of O (140.9 kJ mol^{-1}) is lower than that of S (200.7 kJ mol^{-1}) and the electron affinity of F (322.6 kJ mol^{-1}) is lower than that of Cl (348.5 kJ mol^{-1}). It may be noted that **chlorine has the highest electron affinity**.

1.3.7 Electronegativity

The relative tendency (or ability) of an atom of an element to attract a bonded pair of electrons to itself in a molecule is called **electronegativity**. It is dimensionless quantity and does not, have any unit. It is a relative property.

Variation in a Period: As we move across a period from left to right, the atomic size goes on decreasing, the valence electrons come nearer to the nucleus. The attraction between the nucleus and valence electrons goes on increasing. So also the tendency of the nucleus and, hence the atom to attract, bonded pair of electrons goes on increasing. Hence, electronegativity goes on increasing from left to right i.e., smaller the atomic size greater the electronegativity.

Elements	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0

Variation in a Group: It is seen that electronegativity goes on decreasing down a group. Atomic size goes on increasing down the group. Due to increase in atomic size down the group, the valence electrons are further away from the nucleus. The attraction between the nucleus and valence electrons goes on decreasing. So, tendency of the nucleus and hence the atom to attract the bonded pair of electrons goes on decreasing and hence electronegativity decreases on moving down the group.

Elements	Li	Na	K	Rb
Electronegativity	1.0	0.9	0.8	0.8
Elements	F	Cl	Br	I
Electronegativity	4.0	3.0	2.8	2.5

SELF EVALUATION

I. Among the elements of the second period Li to Ne pick out the element:

(*a*) with the highest first ionization energy.
(*b*) with the highest electronegativity.
(*c*) with the largest atomic radius.
(*d*) the most reactive non-metal.
(*e*) the most reactive metal.

II. Which of the following pairs would have a larger size? Explain your answer.

(*a*) K or K^+ (*b*) Br or Br^- (*c*) O^{2-} or F^- (*d*) Li^+ or Na^+

III. The elements, Na, Mg, Al, Si, P, S, Cl and Ar are arranged in the periodic table in the increasing order of their atomic numbers:

(*i*) What is meant by atomic number?
(*ii*) Which element is most electronegative?
(*iii*) Which element is most electropositive?
(*iv*) Which element is least reactive?
(*v*) Which elements are gases at room temperature?

1.4 GROUP-1 ELEMENTS–THE ALKALI METALS

The elements in the first group of the periodic table are: *lithium* (Li), *sodium* (Na), *potassium* (K), *rubidium* (Rb), *caesium* (Cs) and *francium* (Fr). All these elements are called **alkali metals**, because their hydroxides are strong bases or alkalies. The last element of this group, francium, is radioactive and has a very short lifespan. Therefore, not much is known about it. All the alkali metals have one electron in their valency shell. The alkali metals show a marked similarity and their properties change in a graded manner as discussed below. Though hydrogen (H) is listed in Group 1 because of its electronic configuration, but it is not considered as an alkali metal as it rarely shows any similar behaviour.

Electronic Configuration: The general electronic configuration of alkali metals are given in Table 1.3.

TABLE 1.3 Electronic Configuration of Alkali Metals

Elements	Atomic Number	Electronic Configuration
Lithium (Li)	3	2, 1
Sodium (Na)	11	2, 8, 1
Potassium (K)	19	2, 8, 8, 1
Rubidium (Rb)	37	2, 8, 18, 8, 1
Caesium (Cs)	55	2, 8, 18, 18, 8, 1
Francium (Fr)	87	2, 8, 18, 32, 18, 8, 1

1.5 VARIATION IN PROPERTIES OF ALKALI METALS DOWN THE GROUP

1.5.1 Physical Properties

The trends observed in the variation of different physical properties are discussed as under:

1. **Density:** The densities of alkali metals are low and increase down the group (Table 1.4).

   Explanation: *The low density of alkali metals is due to their large atomic size. However, in going down the group atomic size as well as atomic mass increase but the mass is not neutralised by the increase in atomic volume.*

TABLE 1.4 **Densities of Alkali Elements**

Elements	Electronic Configuration	Melting Points	Density (g/cm^3)	Atomic Radius
Lithium (Li)	2, 1	181	0.53	1.52
Sodium (Na)	2, 8, 1	98	0.97	1.56
Potassium (K)	2, 8, 8, 1	63	0.86	2.27
Rubidium (Rb)	2, 8, 18, 8, 1	39	1.53	2.47
Caesium (Cs)	2, 8, 18, 18, 8, 1	25	1.88	2.65

2. **Melting Points and Boiling Points:** All alkali metals have large size, which increases down the group as a result of which the intermetallic bonds in them are quite weak. Hence they have low melting and boiling points, which decrease down the group with the increase in the atomic size.

   Explanation: *The low melting points are ascribed to their large atomic size, due to which the binding energies of their atoms in the crystal lattice are low. In moving down the group, atomic size increases and the strength of metallic bond decreases which results in decrease in melting points.* The boiling points of the members of this group also follow the same order due to similar reasons.

3. **Atomic and Ionic Radii:** The atoms of alkali metals have the biggest size in their respective periods. The atomic radii increase on moving down the group of alkali metals.

   Explanation: *On moving down the group, there is a progressive addition of new energy shells, which are situated farther from the nucleus. Although increased nuclear charge attracts the outer electrons with great force, the atomic size dominates due to increase in the number of shells.*

   Alkali metals change into positively charged ions by losing their lone valence electron. The radii of positive ion is smaller as compared to that of parent atom. Ionic radius increases with increase in atomic number down the group.

   Na^+ cation is smaller than Na atom because outer shell (the third shell) has been removed from the atom.

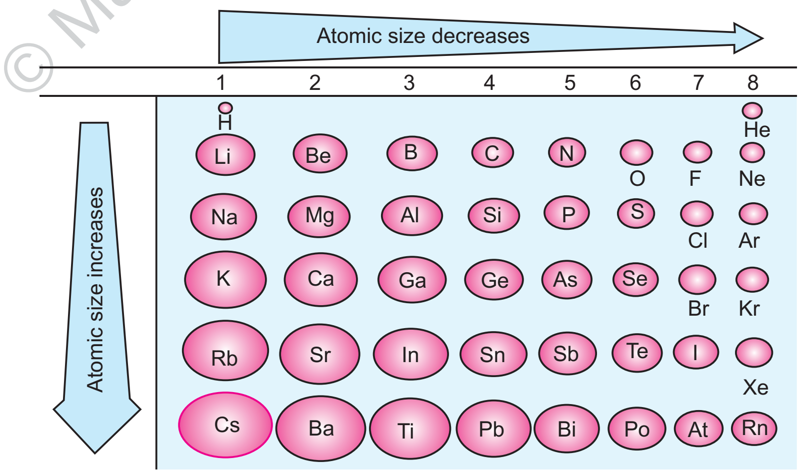


Fig. 1.1 *Atomic and ionic radii*

4. **Ionisation Enthalpy or Ionisation Energy and Electropositive Character:** Ionisation enthalpy of an element is the energy needed to remove the most loosely bound electron from the outermost orbit of the atom in the gaseous state. The first ionisation enthalpies of alkali metals are quite low as compared with the elements of the other group belonging to the same period. However, within the group, the ionisation enthalpies of the alkali metals decrease down the group.

   Explanation: *The elements of alkali metals have the largest atomic size in their respective periods, therefore, the outermost electrons which are far away from the nucleus experience a smaller force of attraction with the nucleus and, hence, can be easily removed.*

   Decrease in ionisation enthalpy (energy) on moving down the group is ascribed to increase in size of the atoms of alkali metals and increase in the magnitude of screening effect by virtue of increase in number of intervening electrons. This will reduce progressively the force of attraction between the nucleus and the outermost electron, resulting in decrease in ionisation enthalpy (I.E.).

   Due to low ionisation enthalpies, these metals have a strong tendency to lose their valence electrons and thus, change into positive ions. Consequently, **alkali metals are strongly electropositive or metallic in character.** As this tendency for losing electrons increases down the group, the electropositive character increases.

5. **Oxidation State:** The alkali metals have a strong tendency to lose their valence electron and change into unipositive ions. Thus, alkali metals exhibit +1 oxidation state in their compounds.

   Explanation: *After losing the valence electron, these metals acquire the stable configuration of a noble gas, therefore, they have no tendency to form the M^{2+} ions. The magnitude of second ionisation energy is very high and, therefore, the second electron is seldom lost.*

6. **Flame Colouration:** All the elements impart characteristic colour to Bunsen flame as described below:

Li	Na	K	Rb	Cs
Red	Golden yellow	Pale violet	Violet	Blue

   Explanation: *When energy is supplied to these elements in their salts, the electrons in their valence shells get excited to higher energy states. When these electrons fall back to their original states they emit energy in the form of radiations which falls in the visible region, imparting a characteristic colouration to the flame.*

7. **Photoelectric Effect:** Alkali metals (except Li) show the phenomenon of photoelectric effect (Fig. 1.2). (A phenomenon of emission of electrons from metal surface by exposing it to visible light).

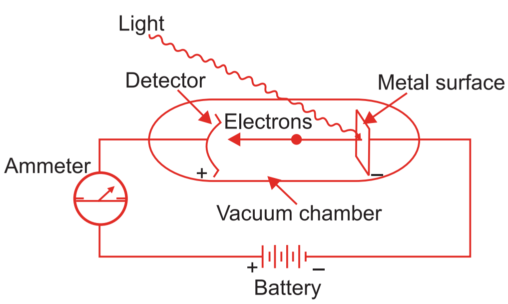


Fig. 1.2 *Explanation of photoelectric effect*

A 1.2

Flame Tests for Li, Na, K.

Clean a platinum or nichrome wire by dipping it into concentrated hydrochloric acid, and then holding it in a hot Bunsen flame. Repeat this until the wire doesn't produce any colour in the flame. When the wire is clean, moisten it again with some of the acids and then dip it into a small amount, of the solid you are testing so that some sticks to the wire. Place the wire back in the flame again. If the flame colour is weak, it is often worthwhile to dip the wire back in the acids again, and put, it back, you often get a very short but intense flash of colour by doing that.

Fig. 1.3 *A yellow flame is obtained with sodium ions*

Note the colour and the corresponding element it represents.

Explanation: *The ability to exhibit photoelectric effect is attributed to low values of their ionisation energies. Lithium does not emit photoelectrons when exposed to light due to high I.E. value. Other elements of the alkali group show this property because of low ionisation energies. When the metal surface is exposed to radiations, electrons are easily ejected out due to low ionisation energy of the elements.*

SELF EVALUATION

I. Explain why alkali metals are good reducing agents.

II. Explain why:

(*i*) Potassium is more reactive than sodium.

(*ii*) Alkali metals do not form divalent ions.

III. Explain the following:

(*i*) Alkali metals are used in photoelectric cells.

(*ii*) Each alkali metal impart a characteristic colour to the flame.

1.5.2 Chemical Properties

The alkali metals are highly reactive elements. High chemical reactivity of alkali metals is due to *low values of first ionisation energy.*

Value of ionisation energy decreases down the group from lithium to caesium, therefore, **the reactivity of alkali metals increases from Li to Cs.** These elements are highly reactive especially towards the more electronegative elements such as chlorine and oxygen. The important chemical characteristics of alkali metals are discussed as under:

1. **Reaction with Hydrogen (Hydrides):** Alkali metals react with hydrogen to form colourless crystalline hydrides which are ionic in nature.

$$2M + H_2 \xrightarrow{\text{Heat}} 2M^+H^-$$

2. **Reaction with Oxygen (Oxides):** Alkali metals on combustion form different kinds of oxides. For example, the alkali metals on reaction with limited quantity of oxygen form normal oxides of the formula, M_2O.

$$2M + \frac{1}{2}O_2 \longrightarrow M_2O$$

(where M stands for Li, Na, K, Rb, Cs).

With excess of air, lithium forms *normal oxide*, Li_2O; sodium forms *peroxide*, Na_2O_2, while potassium, rubidium and caesium form *superoxides* having general formula MO_2. The reactions are given as under:

$$\underset{\text{Lithium}}{2Li} + \frac{1}{2}O_2 \longrightarrow \underset{\text{Lithium oxide}}{Li_2O}$$

$$\underset{\text{Sodium}}{2Na} + O_2 \longrightarrow \underset{\text{Sodium peroxide}}{Na_2O_2}$$

$$\underset{\text{Potassium}}{K} + O_2 \longrightarrow \underset{\text{Potassium superoxide}}{KO_2}$$

Explanation: *Small Li^+ ion has a strong positive field around it. When it combines with O^{2-} the electron cloud is strongly held by it and the spread of negative charge towards another oxygen is prevented. Thus, it does not form peroxide. It may be noticed that formation of peroxide would involve the transmission of negative charge from one* *oxygen to another. Na^+ ion being bigger in size has a weaker positive field around it. It does not have that much hold in the electron charge on O^{2-} and permits the negative charge to transmit further and combine with another oxygen resulting in the formation of peroxide, Na_2O_2.*

Still bigger K^+, Rb^+ ions have still weaker positive fields and, thus, the process does not stop at peroxides (O_2^{2-}). Here the formation of superoxides (O_2^-) like K_2O_2, RbO_2, etc., takes place.

3. **Reaction with Water:** Alkali metals react with water and other compounds containing acidic (active) hydrogen atoms such as hydrogen halides (HX) and acetylene (C_2H_2) and liberate hydrogen gas.

$$\underset{\text{Sodium}}{2Na} + 2H_2O \longrightarrow 2NaOH + H_2$$

$$\underset{\text{Sodium}}{2Na} + \underset{\text{acetylene}}{2HC \equiv CH} \longrightarrow \underset{\text{Sodium acetylide}}{2NaC \equiv CH} + H_2$$

$$\underset{\text{Sodium}}{2Na} + 2HX \longrightarrow 2NaX + H_2 \qquad \text{where X = halogen}$$

On moving down the group from lithium to caesium, the rate of the reaction increases and in the case of heavier alkali metals, the hydrogen gas liberated catches fire.

All the alkali metals on exposure to oxygen and moisture of the atmosphere form oxides and hydroxides and their surface gets tarnished. To prevent this, these metals are stored under kerosene oil.

4. **Reaction with Chlorine (Chloride):** Alkali metals combine with chlorine to form metal chloride. These are ionic crystalline solids having general formula M^+Cl^-.

$$2M + Cl_2 \longrightarrow 2M^+Cl^-$$

(where M is Li, Na, K).

Reactivity of alkali metals increases from Li to Cs with a particular halogen. Reactivity of halogens with particular alkali metal M decreases from F_2 to I_2.

All the halides of alkali metals are freely soluble in water with the exception of lithium fluoride. Exception of lithium fluoride can be explained on the basis of greater force of attraction between lithium ions and fluoride ions in the crystal lattice, i.e. lattice energy.

A 1.3

Objective: To study the reaction of alkali metals (Li, Na, K) with water.

Apparatus required: Trough, forceps.

Procedure: Take water in a trough and with the help of the forceps drop the small piece of alkali metal (Li, Na, K) in it. A gas will be evolved with a pop sound.

Observations:

1. A gas evolves with a pop sound.
2. Vigorous reaction takes place.
3. On moving down the group reactivity increases.

Result: Alkali metals such as lithium, sodium and potassium react with water to produce an alkali and hydrogen gas.

- $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$
- $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$
- $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

The reactivity of metals increases down Group 1.

Precautions:

1. Handle elements with care.
2. Do not carry out reactions close to your body.
3. Keep your face in a safe distance.

SELF EVALUATION

I. Give the answers of the following:

(*i*) Why does table salt get wet in rainy season?

(*ii*) Lithium halide is covalent in nature.

(*iii*) Li^+ ion is highly hydrated.

1.6 GROUP-18 ELEMENTS—THE NOBLE GASES

The elements **helium (He)**, **neon (Ne)**, **argon (Ar)**, **krypton (Kr)**, **xenon (Xe)** and **radon (Rn)** are the members of **Group-18** of the periodic table. All of them are gases under ordinary conditions of temperature and pressure. The last member of the group, i.e., radon is radioactive. They are also known as *rare* gases because they are found in very small quantities in nature. They are highly non-reactive and show no tendency to react with other elements and are therefore also called **inert gases.**

1.6.1 Occurrence

Because of their inert nature, noble gases always occur in free state. All the noble gases except radon are present in the air in small amounts. Helium is present in **natural gas** upto the extent of 10 per cent. It is also present in small quantities in radioactive minerals. Water from certain springs also contains rich quantities of helium and argon.

Helium atom contains two electrons in its outermost shell. While other noble gases contain 8 electrons in their outermost shell.

Since, atoms of noble gases have 8 electrons in their valence shell they do not have the tendency to lose or gain electrons. Thus, they do not participate in chemical reactions.

There is regular and smooth gradation in physical properties of this group elements.

1.7 PHYSICAL PROPERTIES OF NOBLE GASES

1. **Physical State:** All of them are monoatomic, colourless, odourless and tasteless gases.
2. **Boiling and Melting Points:** Due to weak intermolecular van der Waals' forces between them they possess very low melting and boiling points in comparison to those of other substances of comparable masses. However, melting and boiling points increase with the increase in atomic number because van der Waals' forces become stronger with increase in size of the atoms or molecules.
3. **Solubility:** They are sparingly soluble in water. The solubility increases with increase in atomic number.

1.7.1 Periodic Properties of Noble Gases

1. **Atomic Radii:** The atomic radii of noble gases are quite large as compared with atomic radii of the other atoms belonging to the same period. This is because of the completely filled octet. The electrons repel each other and their atomic radii increase. On moving down the group the atomic radii of noble gases increase due to addition of new shells.
2. **Electron Affinities:** Due to the stable electronic configurations, noble gas atoms show no tendency to accept additional electron. Therefore, their electron affinities are almost zero.
3. **Ionisation Enthalpy:** The ionisation enthalpies of noble gases are very high. This is due to the stable configurations of noble gases. However, the ionisation energy decreases with the increase in atomic number from He to Rn due to increase in atomic size and decrease in the force of attraction between the nucleus and outer electrons.
4. **Liquefaction:** It is extremely difficult to liquefy these gases as there are only weak van der Waals' forces which hold atoms together. Since these forces increase with increase in atomic size and population of electrons, *ease of liquefaction increases down the zero group from He to Rn.*

1.8 CHEMICAL PROPERTIES OF NOBLE GASES

The noble gases are chemically inert and do not participate in the reactions easily. Following factors are responsible for the chemical inactivity of noble gases, they have stable electronic configuration, low affinity for electrons and exceptionally high ionization enthalpies.

Prior to 1962, it was thought that noble gases do not combine at all with other elements. However, in 1962, it was observed by N. Bartlett that O_2 reacts with PtF_6 to form the compound $[O_2^+][PtF_6]^-$ [Oxygen hexafluoroplatinate (V)]. In this compound, O_2 is oxidized to O_2^+ by PtF_6. Since ionization energies of O_2 and xenon are comparable, he thought that a similar reaction should be possible with xenon. He tried the reaction between Xe and PtF_6 by direct interaction and obtained the first ever compound of noble gases.

$Xe^+ [PtF_6]^-$ in the form of red crystalline solid.

$$Xe + PtF_6 \longrightarrow \underset{\text{Xenon hexafluoroplatinate (V)}}{Xe^+[PtF_6]^-}$$

After that, more attempts were made to prepare more compounds of noble gases. Now many compounds of xenon and krypton are known with fluorine or oxygen. The compounds of krypton are fewer, only the difluoride, (KrF_2) has been studied in details. It has not been possible to prepare the compounds of He, Ne or Ar.

Other examples

$$Xe(g) + F_2(g) \xrightarrow{5\,atm,\ 400°C} XeF_2$$

$$XeF_2 + AsF_5 \longrightarrow [XeF]^+ [AsF_6]^-$$

$$6XeF_4(s) + 12H_2O(l) \longrightarrow 2XeO(aq) + 4Xe(g) + 5O_2(g) + 24HF(aq)$$

1.9 USES OF NOBLE GASES

A few important uses of noble gases are given below:

1. Helium:

(*i*) Due to its very low density and non-inflammable nature it is used for filling airships and weather balloons and for inflating the tyres of aeroplanes.

(*ii*) Unlike nitrogen, it is not soluble in blood even under high pressure. Due to this property, a mixture of helium and oxygen is used for respiration by deep-sea divers.

(*iii*) It is used for providing inert atmosphere in the welding of metal or alloys that are easily oxidised.

(*iv*) As its boiling point is low it is used for producing low temperature for research purposes.

2. Neon:

(*i*) On passing an electric discharge through a tube containing neon gas at low pressure, an orange red glow is produced. Due to this property neon is used in neon lamps and tubes to produce brilliant coloured glow. Neon tubes are used in advertising signs.

(*ii*) Neon lamps are used to serve as beacon lights for ships and planes, as it can penetrate through mist and fog.

3. Argon:

(*i*) Mixture of argon and nitrogen is used in gas filled electric lamps and also bulbs and tubes.

(*ii*) It is used in providing inert atmosphere during welding operations.

(*iii*) It is used in the production of titanium and in gas chromatography.

4. Krypton and Xenon:

(*i*) Krypton is used in miner's cap lamp.

(*ii*) Krypton and xenon are also used to a small extent in electric bulbs.

(*iii*) Xenon is used in photographic flash tubes.

5. Radon:

(*i*) Radon is used in radiotherapy for the treatment of cancer and other malignant growths.

(*ii*) The radioactive radon is used in place of X-rays in radiography.

SELF EVALUATION

I. Argon is a gas which was discovered by William Ramsay in 1895. He found that when he removed all the known gases from the air he was left with a small volume of unreactive gas. When all the elements were arranged in order of their atomic number, elements with similar properties to argon appeared in a periodic pattern. The group of similar elements included helium and neon.

(*i*) Explain what is meant by terms

(*a*) atomic number

(*b*) periodic pattern

(*ii*) (*a*) State one physical property and one chemical property shared by helium, argon and neon.

(*b*) What is the name given to this group of elements?

II. Can we use neon gas for advertising purpose?

III. Ugyen is a grade IX student. He wishes to celebrate his 14th birthday with his friends. For this he wanted to buy some balloons and fill them with gas lighter than air. The shopkeeper showed them two cylinders containing helium and hydrogen; which are used to fill balloons, hydrogen and helium cylinder. Out of these helium is more expensive.

(*i*) Which gas would Ugyen choose to fill balloons?

(*ii*) Give one more use of helium or hydrogen.

(*iii*) Why did Ugyen choose that particular gas?

RECAPITULATION

1. Classification of elements depends upon the similarities in their properties.
2. Mendeleev arranged the elements according to their chemical properties and in increasing order of their atomic weights.
3. Mendeleev even predicted the existence of some undiscovered elements on the basis of gaps in his periodic table.
4. Advantages of Mendeleev's Periodic Table:

   (*i*) Systematic study of elements became convenient.

   (*ii*) More of new elements were discovered, for example, Ga, Ge, Sc, etc.

(*iii*) Doubtful atomic weights of certain elements were corrected, for example, beryllium and indium.

(*iv*) Classification of elements on the basis of their fundamental property.

5. Anomalies in the arrangement of elements in order of their increasing mass could be removed when the elements were arranged on the basis of their increasing atomic number—a fundamental property of the element, discovered by Moseley.
6. In modern periodic table elements are arranged in 7 horizontal rows called periods and 18 vertical columns called groups. This arrangement of elements show periodicity of properties consisting atomic size, valency, or combining capacity, metallic and non-metallic characters.
7. Lanthanoids and actinoids are also called inner transition elements as they lie inside the transition elements actually.
8. The elements of group 3 to 12 are known as transition elements. These are so named because there is a transition from highly electropositive elements of group 1 and 2 to highly electronegative elements of group 17.
9. The combining capacity of an element is known as valency. The electrons present in valence shell are called valence electrons. These electrons determine the valency of the atom. The orbitals present in the valence shell are called valence orbitals.
10. Mendeleev's Periodic Table have vertical columns called groups and horizontal rows called periods.
11. Modern periodic law can be stated as "Properties of elements are the periodic functions of their atomic numbers." The atomic number tells us the number of protons in the nucleus of an atom and this number increases by one when we go from one element to the next.
12. Atomic size increases as we go down the group, because as we go down the group new shells are being added.
13. The metallic character across a period decreases and down a group increases. On the other hand non-metals are electronegative. They tend to form bonds by gaining electrons.
14. Mendeleev's Periodic Law states that 'the physical and chemical properties of the elements are the periodic functions of their atomic weights'.
15. Elements of second and third periods are called typical elements as they summarise the properties of their respective groups, e.g. Li, Na in group 1 and Be and Mg in group 2.
16. Elements constituting groups 1, 2 and 13 to 18 are called normal or representative elements because the electrons in their valence shells are filled in accordance with the normal rules for filling of electrons in the orbitals.
17. The distance between the centre of the nucleus and the outermost shell of an isolated atom is called atomic size.
18. The tendency of an atom in a molecule to attract shared pair of electrons and achieve the slightly negative charge is called electronegativity.
19. General electronic configuration of alkali metals is ns^1.
20. All alkali metals are highly reactive.
21. The densities of alkali metals are low and increase down the group.
22. All the alkali metals have low melting points. The melting points decrease down the group.
23. Atomic size of alkali metals increases down the group.
24. Smaller is the size of the ion greater the charge density and more degree of hydration.

25. Ionization enthalpy of an element is the energy needed to remove the most loosely bound electrons from the outermost orbit of the atom is the isolated gaseous state.
26. Alkali metals are strongly electropositive or metallic in nature.
27. Alkali metals impart characteristic flame colourations, i.e. Li imparts caramine red, Na imparts golden yellow, potassium imparts pale violet, Rb gives reddish violet while Cs imparts sky blue colour to the flame.
28. All the alkali metals are good reducing agents. Lithium is the strongest reducing agent because of its high hydration energy.
29. Except lithium all alkali metals show the photoelectric effect.
30. A phenomenon of emission of electrons from the metal surface by exposing it to visible light is called photoelectric effect.
31. Reactivity of alkali metals with hydrogen increases down the group.
32. The stability of hydrides of alkali decreases from Li to Cs.
33. Alkali metals also form complex hydrides such as $LiAlH_4$ and $NaBH_4$.
34. Alkali metal lithium form normal oxide, sodium form peroxide, while potassium form superoxide.
35. All the alkali metals on exposure to oxygen and moisture of the atmosphere form oxides and hydroxides.
36. Noble gases have very high I.E. and zero electron affinity.
37. Atomic radius of any noble gas elements is much higher than that of the preceding halogen (due to fully-filled configuration and high interelectronic repulsions).
38. The first noble gas compound obtained was $Xe^+ [PtF_6]^-$ by Bartlett in 1962, by the action of PtF_6 on Xe.
39. Noble gases (He, Ne, Ar, Kr, Xe, Rn) are also called inert gases or rare gases.
40. The melting points and boiling points of noble gases are very low due to weak van der Waals' forces of attraction but increase down the group from He to Rn as the size of the noble gas increases.
41. The more polarizable noble gases such as Kr and Xe form compounds with strong oxidising agents such as oxygen and fluorine.
42. Xenon forms three fluorides: XeF_2, XeF_4 and XeF_6
43. A mixture of helium 80% and oxygen (20%) is used by divers for artificial respiration.
44. Neon gas is used in sign boards since neon light can be seen from a long distance even in mist and fog.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. A modern periodic law may be stated 'as the physical and chemical properties of the elements are the periodic functions of their atomic number'.

2. Electronic configuration of phosphorus is 2, 8, 4.
3. Electron affinity of noble gases is zero.
4. Potassium hydroxide is a stronger alkali than sodium hydroxide.
5. Lithium is the most reactive alkali metal.

II. Fill in the Blanks

Fill in the blanks by using suitable word(s):

1. The electrons present in the outermost shell of an atom are called ____________.
2. The modern periodic table is based on the ____________ of elements.
3. Alkali metals are strong ____________ agents.
4. Mixture of argon and ____________ is used in gas filled electric lamps.
5. The colour of $Xe[PtF_6]$ is ____________.

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. Which of the following elements will form an acidic oxide?
   (*a*) An element with atomic number 7.
   (*b*) An element with atomic number 3.
   (*c*) An element with atomic number 12.
   (*d*) An element with atomic number 19.
2. Which of the following elements does not lose an electron easily?
   (*a*) Na (*b*) F (*c*) Mg (*d*) Al
3. The set representing correct order of ionic radius is
   (*a*) $Li^+ > Be^{2+} > Na^+ > Mg^{2+}$ (*b*) $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$
   (*c*) $Li^+ > Na^+ > Mg^{2+} > Be^{2+}$ (*d*) $Mg^{2+} > Be^{2+} > Li^+ > Na^+$
4. Lithium is the strongest reducing agent among alkali metals due to which of the following factor?
   (*a*) Low ionization energy (*b*) High electron affinity
   (*c*) Low hydration energy (*d*) High lattice energy
5. In flame test, which of the following is the colour of the flame by sodium salt?
   (*a*) Green (*b*) Red (*c*) Violet (*d*) Yellow
6. Which of the following is used for filling balloons?
   (*a*) Helium (*b*) Neon (*c*) Argon (*d*) Xenon
7. The radioactive noble gas is
   (*a*) argon (*b*) helium (*c*) radon (*d*) krypton

IV. Answer the Following Questions

1. Is lithium halide covalent in nature? Explain.
2. Why do alkali metals impart characteristic colour to flame?
3. Explain why alkali metals are good reducing agents.

4. The stability of peroxide and superoxide of alkali metals increases as we go down the group. Explain giving reason.
5. Account for the following:
   (*i*) Alkali metals show an oxidation state of +1 only.
   (*ii*) Degree of hydration of ions decreases from Li^+ to Cs^+.
   (*iii*) Beryllium forms many covalent compounds.
6. State modern periodic law. What is long form of periodic table? What are its main features? How is it superior to Mendeleev's table?
7. Which one of the following pairs of elements would have a higher electron affinity? Explain.
   (*i*) N or O (*ii*) F or Cl
8. The size of an anion is larger than that of its parent atom while that of cation is smaller than that of its parent atom. Explain.
9. Explain why the first ionization energy of Na is lower than that of Mg while the second ionization energy of Na is higher than that of Mg?
10. What prompted Bartlett to discover noble gas compound?
11. Explain the trends in chemical properties of noble gases.
12. Write the uses of the following:
   (*i*) Helium (*ii*) Argon (*iii*) Neon (*iv*) Radon

Web Links:

For more information on the following topics, visit the websites listed below:

Modern Periodic Table

http://byjus.com/chemistry/modern-periodic-table/

Group – 1 Element

https://chem.libretexts.org/Core/Inorganic Chemistry/Descriptive Chemistry/Elements Organized by Block 1 s-Block Elements/Group 1%3A The Alkali Metals

Group – 18 Elements

https://en.wikipedia.org/wiki/Noble gas

CHAPTER 2

CHEMICAL BONDING

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Describe the structure and bonding in metal elements.
- Explain duplet and octet rules.
- Explain the formation of ions when atoms gain or lose electrons to form ionic compounds.
- Explain that the ionic lattices are held together by the attraction between oppositely charged ions.
- Explain formation of covalent bonds when atoms share electrons.
- Explain that substances which have covalent bonding can be elements (e.g. H_2, O_2, N_2, C) or compounds (e.g. CH_4, CO_2, SiO_2).
- Describe that substances with covalent bonds can form simple molecular structures or giant structures.
- Outline the physical properties of substances with structures (metallic, ionic or covalent) and those with simple molecular structures.
- Explain that the physical properties of substances with structures (metallic, ionic or covalent) differ from those with simple molecular structures.

2.1 INTRODUCTION

Chemical bonding may be defined as the combination of two or more atoms through rearrangement of electrons in their outermost shells, by sharing or transferring of electrons among themselves, so that all the atoms attain a stable configuration of nearest inert gas.

Table 2.1: Electronic Configuration of Inert Gases

Element	Atomic No.	Electronic Configuration
Helium	2	2
Neon	10	2, 8
Argon	18	2, 8, 8
Krypton	36	2, 8, 18, 8
Xenon	54	2, 8, 18, 18, 8
Radon	86	2, 8, 18, 32, 18, 8

The atoms of different elements combine with each other in order to complete their respective octets (i.e., 8 electrons in their outermost shell) or duplet (i.e., outermost shell having 2 electrons) in case of H, Li and Be to attain stable inert gas configuration.

2.2 DUPLET RULE

Helium, a rare gas atom is also stable and exists as a monoatomic gas. Its atom does not have the octet structure but has two electrons in the valence shell. This structure of two electrons in the first and outermost orbit is called the duplet structure. Helium is stable because the capacity of the first shell is two electrons and it is fully filled. Other atoms whose atomic numbers are close to that of helium, such as hydrogen and lithium become stable by acquiring the duplet structure.

2.3 OCTET RULE

As mentioned above, the noble gases (except He) possess 8 electrons in their valence shells and exist as monoatomic gases. They do not enter into chemical combination under normal conditions. These observations depict that a valence shell containing eight electrons is particularly stable. Since all other atoms have a basic urge to attain the noble gas configuration, they gain, lose or share electrons until their valence shells contain eight electrons. This has led **G.N. Lewis** in 1916 to put forward the octet rule which was the basis of the electronic theory of valency. The rule can be stated as *"During the formation of a molecule, an atom of a particular element gains, loses or shares electrons until it acquires a stable configuration of eight electrons in its valence shell"*.

The octet rule is found to be very useful in explaining the normal valency of elements and in the study of the chemical combination of atoms leading to the formation of molecules.

2.4 CHEMICAL BOND

It is defined as the attractive forces which hold the various chemical constituents (atoms, ions, etc.) together in a molecule.

2.5 TYPES OF CHEMICAL BONDS

On the basis of electronic theory, the following types of combinations are possible between atoms:

(*i*) **Electrovalent** or **ionic bond** which involves the transfer of one or more electrons from one atom to another.

(*ii*) **Covalent bond** which involves mutual sharing of electrons between two atoms having contributed equal number of electrons.

(*iii*) **Coordinate bond** which also involves sharing of electrons by two atoms but the shared pair is contributed by only one of the atoms.

2.5.1 Formation of Cation and Anion

Atoms of certain elements tend to complete their outermost shell (duplets or octets) by losing or by gaining electrons.

The chemical species obtained when an atom loses or gains electrons is called an ion.

1. **Cation:** A cation is formed when an atom loses one or more electrons from its outermost shell and achieve positive charge. For example,

$$\underset{2,\,8,\,1}{{}_{11}Na} \xrightarrow{-1e^-} \underset{2,\,8}{{}_{10}Na^+}$$

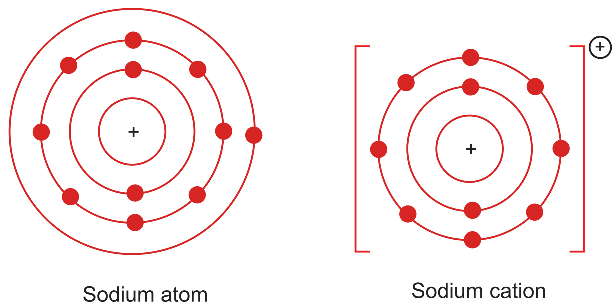


Fig. 2.1 *Formation of cation*

2. **Anion:** An anion is formed when an atom gains one or more electrons in its outermost shell and achieve negative charge. For example,

$$\underset{2,\,7}{F_9} + e^- \longrightarrow \underset{2,\,8}{F^-_{10}}$$

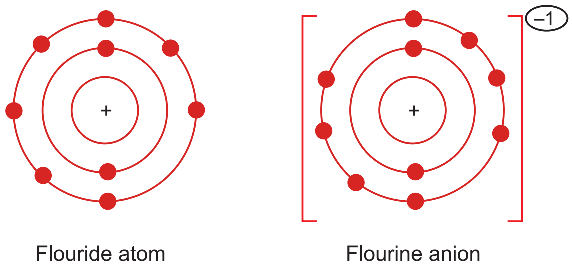


Fig. 2.2 *Formation of anion*

2.6 ELECTROVALENT BOND OR IONIC BOND

Atoms are electrically neutral. If an atom loses electrons, it will become positively charged and is called a **cation**. On the other hand, the atom which gains electrons will become negatively charged and is called an **anion**. The cation and anion are held by electrostatic force of attraction which is known as the *ionic* or *electrovalent bond.* It is shown in Fig. 2.3.

An **electrovalent** bond involves the complete transference of one or more electrons from one atom to another dissimilar atom so that both the atoms acquire stable inert gas configuration. For example, formation of sodium chloride from sodium and chloride.

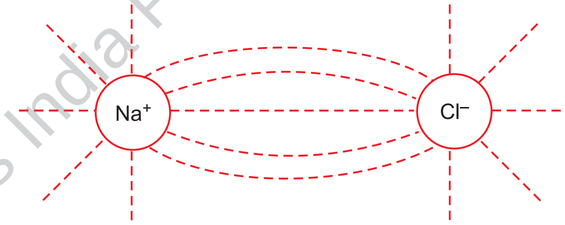


Fig. 2.3 *Electrostatic forces in NaCl*

2.6.1 Formation of Sodium Chloride (NaCl)

The electronic configuration of sodium (At. No. 11) is 2, 8, 1. It loses one electron from its outermost shell, and attains the configuration of inert gas neon 2, 8. The electrons of sodium are shown in Fig. 2.4.

On the other hand, the electronic configuration of chlorine (At. No. 17) is 2, 8, 7. It gains one electron and attains the configuration of argon, i.e., 2, 8, 8.

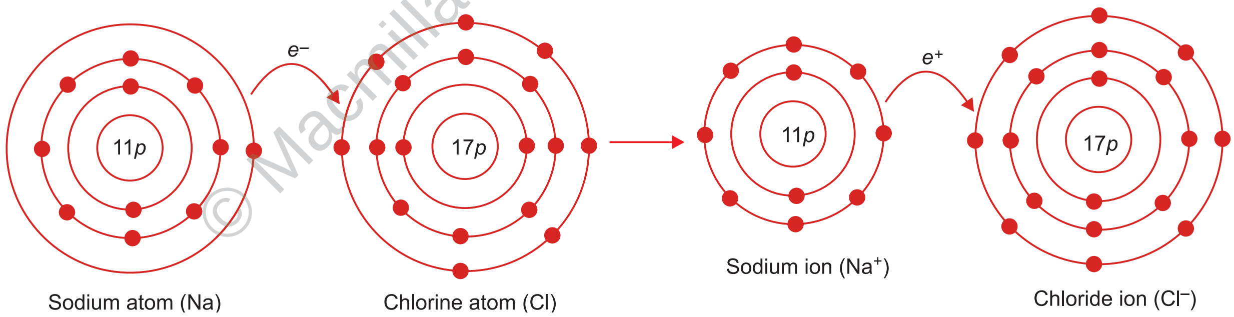
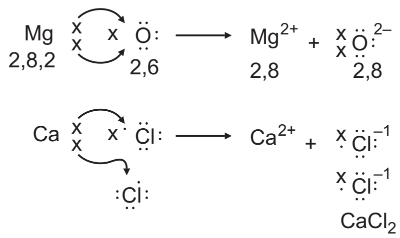


Fig. 2.4 *Ionic bond between* Na^+ *and* Cl^-

In this way, by the transfer of one electron from sodium to chlorine, both the atoms attain stable configuration. Sodium atom is converted into cation (written as Na^+) while chlorine atom is converted into anion (written as Cl^-), the oppositely charged ions are held by electrovalent bond. The formation of sodium chloride can easily be depicted by using **Electron dots** or **Cross Symbols** or **Lewis Symbols** in which only the outer shell electrons are shown as dots surrounding the symbols of the elements. In terms of Lewis symbols, the reaction between sodium and chlorine, can be written as shown below.

$$\underset{(2,8,1)}{Na\times} + \underset{(2,8,7)}{\cdot\ddot{\underset{\cdot\cdot}{Cl}}:} \longrightarrow \underset{(2,8)}{Na^+} + \underset{(2,8,8)}{{}^{\times}_{\cdot}\ddot{\underset{\cdot\cdot}{Cl}}:^-} \quad \text{or} \quad NaCl$$

Some more examples:



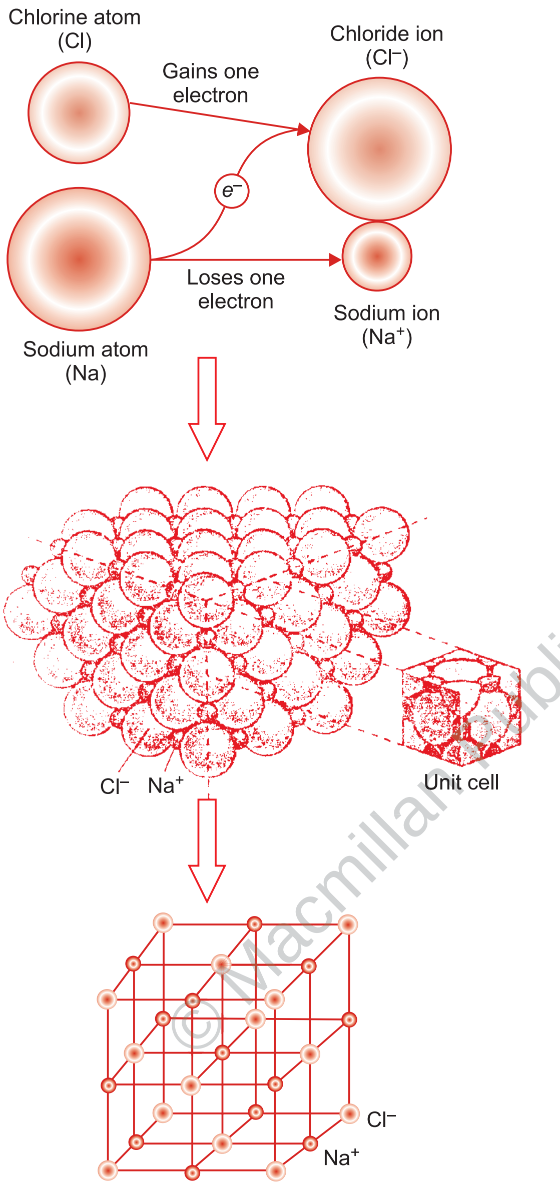
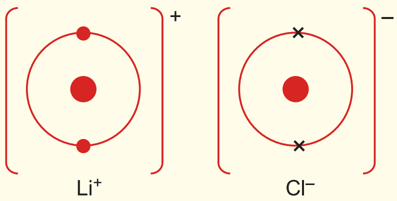


Fig. 2.5 *The arrangement of Na^+ and Cl^- in a sodium chloride crystal*

A 2.1

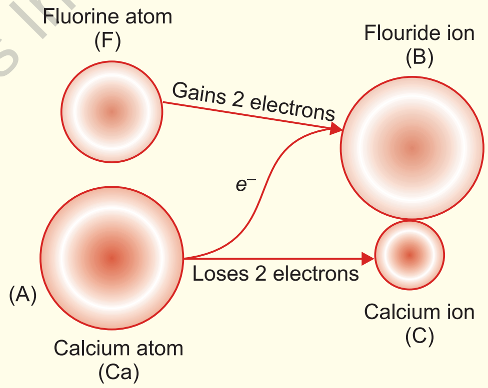
(*a*) Complete the diagrams below for the ions in the structure to show their electron arrangement. Draw the missing shells with appropriate number of electrons, showing clearly the origin of the electrons involved.



(*b*) Draw an ionic diagram similar to the one above for the structure of magnesium oxide de.

A 2.2

Represent A, B and C in the given activity.



Crystal Lattice: A regular arrangement of the constituents of a crystal in three dimensional space is called space lattice or crystal lattice.

Unit cell: It is the smallest part of the crystal lattice that will show the symmetry of the crystal as a whole.

2.7 GENERAL PROPERTIES OF IONIC COMPOUNDS

The ionic compounds possess the following properties:

> The ions may combine with the solvent to liberate energy called the hydration energy which is sufficient to overcome the attractive forces between the ions.

1. **Physical State:** These compounds usually exist in the solid state.
2. **Solubility:** Ionic compounds are soluble in water which is polar in nature.
3. **Ionic Reactions:** The ionic compounds undergo ionic reactions which are very fast.

   For example,

   $$Na^+Cl^-(aq) + Ag^+NO_3^-(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

   Precipitation of silver chloride takes place in a fraction of second.
4. **High Melting and Boiling Points:** Ionic compounds possess high melting and boiling points.
5. **Crystal Structure:** In ionic compounds ions are arranged in a regular pattern to form a crystal lattice.
6. **Electrical Conductivity:** Ionic compounds are good conductors of electricity in solution or in the molten state. In solution or molten state, their ions are free to move.
7. **Highly Brittle:** Ionic solids are highly brittle.

SELF EVALUATION

I. Look at the table. Use the information to answer the questions.

		Valency		
		1	**2**	**3**
Positive ions (cations)	**Metals**	sodium (Na^+) potassium (K^+) silver (Ag^+)	magnesium (Mg^{2+}) copper (Cu^{2+}) zinc (Zn^{2+}) iron (Fe^{2+})	aluminium (Al^{3+}) iron (Fe^{3+}) chromium (Cr^{3+})
	Compound ions	ammonium (NH_4^+)		
Negative ions (anions)	**Non-metals**	chloride (Cl^-) bromide (Br^-) iodide (I^-)	oxide (O^{2-}) sulfide (S^{2-})	nitride (N^{3-})
	Compound ions	nitrate (NO_3^-) hydroxide (OH^-)	carbonate (CO_3^{2-}) sulfate (SO_4^{2-})	phosphate (PO_4^{3-})

Use the information in the table to work out the formulae of the following ionic compounds.

(*i*) Copper oxide ....

(*ii*) Sodium carbonate ....

(*iii*) Zinc sulphate ....

(*iv*) Silver nitrate ....

(*v*) Magnesium bromide ....

(*vi*) Ammonium sulphate ....

(*vii*) Magnesium nitride ....

(*viii*) Potassium phosphate ....

(*ix*) Iron (III) hydroxide ....

(*x*) Chromium (III) chloride ....

II. When magnesium forms compounds it usually does so as cation (positive ion) with two unit charge.

(*i*) What change in electron structure occurs when magnesium atom becomes a magnesium ion?

(*ii*) Why the formula of the compound formed when magnesium and chlorine react is $MgCl_2$ and not $MgCl$?

2.8 COVALENT BOND

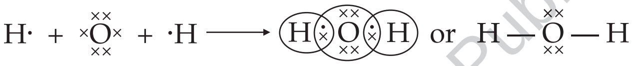
A covalent bond is formed when two atoms combine by mutual sharing of electrons between them so as to complete their duplets or octets configuration. For example, in the H_2 molecule, the two hydrogen atoms share the two electrons through covalent bonding. Here, the atoms involved are similar. But it is not necessary that the atoms have to be of the same element. For example, in a water molecule, the covalent bond is between the oxygen atom and each hydrogen atom. Each bond contains two electrons, one from the hydrogen atom and the other from the oxygen atom. In such cases, gain or loss of electrons is not possible.

(i) *Formation of Hydrogen Molecule:*

$$H\cdot + {}_\times H \longrightarrow (H \overset{\times}{\cdot} H) \text{ or } H-H$$

(Hydrogen tends to have two electrons in its outermost orbit)

(ii) *Formation of Water Molecule:*



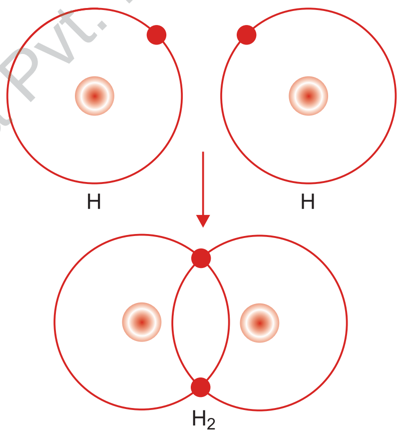


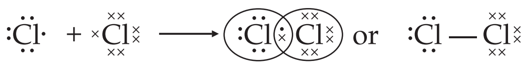
Fig. 2.6 *Representation of covalent bond formation (H_2)*

2.8.1 Types of Covalent Bonds

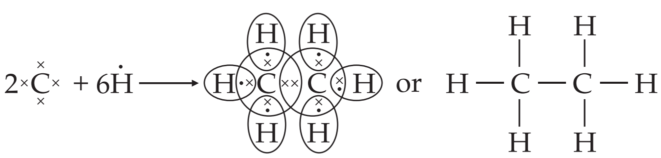
Based on the number of shared pair of electrons covalent bond can be divided in three types.

1. **Single Covalent Bond:** A covalent bond formed by mutual sharing of one pair of electrons is called a single covalent bond. A single covalent bond is represented by one line (—) between the two atoms. For example

   (i) *Formation of Chlorine Molecule (Cl_2):* Chlorine atom has seven valence electrons. Thus each Cl atom requires one more electron to achieve the nearest noble gas configuration. This lead to the formation of single covalent bond between the two chlorine atoms.

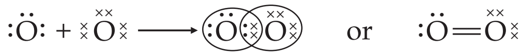


   (ii) *Formation of Ethane (C_2H_6):* Ethane molecule is formed as follows:

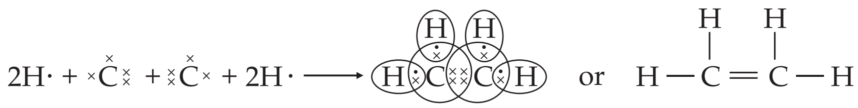


2. **Double Covalent Bond:** The bond formed between two atoms due to the sharing of two electron pairs is called a double covalent bond or simply a double bond. It is denoted by two lines (═) between the two atoms. For example,

(i) *Formation of Oxygen Molecule (O_2):* Each oxygen has 6 electrons in its outermost shell. So 2 more electrons are required to achieve the nearest noble gas configuration.

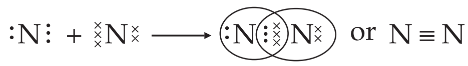


(ii) *Formation of Ethene Molecule (C_2H_4):* Ethene molecule is formed as follows



3. **Triple Covalent Bond:** The bond formed between two atoms due to the sharing of three electron pairs is called triple covalent bond or simply a triple bond. For example

(i) *Formation of Nitrogen Molecule:*



2.9 CHARACTERISTICS OF COVALENT COMPOUNDS

Covalent compounds possess the following general characteristics:

1. **Physical State:** Unlike ionic compounds, the covalent compounds exist in all the three states, *viz.*, solid, liquid and gaseous.
2. **Non-ionic Reactions:** Since these compounds are molecular in nature and not ionic, their reactions are molecular and proceed at a much slower rate than those of ionic compounds.
3. **Melting and Boiling Points:** Covalent compounds have low melting and boiling points because the molecules in covalent compounds are held together less rigidly (by weak van der Waals' forces), than in case of ionic compounds which involve electrostatic forces of attraction.
4. **Crystal Structure:** The crystal structure of covalent compounds differs from that of ionic compounds. They usually consist of molecules rather than ions.
5. **Solubility:** Covalent compounds are generally soluble in organic (non-polar or weakly polar) solvents but insoluble in water and other polar solvents.
6. **Isomerism:** As the atoms in covalent compounds are held together by the shared electrons and not by the electrostatic force, it is rigid and directional and hence shows isomerism in molecules.

A 2.3

Complete the table about the properties of covalent and ionic bond.

Covalent bond	Ionic bond
(*a*) Formed by sharing of electrons between the two atoms	(*a*) ....
(*b*)	(*b*) They are fast reactions
(*c*) Possess low melting and boiling points	(*c*) ....
(*d*)	(*d*) Soluble in polar solvents like water
(*e*)	(*e*) Generally solids
(*f*) Forces involved are weak van der Waal's forces of attraction	(*f*) ....

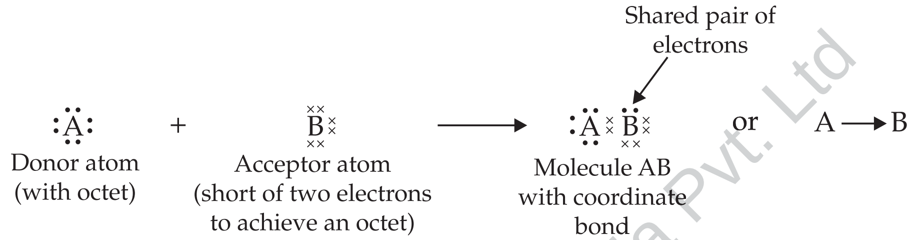
Note: Isomers have the same molecular formula, but have different arrangements of atoms. This phenomenon is termed as isomerism.

7. **Electrical Conductivity:** Since there are no free ions in covalent compounds, they are bad conductors of electricity.

2.10 COORDINATE BOND

The bond formed between two atoms by sharing a pair of electrons, provided entirely by one of the combining atoms, is called a coordinate bond or dative bond.

A coordinate bond is formed between two atoms (or ions) one of which has acquired an octet and the other is short of two electrons to complete its octet or duplet. The former donates a pair of electrons to the latter. The atom which donates a pair of electrons is referred to as donor while which takes the electrons is known as acceptor. A coordinate bond is represented by an arrow (⟶) pointing from donor to the acceptor atom.

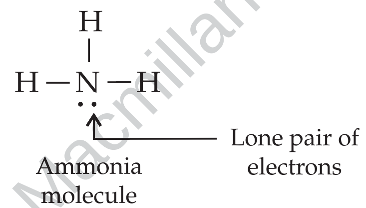


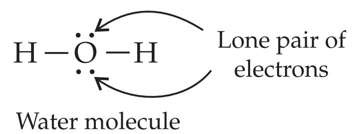
Examples of the coordinate bonds are found in ammonium (NH_4^+) ion, hydronium ions (H_3O^+), SO_3, etc.

2.10.1 Conditions for the Formation of a Coordinate Bond

(*a*) One of the two atoms must have at least one lone pair of electrons. A pair of electrons which is not shared with any other atom is known as the lone pair of electrons but it can be provided to other atom for the formation of coordinate bond.

For example, in ammonia (NH_3) molecule the nitrogen has one lone pair of the electrons. Similarly, in water (H_2O) molecule, the oxygen atom has two lone pairs of electrons.

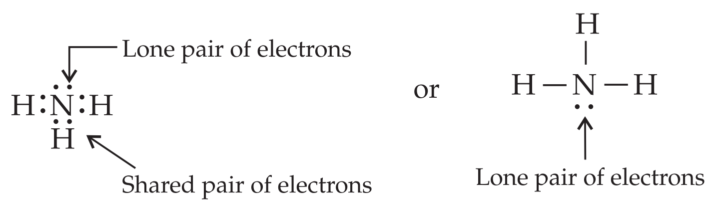




(*b*) Another atom should be short of at least a lone pair of electrons.

2.10.2 Formation and Structure of Ammonium Ion

In Ammonia (NH_3) molecule, nitrogen atom has a lone pair of electrons which means nitrogen atom (electronic configuration 2, 5) has five electrons in its valence shell, out of which only three electrons are involved in the formation of three N—H bonds as described earlier.



Due to the presence of a lone pair of electrons, nitrogen atom in NH_3 molecule can form a coordinate bond with some other atom or ion which is short of two electrons to acquire stable electronic configuration.

Hydrogen ion (H^+) does not possess any electron and is short of two electrons to achieve stable electronic configuration (a duplet) similar to that of helium which is its nearest noble gas. Thus, nitrogen atom in ammonia shares a lone pair of electrons with hydrogen ion which tends to acquire the configuration of helium atom. Thus, a coordinate bond results between nitrogen atom and the hydrogen ion in ammonium ion (NH_4^+). The formation and structure of NH_4^+ ion is shown in Fig. 2.7.

Electron dot structural diagram:

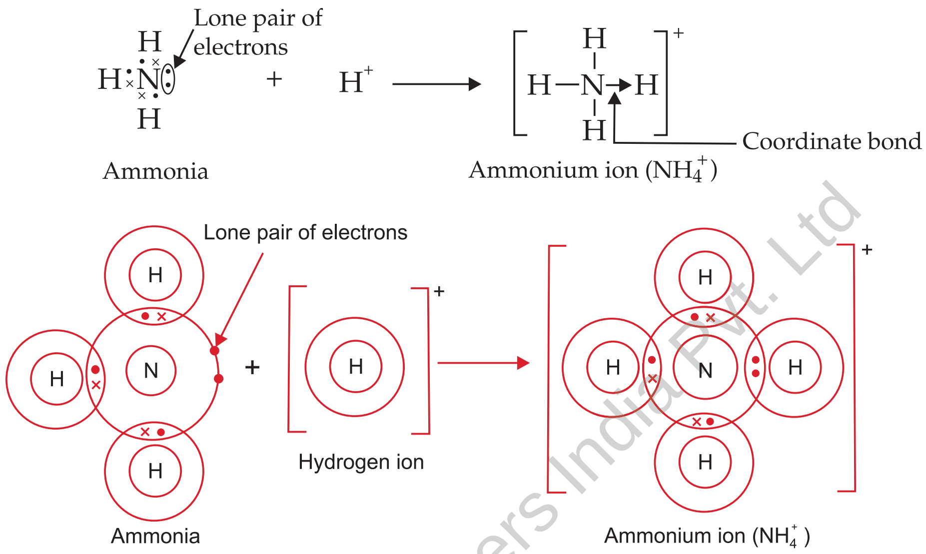


Fig. 2.7 *Formation of NH_4^+ ion*

2.11 TYPES OF CRYSTALLINE SOLIDS

Solids can be classified as crystalline or amorphous solids on the basis of the nature of order present in the arrangement of their constituent particles.

A **crystalline solid** usually consists of a large number of small crystals, each of them having a definite geometry. For example: iron, copper, silver, sodium chloride, zinc sulphate and naphthalene, etc.

An **amorphous solid** consists of particles of irregular shape. For example: glass, rubber and plastics, etc.

A covalent bond is formed between two similar or identical atoms. The electron pair forming the bond is shared equally between the two atoms and thus the bonds formed will be non-polar and molecule will be called **non-polar molecule**.

When two atoms forming a covalent bond, are dissimilar, i.e. their electronegativities are different, the pair forming the bond will be closer to the more electronegative atom and thus will acquire partial charges. Such molecules are called **polar molecules**.

Dipole-dipole forces act between the molecules possessing permanent dipole. Ends of the dipole possess "partial charges" and they are shown by Greek letter delta (δ).

Whenever a molecule contains a hydrogen atom linked to a highly electronegative atom (such as F, N, O), this atom attracts the shared pair of electrons more and so this end of molecules becomes slightly negative and other end (H end) becomes slightly positive resulting in the formation of a weak bond. This bond is called **hydrogen bond**.

Dipole-dipole forces are attractive forces between the positive end of one polar molecule and the negative end of another polar molecule. They are much weaker than ionic or covalent bonds and have a significant effect only when the molecules involved are closed together.

On the basis of the arrangement of constituent particles and the nature of attractive forces operating between them, crystalline solids are classified into various categories. Let us now learn about these categories:

1. Molecular solids
2. Ionic solids
3. Covalent solids
4. Metallic solids.

2.11.1 Molecular Solids

In these solids, the constituent particles which pack up together are molecules of the substance. These molecules may be **non-polar** such as CO_2, I_2, CCl_4, etc. or they may be **polar** such as H_2O, HCl, NH_3, etc.

(*i*) In case of **non-polar molecules,** the attractive forces operating between the molecules are *van der Waal's forces*. The examples of such solids are: *dry ice* (solid CO_2), *iodine, etc.*

(*ii*) In case of **polar molecules**, the attractive forces operating between the molecules in solid state are **dipole-dipole forces.** For example: *solid* SO_2, *solid HCl*. However, in some solids with polar molecules, the interparticle forces are **hydrogen bonds.**

Characteristics of Molecular Solids

Some of the general characteristics of molecular solids are:

(*i*) They are generally soft.

(*ii*) Their melting points are low to moderately high.

(*iii*) They have generally low density.

(*iv*) They are generally bad conductors of heat and electricity.

2.11.2 Ionic Solids

In ionic solids, the constituent particles are ions of opposite charges. Each ion is surrounded by a definite number of ions of opposite charge. The number of ions that surrounds a particular ion of opposite charge is called **coordination number** of the ion.

Coordination Number

It is the number of atoms, ions or molecules that surround a central atom in a crystal. The coordination number may vary from 2 to 16.

Sodium chloride (rock salt NaCl) is an ionic solid. When sodium metal reacts with chlorine gas, a redox reaction occurs, and electrons are transferred from sodium to chlorine to give Na^+Cl^-. The coulombic interaction between the oppositely charged ions hold an ionic solid together. In sodium chloride crystal each sodium ion (Na^+) is surrounded by six chloride ions. Hence coordination number of Na^+ is 6. At the same time each chloride ion is surrounded by six Na^+ ions. Therefore, the coordination number of Cl^- ion is also 6. However, in calcium fluoride crystal each Ca^{2+} ion is surrounded by eight fluoride (F^-) ions and each F^- ion is surrounded by four Ca^{2+} ions. Thus, in CaF_2 crystal, coordination numbers of Ca^{2+} and F^- ions are respectively 8 and 4. Al^{3+} has coordination number 4 in $[AlCl_4]^-$ but 6 in $[AlF_6]^{3-}$.

The interparticle forces in ionic solids are **ionic bonds** operating between the ions of opposite charges. Some examples of ionic solids are : sodium chloride (NaCl); caesium chloride (CsCl), zinc sulphide (ZnS), calcium fluoride (CaF_2), etc.

Characteristics of Ionic Solids

Some common characteristics of ionic solids are as follows:

(*i*) They are hard and brittle.

(*ii*) They have high melting points and low volatility.

(*iii*) They are poor conductors of electricity in solid state, however, they become good conductors of electricity in molten state or in dissolved state.

(*iv*) They are generally soluble in polar solvents like water.

2.11.3 Covalent or Network Solids

In these type of solids the constituent particles are atoms of same or different elements connected to each other by covalent bond network. For example, in **diamond** only carbon atoms constitute the covalent network while in **carborundum** covalent bond network is constituted by silicon and carbon atoms. Obviously, the interparticle forces operating in these solids are **covalent bonds.** These solids are also called **network solids** because the covalent bonds extend in three dimensions forming a giant interlocking structure. Some examples of covalent solids are:

Diamond, silicon carbide, silica, etc.

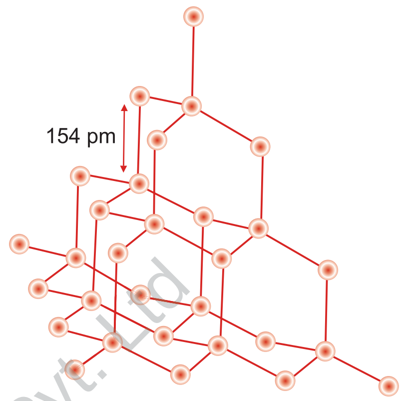


Fig. 2.8 *Network structure of diamond*

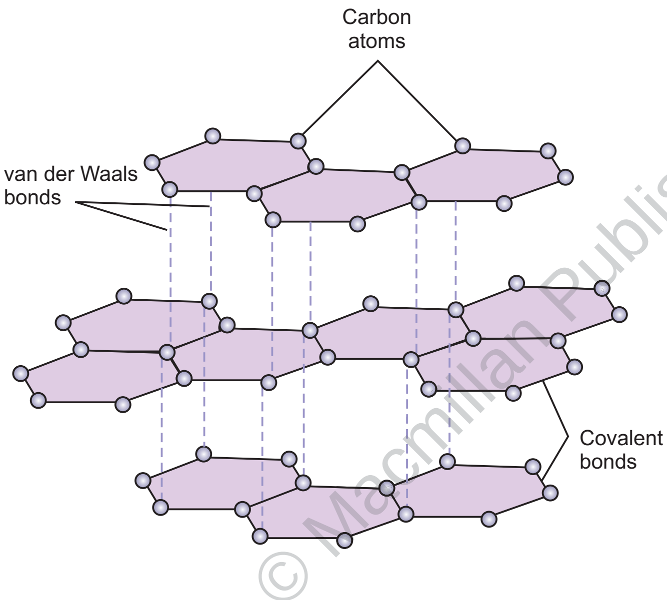


Fig. 2.9 *Structure of graphite*

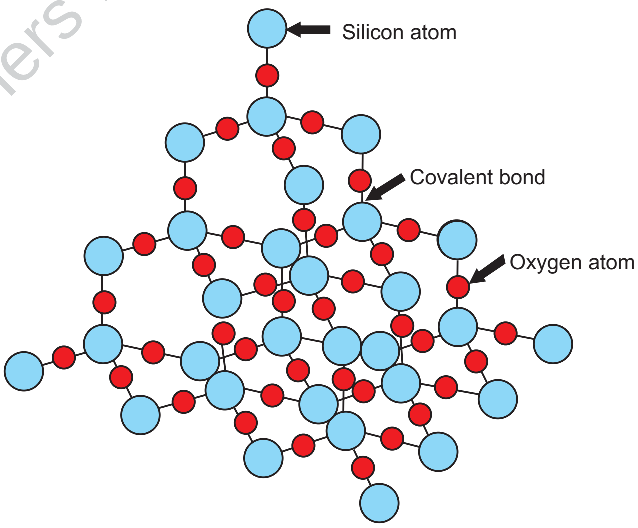


Fig. 2.10 *Three-dimensional structure of silica*

Silica is the name of the group of minerals which is composed of silicon and oxygen. It is found mostly in the crystalline state. Silica has a giant covalent structure. **The sharing of four oxygen atoms per tetrahedron forms an indefinite three-dimensional structure.** Each oxygen atom is bonded to two silicon atoms. Each silicon atom is surrounded tetrahedrally by four oxygen atoms and each oxygen atom is linked covalently to two silicon atoms which results in an infinite three-dimensional structure. Silica occurs in earth's crust as crystalline **quartz, tridymite** and **crystabollite** and as amorphous variety **flint.**

Characteristics of Covalent Solids

Some common characteristics of covalent solids are:

(*i*) They are very hard and brittle. Diamond is the hardest substance that occurs in nature.

(*ii*) Some compounds have very high melting points.

(*iii*) They are poor conductors of heat and electricity.

(*iv*) They have high enthalpies of fusion.

(*v*) Covalent solids like diamond and graphite have very high melting and boiling points because strong covalent bonds between atoms are broken during melting.

Enthalpy: It is the amount of heat content absorbed or released in a system at constant pressure. It is usually expressed as the change in enthalpy.

2.11.4 Metallic Solids

In these type of solids, the constituent particles are metal atoms. The interparticle forces in these solids are **metallic bonds.** In the metallic crystals the metal atoms occupy the fixed positions but their valence electrons are mobile. The close packed assembly of metal kernels (part of metal atom without valence electrons) remain immersed in the sea of mobile valence electrons. The attractive force between the kernels and mobile valence electrons is termed as metallic bond. For example, Cu (Copper).

Characteristics of Metallic Solids

The common characteristics of metallic solids are as follows:

(*i*) They generally range from soft to very hard.

(*ii*) They are malleable and ductile.

(*iii*) They are good conductors of heat and electricity.

(*iv*) They possess bright lustre.

(*v*) They have high melting and boiling points.

(*vi*) They have moderate enthalpies of fusion.

A 2.4

Conduct an activity in the laboratory and demonstrate the solubility of different samples such as NaCl, graphite, iron, $CuSO_4$, brass, sugar in water. On the basis of their properties, ask students to identify the substances as ionic, covalent, molecular and metallic solids.

The summary of classification of solids on the basis of interparticle forces is given in Table 2.2.

Table 2.2: Comparison of Physical Properties of Different Types of Crystalline Solids

S. No.	Crystal Classification	Constituent Particles	Binding Forces	Properties	Examples
1.	Molecular	Polar or non- polar molecules	van der Waal's forces (Dipole-dipole forces)	Fairly soft, low to moderately high melting points, poor thermal and electrical conductors	Dry ice (solid CO_2), methane (CH_4) argon, chloroform ($CHCl_3$), iodine (I_2), H_2O, NH_3.
2.	Ionic	Positive and negative ions	Ionic bonds	Hard and brittle, high melting points, high enthalpies of fusion, poor thermal and electrical conductors	NaCl, ZnS, MgO, CaF_2, KNO_3.
3.	Covalent	Atoms that are connected in covalent bond network	Covalent bonds	Very hard, very high melting points, poor thermal and electrical conductors	Diamond, quartz silica, silicon carbide.
4.	Metallic	Cations with free moving electrons from the outer shells of metal atoms	Metallic bonds	Soft to very hard, low to very high melting points, excellent thermal and electrical conductors, malleable and ductile	All metallic elements, for example, Cu, Fe, Zn, Ag, Au.

SELF EVALUATION

I. Classify the following solids in different categories based on the nature of intermolecular forces operating in them.

potassium sulphate, silver, diamond, silica, ammonia, water, zinc sulphate, silicon carbide

II. Explain why ionic solids conduct electricity in molten state and not in solid state?

RECAPITULATION

1. Atoms of different elements combine with each other to complete their octets (duplet in case of H, Li).
2. The electrovalency of an element is the number of electrons lost or gained during the formation of electrovalent or ionic compound.
3. Coordination number of an ion is the number of oppositely charged ions present around that ion in the ionic compound, e.g., for Na^+ and Cl^- in NaCl, it is 6.
4. Ionic compounds are made up of ions whereas covalent compounds are molecular.
5. Ionic compounds are solids whereas covalent compounds are solid, liquid or gaseous.
6. Octet rule states that in the formation of a molecule, an atom of a particular element gains, loses or shares electrons until it acquires a stable configuration of possessing eight electrons in its valence shell.
7. Electrovalency is equal to the number of electrons lost or gained by the atom to form an ion.
8. Covalency is equal to the number of electrons contributed by an atom, for sharing between the combining atoms.
9. Polar covalent bonds are formed between atoms of different elements which have a large difference of electronegativity. The pair of electrons in such cases are not shared equally between the combining atoms, but get drawn towards the more electronegative atom.
10. Non-polar covalent bonds are formed between atoms of same elements of equal electronegativity. The pair of electrons in such cases are shared equally between the combining atoms.
11. Coordinate bonds are formed when the shared pair of electrons is contributed by only one of the combining atoms. For the formation of this bond, the donor atom must have achieved octet and must have at least one lone pair of electrons. The acceptor atom must be short of two electrons to achieve stable electronic configuration.
12. Lone pair or unshared pair of electrons is a pair of valence electrons which is not shared with any other atom.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. Atoms of all elements follow octet rule.
2. All the electrons in an atom are termed as valence electrons.
3. The ionic bond is non-directional.
4. Ionic compound exists in crystalline form.
5. Covalent compounds have low melting points.
6. All covalent compounds exist as giant aggregates.
7. Covalent bond is formed due to mutual sharing of electrons.
8. Covalent compounds are good conductors of electricity.

II. Fill in the Blanks

Fill in the blanks by using suitable words:

1. In electron dot structure, the dots represent __________ electrons.
2. Ionic compounds are highly soluble in __________.
3. The bond formed by the mutual sharing of two electrons is called __________ bond.
4. There are __________ bonds in nitrogen molecule.
5. The __________ bonds are non-directional.
6. NaCl is __________ in molten state.
7. Formula of common salt is __________.
8. Oxygen gas is a __________ molecule.

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. In hydrogen molecule, the two hydrogen atoms are held together by
   (*a*) coordinate bond (*b*) covalent bond (*c*) ionic bond (*d*) hydrogen bond
2. Which of the following is an ionic compound?
   (*a*) HCN (*b*) CCl_4 (*c*) KCl (*d*) CO_2
3. When double bond is connecting two atoms, there is a sharing of
   (*a*) 2 electrons (*b*) 4 electrons (*c*) 1 electron (*d*) none of these
4. The maximum valency of an element with atomic number 7 is
   (*a*) 7 (*b*) 3 (*c*) 4 (*d*) 5
5. Which of the following contains a coordinate bond?
   (*a*) O_3 (*b*) $BaCl_2$ (*c*) HCl (*d*) H_2O
6. Molten NaCl conducts electricity due to the presence of
   (*a*) free electrons (*b*) free ions
   (*c*) free molecules (*d*) atoms of sodium and chloride
7. The nature of chemical bonding in diamond is
   (*a*) ionic (*b*) covalent (*c*) coordinate (*d*) none of these
8. Which force binds together the carbon atom in the diamond?
   (*a*) Ionic (*b*) Covalent (*c*) van der Waal's (*d*) none of these
9. The octet rule is not valid for the molecule
   (*a*) CO_2 (*b*) SF_4 (*c*) O_2 (*d*) CO

10. Which is soluble in water?
   (*a*) CS_2 (*b*) C_2H_5OH (*c*) CCl_4 (*d*) $CHCl_3$
11. The electronic configuration of sodium is
   (*a*) 2, 8, 2 (*b*) 2, 8, 1 (*c*) 2, 8, 3 (*d*) 2, 8, 4

IV. Matching

Match the items of Column I with the corresponding items of Column II:

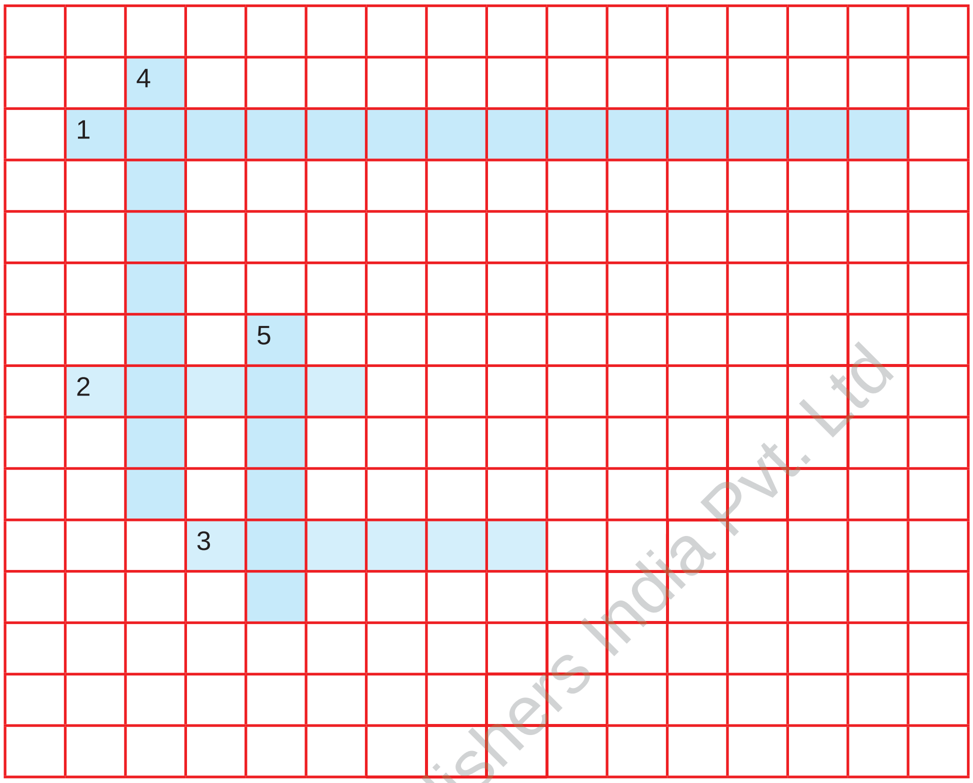
Column I	Column II
1. Ionic compounds usually exist	(*a*) Formation of cation
2. Loss of electron indicates	(*b*) Non-polar molecules
3. CCl_4 and CH_4 are	(*c*) Bad conductor of electricity
4. Graphite is used as	(*d*) Solid state
5. Covalent compounds are	(*e*) Lubricant

V. Answer the Following Questions

1. Why is argon monoatomic in nature?
2. Indicate the type of bonding in CaO, CO_2 and NH_3.
3. Why every molecule has a definite geometry?
4. Why water molecules remain associated in liquid water?
5. Why do atoms combine?
6. Solid A is a very hard electrical insulator in solid as well as in molten state and melts at extremely high temperature. What type of solid is it?
7. What do you understand by lone pair and shared pair of electrons?
8. What are polar and non-polar molecules?
9. What are the necessary conditions for the formation of a covalent bond?
10. Explain why He_2 is not formed.
11. What important characteristics should M and X have so that they form a predominantly ionic compound MX?
12. Define an electrovalent bond. What are the necessary conditions for the formation of ionic bond? Give two examples of ionic compounds.
13. HCl gas is a covalent compound but its aqueous solution conducts electricity. Why?
14. The melting points of electrovalent compounds are higher than the melting points of covalent compounds. Why?
15. How will you account for the fact that in metals the atoms are held neither by ionic bonds nor by covalent bonds?
16. Explain why NaCl is a better conductor in the molten state than in the solid state?
17. Draw the electron dot representation of (*i*) Ethane and (*ii*) carbon tetrachloride.
18. When an element in the first group of the periodic table reacts with an element in the seventh group, what is the type of bond that is likely to be formed?
19. NaCl and urea (H_2NCONH_2) are soluble in water. NaCl solution is a conductor of electricity while urea is not. Explain.

VI. Puz e

Read the given clues carefully and write the appropriate word inside the boxes to solve the puz e.



Across:

1. A bond, when the shared electrons are contributed by only one of the atoms.
2. It is formed between the two dissimilar atoms or atoms with different electronegativity.
3. An element which is the constituent of common salt.

Down:

4. A bond is formed by complete transfer of electrons.
5. An element lose electron.

[Hints: 1. Coordinate bond 2. Polar 3. Sodium 4. Ionic bond 5. Cation]

Web Links:

For more information on the following topics, visit the websites listed below:

Chemical Bond

https://en.wik pedia.org/wik /Chemical bond

https://www.visionlearning.com/en/library/Chemistry/1/Chemical-Bonding/55

Covalent Bond

https://en.wik pedia.org/wik /Covalent bond

Crystalline Solids

https://en.wik pedia.org/wik /Crystal

CHAPTER

3

REACTIVITY OF METALS

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Describe using balanced chemical equations, the reactions of common metals (e.g. Ca, Mg, Zn, Fe, Pb, Cu) with acids, oxygen and water.
- Construct a reactivity series for metals (e.g. Ca, Mg, Zn, Fe, Pb, Cu) by using the knowledge of the reactions of these metals with acids, water and oxygen.
- Apply the reactivity series to predict the reactions of other metals.

3.1 INTRODUCTION

The elements can be broadly divided into metals, metalloids and non-metals according to their physical and chemical properties. Some metals are more reactive and some are less reactive. Names of the metals, metalloids and non-metals are given in Table 3.1 which is called Periodic Table.

3.2 METALS

Metals are the elements (except hydrogen) which form positive ions by losing electrons or donating electrons.

For example, sodium, magnesium and aluminium are metals because they form positive ions (cations) by losing 1, 2 and 3 electrons respectively.

$$Na \longrightarrow Na^+ + e^-$$
$$Mg \longrightarrow Mg^{2+} + 2e^-$$
$$Al \longrightarrow Al^{3+} + 3e^-$$

Some of the examples of metals are iron, aluminium, copper, silver, gold, platinum, zinc, tin, lead, mercury, sodium, potassium, calcium, magnesium, etc. All the metals are solid, except mercury which is a liquid metal.

Elements of group 1 (IA) — Li, Na, K, Rb, Cs and Fr are known as alkali metals. These elements have low density, low melting and boiling points, low ionisation energy, highly electropositive and highly reactive (usually kept in kerosene to prevent them from coming in contact with moisture and oxygen present in the air). They get tarnished in air. They are strong reducing agents.

Elements of group 2 (IIA) — Be, Mg, Ca, Sr, Ba and Ra are known as alkaline earth metals. They are quite reactive but less reactive than alkali metals.

Elements of group 3 to 12 are known as transition metals.

Table 3.1: The Periodic Table of Elements

	Non-metal	Metalloid	Metal

	Gas	Solid	Liquid

Periods	Normal Elements		Transition Metals										Normal Elements					
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
	IA	IIA	IIIB	IVB	VB	VIB	VIIB		VIII		IB	IIB	IIIA	IVA	VA	VIA	VIIA	Zero
1	1 H Hydrogen																	2 He Helium
2	3 Li Lithium	4 Be Beryllium											5 B Boron	6 C Carbon	7 N Nitrogen	8 O Oxygen	9 F Fluorine	10 Ne Neon
3	11 Na Sodium	12 Mg Magnesium											13 Al Aluminium	14 Si Silicon	15 P Phosphorus	16 S Sulphur	17 Cl Chlorine	18 Ar Argon
4	19 K Potassium	20 Ca Calcium	21 Sc Scandium	22 Ti Titanium	23 V Vanadium	24 Cr Chromium	25 Mn Manganese	26 Fe Iron	27 Co Cobalt	28 Ni Nickel	29 Cu Copper	30 Zn Zinc	31 Ga Gallium	32 Ge Germanium	33 As Arsenic	34 Se Selenium	35 Br Bromine	36 Kr Krypton
5	37 Rb Rubidium	38 Sr Strontium	39 Y Yttrium	40 Zr Zirconium	41 Nb Niobium	42 Mo Molybdenum	43 Te Technetium	44 Ru Ruthenium	45 Rh Rhodium	46 Pd Palladium	47 Ag Silver	48 Cd Cadmium	49 In Indium	50 Sn Tin	51 Sb Antimony	52 Te Tellurium	53 I Iodine	54 Xe Xenon
6	55 Cs Caesium	56 Ba Barium	57-71 Lanthanide Series	72 Hf Hafnium	73 Ta Tantalum	74 W Tungsten	75 Rc Rhenium	76 Os Osmium	77 Ir Iridium	78 Pt Platinum	79 Au Gold	80 Hg Mercury	81 Tl Thallium	82 Pb Lead	83 Bi Bismuth	84 Po Polonium	85 At Astatine	86 Rn Radon
7	87 Fr Francium	88 Ra Radium	89-103 Actinide Series	104 Rf Ruther-fordium	105 Db Dubnium	106 Sg Seaborgium	107 Bh Bohrium	108 Hs Hassium	109 Mt Meitnerium	110 Ds Darmstadtium	111 Rg Roentgenium	112 Cn Copernicium	113 Uut Ununtrium	114 Fl Flerovium	115 Uup Ununpentium	116 Lv Livermorium	117 Uus Ununseptium	118 UuO Ununoctium

INNER TRANSI-TION METALS															
Lanthanide Series	57 La Lanthanum	58 Ce Cerium	59 Pr Praseody-mium	60 Nd Neodymium	61 Pm Promethium	62 Sm Samarium	63 Eu Europium	64 Gd Gadolinium	65 Tb Terbium	66 Dy Dysprosium	67 Ho Holmium	68 Er Erbium	69 Tm Thulium	70 Yb Ytterbium	71 Lu Lutetium
Actinide Series	89 Ac Actinium	90 Th Thorium	91 Pa Protactinium	92 U Uranium	93 Np Neptunium	94 Pu Plutonium	95 Am Americium	96 Cm Curium	97 Bk Berkelium	98 Cf Californium	99 Es Einsteinium	100 Fm Fermium	101 Md Mendelevium	102 No Nobelium	103 Lr Lawrencium

3.3 NON-METALS

Non-metals are the elements which form negative ions by accepting (or gaining) electrons. For example,

$$Cl + e^- \longrightarrow Cl^-$$
$$O + 2e^- \longrightarrow O^{2-}$$

Hydrogen is an exception, it is a non-metal which loses electron to form positive ion.

The most abundant non-metal in the earth's crust is oxygen. The major non-metals in the earth's crust in the decreasing order of their abundance are: oxygen, silicon, phosphorus and sulphur.

Metalloids show the properties of metals and non-metals, e.g. arsenic, antimony etc.

3.4 PROPERTIES OF METALS AND NON-METALS

Table 3.2: Physical Properties

S. No.	Property	Metals	Non-metals
1.	State	Usually solid – except mercury which is a liquid at room temperature. The melting point of Gallium and Caesium is 29.76°C and 28.44°C respectively. So, if room temperature is around 29.76ºC, they may also be in liquid state.	Exist in all the three states. Bromine is the only liquid.
2.	Density	Usually high, except for sodium, potassium, calcium.	Usually low (Diamond has high density).
3.	Melting point	Usually high except mercury, gallium, caesium, tin, lead, alkali and alkaline earth metals.	Low (exceptions are carbon, boron and silicon).
4.	Boiling point	Usually high	Low
5.	Hardness	Usually hard, except mercury, sodium, calcium, potassium, and lead.	They are generally soft except diamond, which is very hard.
6.	Malleability	Malleable i.e., can be beaten into thin sheets (except zinc and mercury).	Not malleable – generally brittle.
7.	Ductility	Ductile i.e., can be drawn into thin wires, except sodium, potassium, calcium, and mercury.	Not ductile cannot be drawn into thin wires.
8.	Conduction of heat	Good	Poor
9.	Conduction of electricity	Good (silver is the best conductor of electricity followed by copper).	Non-conductor, except carbon in the form of graphite and gas carbon.
10.	Lustre	Newly cut metals have high lustre. Some get tarnished immediately.	Usually not lustrous, except iodine and diamond, which is the most lustrous of all the substances.
11.	Alloy formation	Form alloys	Generally do not form alloys. But carbon, phosphorus, sulphur etc., can be present in some alloys.
12.	Tenacity	Usually have high tensile strength except sodium, potassium, calcium, lead, etc.	Low tensile strength.

Table 3.3: Chemical Properties

S. No.	Property	Metals	Non-metals
1.	Electronic configuration	Usually have 1, 2 or 3 electrons in their valence shell. Greater the number of shells, and lesser the number of valence electrons greater is the reactivity of the metal.	Usually have 5, 6 or 7 electrons in their valence shell. If any element has 4 electrons, it is also considered as a non-metal. Element that has 8 electrons in its valence shell is called a noble gas. Lesser the number of shells and greater the number of valence electrons, greater is the reactivity of the non-metal.
2.	Ionization and charge of ions	Always ionize by losing electrons: $Na - e^- \longrightarrow Na^+$ Positive.	Always ionize by gaining electrons: $Cl + e^- \longrightarrow Cl^-$ Negative.
3.	Deposition during electrolysis	Always at the cathode.	Always at the anode.
4.	Type of valency	Always electrovalency.	Electrovalency or covalency.
5.	Redox reaction	Lose electrons hence get oxidized. $Na \longrightarrow Na^+ + e^-$	Gain electrons hence get reduced. $Cl + e^- \longrightarrow Cl^-$
6.	Redox agents	Reducing agents.	Oxidizing agents.
7.	Nature of oxides	Generally form basic oxides (Na_2O, CaO, MgO, Fe_2O_3, CuO). Some of which are also amphoteric, such as aluminium oxide, zinc oxide, lead oxide, etc.	Generally form acidic oxide (CO_2, SO_2, SO_3, NO_2, P_2O_5). Some oxides are neutral, such as nitrous oxide, nitric oxide, carbon monoxide, water, etc.
8.	Atomicity	Always monoatomic.	Mono, di, tri or polyatomic Noble gases – monoatomic H_2, N_2, O_2 – diatomic, O_3 – triatomic, P_4, S_8 – polyatomic.
9.	Action with chlorine	Produce chlorides which are electrovalent.	Produce chlorides which are covalent.
10.	Action with dilute acids	Give respective salt and hydrogen.	No reaction.

3.5 ACTIVITY SERIES OR REACTIVITY SERIES OF METALS

The reactivity of metals, depends upon the ease with which they can loose electrons. Hydrogen is not a metal but it has been assigned a place in the activity series because it also loses, one electron more easily than some other metals and forms positive ion. Thus, it finds a place in the reactivity series of metals. Some metals are chemically very reactive whereas others are less reactive or unreactive. For example, sodium and potassium react very rapidly with cold water so they are very reactive metals. Zinc and iron react only with steam, so they are less reactive metals. On the other hand, copper and silver do not react even with steam, so they are unreactive metals. On the basis of reactivity of various metals with oxygen, water and acids, the metals have been arranged in a group or series according to their chemical reactivity.

The arrangement of metals in a vertical column in the order of decreasing reactivities is called **reactivity series of metals** or **activity series of metals**.

In activity series, the most reactive metal is placed at the top whereas the least reactive metal is placed at the bottom. The activity series of common metals is given below:

Activity Series or Reactivity Series of Metals

	Metal	Symbol	
These metals are more reactive than hydrogen	**Potassium**	**K**	← Most reactive
	Sodium	Na	↓
	Calcium	Ca	↓
	Magnesium	Mg	↓
	Aluminium	Al	Decreasing chemical reactivity
	Zinc	Zn	↓
	Iron	Fe	↓
	Tin	Sn	↓
	Lead	Pb	↓
	Hydrogen	**H**	↓
These metals are less reactive than hydrogen	Copper	Cu	↓
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	← Least reactive

3.5.1 Reaction of Metals with Air

Metal	Reaction with air
Ca	Tarnishes in moist air. Burns in air when heated.
Mg	Tarnishes in ordinary air. Burns when heated in air. $2Mg + O_2 \longrightarrow 2MgO$
Al, Zn	Get tarnished in moist air. Burn when strongly heated. $4Al + 3O_2 \longrightarrow 2Al_2O_3(s)$
Fe	Rusts in moist air. Burns when heated strongly in air. $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$
Pb	Tarnishes in moist air. Burns when heated strongly.
Cu	Tarnishes in moist air. Forms oxides when heated strongly. $2Cu + O_2 \longrightarrow 2CuO$

3.5.2 Reaction of Metals with Water

Metal	Reaction with water
Ca	Reaction of calcium is not so vigorous as the highly reactive metals like Na or K. If water is hot, K and Na may explode.
Mg	Reacts very slowly with cold water. Burning magnesium continues to burn in steam to form oxide and hydrogen. When steam is passed over red hot metal, it reacts similarly. $Mg + H_2O \longrightarrow MgO + H_2$
Al	Red hot metal reacts with steam to form oxide and hydrogen.
Zn	$Zn + H_2O \longrightarrow ZnO + H_2$
Fe	Red hot metal reacts with steam to form tri-iron tetroxide and hydrogen. The reaction is reversible. $3Fe + 4H_2O \rightleftharpoons Fe_3O_4 + 4H_2$
Pb	No action with boiling water or steam.
Cu	

3.5.3 Reaction of Metals with Acids

Metal	Reaction with acids
Ca	Reacts vigorously with acids forming hydrogen, but less vigorous than K and Na. $Ca + 2HCl \longrightarrow CaCl_2 + H_2$
Mg	React with gradually decreasing vigour, forming hydrogen. $Mg + 2HCl \longrightarrow MgCl_2 + H_2$
Zn	$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$
Fe	$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$
Pb [H] Cu Hg Ag	No action with dilute acids. Lead combines with concentrated hydrochloric acid but the layer of lead chloride formed slows down the reaction and finally arrests the reaction. Hot concentrated sulphuric acid yields sulphur dioxide, and concentrated nitric acid yields nitrogen dioxide.
Au	Dissolves in 'aqua regia'.
Pt	No action with dilute or concentrated acid.

Potassium is the most reactive metal here, so it has been placed at the top. As we move down, the chemical reactivity of metals decreases. Gold being least reactive metal has been placed at the bottom. Though hydrogen is not a metal but even then it has been placed in the activity series of metals. This is due to the fact that like metals, hydrogen also loses electron and forms positive ion, H^+.

Aqua regia (translates to Royal water in Latin) is a chemical substance. It is made by mixing one part nitric acid and three parts hydrochloric acid. It is one of the few substances that can dissolve gold and platinum, and other noble metals. Tantalum, iridium and a few other metals are not dissolved by it.

A 3.1

Using the results of various types of chemical reactions, the metals can be arranged into the reactivity series.

(*a*) Magnesium reacts very slowly indeed with cold water but it does react strongly with steam to give magnesium oxide and a gas. Write the word equation for the reaction between magnesium and steam.

....

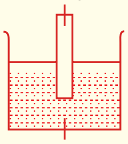
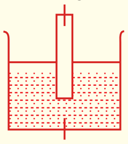
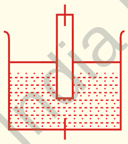
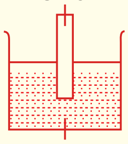
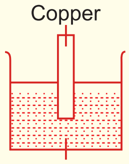
(*b*) Choose one metal from the reactivity series that will not react with steam.

....

(*c*) Choose one metal from the reactivity series that will safely react with dilute sulphuric acid.

....

(*d*) In each of the experiments below, a piece of metal is placed in a solution of a metal salt. Complete the table of observations.

		Zinc  tin(II) chloride solution	Zinc  copper (II) sulphate solution	Tin  copper (II) sulphate solution	Silver  copper (II) sulphate solution	Copper  silver (II) nitrate solution
At the	**Colour of metal**	grey	....	silver-coloured	silver-coloured	....
	Colour of solution	colourless	....	blue	blue	colourless
At the end	**Colour of metal**	coated with silver-coloured crystals	....	coated with brown solid	silver-coloured	coated with silver-coloured crystals
	Colour of solution	colourless	....	colourless	blue	....

(*e*) Use these results, place the metals copper, silver, tin and zinc in order of reactivity (putting the most reactive metal first).

.... > .... > .... > ....

Why some metals are more reactive and others are less reactive? When metals react, they lose electrons to form positive ions. Now, if a metal atom can lose electrons easily to form positive ions, it will react rapidly with some other substances and hence it will be a reactive metal. On the other hand, if a metal atom loses electrons less readily to form a positive ion it will react slowly with other substances. Such a metal will be less reactive.

3.5.4 Uses of Activity Series

1. Metals placed above hydrogen are more electropositive (more reactive) while the metals placed below hydrogen are less electropositive (less reactive).

2. Metals placed above hydrogen i.e., more reactive metals can displace hydrogen from its compounds such as water and dilute acids to form hydrogen gas. The metals placed below hydrogen i.e., less reactive metals cannot displace hydrogen from its compounds.
3. A more reactive metal displaces a less reactive metal from its salt solution. In other words, any metal higher up in the series displaces any metal lower in the series from its salt solution. For example, when a strip of zinc metal is put in the copper sulphate solution (blue), then the blue colour of copper sulphate solution fades gradually due to the displacement of copper ions from the copper sulphate molecules by zinc ions to form colourless zinc sulphate solution. Red brown copper is deposited on the zinc strip.

$$\underset{\text{Blue}}{CuSO_4} + Zn \longrightarrow \underset{\text{Colourless}}{ZnSO_4} + \underset{\text{Reddish brown}}{Cu}$$

   Thus, we see Zn placed above Cu in the activity series displaces copper from its salt solution. If we put silver (Ag) metal in copper sulphate solution, no reaction will take place.
4. The reducing power of metals decreases down in the series. This is the reason why sodium and potassium are very strong reducing agents. Hence, oxides of their metals cannot be reduced by hydrogen or carbon.
5. Thermal stability of the oxides of metals decreases from top to bottom in the activity series.

3.5.5 Chemical Reactions of Metals

1. Reaction with Acid:

(*i*) $$\underset{\text{Calcium}}{Ca} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Calcium chloride}}{CaCl_2} + H_2$$

(*ii*) $$\underset{\text{Magnesium}}{Mg} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Magnesium chloride}}{MgCl_2} + H_2$$

(*iii*) $$\underset{\text{Zinc}}{Zn} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2} + H_2$$

(*iv*) $$\underset{\text{Iron}}{Fe} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Ferrous chloride}}{FeCl_2} + H_2$$

(*v*) $$\underset{\text{Zinc}}{Zn} + \underset{\text{Sulphuric acid}}{H_2SO_4} \longrightarrow \underset{\substack{\text{Zinc sulphate}\\\text{(Slow reaction)}}}{ZnSO_4} + H_2$$

(*vi*) $$\underset{\text{Copper}}{Cu} + \underset{\substack{\text{Sulphuric acid}\\\text{(Conc.)}}}{2H_2SO_4} \xrightarrow{\text{Heat}} \underset{\text{Copper sulphate}}{CuSO_4} + 2H_2O + SO_2$$

2. Reaction with Oxygen:

(*i*) $$\underset{\text{Calcium}}{2Ca} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Calcium oxide}}{2CaO}$$

(*ii*) $$\underset{\text{Magnesium}}{2Mg} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Magnesium oxide}}{2MgO}$$

(*iii*) $$\underset{\text{Zinc}}{2Zn} + \underset{\text{Oxygen}}{O_2} \xrightarrow{\text{Heat}} \underset{\text{Zinc oxide}}{2ZnO}$$

A 3.2

Reaction of a Metal with Acid

Take about 5 mL of dil. H_2SO_4 in a test tube and add a few granules of zinc to it. Observe on the surface of zinc granules. Pass the gas being evolved through the soap solution. Take a burning candle near a gas filled bubble and observe. Repeat this activity with some more acids like HCl, HNO_3 and CH_3COOH.

Observations:

(*a*) Colourless, odourless gas is evolved from the surface of zinc granules which form bubbles in the soap solution.

(*b*) The gas in soap bubbles burn with a 'pop' sound.

Precautions:

(*a*) Handle acids carefully.

(*b*) Use test tube holder when required.

(iv) $4Fe + 3O_2 \xrightarrow{\text{Heat}} 2Fe_2O_3$
Iron + Oxygen → Iron oxide

$3Fe + 2O_2 \xrightarrow{\text{Heat}} Fe_3O_4$
Iron + Oxygen → Iron oxide

(v) $2Pb + O_2 \longrightarrow 2PbO$
Lead + Oxygen → Lead oxide

(vi) $4Cu + O_2 \longrightarrow 2Cu_2O$
Copper + Oxygen → Copper oxide

3. Reaction with Water:

(i) $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$
Calcium + Water → Calcium hydroxide

(ii) $Mg + 2H_2O \xrightarrow{\text{Heat}} Mg(OH)_2 + H_2$
Magnesium + Water → Magnesium hydroxide

(iii) $Zn + H_2O \longrightarrow$ No reaction
Zinc + Water

(iv) $4Fe + H_2O + 3O_2 \longrightarrow 2Fe_2O_3.xH_2O$
Iron + Water → Hydrated iron (II) Oxide (Rust)

(v) $Pb + H_2O \longrightarrow$ No reaction under normal condition
Lead + Water

(vi) $Cu + H_2O \longrightarrow$ No reaction under normal condition
Copper + Water

A 3.3

Sodium Reacts with Water

Take a beaker half-filled with water. Add a small piece of dry sodium into the water. Observe what happens. Sodium reacts vigorously with water and a gas is evolved and the beaker becomes hot. Dip a red litmus paper in the beaker and note the change in colour of litmus. It turns blue. This is due to the formation of sodium hydroxide solution. The gas evolved is hydrogen. Do the same test with Mg, Ca, Zn and Cu and record your observations.

Precautions:

(*a*) Use only very small quantity of sodium.
(*b*) Wear safety goggles.
(*c*) Stand at a distance to observe the reaction.

A 3.4

Investigating Exothermic Nature of Acids in Water

Take about 50 ml of water in a 200 ml beaker, add 10 – 15 drops of concentrated H_2SO_4 to it. During addition, stir the solution with the help of a glass rod. Now touch the beaker from outside to feel the change is temperature.

Observation:

Heat is evolved during this reaction, so the beaker feels hot.

Precautions:

(*a*) Handle acid carefully.
(*b*) Use small quantity of acid.

SELF EVALUATION

1. Magnesium reacts slowly with cold water but it reacts strongly with steam to give magnesium oxide and a gas. Write the word equation for the reaction between magnesium and steam.
2. Choose one metal from the reactivity series that will not react with steam.
3. Choose one metal from the reactivity series that will react with sulphuric acid.
4. In some parts of the world, copper is found in its 'native state'. Why is zinc never found in its native state?
5. The following metals are listed in order of reactivity: sodium > magnesium > zinc > copper

   Describe the reactions of these metals with (*a*) water (*b*) dil. HCl. Point how the reactions illustrate the change in reactivity.

RECAPITULATION

1. Elements can be divided into metals, non-metals, metalloids and noble gases.
2. Metals are solids (except mercury), have bright metallic lustre, high densities. They are hard, good conductors of heat and electricity, and have high tensile strength, etc.
3. Metals are electropositive elements and have a tendency to form positive ions by losing their valence electrons. The number of electrons lost by an atom of a metal to form a positive ion is the valency of that metal.
4. Non-metals exist in all the three states of matter i.e., solid, liquid and gas. They have no lustre, generally have low densities, are poor conductors of heat and electricity, and have low tensile strength, etc.
5. Non-metals are electronegative elements and have a tendency to form negative ions by gaining electrons. The number of electrons gained by an atom of a non-metal is the valency of that non-metal.
6. Metals are placed at the left, in the middle and at the bottom of the periodic table, whereas non-metals are placed on the right hand side of the periodic table.
7. Elements of group 1 (IA) — Li, Na, K, Rb, Cs and Fr are known as alkali metals.
8. Elements of group 2 (IIA) — Be, Mg, Ca, Sr, Ba and Ra are known as alkaline earth metals.
9. Elements of group 3 to 12 are known as transition metals.
10. Metalloids show the properties of metals as well as non-metals, e.g. arsenic, antimony.
11. Activity series of metals can be represented as Li > K > Na > Ca > Mg > Al > Zn > Fe > Pb > H > Cu > Hg > Ag > Pt > Au, on moving from top to bottom, the reactivity of their oxides goes on decreasing.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. Reactive metals occur in their native state.
2. Metalloids show properties of metals as well as non-metals.
3. The reducing power of metals decreases down in the series.
4. Metals react with acids and liberate oxygen gas.
5. Metals are generally ductile in nature.

II. Fill in the Blanks

Fill in the blanks by using suitable word(s):

1. Metals placed above hydrogen are more __________.
2. Metals are placed at the __________, in the __________ and at the __________ of the periodic table.
3. Alkali metals are strong __________ agents.
4. All metals are __________ except Hg.
5. Metals are the elements that conduct __________ and __________ .

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. Elements of groups 3 to 12 are known as
   (*a*) metalloids (*b*) transition metals (*c*) minerals (*d*) flux
2. They exist in all the three states of matter.
   (*a*) Metals (*b*) Non-metal (*c*) Metalloid (*d*) All of these
3. Metals react with acids and liberate
   (*a*) hydrogen (*b*) oxygen (*c*) carbon dioxide (*d*) sulphur dioxide
4. Which group of elements is known as alkali metals?
   (*a*) Group 1 (IA) (*b*) Group (IIA) (*c*) Group (III B) (*d*) Group VIII
5. Which metal is the alkaline earth metal?
   (*a*) Aluminium (*b*) Gold (*c*) Magnesium (*d*) Sodium
6. $Ca + 2H_2O \longrightarrow$
   (*a*) $Ca(OH)_2 + H_2$ (*b*) H_2 (*c*) $Ca(OH)_2$ (*d*) None of these
7. $Ca + 2HCl \longrightarrow$
   (*a*) $CaCl_2$ (*b*) H_2 (*c*) $CaCl_2 + H_2$ (*d*) None of these

IV. Matching

Match the items of Column I with the corresponding items of Column II:

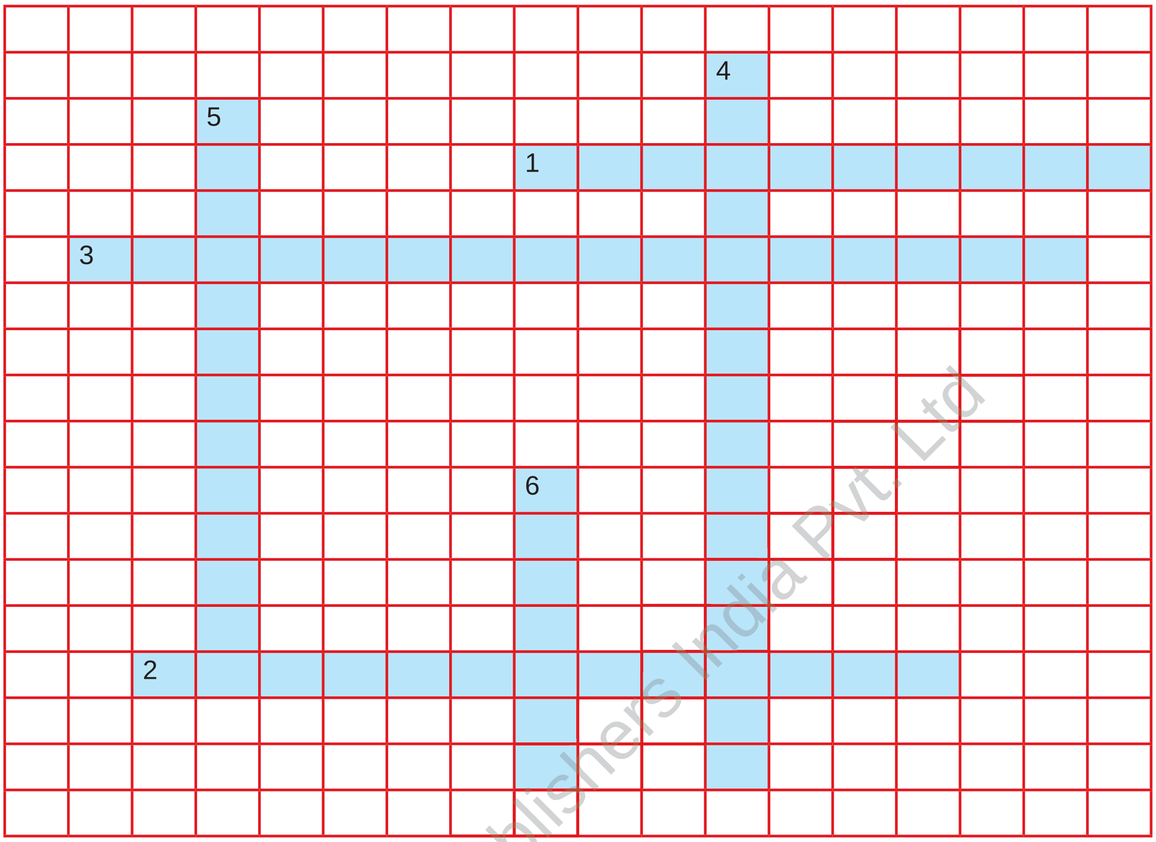
Column I	Column II
1. Potassium	(*a*) Decreases from top to bottom in the activity series
2. Thermal stability of oxides of metal	(*b*) Metalloids
3. Arsenic	(*c*) Transition metals
4. Forms negative ions by gaining of electrons	(*d*) Most reactive metal
5. Elements of group 3 to 12	(*e*) Non-metals

V. Answer the Following Questions

1. Name one metal and one non-metal which exists in liquid state at room temperature.
2. What would you observe in the following cases:
   (*i*) Sodium and potassium are placed in hot water
   (*ii*) Burning magnesium is introduced into steam?
3. Compare the activity of sodium and iron with water and dilute acids.
4. Define the terms:
   (*i*) Alkali metals (*ii*) Transition metals (*iii*) Metalloids
5. Give equations for the following conversions:
   (*a*) Sodium to sodium hydroxide using water
   (*b*) Steam to hydrogen using burning magnesium.
6. Explain placement of hydrogen atom in metal activity series.
7. What is the importance of activity series?
8. Explain reaction of metals with the following with the help of chemical equations:
   (*i*) with acids (*ii*) with oxygen (*iii*) with water
9. Why are some metals more reactive than others?
10. Give a brief look of the physical properties of metals.

VI. Puzzle

Read the given clues carefully and write the appropriate word inside the boxes to solve the puzzle.



Across:

1. Show properties of both metals and non-metals.
2. Elements of group (IIA) — Be, Mg, Ca, Sr, are known as ______ metals.
3. Arrangements of metals in order of decreasing reactivities.

Down:

4. Elements of group 3 to 12.
5. Elements of group 1 (IA) — Li, Na, K, Rb, Cs, Fr.
6. It is equivalent to the number of electrons lost by an atom of a metal or gained by an atom of a non-metal.

[Hints: 1. Metalloids 2. Alkaline Earth 3. Reactivity series 4. Transition metals 5. Alkali metals 6. Valency]

Web Links:

For more information on the following topics, visit the websites listed below:

Reactivity of Metals

http://www.dynamicscience.com.au/tester/solutions1/chemistry/reactivity%20of%20metals.htm

Properties of Metals and Non-metals

https://en.wikipedia.org/wiki/Properties_of_metals_metalloids_and_nonmetals

CHAPTER 4

CHEMICAL REACTIONS, CONSERVATION OF MASS AND STOICHIOMETRY

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Understand the concepts of physical and chemical changes.
- Learn about the chemical equation.
- Explain the rules for writing chemical equations.
- Represent chemical reactions by balanced symbol equations.
- Discuss the law of conservation of mass.
- Understand the stoichiometry.

4.1 INTRODUCTION

Everything in the universe changes with time. A large number of changes take place in our daily life. If an iron nail is left in the moist air, it rusts. White phosphorus when exposed to air catches fire. When you light the wick of a wax candle, it burns. Do you know what is happening in each of these reactions? Each substance, whether iron, phosphorus or wax, is reacting with oxygen. But you will notice that the rate of each reaction is different. Rusting of iron is a **slow reaction,** whereas burning of wax is a very **fast reaction**. Thus, it can be said that there are slow reactions, fast reactions and reactions which proceed with moderate rates. Chemical reactions are the process in which new substances with new properties are formed. The substances which take part in any chemical reaction are called **reactant** and the substances formed are called **product**.

$$\text{Reactant} \xrightarrow[\text{Catalyst, Light}]{\text{Temp., Pressure}} \text{Product}$$

For example,

$$\underset{\text{Nitrogen}}{N_2\,(g)} + \underset{\text{Hydrogen}}{3H_2\,(g)} \longrightarrow \underset{\text{Ammonia}}{2NH_3\,(g)}$$

In this reaction, nitrogen and hydrogen are reactants while ammonia is a product. When a chemical reaction occurs, one or more of the following changes take place:

1. Change in state
2. Formation of precipitate
3. Change in colour
4. Evolution of gas
5. Change in temperature

4.2 CHEMICAL CHANGE

A change in which one or more new substances are produced is known as a chemical change. Chemical change cannot be reversed easily, so chemical change is also called **permanent change**. It is also known as **chemical reaction**. For example,

(*i*) **Decomposition of Calcium Carbonate:** Calcium carbonate on heating produce calcium oxide and carbon dioxide.

$$\underset{\text{Calcium carbonate}}{CaCO_3\,(s)} \xrightarrow{\text{Heat}} \underset{\text{Calcium oxide}}{CaO\,(s)} + \underset{\text{Carbon dioxide}}{CO_2\,(s)}$$

(*ii*) **Mixing of Ammonia and Hydrochloric Acid:** The reaction between HCl and NH_3 to produce ammonium chloride.

$$\underset{\text{Ammonia}}{NH_3} + \underset{\text{Hydrochloric acid}}{HCl} \longrightarrow \underset{\text{Ammonium chloride}}{NH_4Cl}$$

(*iii*) **Mixing of Coke with Steam:** Coke reacts with steam as a result carbon monoxide is produced along with hydrogen gas.

$$\underset{\text{Coke}}{C\,(s)} + \underset{\text{Steam}}{H_2O\,(g)} \longrightarrow \underset{\text{Carbon monoxide}}{CO\,(g)} + \underset{\text{Hydrogen gas}}{H_2(g)}$$

4.3 CHEMICAL EQUATION

A **chemical equation** is a shorthand representation of a chemical reaction using the symbols and formulae of substances involved in the chemical reaction.

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

The above equation is long and not convenient for describing chemical reaction. So, such a long worded equation can be described by using symbols and formulae of the substances involved in the reaction. For a complete chemical equation, the number of atoms of various elements on both sides are made equal, i.e. the equation is balanced.

$$\underset{\text{Carbon}}{C\,(s)} + \underset{\text{Oxygen}}{O_2\,(g)} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2\,(g)}$$

4.3.1 Types of Chemical Equations

There are two types of chemical equations:

1. **Unbalanced Chemical Equation:** An unbalanced chemical equation is one in which the number of atoms of the elements on both sides of the equation is not the same.

$$\underset{\text{Hydrogen}}{H_2\,(g)} + \underset{\text{Oxygen}}{O_2\,(g)} \longrightarrow \underset{\text{Water}}{H_2O\,(l)}$$

2. **Balanced Chemical Equation:** A balanced chemical equation is one which contains an equal number of each element on both sides of the equation.

$$\underset{\text{Hydrogen}}{H_2\,(g)} + \underset{\text{Chlorine}}{Cl_2\,(g)} \longrightarrow \underset{\text{Hydrogen chloride}}{2HCl\,(g)}$$

4.3.2 Rules for Writing Chemical Equations

Reaction must be written in a systematic manner using the following certain rules:

1. The reactants taking part in the reaction are written in terms of their symbols or molecular formulae on the left-hand side of the equation.
2. A sign of plus (+) is added between the symbols or formulae of the reactants and products.
3. The products of reaction are written in terms of their symbols or molecular formulae on the right-hand side of the equation.
4. In between the reactants and the products an arrow sign ($\longrightarrow$) is marked. Tail towards reactants and head towards products. The arrow means 'reacts to give'.

$$Mg + 2HCl \longrightarrow MgCl_2 + H_2$$

In this chemical equation Mg and HCl are the reactants, and $MgCl_2$ and H_2 are the products. The arrow indicates that the Mg reacts with HCl to produce $MgCl_2$ and H_2.

4.3.3 Steps for Balancing Chemical Equations

Hit and Trial Method:

Step I: First of all write the skeleton equation. A skeletal equation is an equation in which various reactants and products are represented by their respective symbols or formulae.

$$H_2 + O_2 \longrightarrow H_2O$$

Step II: Count the number of atoms present on reactant side and product side in the table.

Name of atom	Right side	Left side
H	2	2
O	1	2

Step III: Balancing starts with the molecules or the compounds that contain maximum number of atoms whether of reactant or product. Here, H_2O contains maximum number of atoms. It contains one short of oxygen atom and so it is multiplied by 2 as shown below.

$$H_2 + O_2 \longrightarrow 2H_2O$$

Step IV: Now examine the effect of multiplication of the molecule on the balance of other atoms. As is seen, the oxygen atoms balance but now hydrogen atoms on left is less by 2. So, multiply H_2 on the left by 2.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Step V: Further count the number of atoms of each type on both reactant and product sides in the above equation.

Name of atom	Right side	Left side
H	4	4
O	2	2

So, the equation is balanced. If the number of atoms on the both sides is not equal, continue with the above steps till balanced equation is obtained. This is called **hit and trial method.**

A **4.1**

Complete the following spaces.

Step I: Nitrogen and hydrogen combine to form ammonia. Write the chemical equation.

$$N_2 \quad + \quad H_2 \longrightarrow NH_3$$

Step II: Count the number of atoms present on reactant side and product side.

Name of atom	Right side	Left side
N	1	-----
H	-----	2

Step III: Balancing start with the molecules of the compound that contain maximum number of atoms.

$$N_2 \quad + \quad H_2 \longrightarrow NH_3$$

Step IV: Examine the effect of multiplication and make further atoms/molecules equal in number on both sides.

$$N_2 \quad + \quad \_\_H_2 \longrightarrow \_\_NH_3$$

Step V: Further count the atoms.

Name of atom	Right side	Left side
N	-----	2
H	6	-----

4.3.4 Significance of a Chemical Equation

A chemical equation has qualitative as well as quantitative significance.

Qualitatively, it represents the names of the reactants and products. The conditions required for the reaction are specified above and below the arrow.

Quantitatively, it represents:

(*i*) the relative number of molecules of reactants and products,

(*ii*) the relative number of moles of reactants and products,

(*iii*) the relative masses of reactants and products.

A **4.2**

Balance the following chemical equation:

(*a*) $HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + H_2O$

(*b*) $NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

(*c*) $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$

(*d*) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + HCl$

4.3.5 Limitations of a Chemical Equation

1. A chemical equation does not indicate the physical states of the reactants and products. So, in a balanced chemical equation while stating the physical states (*s*) for solid, (*l*) for liquid, (*g*) for gas and (*aq*) for aqueous solution can be added. Sometimes, gas is indicated by upward pointing arrow ($\uparrow$). Similarly, for a precipitate, it is indicated by downward pointing arrow ($\downarrow$) placed in front of a molecule.
2. It does not indicate whether heat is evolved or absorbed in a reaction.
3. It does not indicate whether the reaction is fast or slow.

4. Chemical equation does not indicate the concentration of the reactants and the products.
5. A chemical equation does not tell about the conditions of the reaction. Therefore, in a balanced chemical equation, the conditions of the reaction such as temperature, pressure, and catalyst are written above or below the arrow. For example,

$$\underset{\text{Ethene}}{CH_2 = CH_2} + H_2(g) \xrightarrow[300°C]{Ni(Catalyst)} \underset{\text{Ethane}}{CH_3 — CH_3}$$

SELF EVALUATION

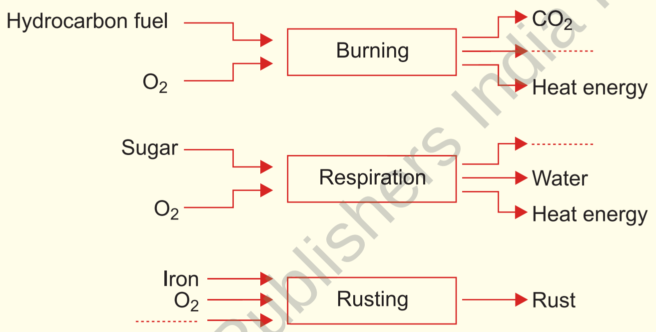
I. When burning magnesium is placed in a glass jar of carbon dioxide, it continues to burn and a black solid is produced.

(*i*) How would you show that black jar contains CO_2?

(*ii*) Suggest what the black solid is.

(*iii*) Suggest the name of the other product of the reaction.

II. Copy out and complete the diagram to show what is used and what is produced in (*a*) burning (*b*) respiration, (*c*) rusting.



Which type of chemical change is involved in all of the reactions (*a*), (*b*) and (*c*)?

4.4 STOICHIOMETRY AND STOICHIOMETRY CALCULATIONS

Stoichiometry deals with the calculation of masses (sometimes volume also) of the reactant and the products involved in a chemical reaction. The interpretation of the coefficients as the number of moles is the basis of all stoichiometric calculations. For the stoichiometric calculation, first of all write the chemical equation for the reaction and then balance it. The balanced chemical equation gives the stoichiometric coefficient which gives the proportion by moles.

The quantitative information conveyed by a chemical equation helps in a number of calculations. The problems involving these calculations may be classified into the following types:

1. **Mole to mole relationship:** In such type of problems, the moles of one of the reactants or products is to be calculated if that of other reactants and products are given.
2. **Mass-mass relationship:** In such type of problems, the mass of one of the reactants or products is to be calculated if that of the other reactants or products given.
3. **Mass-volume relationship:** In such type of problems, the mass or volume of one of the reactants or products is calculated from the mass or volume of other substances.
4. **Volume-volume relationship:** In such type of problems, the volume of one of the reactants or products is given and that of the other is to be calculated.

The general method of calculation of all the problems consists of some important steps:

(*i*) Write down the balanced chemical equation.

(*ii*) Write the relative number of moles or the relative masses of the reactants and products below the formulae.

(*iii*) For gaseous substances, write down 22.4 litre at STP below the formula in place of 1 mole.

(*iv*) Calculate the results by a unitary method.

SOLVED EXAMPLES

Example 1. How many moles of methane are required to produce 22 gram of CO_2 after combustion?

Sol. $\underset{1\text{ mole}}{CH_4} + 2O_2 \longrightarrow \underset{44\text{ g}}{CO_2} + 2H_2O$

$\because$ 44 gram of CO_2 is produced from 1 mole of CH_4

$\therefore$ 22 gram of CO_2 is produced from $CH_4 = \dfrac{22}{44}$

$= 0.5$ mole.

Example 2. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid (HCl) according to the reaction:

$$\mathbf{MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_2(aq) + 2H_2O(l) + 5Cl_2(g)}$$

How many grams of HCl react with 5.0 g of manganese dioxide?

Sol. The balanced chemical equation is:

$$\underset{1\text{ mol}}{MnO_2(s)} + \underset{4\text{ mol}}{4HCl(aq)} \longrightarrow MnCl_2(aq) + 2H_2O(l) + 5Cl_2(g)$$

$55 + 2 \times 16$	$4 \times (1 + 35.5)$
$= 87$ g	$= 146.0$ g
5 g	?

$\because$ 87 g of MnO_2 reacts with 146.0 g of HCl

$\therefore$ 5 g of MnO_2 will react with $= \dfrac{146.0}{86} \times 5$

$= \mathbf{8.39\ g}$

Example 3. What weight of zinc would be required to produce enough hydrogen to reduce completely 8.5 g of copper oxide to copper?

Sol. (*i*) Calculation of H_2 required to reduce CuO to Cu.

$$\underset{\substack{1\text{ mol}\\63.5+16=79.5\text{ g}}}{CuO} + \underset{\substack{1\text{ mol}\\2\text{ g}}}{H_2} \longrightarrow Cu + H_2O$$

One mole of CuO requires 1 mole of H_2 for reduction to copper.

$\because$ or 79.5 g of CuO require H_2 for reduction to copper = 2 g

$\therefore$ 8.5 g of CuO require H_2 for reduction to copper $= \dfrac{2}{79.5} \times 8.5$

$= 0.214$ g

(*ii*) Calculation of weight of zinc required to produce H_2

$$\underset{\substack{1\text{ mol}\\65\text{ g}}}{Zn} + H_2SO_4 \longrightarrow ZnSO_4 + \underset{\substack{1\text{ mol}\\2\text{ g}}}{H_2}$$

$\because$ 2 g of H_2 is produced from zinc = 65 g

$\therefore$ 0.214 g of H_2 would be produced from zinc = $\frac{65}{2} \times 0.214$

$= \mathbf{6.955\ g}$

Example 4. **$KClO_3$ on heating gives KCl and O_2. What is the volume of O_2 at NTP liberated by 0.1 mole of $KClO_3$?**

Sol.

$$\underset{\text{2 mole}}{2KClO_3} \xrightarrow{\text{Heat}} 2KCl + \underset{\substack{\text{3 mole} \\ (3 \times 22.4\ L) \\ = 67.2\ L}}{3O_2}$$

$\because$ 2 moles of $KClO_3$ evolve O_2 at NTP = 67.2 L

$\therefore$ 0.1 mole of $KClO_3$ evolve O_2 at NTP = $\frac{67.2 \quad 0.1}{2}$

$= \mathbf{3.36\ L}$

Example 5. **Calculate the volume of oxygen at NTP that would be required to convert 5.2 L of carbon monoxide to carbon dioxide.**

Sol. The balanced chemical equation is :

$$\underset{\substack{\text{2 mol} \\ 2 \times 22.4\ L}}{2CO} + \underset{\substack{\text{1 mol} \\ 22.4\ L}}{O_2} \longrightarrow 2CO_2$$

$\because$ Volume of oxygen required to convert 2 × 22.4 L of CO at NTP = 22.4 L

$\therefore$ Volume of oxygen required to convert 5.2 L of CO at

$$\text{NTP} = \frac{22.4}{2 \quad 22.4} \times 5.2$$

$= \mathbf{2.6\ L}$

SELF EVALUATION

1. Calculate the amount of water in gram produced by combustion of 16 g methane (CH_4).
2. What mass of calcium oxide will be obtained by heating 3 mole of $CaCO_3$?
3. Calculate the amount of $KClO_3$ needed for burning 112 L of O_2 gas at NTP.
4. What volume of oxygen at STP is required to react completely with 200 cm^3 of acetylene?

4.5 TYPES OF CHEMICAL REACTIONS

There are different types of chemical reactions. Actually chemical reactions involve the breaking and making of bonds between atoms to produce new substances. There are rearrangements taking place or a regrouping of atoms in various ways to give products. Some important types of chemical reactions are:

1. Combination reaction or addition reaction
2. Decomposition reaction
3. Displacement reaction
4. Double displacement reaction
5. Oxidation and reduction reaction
6. Neutralisation reaction

Now, we will discuss all of these one by one in detail.

1. **Combination Reaction:** The reaction of two or more substances (either elements or compounds) to produce a new substance is called a combination reaction. For example,

   (*i*) *Calcium oxide reacts with water to produce slaked lime releasing a large amount of heat*:

$$\underset{\substack{\text{Quicklime}\\\text{(Calcium oxide)}}}{CaO(s)} + \underset{\text{Water}}{H_2O(l)} \longrightarrow \underset{\substack{\text{Slaked lime}\\\text{(Calcium hydroxide)}}}{Ca(OH)_2\,(aq)}$$

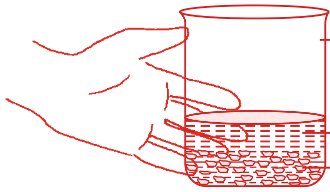


Fig. 4.1 *Formation of slaked lime by the reaction of calcium oxide with water*

In this reaction, calcium oxide and water combine to form a single product, i.e. calcium hydroxide. Such a reaction in which a single product is formed from two or more reactants is known as a combination reaction.

   (*ii*) *When hydrogen and oxygen combine, water is formed:*

$$\underset{\text{Hydrogen}}{2H_2\,(g)} + \underset{\text{Oxygen}}{O_2\,(g)} \longrightarrow \underset{\text{Water}}{2H_2O\,(l)}$$

2. **Decomposition Reaction:** The reaction in which single compound splits into two or more simpler substances is called decomposition reaction. For example,

   (*i*) *Decomposition of ferrous sulphate:*

$$\underset{\text{Ferrous sulphate}}{2FeSO_4\,(s)} \xrightarrow{\text{Heat}} \underset{\text{Ferric oxide}}{Fe_2O_3\,(s)} + \underset{\text{Sulphur dioxide}}{SO_2\,(g)} + \underset{\text{Sulphur trioxide}}{SO_3\,(g)}$$

In this reaction, single reactant splits into simpler products. Ferrous sulphate crystals ($FeSO_4.7H_2O$) lose water when heated and the colour of the crystals changes (light green changes to white). It then decomposes to ferric oxide (Fe_2O_3), sulphur dioxide (SO_2) and sulphur trioxide (SO_3). Ferric oxide is a solid, while SO_2 and SO_3 are gases.

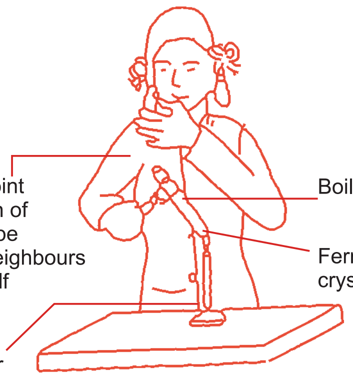


Fig. 4.2 *Correct way of heating the test tube containing crystals of ferrous sulphate and of smelling the odour*

   (*ii*) *Decomposition of lead nitrate:* Lead nitrate decomposes into lead oxide, nitrogen dioxide along with oxygen on heating. Lead nitrate is a white crystalline substance which on heating gives reddish brown fumes with pungent smell. These fumes are of nitrogen dioxide.

$$\underset{\text{Lead nitrate}}{2Pb(NO_3)_2\,(s)} \xrightarrow{\text{Heat}} \underset{\text{Lead oxide}}{2PbO} + \underset{\substack{\text{Nitrogen}\\\text{dioxide}}}{4NO_2\,(g)} + \underset{\text{Oxygen}}{O_2}$$

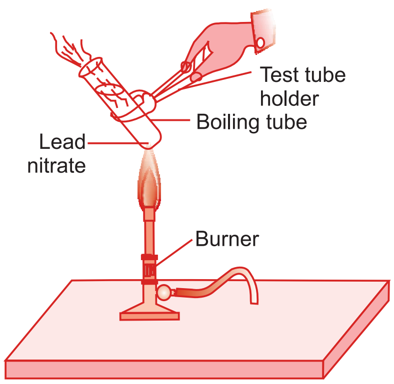


Fig. 4.3 *Heating of lead nitrate and emission of nitrogen dioxide*

3. **Displacement Reaction:** The reaction in which the atom of one element takes the place of the atom of another element in a compound is called displacement reaction. A displacement reaction occurs because a more reactive element displaces a less reactive one from a solution of one of its compounds.

Those reactions in which one element takes the place of another element in a compound are also called **single displacement reaction**. These reactions are very common in metals and often known as **metal displacement reactions**. For example,

(*i*) *Displacement of a less active metal by a more active metal:* If an iron nail is dipped in a copper sulphate solution, it is coated with copper.

Fe	+	$CuSO_4$	$\longrightarrow$	$FeSO_4$	+	Cu
Iron		Copper sulphate		Iron sulphate		Copper

We can say, iron displaces copper from copper sulphate solution. From this reaction, we also conclude that Fe is more reactive metal than Cu.

(*ii*) *Displacement of a less active non-metal by a more active non-metal:* When chlorine gas is passed through sodium bromide solution, sodium chloride and bromine are formed.

$2NaBr$	+	Cl_2	$\longrightarrow$	$2NaCl$	+	Br_2
Sodium bromide		Chlorine		Sodium chloride		Bromine

In this reaction chlorine has displaced bromine from sodium bromide. Therefore, chlorine is more reactive than bromine. The liberated Br_2 has imparted its colour to the solution.

A 4.3

Objective: To study the nature of the reaction between iron nail and copper sulphate solution.

Apparatus and materials required: Test tube, thread, copper sulphate solution, iron nail.

Procedure: First of all we take two iron nails and clean them by rubbing with sand paper. After that take two test tubes marked as (A) and (B). In each test tube take about 15 mL copper sulphate solution. Tie one iron nail with a thread and carefully immerse in the solution of copper sulphate of test tube B for about 15 minutes. Keep the second clean iron nail on a clean sheet of paper for comparison. After 15 minutes, take out the iron nail from the copper sulphate solution. Now, compare the intensity of the blue colour of copper sulphate solution in test tubes A and B.

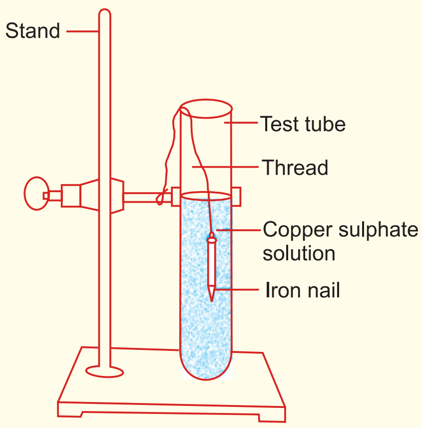


Fig. 4.4 *Iron nail placed in copper sulphate solution*

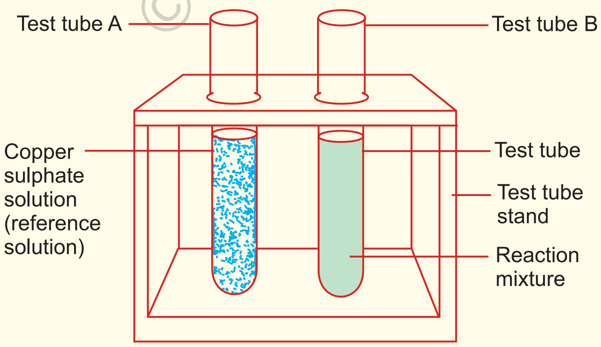


Fig. 4.5 *Comparison of colour of copper sulphate solution before and after the reaction*

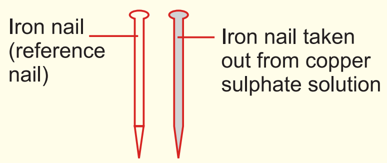


Fig. 4.6 *Comparison of colour of nail before and after the reaction*

Observations: When iron nail is dipped in the solution of test tube B, it turns brown and the solution in the test tube turns light green.

Conclusions: A brown coating on iron nail after the experiment shows that iron has displaced copper from copper sulphate solution.

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

Iron	Copper sulphate	Copper	Ferrous sulphate
(Grey)	(Blue)	(Brown)	(Light green)

A light green solution shows that ferrous ions are present in the solution. On the basis of above experiment, we can say iron is more reactive than copper.

4. **Double Displacement Reaction:** The reaction in which two reacting molecules exchange their partner ions in aqueous solution and form two new compounds is called double displacement reaction. For example,

(*i*) *When sodium sulphate reacts with barium chloride, two new products, barium sulphate and sodium chloride are obtained:*

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

Sodium sulphate	Barium chloride	Barium sulphate (white ppt.)	Sodium chloride

(*ii*) *When silver nitrate solution is added to sodium chloride solution, a white precipitate of silver chloride is formed with sodium nitrate solution:*

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

Silver nitrate	Sodium chloride	Silver chloride (white ppt.)	Sodium nitrate

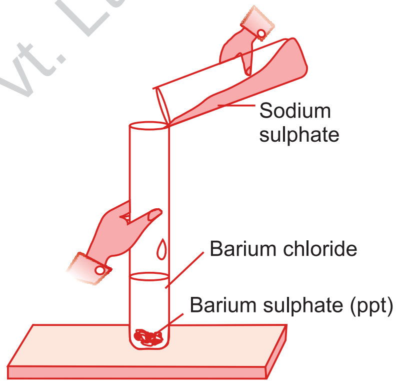


Fig. 4.7 *Formation of barium sulphate and sodium chloride*

5. **Oxidation and Reduction Reaction:** According to the classical concept, oxidation may be defined as *a process which involves addition of oxygen or any electronegative element, or as a process which involves the removal of hydrogen or any other electropositive element.* Examples of oxidation reaction are given below:

(*i*) *Addition of oxygen:*

$$2Mg + O_2 \longrightarrow 2MgO$$

Magnesium	Oxygen	Magnesium oxide

(*ii*) *Addition of electronegative element:*

$$2FeCl_2\ (aq) + Cl_2\ (g) \longrightarrow 2FeCl_3\ (aq)$$

Ferrous chloride	Chlorine	Ferric chloride

(*iii*) *Removal of hydrogen:*

$$2H_2S\ (g) + O_2\ (g) \longrightarrow 2S\ (s) + 2H_2O\ (l)$$

Hydrogen sulphide	Oxygen	Sulphur	Water

(*iv*) *Removal of electropositive element:*

$$2KI\ (aq) + H_2O\ (l) + O_3\ (g) \longrightarrow 2KOH\ (aq) + I_2\ (s) + O_2\ (g)$$

Potassium iodide	Water	Ozone	Potassium hydroxide	Iodine	Oxygen

In all these reactions, the element or the compound has undergone oxidation.

The substance which helps in oxidation is called **oxidising agent** and after affecting oxidation, it is reduced in a chemical reaction. For example, in the above reactions, oxygen, chlorine and ozone are the oxidising agents. Other important oxidising agents are $K_2Cr_2O_7$, $KMnO_4$, H_2O_2, HNO_3, etc.

Reduction Reaction: According to the classical concept, reduction may be defined as *a process which involves addition of hydrogen or any other electropositive element or removal of oxygen or any other electronegative element.* For example,

(*i*) *Addition of hydrogen:*

$$\underset{\text{Bromine}}{Br_2\,(g)} + \underset{\text{Hydrogen sulphide}}{H_2S\,(g)} \longrightarrow \underset{\text{Hydrogen bromide}}{2HBr\,(g)} + \underset{\text{Sulphur}}{S\,(s)}$$

(*ii*) *Addition of electropositive element:*

$$\underset{\substack{\text{Mercury (II)}\\ \text{chloride}}}{2HgCl_2\,(aq)} + \underset{\substack{\text{Tin (II)}\\ \text{chloride}}}{SnCl_2\,(aq)} \longrightarrow \underset{\substack{\text{Mercury (I)}\\ \text{chloride}}}{Hg_2Cl_2\,(s)} + \underset{\substack{\text{Tin (IV)}\\ \text{chloride}}}{SnCl_4\,(aq)}$$

(*iii*) *Removal of oxygen:*

$$\underset{\text{Copper oxide}}{CuO\,(s)} + \underset{\text{Hydrogen}}{H_2\,(g)} \longrightarrow \underset{\text{Copper}}{Cu\,(s)} + \underset{\text{Water}}{H_2O\,(l)}$$

(*iv*) *Removal of electronegative element:*

$$\underset{\substack{\text{Ferric}\\ \text{chloride}}}{2FeCl_3\,(aq)} + \underset{\substack{\text{Sulphur}\\ \text{dioxide}}}{SO_2\,(g)} + \underset{\text{Water}}{2H_2O\,(l)} \longrightarrow \underset{\substack{\text{Ferrous}\\ \text{chloride}}}{2FeCl_2\,(aq)} + \underset{\substack{\text{Sulphuric}\\ \text{acid}}}{H_2SO_4\,(aq)} + \underset{\substack{\text{Hydrochloric}\\ \text{acid}}}{2HCl\,(aq)}$$

Reducing agent may be defined as a substance which provides hydrogen or any other electropositive element, or removes oxygen or any other electronegative element. After causing reduction, it is oxidised in a chemical reaction.

In short we can say that, an oxidising agent itself undergoes reduction whereas a reducing agent undergoes oxidation in all their respective reactions.

These reactions are also called **redox reactions**.

6. **Neutralisation Reaction:** When acids react with bases or alkalis their acidity is destroyed and a salt is produced. Such reactions are known as neutralisation reactions. Examples of neutralisation reactions are also double displacement reactions. For example,

(*i*) *Combination of sulphuric acid with copper oxide:*

$$\underset{\substack{\text{Sulphuric acid}\\ \text{(Acid)}}}{H_2SO_4} + \underset{\substack{\text{Copper oxide}\\ \text{(Base)}}}{CuO} \longrightarrow \underset{\substack{\text{Copper sulphate}\\ \text{(Salt)}}}{CuSO_4} + \underset{\text{Water}}{H_2O}$$

(*ii*) *Hydrochloric acid reacts with sodium hydroxide:*

$$\underset{\substack{\text{Hydrochloric acid}\\ \text{(Acid)}}}{HCl} + \underset{\substack{\text{Sodium hydroxide}\\ \text{(Base)}}}{NaOH} \longrightarrow \underset{\substack{\text{Sodium chloride}\\ \text{(Salt)}}}{NaCl} + \underset{\text{Water}}{H_2O}$$

SELF EVALUATION

1. **Complete the following passage about chemical reaction by filling in the spaces:**

   The reaction in which single compound splits into two or more similar substances is called _________ reaction.

   $2FeSO_4\ (s) \xrightarrow{\text{Heat}}$ _________ $+ SO_3\ (g)$. During burning of coal _________ is formed.

   Calcium oxide and water combine to form a _________. Such a reaction in which a single product is formed from two or more reactants is known as _________.

2. Take three glass bottles with wide mouths. Label them A, B and C. Fill about half of the bottle A with tap water. Fill bottle B to the same level as in A with water which has been boiled for several minutes. In bottle C, take the same level of boiled water as in other bottles. In each bottle put a few similar iron nails so that they are completely under water. Add a small amount of cooking oil to the water in bottle C, so that it forms a film on its surface. Put the bottles away for a few days. Take out the nails from each bottle and observe them. Explain your observations.

4.6 MOLE

Mole as a unit of measurement of amount of substance. In a chemical investigation, we are often interested in knowing the number of atoms and molecules we are dealing with. For example, let us consider the following reaction:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

In this reaction, one molecule of nitrogen reacts with three molecules of hydrogen. So, it would be desirable to take the molecules of nitrogen and hydrogen in the ratio 1 : 3, so that the reactants are completely consumed during the reaction. But we know atoms and molecules are so small in size that it is not possible to count them individually. Moreover, the smallest sample of the substance that we may take contains a very large number of atoms or molecules. In order to overcome these difficulties, the *concept of mole* was introduced. According to this concept, number of particles of the substance is related to mass of the substance.

The **mole** may be defined as *"the amount of the substance that contains as many specified elementary particles (atoms, molecules, ions or electrons, etc.) as the number of atoms in 12 g of carbon-12 isotope."*

Mole is thus a unit for counting atoms, molecules, ions, etc. It stands for 6.023×10^{23} particles irrespective of their nature. *The term mole should not be confused with the physical quantity or amount of the substance because amount of the substance is not a number.*

ONE MOLE			
6.02×10^{23} atoms	6.02×10^{23} molecules	6.02×10^{23} formula units	**PARTICLES**
GAM or 1 g-atom	GMM or 1 g-molecule	GFM or 1 g-formula mass	**MASS**
22.4 L (*if gas*)	GMM/Density (*if solid or liquid*)		**VOLUME**

IMPORTANT FORMULAE ABOUT MOLE CONCEPT

The following formulae are quite helpful in solving the problems on mole concept. In these formulae N_A represents **6.02 × 10²³**.

Mass of **1 atom** of element = $\dfrac{\mathbf{GAM}}{\mathbf{N_A}}$

Mass of **1 molecule** of substance = $\dfrac{\mathbf{GMM}}{\mathbf{N_A}}$

Number of molecules in **W** g of substance = $\dfrac{\mathbf{WN_A}}{\mathbf{GMM}}$

Number of molecules in **V** litre of gas at S.T.P. = $\dfrac{\mathbf{VN_A}}{\mathbf{22.4}}$

Number of molecules in **n** moles of substance = $\mathbf{n \times N_A}$

1 **u** or (1 **amu**) = $\mathbf{1/N_A}$ **g** and is also called **1 avogram** or **1 dalton.**

The interconversion of mol, number of molecules and mass in gram is represented diagrammatically in the figure below.

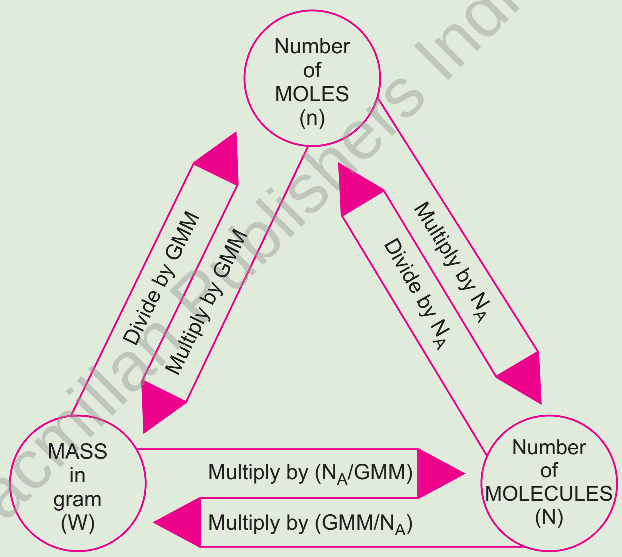


Fig. 4.8 *Interconversion of mol, mass and number of molecules*

Avogadro's Constant, N_A

From measurements based on the carbon-12 scale the number of atoms in 0.012 kg, or 12.00 g, of C–12 is 6.02×10^{23}. From the definition of the mole, *one mole of every substance contains this same number of entities.*

This number is called the *Avogadro constant.* It has the symbol N_A.

The Avogadro constant has units mol^{-1} and can be quoted as $N_A = 6.02 \times 10^{23}$ entities mol^{-1}.

From the definition of Avogadro's constant:

1 mol of any substance contains N_A entities.

½ mol of any substance contains ½ × N_A entities.

3 mol of any substance contains $3 \times N_A$ entities.

X mole of any substance contains $X \times N_A$ entities.

$\therefore$ Number of entities, N, in a given amount of substance, n, is given by:

$$N = n \times N_A$$

$$n = \frac{N}{N_A}$$

To conclude:

6.023×10^{23} atoms of oxygen weigh = 16 gram = GAM of oxygen or 1 g atom

6.023×10^{23} atoms of nitrogen weigh = 14 gram = GAM of nitrogen or 1 g atom

6.023×10^{23} molecules of oxygen weigh = 32 gram = GMM of oxygen or 1 g molecule

6.023×10^{23} molecules of water weigh = 18 gram = GMM of water of 1 g molecule.

SOLVED EXAMPLES

Example 1. **(*i*) Calculate the number of atoms in 16 g of Cu.**

(*ii*) How many moles are there in 2.7 g of Al?

(Given: Atomic mass of Cu = 64*u*, Al = 27*u*).

Sol. (*i*) Atomic mass of Cu = 64

$$\text{Number of moles in 16 g Cu}(n) = \frac{\text{Mass in g}}{\text{Atomic mass}} = \frac{16}{64} = \frac{1}{4}$$

$$= 0.25 \text{ mol}$$

$$\therefore \quad \text{Number of atoms} = \text{mol} \times N_A = 0.25 \times 6.02 \times 10^{23}$$

$$= \mathbf{1.50 \times 10^{23}}$$

Example 2. **What will be the mass of one ^{12}C atom in gram?**

Sol. 1 mole of ^{12}C atom = 6.022×10^{23} atom = 12 g

$$\therefore \quad 6.022 \times 10^{23} \text{ atoms} = 6.022 \times 10^{23} \text{ atoms} = 12 \text{ g}$$

$$\because \text{ 1 atom of } ^{12}C \text{ will have mass} = \frac{12}{6.022 \times 10^{23}}$$

$$= \mathbf{1.99 \times 10^{-23} \text{ g}}$$

Example 3. **Calculate the number of atoms in 52 moles of Ar.**

Sol. 1 mole of Ar contains atoms = 6.022×10^{23}

$$\text{52 moles of Ar contain atoms} = 6.022 \times 10^{23} \times 52$$

$$= \mathbf{3.13 \times 10^{25}}$$

Example 4. **Calculate the: (*i*) amount of CO_2, (*ii*) mass of gas, and (*iii*) number of molecules of gas in 2.8 dm^3 of CO_2 at STP (V_m = 22.4 dm^3 mol^{-1}, atomic mass; C = 12, O = 16).**

Sol. (*i*) Volume of CO_2, V = 2.80 dm^3

$$\text{molar volume of } CO_2, V_m = 22.4 \text{ dm}^3 \text{ mol}^{-1}$$

$$\text{amount of } CO_2, n = \frac{V(\text{dm}^3)}{V_m}$$

$$\therefore \quad \text{amount of } CO_2, n = \frac{2.80 \text{ dm}^3}{22.4 \text{ dm}^3 \text{ mol}^{-1}}$$

$$= 0.125 \text{ mol}$$

(*ii*) molar mass of CO_2, $M = 44 \text{ g mol}^{-1}$

$$\text{amount of } CO_2, n = 0.125 \text{ mol}$$

But $m = n \times M$

$\therefore$ $\text{mass of } CO_2 = 0.125 \text{ mol} \times 44 \text{ g mol}^{-1} = 5.5 \text{ g}$

(*iii*) $N = n \times L$, $n = 0.125$ mol, $L = 6.02 \times 10^{23} \text{ mol}^{-1}$

$\therefore$ number of molecules of $CO_2 = 0.125 \text{ mol} \times (6.02 \times 10^{23} \text{ molecules mol}^{-1})$

$= \mathbf{7.53 \times 10^{22}}$ **molecules.**

SELF EVALUATION

1. What is the mass of 0.600 mol of K_2CO_3?
2. How many molecules of O_2 gas are present in 16 g of oxygen gas?
3. Calculate the number of atoms in 52 g of He.

4.7 LAW OF CONSERVATION OF MASS

This law deals with the masses of the reactants and the products of a chemical reaction (or a physical change) and was studied by the great French chemist **Antoine Lavoisier** in 1774. This law may be stated as follows:

In all physical and chemical changes, the total mass of the reactants is equal to that of the products.

Thus, according to this law, there is no increase or decrease in the total mass of matter during a chemical or a physical change. In other words, matter can neither be created nor destroyed.

Hence, this law is also called the **law of indestructibility of matter**. For example,

$$\begin{array}{ccc} C + O_2 & \longrightarrow & CO_2 \\ \underbrace{12 + (2 \times 16)}_{44 \text{ g}} & \longrightarrow & \underbrace{12 + 32}_{44 \text{ g}} \end{array}$$

So, the mass of the reactants remains equal to the mass of the products.

If 32 g of sulphur reacts with 56 g of iron, 88 g of iron sulphide is produced.

$$\begin{array}{ccccc} Fe & + & S & \longrightarrow & FeS \\ 56 \text{ g} & & 32 \text{ g} & & 88 \text{ g} \end{array}$$

Hence we can say, mass of reactants before reaction will be same as mass of products after reaction.

SELF EVALUATION

Complete the following blanks on the basis of law of conservation of mass.

1. $2Mg + O_2 \longrightarrow 2MgO$

   $12.2 \text{ g} + 8.0 \text{ g} \longrightarrow$ ______
2. $H_2 + Cl_2 \longrightarrow 2HCl$

   $18 \text{ g} + 633 \text{ g} \longrightarrow$ ______
3. $CaCO_3 \longrightarrow CaO + CO_2$

   $200 \text{ g} \longrightarrow$ ______ $+ 88 \text{ g}$

A 4.4

Objective: To verify the law of conservation of mass.

Apparatus required: Laboratory balance, weighing paper, 600 mL beaker, 1 large 12 inch balloon, soda, juice bottle with cap, laboratory funnel and 250 mL flask.

Procedure:

1. Measure 30 mL of vinegar using a graduated cylinder. Pour the vinegar into a beaker. Place 10 g of baking soda onto a piece of weighing paper and fold the contents into a neat little packet. Place the beaker of vinegar and the packet of baking soda on the pan of laboratory balance. Measure and record the mass.
2. Carefully place the packet into the beaker and allow the reaction to complete. Measure and record the mass of the remaining products.
3. Pour another 30 mL of vinegar into a bottle using a funnel. Prepare another packet of baking soda as in step 1 and measure and record the mass of the reactants before mixing.
4. Tilt the bottle, insert the packet of baking soda into its neck and tightly secure the cap.
5. Tip the bottle upright and allow the contents to fully react. Measure and record the mass of the products.
6. Pour another 30 mL of vinegar into a 250 mL of flask using the funnel to make certain that the neck of the flask does not get wet. Prepare the last packet and record the mass of the reactants.
7. Tilt the flask (as in step 4) and slide the packet just into its neck. While one student holds the tilted flask, another must slip the open end of the balloon over the mouth of flask. Measure and record the mass of the products.

Result: Draw the mass of the products and reactants for each trial in a bar graph.

4.8 GAS LAWS AND MOLE CONCEPT

Matter may be defined as anything that possesses mass and occupies space. Matter exists in three states—*solid, liquid and gas*. These states can be changed into one another by change of temperature, pressure or both. When the forces of attraction between the component particles of matter are minimum, it exists in gaseous state. In this state, the particles are far apart from one another and their positions are not fixed in space. *Hence, gases do not have definite shape and volume.* They have a tendency to fill all the space available to them and take on the shape and volume of the vessel in which they are kept. Other important characteristics of the gases are as follows:

(*i*) Gases or the gaseous mixtures are quite homogeneous in composition.

(*ii*) Gases possess high compressibility.

(*iii*) Gases are converted into liquid state by cooling under high pressure.

(*iv*) Gases have the property of diffusion.

(*v*) Density of matter is very low in gaseous state.

Measurable properties of gases

(*i*) ***Volume:*** It is equal to the volume of the container and it is expressed in terms of litres (L), millilitres (mL) or cubic centimeters (cm^3), cubic meters (m^3) or cubic decimeters (dm^3).

$$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$$

(*ii*) ***Pressure:*** It is expressed in the units such as atmosphere, millimetres (mm) of Hg, torr, bar etc. SI unit of pressure is Pascal (Pa).

$$1 \text{ atm} = 760 \text{ mm of Hg} = 760 \text{ torr}$$

$$1 \text{ atm} = 101325 \text{ Pa} = 1.01325 \text{ bar}$$

(*iii*) ***Temperature:*** It is measured in celsius scale (°C) or in Kelvin scale (K). SI unit of temperature is Kelvin (K).

$$T(K) = t°C + 273$$

Gas Laws

1. Boyle's Law: The volume of a given mass of a gas is inversely proportional to its pressure at constant temperature.

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

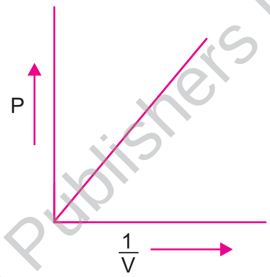
$$VP = K$$

K is a constant and its value depends on mass, temperature and nature of gas.

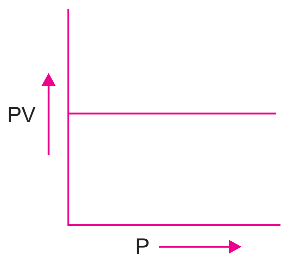
$$\therefore \quad P_1V_1 = P_2V_2$$

Graphical Representation of Boyle's Law:

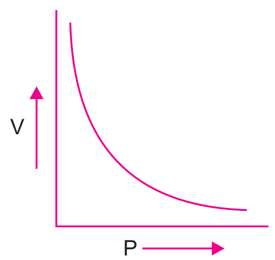
(*i*) Plot of Pressure P against $\frac{1}{V}$ is a straight line passing through the origin as shown below:



(*ii*) Plot of P × V against P is a straight line parallel to the X-axis as shown below:



(*iii*) Plot of V against P is a branch of hyperbola as show below:



2. Charles's Law: This law states that, *"for a given mass of a gas, at constant pressure, its volume is directly proportional to absolute temperature."*

Mathematically,

$V \propto T$ (for fixed mass of a gas at constant pressure)

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

or $$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

or $$d_1 T_1 = d_2 T_2$$

where V_1 is the volume of a given mass of gas at temperature T_1 and V_2 becomes the volume of same mass of gas at temperature T_2. d_1 and d_2 are the densities of the gas at T_1 and T_2.

Charles's law may be alternatively stated as, *"If pressure of a given mass of a gas is kept constant, the volume of the gas increases or decreases by* $\frac{1}{273}$ *of its volume at 0°C for one degree rise or fall in its temperature."*

Mathematically, $V_t = V_0 + \frac{V_0}{273} \times t°C$

where V_t is the volume at any temperature t°C and V_0 is the volume at 0°C.

A 4.5

To study the effect of increase of temperature on the volume of gas at constant pressure.

Apparatus required: Round bottom flask, beaker, delivery tube.

Procedure:

1. Set up the apparatus as shown in Fig. 4.9

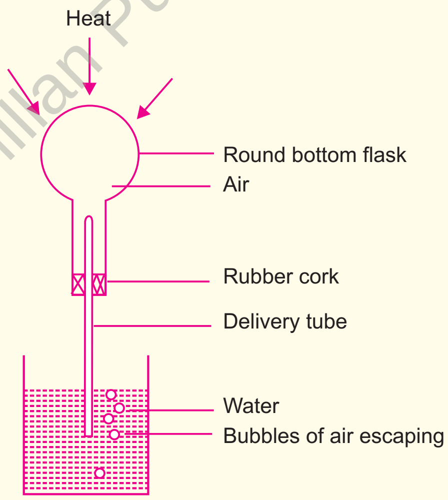


Fig. 4.9

2. Now, warm the air inside the flask gently by placing the hands on the flask. Note your observation.
3. Warm the flask gently with a low bunsen flame. Again note your observation.

   Try to explain your observations.

   Suggest a conclusion for the activity.

3. The Gas Equation: *This expresses a relation between pressure, volume and temperature of a gas.* This equation is obtained on combining the Boyle's and Charles's laws.

According to Boyle's law, $V \propto \frac{1}{P}$ (constant T and mass of gas)

According to Charles's law, $V \propto T$ (constant P and mass of gas)

On combining these two laws, we get

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

or $$V \propto \frac{T}{P}$$

or $$\frac{PV}{T} = \text{constant (R)}$$

where R is gas constant

or $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad ...(i)$$

Gas equation

$$PV = RT \quad ...(ii)$$

is applicable to one mole of a gas. For *n* moles of a gas, this equation takes the form,

$$\boldsymbol{PV = nRT} \quad ...(iii)$$

This is called ***ideal gas equation,***

or $$PV = \frac{w}{M}RT \quad ...(iv)$$

where w = mass of the gas in grams

M = Molecular weight of the gas

4. Dalton's Law of Partial Pressures: This law states that, *"the total pressure developed by a mixture of non-reacting gases is equal to the sum of the partial pressure exerted by the individual gases."*

Mathematically, $P = p_1 + p_2 + p_3 + ...$

where P is the total pressure and p_1, p_2, p_3, etc. are the partial pressures of the different non-reacting gases.

The partial pressure of a gas means *the pressure exerted by that gas when it is the present alone in the same vessel at the same temperature.*

$$\text{Partial pressure} = \frac{\text{Initial volume of the gas} \times \text{Initial pressure}}{\text{Total volume}}$$

$$= \text{Mole fraction of the gas} \times \text{Total pressure}$$

Dalton's law may be employed for calculating the percentage composition of a gaseous mixture according to the following relation.

$$\%\text{ of a gas in the mixture} = \frac{\text{Partial pressure of the gas} \times 100}{\text{Total pressure}}$$

In terms of the number of moles,

$$P = \frac{(n_1 + n_2 + n_3 ...)}{V} RT$$

where n_1, n_2, n_3 are the number of moles of different gases present in the mixture.

When a gas is collected by downward displacement of water, it becomes moist due to the water vapours present in it. These water vapours also exert their own pressure at the same temperature of the gas. **"This partial pressure of the water vapours is known as aqueous tension."** So, in order to get the pressure of the dry gas, aqueous tension should be subtracted from the observed pressure of the gas.

P wet gas = P dry gas + P water vapour

so P dry gas = P wet gas – P water vapour

= Observed pressure – Aqueous tension.

5. Graham's Law of Diffusion: It states that, "*under similar conditions of temperature and pressure the rate of diffusion of a gas "r" is inversely proportional to the square root of its density 'd' or molecular weight "M"*.

$$\text{Mathematically,} \quad r \propto \frac{1}{\sqrt{d}} \qquad \text{or} \qquad r \propto \frac{1}{\sqrt{M}}$$

If r_1 and r_2 are the rates of diffusion of two gases and their densities are d_1 and d_2 respectively, then

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$$

If M_1 and M_2 are the respective molecular weights of the gases, then vapour density = Mol. wt/2

$$\therefore \qquad \frac{r_1}{r_2} = \sqrt{\frac{M_2/2}{M_1/2}} = \sqrt{\frac{M_2}{M_1}}$$

4.6

To study the relationship between the rate of diffusion of a gas and its molar mass by using H_2 and CO_2 gases.

Diffusion of Hydrogen and Air:

1. Set up the apparatus as shown in Fig. 4.10 but without the dilute hydrochloric acid. Note the levels of the *coloured water* in arms A and B of the U-tube.

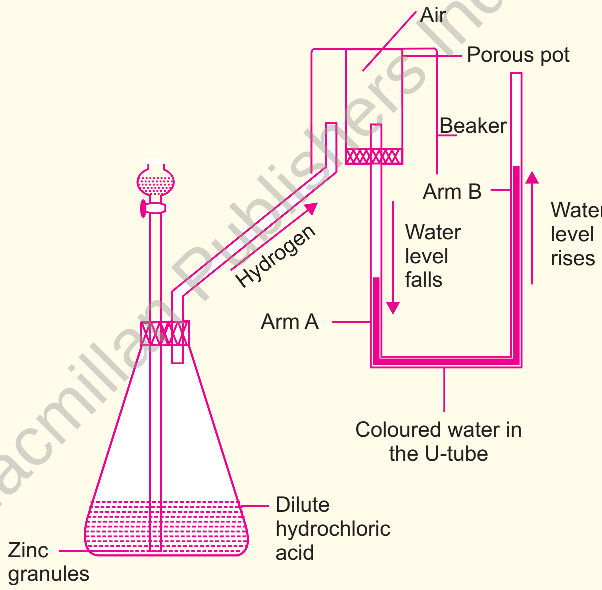


Fig. 4.10 *H_2 diffuses faster than air*

2. Add some dilute hydrochloric acid to the zinc metal. Note the levels of the coloured water in the U-tube as hydrogen gas is produced.
   (*i*) Explain your observation.
3. Stop the hydrogen supply by removing the conical flask. Then remove the beaker. Note the levels of the *coloured water* again.
   (*ii*) Explain your observation.
   Molar mass of hydrogen z g mol^{-1}. Gas is lighter than air. Lowering of water level indicates that diffusion of hydrogen into the porous pot is faster than the diffusion of air out of the porous pot. The more accumulation of hydrogen increases the pressure inside the pot and the water levels is pushed down.

6. Gay-Lussac's Law: *Whenever gases take part in a reaction, as reactants or products, they do so in volumes which bear a simple ratio to one another provided all the volumes are measured at the same temperature and pressure, e.g.*

$$\begin{array}{ccccc} H_2 & + & Cl_2 & \longrightarrow & 2HCl \\ \text{1 volume} & & \text{1 volume} & & \text{2 volume} \end{array}$$

Ratio = 1 : 1 : 2 (a simple ratio).

The mathematical relationship between pressure and temperature in known as Gay-Lussac's Law. It state that at constant volume, pressure of a fixed amount of a gas directly proportional to the temperature.

$$P \propto T$$

$$\frac{P}{T} = K$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

7. Avogadro's Law: According to this law, "*under similar conditions of pressure and temperature, equal volumes of all gases contain equal number of molecules.*"

Mathematically, $V \propto n$ (at constant temperature and pressure) where n is the total number of molecule in volume V.

Molar Volume: 1 gram-mole (Molecular weight in grams) of a gaseous substance occupies 22.4 litres or d (22400×10^{-6} m^3) at NTP. It is known as **molar volume.**

Avogadro's Number: It is the number of formula units present in one gram mole (mol. wt. in grams) of a substance. It value is 6.023×10^{23}. It is denoted by N or N_A.

One gram-mole of any substance contains 6.023×10^{23} molecules.

Similarly 1 gram-atom (atomic weight in grams) of any element contains 6.023×10^{23} atoms.

SOLVED EXAMPLES

Example 1. A sample of gas occupies 100 dm^3 at 1 bar pressure and at °C. If the volume of the gas is reduced to 5 dm^3 at the same temperature, what additional pressure must be applied?

Sol. From the given data: $P_1 = 1$ bar, $P_2 = ?$

$V_1 = 100\ dm^3$, $V_2 = 5\ dm^3$

Since the temperature is constant, Boyle's law can be applied

$$P_1V_1 = P_2V_2 \qquad P_2 = \frac{P_1V_1}{V_2}$$

$$P_2 = \frac{(1\ \text{bar}) \times (100\ dm^3)}{(5\ dm^3)} = 20\ \text{bar}$$

$\therefore$ Additional pressure applied $= 20 - 1 = 19$ bar

Example 2. A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 2.27 L volume, up to what volume can the balloon be expanded?

Sol. According to Boyle's law $P_1V_1 = P_2V_2$

If P_1 is 1 bar, V_1 will be 2.27 L

If $P_2 = 0.2$ bar,

then $V_2 = \frac{P_1V_1}{P_2}$

$\Rightarrow$ $V_2 = \frac{1 \text{ bar} \times 2.27 \text{ L}}{0.2 \text{ bar}}$

$= 11.35 \text{ L}$

Since balloon bursts at 0.2 bar pressure, the volume of balloon should be less than 11.35 L.

Example 3. On a ship sailing in pacific ocean where temperature is 23.4°C, a balloon is filled with 2 L air. What will be the volume of the balloon when the ship reaches Indian ocean, where temperature is 26.1°C?

Sol. $V_1 = 2 \text{ L}$ $\quad T_2 = 26.1 + 273$

$T_1 = (23.4 + 273) \text{ K}$ $\quad = 299.1 \text{ K}$

$= 296.4 \text{ K}$

From Charles's law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$\Rightarrow$ $V_2 = \frac{V_1T_2}{T_1}$

$\Rightarrow$ $V_2 = \frac{2\text{L} \times 299.1 \text{ K}}{296.4 \text{ K}}$

$= 2 \text{ L} \times 1.009$

$= 2.018 \text{ L}$

Example 4. A flask has a capacity of 500 ml. What volume of air escapes if the flask is heated from 25°C to 35°C and pressure remains constant?

Sol. Initial volume = 500 ml, Initial temperature = 25 + 273 = 298° K.

Final temperature = 35 + 273 = 308°K

Final volume = ?

According to Charles's law, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

$\therefore$ $V_2 = \frac{T_2 \times V_1}{T_1} = \frac{500 \times 308}{298}$

$= 516.8 \text{ ml}$

Out of this, 500 ml air can remain in the flask and the remaining 16.8 ml will escape out.

Example 5. What is the volume of 6 grams of hydrogen at 1.5 atm and 273°C?

Sol. Pressure, $P = 1.5 \text{ atm}$ $\quad V = ?$

$R = 0.082$ $\quad T = 273 + 273 = 546 \text{ K}$

From the ideal gas equation,

$PV = nRT$

n = number of moles of hydrogen

$= \frac{\text{mass of hydrogen}}{\text{molecular wt.}} = \frac{6}{2} = 3 \text{ moles}$

Substituting these values in $PV = nRT$

$1.5 \times V = 3 \times 0.082 \times 546$

$$V = \frac{3 \times 0.082 \times 546}{1.5}$$

$$= 89.544 \text{ litres.}$$

Example 6. **Calculate the temperature of 2.0 moles of a gas occupying a volume of 5.0 litres of 2.46 atmospheres.**

Sol. Given,

$$P = 2.46 \text{ atm} \qquad V = 5.0 \text{ litres}$$

$$R = 0.082 \text{ lit-atm degree}^{-1} \text{ mole}^{-1}$$

$$n = 2.0$$

Using the ideal gas equation,

$$PV = nRT$$

$$\therefore \quad T = \frac{PV}{nR} = \frac{2.46 \times 5}{2 \times 0.082} = 75°A$$

$$\therefore \quad \text{Temperature} = 75 - 273$$

$$= \mathbf{-198°C}$$

Example 7. **A gaseous mixture in a cylinder comprises 70% N_2, 25% O_2 and 5% CO_2 by volume. If the pressure of the gaseous mixture is 760 mm Hg, calculate the partial pressure of each gas.**

Sol. Since the volume of a gas at a constant temperature and pressure is proportional to the number of moles of the gas, the volume percentage of a gas in a mixture reflects its mole fraction.

$$\therefore \quad xN_2 = \frac{70}{100} = 0.70, \; xO_2 = \frac{25}{100} = 0.25 \text{ and } x_{CO_2} = \frac{5}{100} = 0.05$$

$$\therefore \quad p_{N_2} = x_{N_2}p = 0.70 \times 760 \text{ mm Hg} = 532 \text{ mm Hg},$$

$$p_{O_2} = x_{O_2}p = 0.25 \times 760 \text{ mm Hg} = 190 \text{ mm Hg and}$$

$$p_{CO_2} = x_{CO_2}p = 0.05 \times 760 \text{ m Hg} = 38 \text{ mm Hg}$$

Example 8. **Determine the partial pressure of each gas and the total pressure of a mixture comprising 6 g of H_2, 32 g of O_2 and 56 g of N_2, if the mixture is present in a 5 L container at 27°C.**

Sol. First, we convert the mass of each gas into the number of moles. Then, we calculate the partial pressure of each, using the ideal gas equation. Lastly, we add the partial pressure to obtain the total pressure.

$$n_{H_2} = 6 \text{ g} \times \frac{1 \text{ mol}}{2\text{g}} = 3 \text{ mol}$$

$$\Rightarrow \quad p_{H_2} = n_{H_2} \times \frac{RT}{V} = 3 \text{ mol} \times \frac{0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{5 \text{ L}}$$

$$= 14.78 \text{ atm}$$

$$n_{O_2} = 32 \text{ g} \times \frac{1 \text{ mol}}{32 \text{ g}} = 1 \text{ mol}$$

$$\Rightarrow \quad p_{O_2} = n_{O_2} \times \frac{RT}{V} = 1 \text{ mol} \times \frac{0.0821 \text{ L atm K}^{-1}\text{mol}^{-1} \times 300 \text{ K}}{5 \text{ L}}$$

$$= 4.93 \text{ atm}$$

$$n_{N_2} = 56 \text{ g} \times \frac{1 \text{ mol}}{28\text{g}} = 2 \text{ mol}$$

$$\Rightarrow \qquad p_{N_2} = n_{N_2} \times \frac{RT}{V} = 2 \text{ mol} \times \frac{0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{5 \text{ L}}$$

$$= 9.85 \text{ atm}$$

$$\therefore \qquad p_T = p_{H_2} + p_{O_2} + p_{N_2}$$

$$= 14.78 \text{ atm} + 4.93 \text{ atm} + 9.85 \text{ atm}$$

$$= 25.56 \text{ atm}$$

Example 9. **Compare the rates of diffusion of $^{235}UF_6$ and $^{238}UF_6$.**

Sol. Let the rate of diffusion of $^{235}UF_6$ be r_1 and rate of diffusion of $^{238}UF_6$ be r_2.

Now, molecular mass of $^{235}UF_6$ $(M_1) = 235 + 6 \times 19 = 349$

Molecular mass of $^{238}UF_6$ $(M_2) = 238 + 6 \times 19 = 352$

According to Graham's law $\quad \dfrac{r_1}{r_2} = \sqrt{\dfrac{M_2}{M_1}}$

$$\frac{r_1}{r_2} = \sqrt{\frac{352}{349}} = 1.004.$$

Thus, $r(^{235}UF_6) : r(^{238}UF_6)$ is 1.004 : 1.

Example 10. **A certain gaseous fluoride of phosphorus has a formula PF_x. Under similar conditions fluorine diffuses 1.82 times faster than the gaseous phosphorus fluoride. Find the value of x and formula of phosphorus fluoride. (Atomic masses, P = 31.0; F = 19.0).**

Sol. According Graham's law

$$\frac{r_{F_2}}{r(PF_x)} = \frac{1.82}{1} = \sqrt{\frac{M_{(PF_x)}}{M_{(F_2)}}}$$

$$M_{(F_2)} = 2 \times 19 = 38$$

$$\frac{1.82}{1} = \sqrt{\frac{M_{(PF_x)}}{38}}$$

or $$M_{(PF_x)} = \frac{1.82 \times 1.82 \times 38}{1}$$

$$= 125.8 \qquad ...(i)$$

Molecular mass of PF_x from formula = $31 + 19x$ ...(*ii*)

Equate (*i*) and (*ii*)

$$31 + 19x = 125.8$$

or $$x = \frac{125.8 - 31}{19}$$

$$= 4.98 = 5$$

$\therefore$ The formula of phosphorus fluoride = $\mathbf{PF_5}$.

Example 11. **An iron cylinder contains helium at a pressure of 250 kPa at 300 K. The cylinder can withstand a pressure of 1×10^6 Pa. The room in which cylinder is placed catches fire. Predict whether the cylinder will blow up before it melts or not, melting point of cylinder = 1800 K.**

Sol. Given, $\quad P_1 = 250$ kPa, $\quad T_1 = 300$ K,

$P_2 = 1 \times 10^6$ Pa, $\quad T_2 = ?$

Since volume of cylinder remains constant.

$$\because \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\therefore \quad \frac{250 \times 10^3}{300} = \frac{1 \times 10^6}{T_2}$$

$$\therefore \quad T_2 = 1200 \text{ K}$$

Example 12. A cylinder containing cooking gas can withstand a pressure of 14.9 atmosphere. The pressure gauge of the cylinder indicates 12 atmosphere at 27°C. Due to sudden fire in the building, temperature starts rising. At what temperature will the cylinder explode?

Sol. Since gas is confined in a cylinder, its volume will remain constant

Initial conditions	*Final conditions*
$P_1 = 12$ atm	$P_2 = 14.9$ atm
$T_1 = 273 + 27 = 300$ K	$T_2 = ?$

At constant volume

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

or

$$T_2 = \frac{P_2 T_1}{P_1}$$

$$= \frac{14.9 \times 300}{12.0}$$

$$= 372.5 \text{ K}$$

or Temperature at which cylinder explodes = 372.5 – 273 = 99.5°C.

SELF EVALUATION

1. A gas occupies 200 ml at a pressure of 0.820 bar at 20°C. How much volume will be occupy when it is subjected to external pressure at 1.025 bar at the same temperature?
2. A sample of gas is found to occupy a volume of 900 cm^3 at 27°C. Calculate the temperature at which it will occupy a volume of 300 cm^3, provided the pressure is kept constant?
3. The gas left in a used aerosol can is at a pressure of 1 atm at 27°C (room temperature). If this can is thrown on to a fire, what is the internal pressure of the gas when its temperature reaches 927°C?
4. At 25°C and 760 mm Hg pressure a gas occupies 600 ml volume. What will be is pressure at a height where temperature is 10°C and volume of the gas is 640 ml?
5. A neon-dioxygen mixture contains 70.6 g dioxygen and 167.5 g neon. If pressure of the mixture of gases in the cylinder is 25 bar. What is the partial pressure of dioxygen and neon is the mixture?
6. Two gases A and B have densities 0.09 g dm^{-3} at 1.45 g dm^{-3} respectively. What is the ratio of their rates of diffusion under similar conditions?

RECAPITULATION

1. A complete chemical equation gives the symbolic representation of the reactants, products and their physical states.
2. A chemical equation is balanced if the number of atoms of each type involved in a chemical reaction on the reactant and product sides of the equation are the same. Equations have to be balanced always. They are balanced according to the law of conservation of mass.
3. In all the physical and chemical changes, the total mass of the reactant is equal to that of the products.
4. When different amounts of the two reactants are mixed, the reactant which reacts completely is called the limiting reactant. The amounts of the products formed depend upon the limiting reactant.
5. A chemical equation has both qualitative as well as quantitative significance.
6. A chemical equation conveys the following:
   (*i*) Number of moles of reactants or products.
   (*ii*) Number of atoms of each element.
   (*iii*) Number of molecules of each reactant and product.
   (*iv*) Number of grams of each reactant and product.
   (*v*) Ratio by weight of different reactants and products.
   (*vi*) Volumes of reactants and products, if gases, under similar conditions.
7. A chemical equation does not tell us about:
   (*i*) Time required for the completion of the reaction.
   (*ii*) The conditions of the reactions.
   (*iii*) Physical state of the reactants and the products.
   (*iv*) Whether heat is absorbed or evolved.
   (*v*) Whether reaction is fast or slow.
8. The combination reactions are opposite to that of decomposition reactions. In a decomposition reaction a single substance gets decomposed to give two or more substances.
9. In a combination reaction, two or more substances react with each other to form a new single substance.
10. Reactions in which energy gets absorbed are called endothermic reactions.
11. When an element displaces another element from its compound, a displacement reaction takes place.
12. In double displacement reactions two different atoms or groups of atoms (ions) are exchanged.
13. Oxidation is the loss of hydrogen or gain of oxygen.
14. Reduction is the gain of hydrogen or loss of oxygen.
15. Stoichiometry is the quantitative study of reactant and products in a chemical reaction. A chemical equation uses chemical symbols to show what happen during a chemical reaction.
16. Chemical reaction is a process in which substance is changed into one or more new substance.
17. One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g of the ^{12}C isotope.

18. Law of conservation of mass is also called the law of indestructibility of matter.
19. **Boyle's law:** The volume of the fixed mass of a gas is inversely proportional to its pressure at constant temperature.
20. **Charles's law:** The volume of the fixed mass of a gas is directly proportional to the temperature at constant pressure.
21. **Avogadro's law:** The equal volume of all the gases under similar conditions of temperature and pressure contain equal number of molecules.
22. **Dalton's law of partial pressure:** At constant temperature, the pressure exerted by a mixture of two or more than two non reacting gases enclosed in a definite volume, is equal to the sum of the individual pressure which each gas would exert it.
23. **Graham's law:** Under similar conditions of temperature and pressure, the rate of diffusion of gases are inversely proportional to the square roots of their densities.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. Rusting of iron is a chemical change.
2. Hydrogen sulphide gas burns in air to give water and sulphur.
3. A balanced chemical equation tells us about the ratio of the reaction.
4. Atomic mass unit is the actual mass of a single atom.
5. $KClO_3 \longrightarrow KCl + O_2$ is a balanced reaction.
6. The mass of one mole of particles is called molar mass.
7. Avogadro's law belong to the relationship between volume and temperature.

II. Fill in the Blanks

Fill in the blanks by using suitable word(s):

1. Burning of magnesium is an ____________ chemical reaction.
2. ____________ is the gain of hydrogen or loss of oxygen.
3. The chemical equation does not indicate the ____________ of the reactants and products.
4. The substances which can bring about the chemical change are known as ____________.
5. A ____________ equation is one which contains an equal number of each element on both sides of the equation.
6. The number of entities contained in molar mass is equal to ____________.
7. The law describing relationship between pressure and volume at constant temperature is called ____________.

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. The substances which take part in any chemical reaction are called
   (*a*) products (*b*) reactants (*c*) metals (*d*) non-metals
2. $NH_3 + HCl \longrightarrow$
   (*a*) N_2 (*b*) H_2 (*c*) NH_4Cl (*d*) No reaction
3. The substance which helps in oxidation is called
   (*a*) reducing agent (*b*) oxidising agent
   (*c*) redox (*d*) none of these
4. Law of conservation of mass is called
   (*a*) law of indestructibility of matter (*b*) Avogadro's law
   (*c*) law of constant proportion (*d*) none of these
5. Rusting of iron is a
   (*a*) fast reaction (*b*) normal reaction (*c*) slow reaction (*d*) none of these
6. 22.4 L of H_2 gas as STP is equivalent to
   (*a*) 1 mole (*b*) 1 gm
   (*c*) 6.022×10^{23} molecules (*d*) 0.01 mol
7. The volume of a fixed mass of the gases a given pressure should become zero when its temperature becomes
   (*a*) –273.15 K (*b*) –273.15°C (*c*) 0°C (*d*) 273.15 K

IV. Matching

Match the items of Column I with the corresponding items of Column II:

Column I	Column II
1. Burning of wax	(*a*) Standard for atomic mass
2. Intermixing of two gases	(*b*) White ppt
3. Quick ime	(*c*) Fast reaction
4. Chemical change	(*d*) Oxidising agent
5. Silver chloride	(*e*) Cannot be reversed
6. Carbon–12	(*f*) CaO
7. $K_2Cr_2O_7$	(*g*) Diffusion

V. Answer the Following Questions

1. What is a balanced equation?
2. What are the limitations of a chemical equation? How are they removed?

3. What is the law called which deals with the ratios of the gaseous reactants and products?
4. State the law of conservation of mass. How is it verified experimentally?
5. (*i*) Differentiate between a chemical reaction and chemical symbol.
   (*ii*) What information can be derived from the symbol C and the formula CO_2?
6. Write the notes on the following:
   (*i*) Redox reaction
   (*ii*) Displacement reaction
   (*iii*) Double displacement reaction
7. State the principle on which balancing of a chemical equation is primarily based.
8. What do the symbols *g*, *l*, *aq* and $\rightleftharpoons$ in the following equation signify?

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

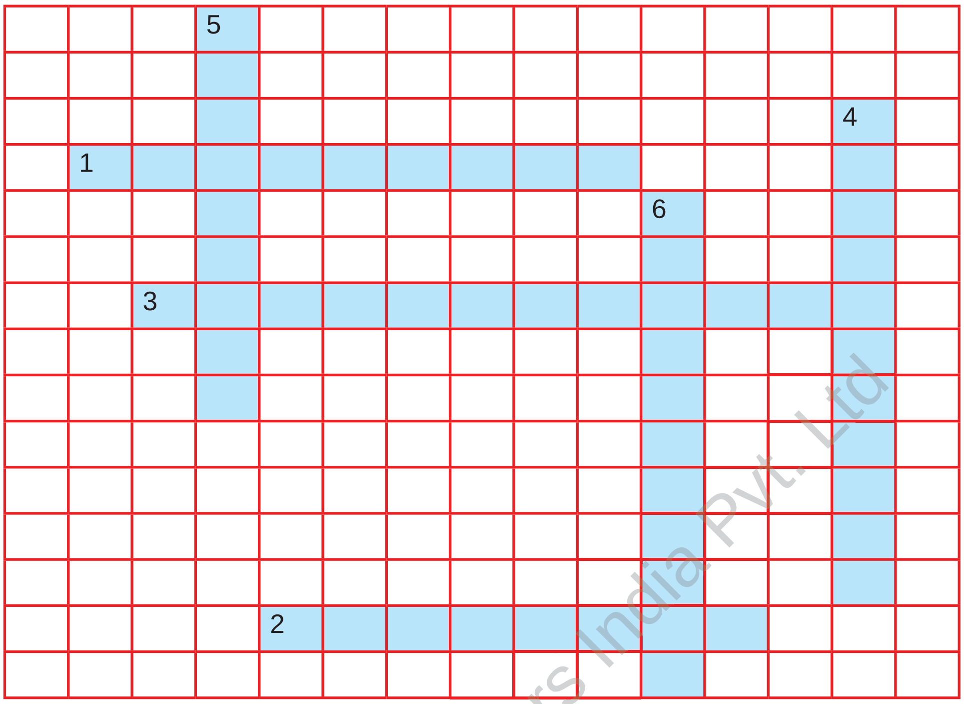
9. Balance the following equations:
   (*i*) $I_2 + HNO_3 \longrightarrow HIO_3 + NO_2 + H_2O$
   (*ii*) $P + HNO_3 \longrightarrow H_3PO_4 + NO_2 + H_2O$
   (*iii*) $Zn + HNO_3 \longrightarrow Zn(NO_3)_2 + NO_2 + H_2O$
10. The law of conservation of mass is not different from the law of conservation of atoms. Explain.
11. What is the significance of the equation?

$$CaCO_3\ (s) + 2HCl\ (l) \longrightarrow CaCl_2\ (aq) + H_2O\ (l) + CO_2\ (g)$$

12. What information is conveyed by a chemical equation?
13. Translate the following into language of chemistry and balance them:
   (*i*) Sodium hydroxide + Aluminium sulphate $\longrightarrow$ Sodium sulphate + Sodium meta aluminate + Water
   (*ii*) Aluminium + Sulphuric acid (conc.) $\longrightarrow$ Aluminium sulphate + Sulphur dioxide + Water
   (*iii*) Lead acetate + Sodium chromate $\longrightarrow$ Lead chromate + Sodium acetate
   (*iv*) Mercuric phosphate + Hydrogen sulphide $\longrightarrow$ Mercuric sulphide + Phosphoric acid
   (*v*) Potassium permanganate + Hydrochloric acid $\longrightarrow$ Potassium chloride + Manganese chloride + Water + Chlorine
14. Define mole with example.
15. What is the mass of one mole electrons? Mass of one electron is 9.1×10^{-31} kg ?
16. Explain the law of mass of conservation with example?
17. Suggest experiment to compare the rate of diffusion of ammonia and hydrogen chloride. What conclusion do you get from the experiment.
18. State Dalton's law of partial pressure. Describe the utility of the law.
19. Explain the ideal gas equation.

VI. Puzzle

Read the given clues carefully and write the appropriate word inside the boxes to solve the puzzle.



Across:

1. Gain of hydrogen or loss of oxygen.
2. Symbolic representation of reactants, products and their physical state.
3. Reaction in which one element displaces another element from its compound.

Down:

4. Reactions in which energy gets absorbed.
5. Loss of hydrogen or gain of oxygen.
6. Reactions opposite to that of decomposition reactions.

[Hints: 1. Reduction 2. Equation 3. Displacement 4. Endothermic 5. Oxidation 6. Combination]

Web Links:

For more information on the following topics, visit the websites listed below:

Chemical Change
http://www.dictionary.com/browse/chemical-change

Chemical Equation
https://www.thoughtco.com/what-is-a-chemical-equation-604026

Balancing Chemical Equation
http://www.wikihow.com/Balance-Chemical-Equations

Types of Chemical Equations
https://www.thoughtco.com/types-of-chemical-reactions-604038

Law of Conservation of Mass
https://www.thoughtco.com/definition-of-conservation-of-mass-law-604412

CHAPTER 5

RATE OF REACTION AND ENERGY TRANSFER

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Understand the rates of reaction
  - Give examples of different chemical reactions where there is a great variation in the rates at which these reactions take place.
  - Describe alteration of the rates of reactions by varying temperature or concentration, or by changing the surface area of a solid reactant, or by adding a catalyst.
- Understand energy transfer in reactions
  - Classify reactions as exothermic reaction or endothermic reaction depending on the temperature change that takes place during the course of the reaction.

5.1 INTRODUCTION

You have already learnt about chemical reactions and their various types in your previous classes. It is interesting to know that all the reactions do not take the same time to complete. In other words, they proceed with different rates. Burning of candle wax is very fast, whereas rusting of iron in moist air takes longer time and is very slow. Thus, a chemical reaction can be a slow reaction, a moderate reaction or a fast reaction, depending on how much time it takes to complete.

In this chapter, you will study about the rates of chemical reactions, factors which affect the rate of a chemical reaction, heat changes during a chemical reaction and energy transfer in a chemical reaction.

5.2 SLOW AND FAST REACTIONS

The reactions which take relatively less time to complete are called **fast reactions**. Burning of fuels such as kerosene, petrol, C.N.G. (compressed natural gas), L.P.G. (liquefied petroleum gas), paper, wood, wax, etc. are very common examples of fast chemical reactions. Let us consider some more examples.

(*i*) Reaction between hydrogen and oxygen under specific conditions is instantaneous, that is, it takes place immediately and takes no time in completion.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

(*ii*) Sodium metal reacts with water to form sodium hydroxide and hydrogen gas. The reaction is very fast.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

So, much heat is evolved in the reaction that hydrogen catches fire. It burns in air to form water vapour, which is also a fast reaction.

(*iii*) Burning of magnesium metal in air is also a fast reaction in which magnesium oxide is formed.

$$2Mg + O_2 \text{ (from air)} \longrightarrow 2MgO$$

There are many more such reactions.

The reactions which take longer time to complete are called **slow reactions**. Formation of coke from dead remains of trees and formation of petroleum from the dead remains of sea animals are examples of extremely slow reaction, as they take several thousand years to complete. Some more examples of slow chemical reactions are given below:

(*i*) Rusting of iron is a slow process. It takes several days for completion, depending on the environmental conditions. When iron is exposed to moist air it is very slowly converted into brown coloured iron oxide ($Fe_2O_3 \cdot xH_2O$), which is called rust. You must have seen rusted iron articles, which get a brownish look.

(*ii*) Fermentation of grape juice into alcohol is another example of slow chemical reaction. It takes several days.

(*iii*) Formation of green coating over copper metal is also an example of slow chemical reaction. When copper articles are exposed to moist air a dull green coating of basic copper carbonate, $CuCO_3 \cdot Cu(OH)_2$ is formed.

Thus, we find there are fast reactions as well as slow reactions. This nature of chemical reactions depends on their rate of reaction. Let us know what is the rate of reaction.

A 5.1

Take a note of changes taking around you and classify them on the basis of slow and fast reactions and discuss it with your teacher.

5.3 COLLISION THEORY

According to collision theory, the reaction between two atoms or molecules takes place because of their collisions. A chemical reaction takes place only if molecules of reactants collide with each other. The total number of collisions taking place among the reacting species in unit time is called **collision frequency.** The collisions which result in the chemical reaction are called effective collisions. So for a chemical reaction two main factors are responsible.

(*i*) **Activation Energy:** The colliding molecules must possess certain minimum energy (threshold energy) to make the collision effective. The additional energy required by the molecules to attain threshold energy is called activation energy.

(*ii*) Proper orientation of the reacting species.

5.4 RATE OF A CHEMICAL REACTION

You have now an idea that different chemical reactions take different time for completion. In other words, they proceed with different rates. When a chemical reaction takes place, the reactants are consumed and the products are formed.

The amount of the reactants consumed, or products formed is described in terms of the rate of reaction, also known as **speed of reaction**. Thus, we can say that, the **rate of reaction** is defined in terms of the change in the amount of the reactants or the change in the amount of the products in a given time interval.

Thus, the rate of reaction is defined as the rate of change of concentration of a reactant or a product.

5.4.1 Rate of Reaction in Terms of Concentration of Reactant

The change in concentration of a reactant with time is the reaction rate.

$$\text{Reactant} \longrightarrow \text{Product}$$

$$A \longrightarrow B$$

$$\text{Rate of Reaction} = -\frac{\text{Change in concentration of reactant}}{\text{Time}}$$

$$= -\frac{\text{Final concentration of reactant} - \text{Initial concentration of reactant}}{\text{Final time} - \text{Initial time}}$$

$$= -\frac{\Delta\,[\text{Reactants}]}{\Delta t} = \frac{-\Delta\,[A]}{\Delta t}$$

Here, it is important to note that concentration of a reactant decreases with time during a chemical reaction and hence final concentration is less than the initial concentration. Since the rate of a chemical reaction cannot be negative, a minus sign (–) is placed in the expression to get the positive value of reaction rate.

5.4.2 Rate of Reaction in Terms of Concentration of Product

The change in concentration of a product with time is the reaction rate for the equation.

$$\text{Reactant} \longrightarrow \text{Product}$$

$$A \longrightarrow B$$

$$\text{Rate of Reaction} = \frac{\text{Change in concentration of product}}{\text{Time}}$$

$$= \frac{\text{Final concentration of product} - \text{Initial concentration of product}}{\text{Time}}$$

$$= \frac{\Delta\,[\text{Product}]}{\text{Time}} = \frac{\Delta\,[B]}{\Delta t}$$

The unit for the rate of reaction depends on two parameters—concentration of the substance and time. Concentration of any reactant or the product may be given as:

(*i*) number of moles per litre of solution, or

(*ii*) number of gram equivalent per litre of solution

The SI unit of time is second(s), but depending upon the nature of reaction (very fast, fast, slow, very slow, etc.) time may be expressed in minutes, hours, days or sometimes in years. For example, if the change in concentration is in moles per litre and time is measured in minutes, then the unit for the rate of reaction is given as moles per litre per minute or mol L^{-1} min^{-1}.

5.5 FACTORS AFFECTING THE RATE OF CHEMICAL REACTIONS

A chemical reaction occurs as a result of collision between the reacting molecules. Any factor which affects the collision rate will influence the reaction rate.

The following factors affect the rate of a chemical reaction.

1. **Nature of the Reactants and the Products:** The rate of a chemical reaction depends upon the nature of the reactants and the products. Ionic reactions involving ions are fast, whereas reactions involving molecules of reactants having covalent bonds are relatively slow. For example, when a solution of silver nitrate is added to a solution of sodium chloride, the reaction is completed immediately and a white precipitate of silver chloride is obtained.

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

This ionic reaction may be given in the following form:

$$AgNO_3 \longrightarrow Ag^+ + NO_3^- \text{ (in solution)}$$

$$NaCl \longrightarrow Na^+ + Cl^- \text{ (in solution)}$$

$$Ag^+\,(aq) + Cl^-\,(aq) \longrightarrow AgCl\,(s) \text{ (fast reaction)}$$

$$6CO_2\,(aq) + 6H_2O(l) \xrightarrow[\text{chlorophyll}]{\text{Sunlight}} \underset{\text{Glucose}}{C_6H_{12}O_6}\,(aq) + 6O_2 \text{ (slow reaction)}$$

2. **Concentration of Reactants:** In general, if the concentration of reactants is more then the rate of reactions will be high. However, there are reactions in which the rate of reaction is independent of the concentration of the reactants. For example,

$$H_2 + I_2 \rightleftharpoons 2HI$$

If more H_2 or I_2 is added to the reaction mixture at equilibrium, the equilibrium will shift in the forward direction till it is re-established. At this new equilibrium, the concentration of the reactants will decrease whereas that of the product will increase than the original concentration.

> Low light intensity lowers the rate of photosynthesis. As the intensity is increased the rate also increases. Since carbon dioxide and water molecules both contain covalent bonds, they require large amount of energy for breaking the bonds. So, photosynthesis is a slow process.

3. **Effect of Temperature and Pressure:** In general, increase in temperature increases the rate of reaction. For example, if you take some calcium carbonate in a test tube and add dilute hydrochloric acid to it, evolution of carbon dioxide is shown by the formation of effervescence. If you heat the test tube the effervescence will increase showing an increase in the reaction rate.

Similarly, it has been observed that the reaction between gaseous molecules takes place at pressure higher than atmospheric pressure. For example, nitrogen gas (N_2) reacts with hydrogen gas (H_2) at higher pressure (200 – 300 atmosphere) and 700 K temperature to form ammonia gas (NH_3).

> The change in temperature alters the state of equilibrium for only those reactions in which either heat is evolved or heat is absorbed. Exothermic reactions are favoured by low temperature whereas endothermic reactions are favoured by high temperature.

4. **Effect of Catalyst:** A catalyst is used to alter the rate of reaction. For example, in the formation of ammonia from nitrogen and hydrogen, as stated above, finely divided iron as catalyst and molybdenum as promoter is used. For example,
   (*i*) Catalyst is a substance which when introduced to a chemical reaction increases the rate of reaction. It neither reacts with the reactants nor with the products. It is obtain as it is.
   (*ii*) Catalyst which increases the rate of reaction is called positive catalyst and the one which decreases the rate of reaction is called negative catalyst.

A 5.2

Both catalysts A and B catalyse the decomposition of hydrogen peroxide. The following figures were obtained at 20°C for the volume of oxygen formed against the time since the start of the reaction.

Time	*0*	*5*	*10*	*15*	*20*	*25*	*30*	*35*
Volume of O_2 with catalyst A (cm^3)	0	4	8	12	16	17	18	18
Volume of O_2 with catalyst B (cm^3)	0	5	10	15	16.5	18	18	18

(*a*) Plot a graph to show both sets of results.
(*b*) Say which is better A or B.
(*c*) Add a line to your graph to show the shape of the graph you would obtain for the uncatalysed reaction.

5. **Surface Area of the Reactant:** Greater the surface area of the reactants, greater will be the rate of reaction. For example, wood shavings burn with greater speed than a log of wood, as the surface area is much more in case of wood shaving.

A 5.3

Objective:

To show that the rate of reaction depends upon temperature as well as on the surface area of the reactants.

Theory:

Reaction of a piece of marble with dilute hydrochloric acid takes place at ordinary temperature with the evolution of carbon dioxide.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

If the reactants are heated then the evolution of carbon dioxide increases, showing that the rate of reaction has increased. Similar observation is found if instead of marble piece, marble powder is taken for the reaction.

Material:

Marble piece, marble powder, test tubes (2), dilute HCl, heating device (may be a candle or spirit lamp, if gas burner is not available).

Procedure:

(*i*) Take a clean test tube and slide a piece of marble in it, taking all the precautions not to break the test tube.
(*ii*) Add about 5 mL dilute HCl and observe the speed with which bubbles of gas are evolving.
(*iii*) Heat the test tube and again observe the speed of the evolution of gas bubbles.
(*iv*) Take another test tube and add marble powder approximately equal to the piece of marble taken in the first test tube. Add 5 mL dilute HCl to it and observe the speed of gas bubbles evolving in the reaction.

Observations:

(*i*) Dilute acid reacts with marble piece to form a gas which comes out in the form of bubbles.
(*ii*) On heating more bubbles are evolved.
(*iii*) If marble powder is taken in place of marble piece then also the speed of evolution of gas increases.

Conclusion:

Increase in temperature and increase in surface area (marble powder) increase the speed of a reaction.

6. **Effect of Light:** Certain chemical reactions take place in presence of sunlight. For example, when chlorine gas reacts with methane in presence of sunlight the hydrogen atoms of methane (CH_4) are replaced one by one to give various products. The ultimate product is carbon tetrachloride (CCl_4).

$$CH_4 + Cl_2 \xrightarrow{\text{Sunlight}} CH_3Cl + HCl$$

$$CH_3Cl + Cl_2 \xrightarrow{\text{Sunlight}} CH_2Cl_2 + HCl$$

$$CH_2Cl_2 + Cl_2 \xrightarrow{\text{Sunlight}} CHCl_3 + HCl$$

$$CHCl_3 + Cl_2 \xrightarrow{\text{Sunlight}} CCl_4 + HCl$$

Photosynthesis in plants is another example, where sunlight plays an important part in the preparation of food for the plants using CO_2 and H_2O molecules.

5.6 REVERSIBLE AND IRREVERSIBLE CHEMICAL REACTIONS

The chemical reactions which take place in both forward as well as in reverse direction are known as **reversible reactions.** In such reactions the products formed can give back the reactants.

For example, when steam (H_2O) is passed over red hot iron (Fe), iron oxide (Fe_3O_4) and hydrogen gas (H_2) are formed.

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

If hydrogen gas is passed over hot iron oxide, then iron and steam are formed.

$$Fe_3O_4 + 4H_2 \longrightarrow 3Fe + 4H_2O$$

The reaction follows a reverse path of the earlier reaction. In other words, the reactants of the first reaction are the products of the second reaction and vice versa. The two reactions together are written as,

$$3Fe + 4H_2O \underset{\text{Reverse reaction}}{\overset{\text{Forward reaction}}{\rightleftharpoons}} Fe_3O_4 + 4H_2 \ (\ell)$$

The double half arrow symbol ($\rightleftharpoons$) shows that the chemical reaction may proceed in both the directions. The upper half arrow represents forward reaction and the lower half arrow represents reverse reaction.

Some more examples of reversible reactions are given here.

(*i*) Formation of hydrogen iodide from hydrogen and iodine.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$

(*ii*) Formation of ammonia from nitrogen gas and hydrogen gas.

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

(*iii*) Hydrolysis of the ester, ethyl ethanoate in water.

$$\underset{\text{Ethyl ethanoate}}{CH_3COOC_2H_5(aq)} + \underset{\text{Water}}{H_2O(\ell)} \rightleftharpoons \underset{\text{Ethanoic acid}}{CH_3COOH(aq)} + \underset{\text{Ethanol}}{C_2H_5OH(aq)}$$

(*iv*) Decomposition of calcium carbonate ($CaCO_3$).

$$\underset{\text{Calcium carbonate}}{CaCO_3(s)} \rightleftharpoons \underset{\text{Calcium oxide}}{CaO(s)} + CO_2(s)$$

This reaction is reversible only when it is carried out in a closed container.

The reactions which take place only in one direction are called **irreversible reactions**. In such reactions the products formed do not give back the original reactants. Such reactions are represented by placing an arrow (→) between the reactants and the products. Arrow shows the direction of the chemical reaction.

Formation of silver chloride (AgCl) by adding the solution of silver nitrate ($AgNO_3$) to the solution of common salt (NaCl) is an example of irreversible chemical reaction. Some more such examples are given here.

(*i*) Burning of carbon in oxygen or air.

$$C + O_2 \longrightarrow CO_2$$

(*ii*) Replacement of copper (Cu) from the blue solution of copper sulphate ($CuSO_4$) by iron metal (Fe).

$$\underset{\text{Copper sulphate}}{CuSO_4\,(aq)} + Fe\,(s) \longrightarrow \underset{\text{Iron (II) sulphate}}{FeSO_4\,(aq)} + Cu\,(s)$$

(*iii*) Formation of calcium carbonate by passing CO_2 gas in lime water.

$$\underset{\text{Lime water}}{Ca(OH)_2} + CO_2 \longrightarrow \underset{\text{Calcium carbonate}}{CaCO_3} + H_2O$$

(*iv*) Formation of a white precipitate of barium sulphate ($BaSO_4$), when a dilute sulphuric acid is added to a solution of barium chloride ($BaCl_2$).

$$\underset{\text{Barium chloride}}{BaCl_2\,(aq)} + H_2SO_4\,(aq) \longrightarrow \underset{\text{Barium sulphate}}{BaSO_4\,(s)} + 2HCl\,(aq)$$

You can find many more reversible and irreversible chemical reactions.

5.7 CHEMICAL EQUILIBRIUM AND ITS DYNAMIC NATURE

In our everyday life, we come across many physical processes such as evaporation of water, melting of ice, dissolution of salts, etc. Let us observe what happens when evaporation of water occurs in a closed vessel at room temperature. It starts evaporating. As the process continues more and more water molecules escape from liquid to vapour state and vapour pressure of water starts increasing. After some time vapour pressure becomes constant and evaporation appears to stop, but actually this does not happen. Now water vapours start condensing which is a reverse process of evaporation.

A stage comes when the rate of evaporation of water becomes equal to the rate of condensation of water vapour. This state of the system is known as equilibrium state. This is also observed in many chemical reactions. In case of chemical reaction the state of equilibrium is known as chemical equilibrium.

In irreversible chemical reaction, the reaction proceeds only in the forward direction. The reaction is completed and no reactants are left behind. However, the reversible reactions do not reach the state of completion. When these reactions are performed in closed containers, a stage is reached when the reaction appears to stop even though some reactants are there. At this stage both reactants and products are present together with some definite concentrations. The system is said to be in the state of chemical equilibrium.

Thus, the chemical equilibrium is defined as the state of a reversible reaction when the concentrations of reactants and the products become constant and do not change with time.

It is also found that this happens when the rate of forward reaction is equal to the rate of reverse reaction. Remember reactions do not stop at equilibrium, but the rates of forward and reverse reactions become equal.

Thus, at the state of chemical equilibrium,

Rate of forward reaction = Rate of reverse reaction

To make the concept more clear let us take the example of formation of hydrogen iodide by the combination of hydrogen (H_2) and iodine (I_2). The reaction is allowed to take place in a closed container at constant temperature.

The forward reaction is shown by the equation

$$H_2\ (g) + I_2\ (g) \longrightarrow 2HI\ (g)$$

The reverse reaction is shown by the equation

$$2HI\ (g) \longrightarrow H_2\ (g) + I_2\ (g)$$

At equilibrium the reaction is shown by the equation

$$H_2\ (g) + I_2\ (g) \rightleftharpoons 2HI\ (g)$$

The change in reaction rates with time is shown in Fig. 5.1. This shows the time (T_e) at which equilibrium is reached.

Iodine vapours are dark violet in colour and the product HI is colourless. If the reaction is not reversible the violet colour of iodine will disappear after some time. But it does not happen. The violet colour persists, though the intensity is diminished. This clearly shows that the reaction is reversible. At the point of equilibrium, formation of HI and its decomposition are taking place with the same rate.

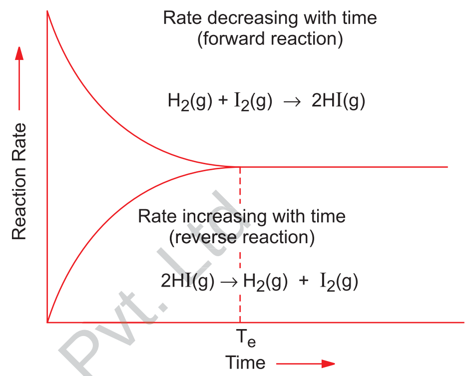


Fig. 5.1 *Change in reaction rate with time*

Similarly, the chemical equilibrium for the reaction between H_2 and N_2 forming ammonia is given as

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

Now we can say that the chemical equilibrium is dynamic in nature. The reactions, both forward and reverse never stop, only their rates become equal at equilibrium.

SELF EVALUATION

1. Write some fast and slow reactions, which are applicable in everyday life.
2. Surjya Pradhan is a student of class IX. He is preparing oxygen gas in laboratory by heating $KClO_3$. His father suggested him to use MnO_2 as a catalyst to get higher yield of oxygen readily.
   (*i*) How catalyst affect the rate of reaction?
   (*ii*) Mention two values exhibited by his father.
3. What would happen if every collision between reactant molecules were to form product?
4. What is reversible reaction? Give one example.

5.8 ENTHALPY CHANGE

Enthalpy is defined as the sum of internal energy (E) plus the product of pressure (P) and volume (V).

$$H = E + P \times V$$

All the reactions taking place are accompanied by either evolution or absorption of heat, i.e. change in enthalpy (Enthalpy change) takes place. It is denoted by 'Δ' pronounced as delta meaning change.

ΔH is negative for exothermic reactions indicating evolution of heat whereas it is positive for endothermic reaction indicating absorption of heat, i.e. heat is supplied to the reactants for reaction to take place.

A 5.4

Answer the following questions based on the graph (Fig. 5.2).

(*a*) ΔH is ____________ for ex thermic reaction and ____________ for endothermic reaction.

(*b*) Heat is released in ____________ reactions and heat is absorbed in ____________ reactions.

(*c*) Energy of reactants is higher in ____________ reactions and energy of products is higher in ____________ reactions.

(*d*) Give examples of endothermic and exothermic reactions based upon enthalpy change (apart from that given in the textbook).

5.9 HEAT OF REACTION

The heat of reaction or enthalpy of reaction is a general term used for heat change accompanying any reaction. However, depending upon the nature of the reaction (i.e., combustion, neutralization, etc.) the heat of reaction is named accordingly.

Let us consider the following two examples:

(*i*) $CH_4\,(g) + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(l);\ \Delta H = -\,890.4\ \text{kJ mol}^{-1}$

(*ii*) $C\,(s) + H_2O\,(g) \longrightarrow CO\,(g) + H_2\,(g);\ \Delta H = +\,131.4\ \text{kJ mol}^{-1}$

The first equation indicates that when 1 mole of methane combines completely with 2 moles of oxygen gas, 890.4 kJ of heat is produced. Similarly, the second equation tells that when 1 mole of solid carbon reacts completely with 1 mole of steam 131.4 kJ of heat is absorbed.

Enthalpy of reaction, ΔH = Sum of enthalpies of product – Sum of enthalpies of reactants

$$= \sum a_1 H\ (\text{Products}) - \sum b_1 H\ (\text{Reactants})$$

The amount of heat evolved or absorbed in a chemical reaction when the number of moles of the reactants as represented by the chemical equation have completely reacted, is called the **heat of reaction.**

5.10 THERMOCHEMICAL EQUATION

A balanced chemical equation not only indicates the quantities of the different reactants and products but also the amount of heat evolved or absorbed is called a **thermochemical equation.**

Fractional coefficients are allowed in writing a thermochemical equation. For example, the formation of water is written as

$$H_2\ (g) + \frac{1}{2}\ O_2\ (g) \longrightarrow H_2O\ (l);\quad \Delta H = -\ 285.8\ \text{kJ}$$

Thus, 285.8 kJ of heat is produced when 1 mole of hydrogen reacts with 0.5 mole of oxygen. If the quantities of reactants are doubled, the heat produced will also be doubled. For example, in the above case, we may write

$$2H_2\ (g) + O_2\ (g) \longrightarrow 2H_2O\ (l);\qquad \Delta H = -\ 571.6\ \text{kJ}$$

5.10.1 Exothermic and Endothermic Reactions

Exothermic reactions are those reactions which are accompanied by the evolution of heat.

The quantity of heat produced is shown along with the products with a 'plus' sign. A few examples of exothermic reactions are given below:

$$C\ (s) + O_2\ (g) \longrightarrow CO_2\ (g) + 393.5\ \text{kJ}$$

$$H_2\ (g) + \frac{1}{2}\ O_2\ (g) \longrightarrow H_2O\ (g) + 285.8\ \text{kJ}$$

$$CH_4\ (g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2H_2O\ (g) + 890.4\ \text{kJ}$$

Endothermic reactions are those reactions which are accompanied by absorption of heat.

Heat absorbed can be written with the products with a minus sign. A few examples of endothermic reactions are given below:

$$N_2\ (g) + O_2\ (g) \longrightarrow 2NO\ (g) - 180.7\ \text{kJ}$$
$$C\ (s) + H_2O\ (g) \longrightarrow CO\ (g) + H_2\ (g) - 131.4\ \text{kJ}$$
$$C\ (s) + 2S\ (s) \longrightarrow CS_2\ (g) - 92.0\ \text{kJ}$$

The enthalpy change (ΔH) accompanying a reaction is given by

$$\Delta H = \text{Heat content of products} - \text{Heat content of reactants}$$
$$= H_P - H_R$$

A reaction is exothermic, if the total heat content of the reactants is more than that of the products, i.e. $H_R > H_P$. Thus, ΔH will be negative for an exothermic process.

Thus, the exothermic reactions given above may be written in terms of ΔH as:

$$C\ (s) + O_2\ (g) \longrightarrow CO_2\ (g); \qquad \Delta H = -393.5\ \text{kJ}$$
$$H_2\ (g) + \frac{1}{2}\ O_2\ (g) \longrightarrow H_2O\ (l); \qquad \Delta H = -285.8\ \text{kJ}$$
$$CH_4\ (g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2H_2O\ (l); \Delta H = -890.4\ \text{kJ}$$

A reaction is endothermic if the total heat content of the reactants is less than that of the products, i.e.

$$H_R < H_P$$

Thus, ΔH will be positive for an endothermic reaction.

Endothermic reactions given above may be written in terms of ΔH as:

$$N_2\ (g) + O_2\ (g) \longrightarrow 2NO\ (g); \quad \Delta H = +180.7\ \text{kJ}$$
$$C\ (s) + H_2O\ (g) \longrightarrow CO\ (g) + H_2\ (g); \quad \Delta H = +131.4\ \text{kJ}$$
$$C\ (s) + 2S\ (s) \longrightarrow CS_2\ (l); \quad \Delta H = +92.0\ \text{kJ}$$

Exothermic and endothermic reactions may be represented graphically as shown in Fig. 5.2(*a*) and (*b*).

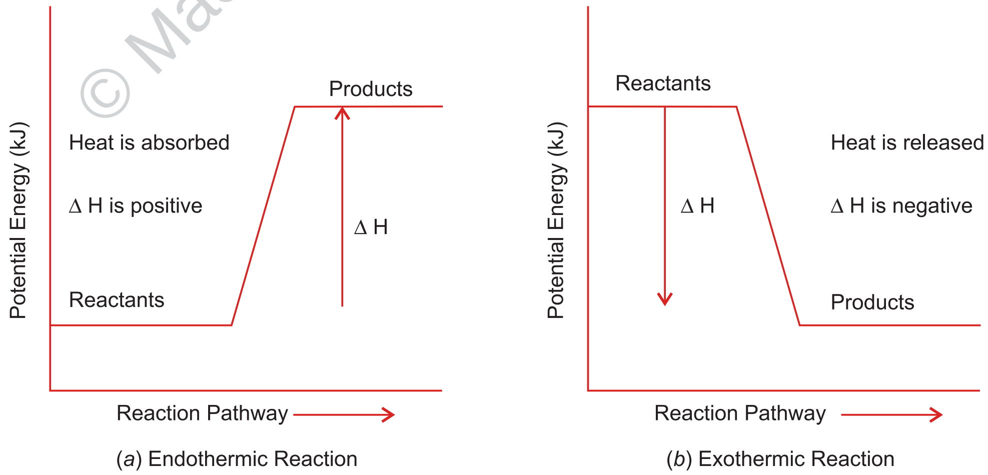


Fig. 5.2 *(a) Endothermic reaction, (b) Exothermic reaction*

A

5.5

I. Objective: To identify exothermic and endothermic reactions.

Apparatus: Test tube.

Procedure:

1. Take water in a beaker and dissolve sodium hydroxide solution in it (swirl to dissolve).
2. Take water in a beaker and dissolve potassium nitrate in it (swirl to dissolve).
3. Take water in a beaker and dissolve calcium oxide in it.
4. Take water in a beaker and dissolve ammonium chloride in it.

Observations:

1. The beaker heats up.
2. The beaker cools down.
3. A large amount of heat is evolved.
4. After some time the beaker cools down.

Conclusion:

1. Heating of beaker shows that it is an exothermic reaction with evolution of heat.
   $NaOH + H_2O \longrightarrow Na^+ + OH^- + H_2O + \text{Heat}$
2. Cooling down of beaker shows that, it is an endothermic reaction as the heat is absorbed.
   $KNO_3 + H_2O \longrightarrow K^+ + NO^{-}_{3} + H_2O$
3. A large amount of heat is evolved showing that it is an exothermic reaction.
   $CaO + H_2O \longrightarrow Ca(OH)_2$
4. It shows that it is an endothermic reaction.
   $NH_4Cl + H_2O \longrightarrow NH_3 + HCl + H_2O$

Precautions:

1. Handle solutions with care.
2. Carry out reactions away from your body.

II. Complete the following:

1. $NaOH$ + ________ $\longrightarrow$ ________ + ________ + H_2O + Heat
2. $CaO + H_2O \longrightarrow$ ____________________
3. ________ + $H_2O \longrightarrow K^+ + NO_3^- + H_2O$
4. $NH_4Cl + H_2O \longrightarrow$ ________ + ________ + H_2O

III.

Activities	*Chemical Reaction*	*Exothermic or Endothermic*
1. Dissolution of sodium hydroxide	$NaOH + H_2O \longrightarrow Na^+ + OH^- + H_2O + \text{Heat}$	____________
2. Dissolution of potassium nitrate	$KNO_3 + H_2O \longrightarrow K^+ + NO_3^- + H_2O - \text{Heat}$	____________
3. Dissolution of calcium oxide	$CaO + H_2O \longrightarrow Ca(OH)_2 + \text{Heat}$	____________
4. Dissolution of ammonium chloride	$NH_4Cl + H_2O \longrightarrow NH_4OH + HCl - \text{Heat}$	____________
5. Reaction of iron fillings with solution of copper sulphate and potassium sulphate	$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu + \text{Heat}$ $Fe + K_2SO_4 \longrightarrow$ No reaction	____________

Question: Give your answer with explanation for each activity. Also, balance the equations.

5.11 ENERGY TRANSFER IN A CHEMICAL REACTION

Thermodynamics is the science of the relation between heat, work and the properties of the system. Just as all currencies are not same, all forms of energy are also not the same.

When our study is confined to chemical changes and chemical substances only, the restricted branch of thermodynamics is known as **chemical thermodynamics**.

5.11.1 Laws of Thermodynamics

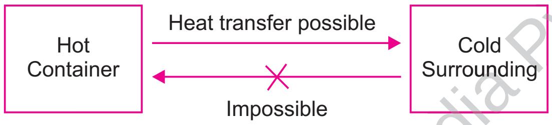
First Law of Thermodynamics

- The first law of thermodynamics is an extension of the law of conservation of energy.
- It states that the change in internal energy of a system is equal to the heat added to the system minus the work done by the system.

$$\Delta U = Q - W$$

Second Law of Thermodynamics

It states that processes occur in a certain direction and that energy has quality as well as quantity.



5.11.2 Importance of Thermodynamics

(*i*) It helps us to predict whether any given chemical reaction can occur under the given set of conditions.

(*ii*) It helps in predicting the extent of reaction before the equilibrium is attained.

5.11.3 Limitations of Thermodynamics

(*i*) It does not tell anything about the rate at which the process takes place.

(*ii*) It deals only with the initial and final states of a system but does not tell anything about the mechanism of the process.

(*iii*) It does not tell anything about the individual atoms and molecules.

Some basic terms used in thermodynamics are explained below:

5.11.4 Open, Closed and Isolated System

System: The part of the universe chosen for thermodynamic consideration, (i.e. to study the effect of temperature, pressure, etc.) is called *system*.

Surroundings: The remaining portion of the universe is called *surroundings*.

Thus, system and surroundings together constitute the universe, i.e. Universe = System + Surroundings.

A system usually consists of a definite amount of one or more substances and is separated from the surroundings by a real or imaginary boundary through which matter and energy can flow the system to the surroundings or vice versa.

(*a*) Open System: A system is said to be open if both mass (matter) and energy (heat or work) can be freely exchanged with the surroundings. For example, if some ice is kept in an open beaker or if some reaction is allowed to take place in a beaker (e.g. between a piece of marble and HCl), where exchange of both matter and energy take place, then the system is known as open system.

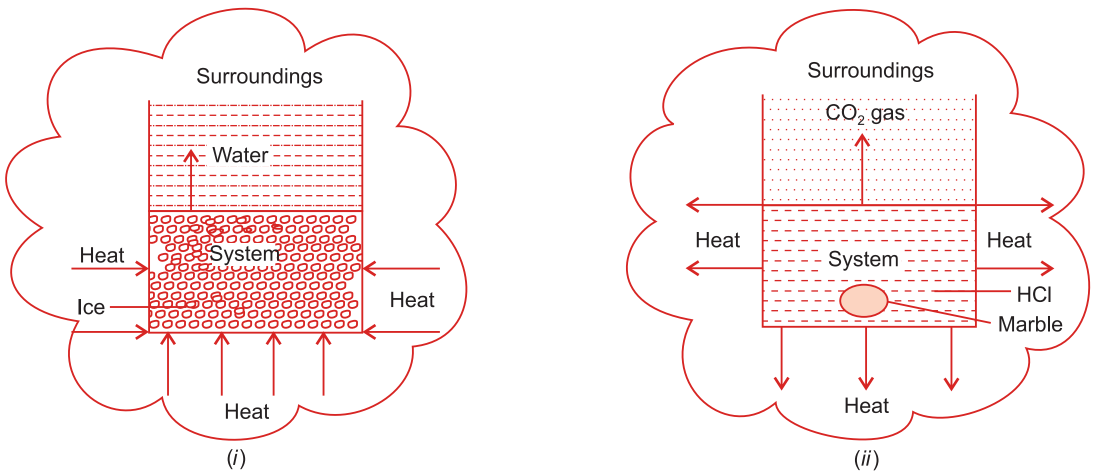


Fig. 5.3 *Open systems*

(*b*) **Closed System:** A system is said to be closed if there is only the exchange of energy (heat or work) with the surroundings and no exchange of mass takes place. For example, if some ice is placed in a closed beaker then as the beaker is closed no exchange of matter between the system and surroundings takes place but the exchange of energy takes place due to conducting walls. Such system is known as closed system.

(*c*) **Isolated System:** A system is said to be isolated if there is neither exchange of mass nor of energy (heat or work) with the surroundings. For example, if ice is placed in a thermos flask which is closed as well as insulated, so no exchange of matter and energy takes place between the system and the surroundings. Such system is known as isolated system.

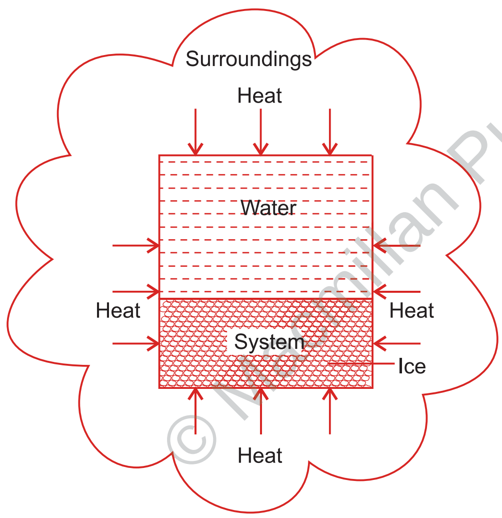


Fig. 5.4 *Closed system*

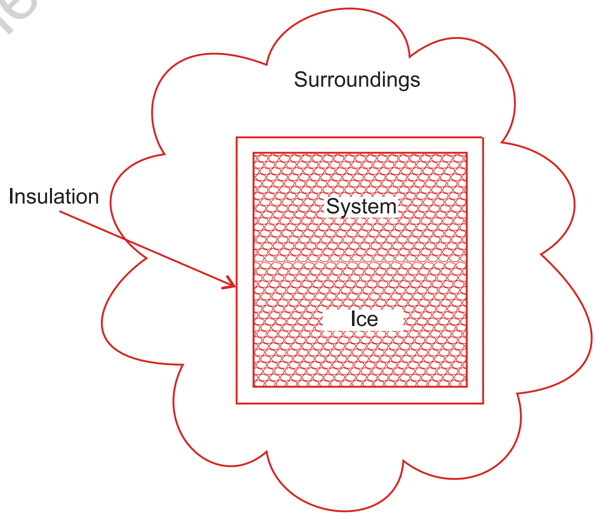


Fig. 5.5 *Isolated system*

5.12 ENTROPY

The degree of randomness is called entropy. It is denoted by "S".

Entropy order in different matter is as below:

Solid → Liquid → Gas

SELF EVALUATION

1. The reaction shown below is exothermic:

   Anhydrous copper (II) sulphate + Water $\longrightarrow$ Copper (II) sulphate·5 water

   Describe or illustrate an experiment by which you could find out whether this statement is true.
2. Is photosynthesis exothermic or endothermic? What are the reactants and what are the products? Illustrate your answer.
3. Give an example of each (*a*) exothermic and (*b*) endothermic reactions which are of vital importance in our everyday life. Explain why the reactions you mentioned are so important.

RECAPITULATION

1. Rate of a chemical reaction is the change in the concentration of the reactants or products.
2. The time taken for completion of reaction decides, whether it is a slow reaction or a fast reaction.
3. Factors which affect the rate of reaction are— nature of reactants and products, concentration of reactants, temperature and pressure, presence of a catalyst, surface area of the reactants, effect of light.
4. In the reversible reactions, the products also react to form the reactants, but in irreversible reactions this does not happen.
5. In case of reversible reaction, a stage comes when the rate of forward reaction becomes equal to the rate of reverse reaction. This is called the state of equilibrium.
6. At the state of equilibrium in a reversible reaction both forward and reverse reactions continue though their rates are same. This shows the dynamic nature of the reversible reactions.
7. The total number of collisions taking place among the reacting species in unit time is called collision frequency.
8. The collisions which result in the chemical reaction are called effective collisions.
9. The minimum energy required to collide the reacting molecules in a chemical reaction is called threshold energy.
10. The additional energy required by the molecules to attain threshold energy is called activation energy.
11. The portion of the universe which is under investigation, e.g. portion of a test tube where reaction takes place, is called system.
12. Everything else in the universe except system is called surroundings, e.g. except the portion of the test tube where reaction takes place is the surrounding, e.g. above and around the test tube.
13. The system which can exchange matter and energy with the surrounding is called an open system, e.g. a spoonful of tea is an open system because it will become cold as well as its taste will change, e.g. it is exchanging energy and matter with the surroundings.
14. Closed system can exchange energy but not matter with the surrounding, e.g. tea kept in a closed kettle. It will become cold but its taste will not change.
15. Isolated system can neither exchange heat nor matter with the surroundings, e.g. tea placed in thermos flask.
16. Temperature is a measure of degree of hotness or coldness. It is measured in Celsius scale, e.g. we have the ice point (0°C) and the steam point (100°C) for water with an interval between them divided into 100 parts called degrees.
17. The process in which heat is given out to the surroundings is called exothermic process. In this process, products are more stable than the reactants because they have lower energy.

18. The process in which heat is absorbed by the system from the surroundings is called endothermic process. In this process, products are less stable than reactants because they have higher energy.
19. Heat of reaction may be defined as the amount of heat evolved or absorbed, e.g. enthalpy change, when the quantities of reactants have reacted completely as indicated by balanced chemical equation.
20. Heat of formation is defined as the amount of heat evolved or absorbed when one mole of the compound is formed from its constituting elements in their standard states.
21. Heat of combustion may be defined as the amount of heat evolved when one mole of a substance is completely oxidised or burnt in excess O_2.
22. Heat of fusion is defined as heat absorbed when 1 mole of a solid changes into liquid completely at its melting point.
23. Heat of vaporisation is defined as heat absorbed when 1 mole of a liquid changes into its vapours completely at its boiling point.
24. Heat of sublimation is defined as the heat absorbed when 1 mole of a solid changes into its vapours completely.
25. Heat of neutralisation is defined as the heat evolved when 1 mole of H^+ of an acid is neutralised completely by 1 mole of OH^- of a base in a very dilute solution. In other words, heat evolved when 1 mole of H^+ combines with one mole of OH^- to form water.
26. A balanced chemical equation together with the designation of its value of heat and the physical states of reactant and product is called thermochemical equation.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. The rate of reaction decreases with time.
2. Catalysts do not alter the state of chemical equilibrium.
3. The rate of reaction depends upon the surface area.
4. Thermodynamics give the information about the rate of reaction.
5. Heat evolved in any chemical reaction is called endothermic.

II. Fill in the Blanks

Fill in the blanks by using suitable word(s):

1. The rate of a chemical reaction depends on the nature of the ____________.
2. During a chemical reaction the concentration of the reactant ____________.
3. Rusting of iron is a ____________.
4. Heat of combustion is always ____________.
5. Solid on heating changes directly into gaseous state. This is called ____________.

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. The rate of a chemical reaction depends upon
   (*a*) time (*b*) pressure (*c*) concentration (*d*) all of these
2. A catalyst
   (*a*) alters the reaction mechanism (*b*) decreases the activation energy
   (*c*) increases the frequency of collision of reacting species
   (*d*) increases the average k netic energy of reacting species

3. Action of catalyst depends on
   (*a*) mass (*b*) solubility (*c*) particle size (*d*) none of these
4. The rate constant depends upon
   (*a*) temperature (*b*) initial concentration
   (*c*) time (*d*) none
5. The energy that favours the dissolution of solute in water is known as
   (*a*) hydration energy (*b*) lattice energy (*c*) affinity energy (*d*) bond energy
6. $N_2(g) + 3H_2(g) \longrightarrow 2NH_3 - 93.7$ kJ is a/an
   (*a*) exothermic reaction (*b*) endothermic reaction
   (*c*) combustion reaction (*d*) displacement reaction
7. The heat of neutralization of a strong acid by a strong base is always
   (*a*) constant (*b*) infinity
   (*c*) less than 13.7 kJ mole^{-1} (*d*) more than 17.5 kJ mole^{-1}
8. Which of the following is endothermic?
   (*a*) Decomposition of water (*b*) Dehydrogenation of ethane to ethene
   (*c*) Graphite to diamond (*d*) All of these

IV. Matching

Match the items of Column I with the corresponding items of Column II:

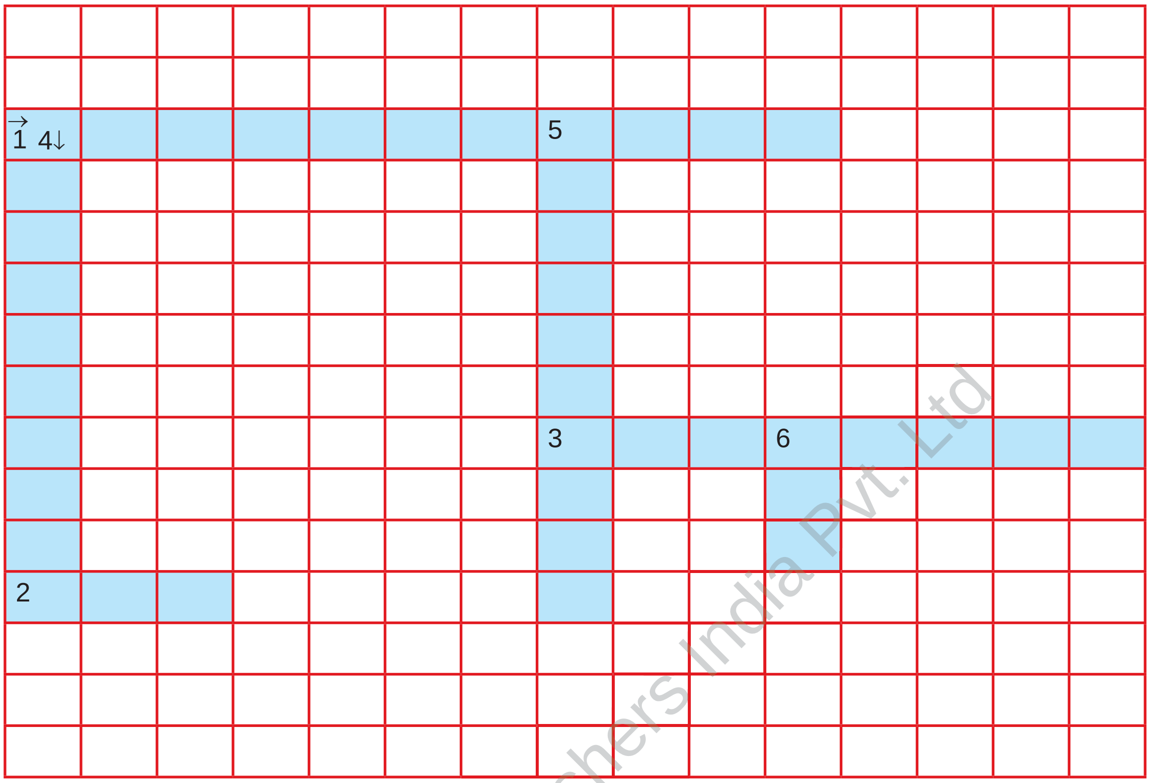
Column I	Column II
1. Mol L^{-1} s^{-1}	(*a*) Endothermic reaction
2. Human body	(*b*) Rate of reaction increases
3. Fermentation of grape juice into alcohol	(*c*) Temperature
4. Degree of hotness and coldness	(*d*) Unit of rate constant
5. Rate of chemical reaction is	(*e*) Open system
6. Surface area	(*f*) Change in concentration

V. Answer the Following Questions

1. What is rate of reaction? Discuss in brief slow and fast reactions.
2. Explain the activation energy.
3. A lump of coal burns at a moderate rate in air while coal dust burns explosively. Why?
4. Why does liquid bromine react slowly as compared to bromine vapour?
5. Explain the effect of sunlight on chemical reaction.
6. The reaction between solution of silver nitrate and solution of sodium chloride is not reversible. Explain.
7. What is an open system?
8. What are exothermic and endothermic process? Give examples.
9. Define the following:
   (*i*) Heat of neutralization
   (*ii*) Heat of sublimation
   (*iii*) Heat of vaporisation
10. If you pump air into cycle tyre, a slight warming effect is noticed at the valve stem, why?

VI. Puzzle

Read the given clues carefully and write the appropriate word inside the boxes to solve the puzzle.



Across:

1. The state when rate of forward reaction becomes equal to the rate of backward reaction.
2. A fuel which is used in automobiles.
3. A system in which no exchange of matter and energy takes place.

Down:

4. A reaction in which heat is evolved.
5. The chemical reactions which take place in both forward as well as reverse direction.
6. It is the short form of household fuel.

[Hints: 1. Equilibrium 2. CNG 3. Isolated 4. Exothermic 5. Reversible 6. LPG]

Web Links:

For more information on the following topics, visit the websites listed below:

Slow and Fast Reactions

http://www.chemicool.com/chemistry/chemical-reaction-rates/

Rate of Chemical Reactions

https://brilliant.org/wiki/rate-of-chemical-reactions/

Factors Affecting Rate of Chemical Reactions

https://www.thoughtco.com/factors-that-effect-chemical-reaction-rate-609200

Chemical Equilibrium

https://www.chemguide.co.uk/physical/equilibmenu.html

Thermochemical Equation

https://en.wikipedia.org/wiki/Thermochemical_equation

CHAPTER

6

GREEN CHEMISTRY

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Describe the nitrogen cycle
  - Describe the different parts of the nitrogen cycle.
  - Explain the importance of converting nitrogen to ammonia for agriculture.
  - Explain the manufacturing of nitrogenous fertilizers and their effect on plant growth.
  - Explain the environmental consequences of the over use of chemical fertilizers.
- Describe the carbon cycle and global climate change
  - Describe that the Earth's atmosphere and oceans have changed over time.
  - Describe the role of carbon cycle in maintaining the atmospheric composition.
  - Explain that the burning of fossil fuels can upset the balance of the carbon cycle resulting to global climate change.
- Understand the problems of plastics
  - State the environmental hazards caused by polymers.
  - State a number of measures to prevent environmental hazards caused by polymers e.g. how polymers can be processed, the development of biodegradable plastics and the removal of toxic products during the disposal of halogenated plastics (e.g. PVC).

6.1 INTRODUCTION

The branch of chemistry that deals with the processes and products that reduce or eliminate the use and generation of hazardous substances is called **green chemistry.**

- Green chemistry is also known as sustainable chemistry.
- It applies to all areas of chemistry.
- It applies innovative scientific solutions to real world environmental problems.
- It reduces the negative impact of chemical products and processes on human health and environment.

All living things are made of carbon. Carbon is also a part of the ocean, air, and even rocks. In the atmosphere, carbon is combined with atmospheric oxygen and forms a gas called carbon dioxide.

Plants use this carbon dioxide and sunlight to make their own food by the process of **photosynthesis**. The carbon becomes part of the plant. Plants that die and get buried may turn into fossil fuels made of carbon such as coal and oil over millions of years. When humans burn fossil fuels, most of the carbon quickly enters the atmosphere as carbon dioxide.

Carbon dioxide is an atmospheric constituent that plays several vital roles in the environment. It plays a crucial role in the weathering of rocks. It is the carbon source for plants. It is stored in biomass, organic matter, in sediments and in carbonate rocks such as limestone.

6.2 NITROGEN CYCLE

Nitrogen is a necessary constituent of all living matter. It is essential for the growth of animals and plants. Animals and plants cannot assimilate free nitrogen. Animals obtain their supply of nitrogen from plants and the plants secure it from nitrogen compounds present in the soil. Thus, nitrogen forms a permanent link between the animal and the plant kingdom.

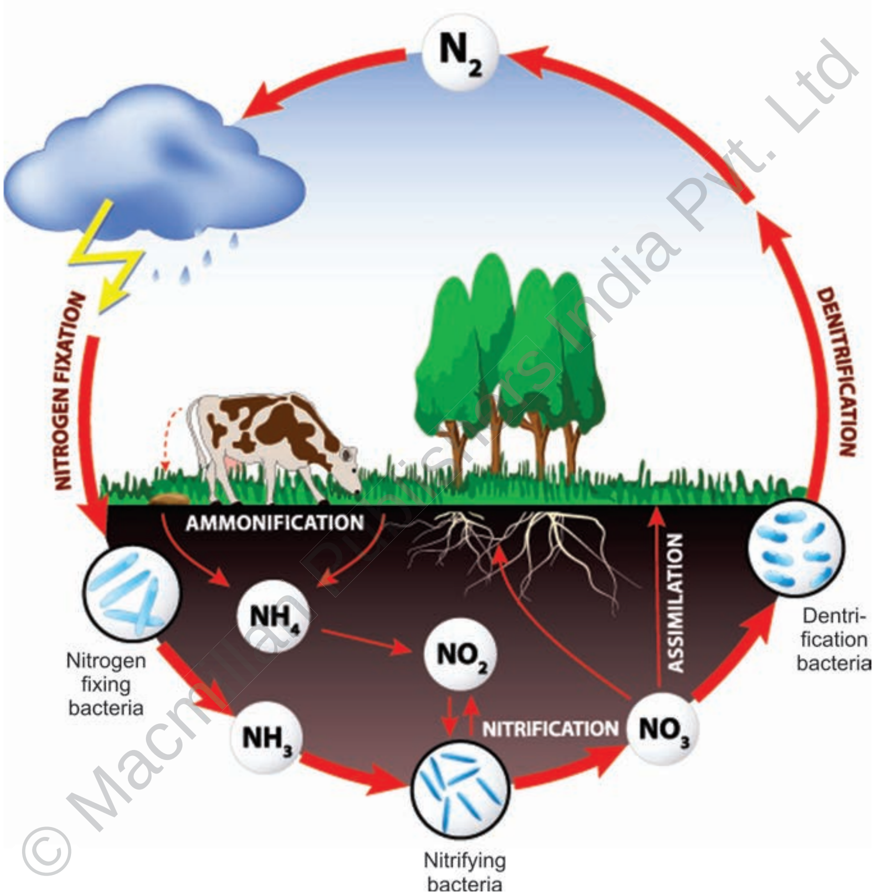


Fig. 6.1 *Nitrogen cycle*

The main source of nitrogen is air and it is to the air that all nitrogen finally returns. The conversion of atmospheric nitrogen into useful compounds for animals and plants and the breaking up of nitrogenous compounds into free nitrogen takes place continuously in nature. This whole process is known as **nitrogen cycle**.

The circulation of nitrogen through the living and non-living components of the biosphere (air, soil, water, plants and animals) is called **nitrogen cycle**. Nitrogen cycle maintains the percentage of nitrogen in the atmosphere, more or less at a constant.

6.2.1 Steps Involved in Nitrogen Cycle

1. The atmospheric nitrogen is fixed into nitrogen compounds such as nitrates by *Rhizobium* bacteria, blue-green algae, and lightning.
2. The plants absorb nitrate compounds from the soil and water and convert them into plant proteins.
3. The plants are consumed up by animals and thus plant proteins are used for making animal proteins.
4. The soil nutrients are drained away from the soil by the action of rain water. This is known as **leaching**. The drained off nutrients accumulate in water bodies such as lakes and ponds, causing an excess of nutrients in these water bodies. This is called **eutrophication**.
5. When the plants and animals die, the putrefying bacteria and fungi present in the soil decompose the proteins of dead plants and animals into ammonia. This process is called **ammonification**.

6.3 NITROGEN FIXATION

The process of converting atmospheric nitrogen into ammonium, nitrates, and nitrites which can be easily used by plants is called **nitrogen fixation**.

$$N_2 \longrightarrow NH_4^+$$

Nitrogen fixation is the process wherein N_2 is converted to ammonium. It is the only way through which organisms can attain nitrogen directly from the atmosphere. Thus, it is very essential.

6.3.1 Denitrification

$$2NO_3^- + 12H^+ \longrightarrow N_2 + 6H_2O$$

Denitrification is an anaerobic process that is carried out by denitrifying bacteria, which convert nitrate to dinitrogen in the following sequence.

$$NO_3^- \longrightarrow NO_2^- \longrightarrow NO \longrightarrow N_2O \longrightarrow N_2$$

In nature, nitrogen of the atmosphere can be fixed by the following methods:

1. **Fixation of Nitrogen by Bacteria:** The root nodules of certain leguminous plants such as peas, beans, etc. contain nitrogen fixing bacteria called *Rhizobium*. These bacteria can directly fix nitrogen gas to nitrogen compounds which can then be utilised by the plants.

   Some non-leguminous plants such as *Alnus* and *Ginkgo* also fix atmospheric nitrogen.
2. **Nitrogen Fixation by Blue-Green Algae:** Blue-green algae such as *Nostoc* and *Anabaena* can also help in nitrogen fixation. These algae are usually found in paddy fields.

   The fixation of nitrogen by bacteria and algae is called **biological fixation** of nitrogen. Most of the nitrogen in nature is fixed by this method.
3. **Nitrogen Fixation by Lightning:** During lightning in the sky, when high temperature is produced, the nitrogen gas of the atmosphere reacts with oxygen to form oxide of nitrogen. This nitrogen oxide dissolves in rainwater to form a very dilute solution of nitric acid. The nitric acid thus formed reacts with alkalis present in the soil to form nitrates which are then absorbed by the plants.

Fig. 6.2 *Cinkgo—a non-leguminous plant*

Fig. 6.3 *Anabaena—a blue-green algae*

$$N_2 + O_2 \xrightarrow{\text{Lightning}} 2NO$$
Free nitrogen, Atmospheric oxygen, Nitric oxide

$$2NO + O_2 \xrightarrow{\text{Oxidation}} 2NO_2$$
Nitrogen dioxide

$$4NO_2 + 2H_2O + O_2 \longrightarrow 4HNO_3$$
Rainwater, Atmospheric oxygen, Nitric acid

$$HNO_3 \longrightarrow H^+ + NO_3^-$$
(Nitrate ion readily absorbed by the plants)

6.3.2 Nitrification

$$NH_4^+ \longrightarrow NO_3^-$$

A process in which ammonium is converted to nitrate by decomposition is called **nitrification**. The bacteria that carry out this reaction gain energy from it. It requires oxygen.

6.3.3 Importance of Converting Nitrogen to Ammonia for Agriculture

Nitrogen is a primary source for the preparation of ammonia. The ammonia produced by nitrogen fixing bacteria is quickly incorporated into protein, which is one of the essential component for the plant growth.

The atmospheric nitrogen which is present as nitrates in the soil, is taken up by the plants. In the presence of sunlight and chlorophyll, it is converted into protein in the plants. The plants are consumed by animals and the proteins present in them are hydrolysed in the animal body to amino acids. These amino acids are used up by the animal body. The death, decay and excreta of animals bring back nitrogenous compounds into the soil. These are mainly present as urea and hippuric acid. These compounds are converted to ammonium salts by some microorganisms.

The ammonium salts are partly converted into nitrites, nitrates and NH_3 by nitrifying bacteria and partly decomposed into free nitrogen by denitrifying bacteria. These nitrites and nitrates are again used by plants while the free nitrogen comes back to the atmosphere. Thus, the cycle is completed. Hence, nitrogen undergoes a never ending cycle in nature.

6.4 FERTILIZERS

Plants require nitrogen, phosphorus, potassium, calcium, magnesium, sulphur, iron, manganese, copper, zinc, boron and chlorine for their proper growth. These elements are present in most types of soils. After repeated cultivation, a stage is reached when the soil becomes deficient of these elements and hence, becomes less productive. Thus, in order to remove this deficiency, certain elements in the form of their compounds are added to the soil to regain its productivity. These substances are called **fertilizers**.

Hence, **fertilizers** are the substances which are added to the soil in order to remove the deficiency of essential elements required for the proper growth of plants.

Every compound containing nitrogen, phosphorous or potassium cannot be used as a fertilizer. The essential requirements of a fertilizer are:

(*i*) The element present in it must be easily available to the plant.
(*ii*) It must be sufficiently soluble in water.
(*iii*) It should be stable so that the element present in it must be available to the plant for a long time.
(*iv*) It should not be harmful to the plant.
(*v*) It should be able to maintain the correct pH of the soil.
(*vi*) It should be cheap also.

6.4.1 Sources of Fertilizers

1. **Natural Sources:**
   (*a*) Chile saltpeter ($NaNO_3$), potassium salts, rock phosphates are some naturally occurring inorganic fertilizers.
   (*b*) Urine decomposing vegetable matter, animal matter and guano are some naturally occurring organic fertilizers.
2. **Artificial Sources:** Fertilizers are manufactured through a chemical process using some chemicals which are required for the growth of the plants.

6.4.2 Examples of Nitrogenous Fertilizers

1. Ammonium sulphate [$(NH_4)_2 \cdot SO_4$].
2. Ammonium sulphate nitrate (ASN), [$(NH_4)_2SO_4 \cdot NH_4NO_3$].
3. Calcium ammonium nitrate (CAN), [$Ca(NO_3)_2 \cdot NH_4NO_3$].
4. Basic calcium nitrate (Norwegian saltpeter) [$Ca(NO_3)_2 \cdot CaO$].
5. Calcium cyanamide [$CaCN_2$].
6. Urea [NH_2CONH_2].

6.4.3 Use of Fertilizers

1. The usage of the enhanced and modified fertilizers leads to a massive increase in agricultural productivity and net output.
2. They strengthen the soil and enhance its fertility.
3. There are numerous crops and plants which have different growing habits and nutrient requirements. The fertilizers help us in understanding the precise needs and requirements of the plants fulfilling which maximum production can be reached.
4. Fertilizers are also essential to meet the global demand for food.
5. They also help fight crop disease and improve quality.

A 6.1

Visit a shop that sells different types of fertilizers. Find out the different types of fertilizers sold by the shopkeeper and write their chemical and market names. Prepare a list about when and where to use these fertilizers. Also, find out the right time, and quanity to use them.

6.4.4 Environmental Consequences of Overuse of Fertilizers

1. While fertilizers are generally good for most plants, but their excessive use can cause damage to the plants.
2. Excessive use of fertilizers in the long run degrades the quality of the soil and disturbs the ecosystem.
3. High doses of fertilizers get washed off through irrigation and rainfall and reach water bodies and pollute them.
4. Root burn is a condition in which the roots of the plants suffer damage due to the overuse of fertilizers.
5. Plants that produce greener, more lush leaves, because of over-fertilization may also attract more pests.
6. Excessive fertilizer that washes into drains eventually makes its way into water bodies such as rivers, lakes, causing pollution. It thus have negative impacts on fish and other aquatic animals.

SELF EVALUATION

1. Process A: Nitrogen fixing bacteria convert nitrogen gas into nitrates.

   Process B: Nitrifying bacteria use oxygen to convert ammonium compounds (from decaying plant and animal matter) into nitrates.

   Process C: Denitrifying bacteria convert nitrates into nitrogen.

   On the basis of above mentioned statements, answer the following questions:

   (*i*) Where does the process A take place?

   (*ii*) What effect the presence of air in the soil will have on process B?

   (*iii*) What effect waterlogged soil which lacks air will have on process C?

   (*iv*) Explain, why do plants grow properly in well drained and aerated soil.

   (*v*) A farmer wants to grow a good crop of wheat without using any of the fertilizers. What should have been planted in the field the previous year to ensure a good crop?

   (*vi*) Explain, why do the garden manure and compost fertilize the soil.

2. Although certain bacteria in the soil convert nitrogen gas into nitrates, other bacteria convert nitrogen into ammonium salts. The ionic equation for this second reaction is:

$$N_2 + 8H^+ + 6e^- \longrightarrow 2NH_4^+$$

   (*i*) Explain why is this a reduction reaction.

   (*ii*) In the presence of hydrogen ions, bacteria of a different type convert nitrate ions into nitrogen gas and water. Give the ionic equation for this reaction.

6.5 CARBON CYCLE

Carbon is the basic building element of all the living organisms. By volume, carbon dioxide in the atmosphere is very low, i.e. about 0.03% v/V. Carbon is absorbed by the plants in the form of CO_2 from the atmosphere. It is absolutely essential for the synthesis of carbohydrates and fats through photosynthesis in green plants. Later on, carbohydrates and fats in green plants are utilized by consumers (herbivores and carnivores) and thus compounds are again digested and converted into other forms. From both producers and consumers, carbon is returned back to the atmosphere during respiration. After the death of producer and consumer certain decomposing agents (such as bacteria and fungi) decompose and degrade the complex organic compounds (carbohydrates, fat, etc.) in their simplest forms (carbon, hydrogen, oxygen, nitrogen, etc.) which are then available for other cycles. Some of the organic carbon is incorporated into the earth's crust in the form of coal, natural gas, etc. These fossils form the most important sources of energy.

CO_2 is released during the combustion of these fossil fuels. Thus, carbon cycle is mainly maintained by the process of photosynthesis, respiration, decomposition and combustion of fossil fuels.

Sea is the second major reservoir of carbon. Sea water contains 50 times, more CO_2 than air. This is in the form of carbonates and bicarbonates. The CO_2 dissolves in sea water to form carbonic acid which converts carbonates into bicarbonates. Thus, the sea regulates the CO_2 contents in the atmosphere.

CO_2 has the unique property of absorbing infrared rays of the sun, thus, keeping the Earth warm. But a huge amount of CO_2 has been released into the atmosphere due to extensive industrialization and cutting of green plants. This has resulted in an excessive absorption of infrared radiations leading to an increase in atmospheric temperature. A greenhouse effect has been experienced which has affected the functioning of living organism adversely.

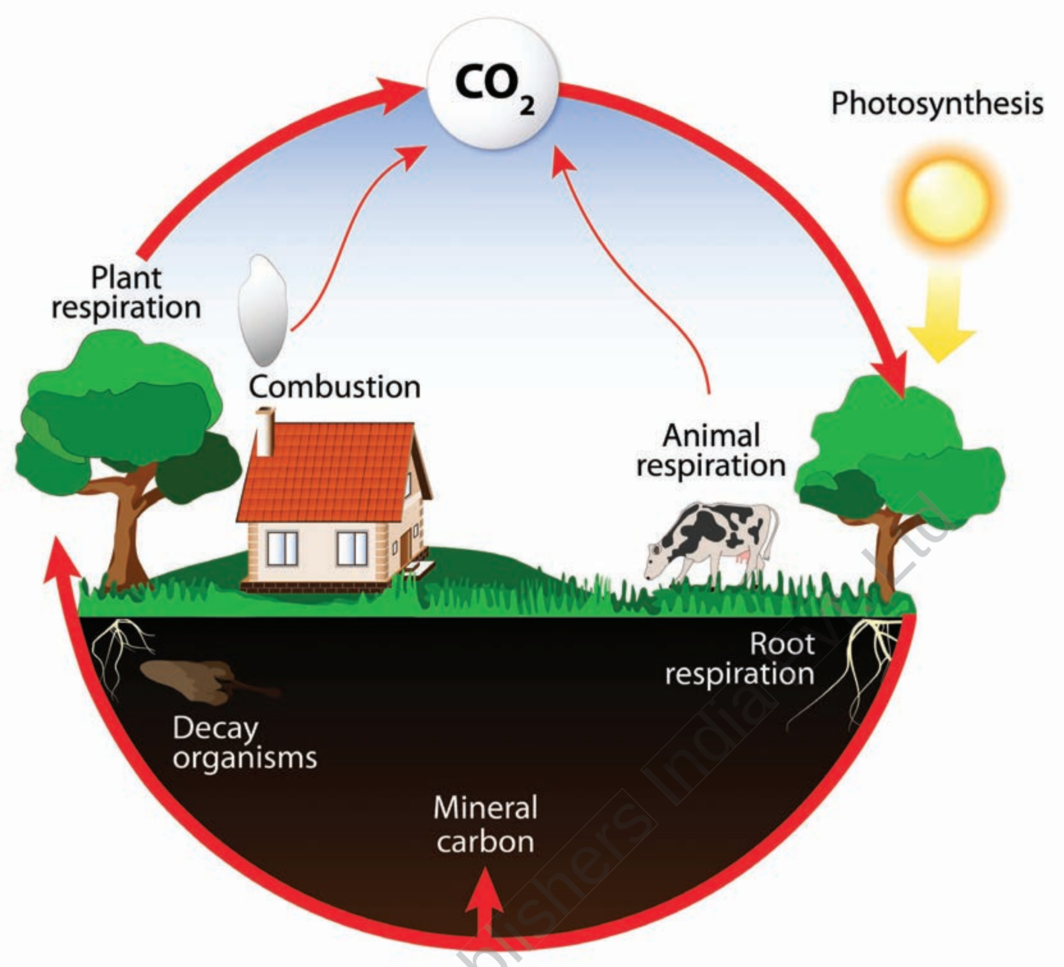


Fig. 6.4 *Carbon cycle*

6.5.1 Weathering

The process of breaking down or dissolving minerals and rocks on the earth's surface is known as weathering. Carbon dioxide and the other atmospheric gases dissolve in the surface water. Dissolved gases are in equilibrium with the gas in the atmosphere. Carbon dioxide reacts with water in solution to form the weak acid, carbonic acid. Carbonic acid dissociates into hydrogen ions and bicarbonate ions. The hydrogen ions and water react with most common minerals (silicates and carbonates) altering the minerals. The products of weathering are predominantly clays (a group of silicate minerals) and soluble ions such as calcium, iron, sodium, and potassium. Bicarbonate ions also remain in solution; a remnant of the carbonic acid that was used to weather the rocks.

$$\underset{\text{Carbon dioxide}}{CO_2} + \underset{\text{Water}}{H_2O} \longrightarrow \underset{\text{Carbonic acid}}{H_2CO_3}$$

$$\underset{\text{Carbonic acid}}{H_2CO_3} \longrightarrow \underset{\text{Hydrogen ion}}{H^+} + \underset{\text{Bicarbonate ion}}{HCO_3^-}$$

Reactions that transform primary silicate minerals into clay minerals.

$$2KAlSi_3O_8 + 2H^+ + 9H_2O \rightleftharpoons Al_2Si_2O_5(OH)_4 + 2K^+ + 4H_4SiO_4$$

$$2NaAlSi_3O_8 + 2H^+ + 9H_2O \rightleftharpoons Al_2Si_2O_5(OH)_4 + 2Na^+ + 4H_4SiO_4$$

6.5.2 Role of Carbon Cycle in Maintaining the Atmosphere Composition

Because of the role of CO_2 in climate, feedbacks in the carbon cycle act to maintain global temperatures within certain bounds so that the climate never gets too hot or too cold to support life on the Earth. This process is a large-scale example of Le Chatelier's Principle. This chemical principle states that "if a reaction at equilibrium is perturbed by the addition or removal of a product or reactant, the reaction will adjust so as to attempt to bring that chemical species back to its original concentration". For example, as carbonic acid is removed from the solution by weathering of rocks, the reaction will adjust by producing more carbonic acid. And since the dissolved CO_2 is in equilibrium with atmospheric CO_2, more CO_2 is removed from the atmosphere to replace that removed from the solution by weathering. For example, if CO_2 concentration increases in the atmosphere global temperature will rise. Rising temperature and more dissolved CO_2 will lead to increased weathering of crystal rocks as a result of faster reaction rates and greater acidity. Enhanced weathering will use up the excess CO_2 thereby cooling the climate.

If global temperature cools as a result of some astronomical forcing or tectonic/ocean circulation effect, the lower temperatures will result in lower rates of chemical weathering. Decreased weathering means less CO_2 being drawn from the atmosphere by weathering reactions, leaving more CO_2 in the atmosphere to increase temperature.

If more rocks become available for rapid weathering as a result of mountain uplift, the enhanced weathering will draw down atmospheric CO_2 and decrease global temperature. But the decreased temperature will slow reaction rates, thereby using less CO_2, thus allowing temperature to become moderate.

6.6 GLOBAL WARMING

One of the important facts on global warming is that global warming is mainly an outcome of the greenhouse effect.

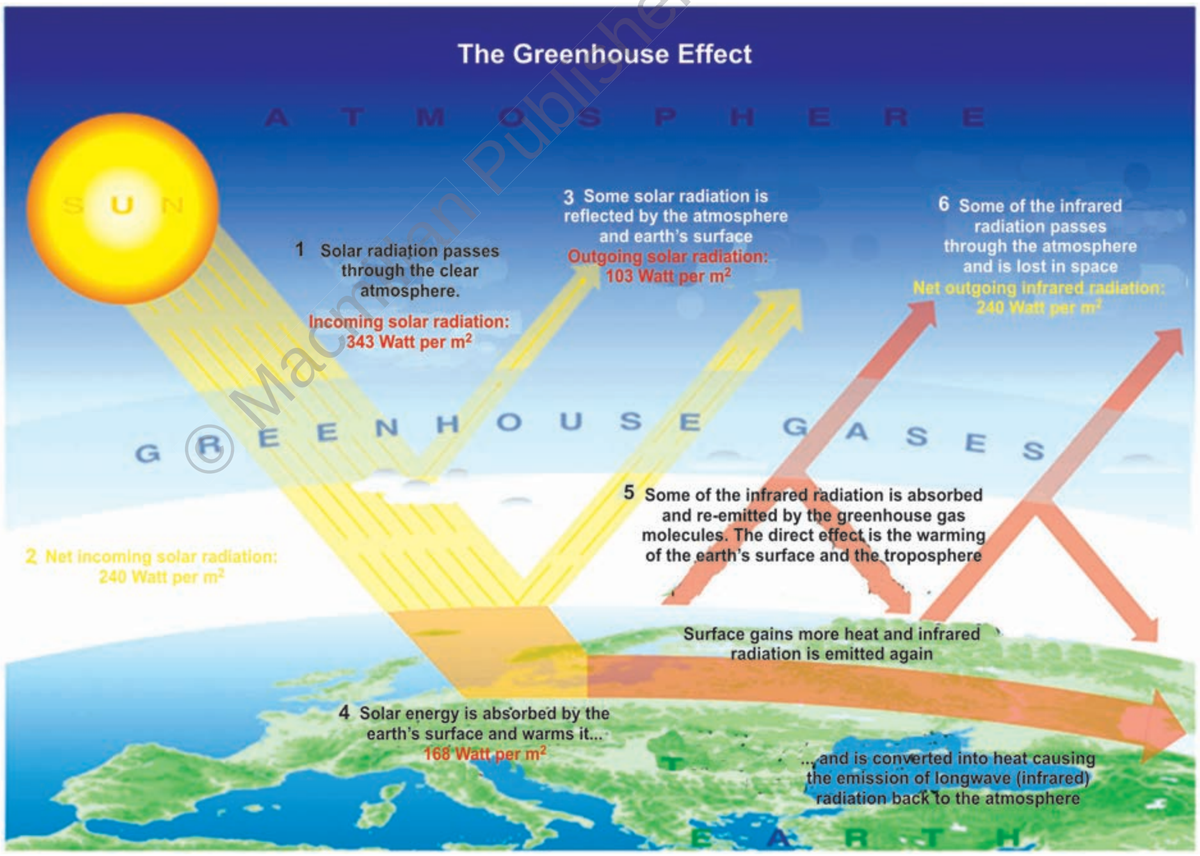


Fig. 6.5 *The greenhouse effect*

The term greenhouse is given to a glasshouse constructed to keep delicate and rare plants that require warmth and protection from weather. The house is constructed by panels (walls) of glass that allow sun rays and heat to pass into the glasshouse. The heat that enters the glasshouse cannot escape out, because the glass panels radiate back the heat to the inside of the glasshouse. This makes the glasshouse warmer than the outside environment. This phenomenon is termed as **greenhouse effect**.

Similar phenomenon occurs on the earth too when the heat is not allowed to radiate back in the space.

The carbon dioxide gas is mainly responsible for greenhouse effect. In addition to carbon dioxide, some other gases like **carbon monoxide, methane** and **chlorofluorocarbons (CFCs)** are also contributing to greenhouse effect. These gases are called greenhouse gases.

The enhanced greenhouse effect increases the evaporation rate of the surface water. The presence of high percentage of water vapours in the atmosphere also contributes to greenhouse and increases the surface temperatures.

GLOBAL WARMING AND GASES

- Carbon dioxide causes up to 57 per cent of global warming.
- Chlorofluorocarbon causes up to 25 per cent of global warming.
- Methane about 12 per cent and oxides of nitrogen cause up to 6 per cent of global warming.

6.6.1 Natural Causes of Global Warming

1. Volcanic Eruptions release various harmful gases and copious amounts of dust into the atmosphere.
2. Decay of organic matter lying under water releases methane gas which is a main participant of greenhouse effect.
3. Changes in solar radiation deregulate the temperature on the Earth surface and cause polar ice caps and glaciers to melt.
4. Forest fire releases soot (carbon particles) and ash in the air which in turn increases content of greenhouse gases in the atmosphere.

6.6.2 Man-made Causes of Global Warming

1. **Burning of Fuels:** Burning of fuels such as wood, cowdung, cake, coal and kerosene in homes increases the content of greenhouse gases, such as carbon monoxide, sulphur dioxide, carbon dioxide, soot, etc.
2. **Vehicles:** Exhaust gases emitted by motor vehicles pollute the air and give rise to global warming by producing harmful pollutants such as sulphur dioxide, nitrogen dioxide (NO_2), carbon monoxide (CO), etc.
3. **Burning of Fossil Fuels:** Fossil fuels are used in factories and industries. The greenhouse gases emitted cause air pollution and in longer prospects cause global warming.
4. **Deforestation:** Deforestation is the clearing of forest by cutting the trees from forest. It causes global warming by increasing the amount of carbon dioxide.
5. **Chlorofluorocarbons (CFCs):** Use of CFCs depletes the ozone layer. As a result of which harmful ultraviolet rays reach the Earth and increase the temperature on the Earth's surface giving rise to global warming.

6.6.3 Effect of Global Warming

The global warming causes following disturbances and adverse effects in the environment:

1. The greenhouse gases increase the atmospheric temperature around the earth.
2. Warming of the earth would make the **polar ice caps** to melt further. The water so produced by the melting of ice will increase the size of the ocean, which may engulf and submerge larger coastal areas of the land.
3. The increase in temperature would alter the climate and rainfall, affecting the flora and fauna.

4. The human beings would be more vulnerable to diseases as increased temperature would substantially increase viruses or parasites and vectors (insects). All these thrive better in warmer conditions.
5. The emission of chlorofluorocarbon have been associated with the potential global warming and depletion of ozone layer in the stratosphere.

A 6.2

Compare the quality of air in a village and a town. At which place would you like to live and why? How air determines global warming? Write a brief report on your observation?

SELF EVALUATION

1. The burning of fossil fuels produces 16000 million tonnes of carbon dioxide per year. Carbon dioxide is thought to increase the average temperature of the air. It is predicted that this increase in temperature will melt some of the ice caps at the North and South poles. Describe the effects which this could have on the life of the people in other parts of the world.
2. The amount of carbon dioxide in the atmosphere is slowly increasing.
   (*i*) Suggest two reasons why this is happening.
   (*ii*) Explain why people call the effect which carbon dioxide has on the atmosphere as the 'greenhouse effect'.
   (*iii*) Why are some people worried about the greenhouse effect?
   (*iv*) Suggest two things which could be done to stop the increase in the percentage of carbon dioxide in the atmosphere.

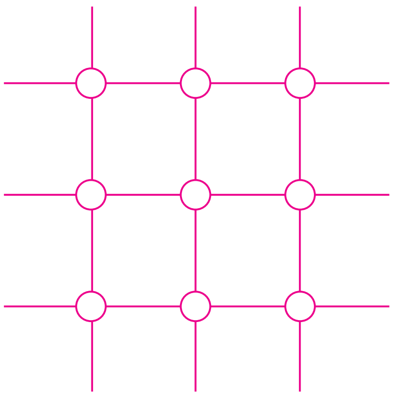
6.7 POLYMERS

Polymers and plastics are high molecular mass substances consisting of large number of repeating structural units called monomers.

Plastics and natural materials such as rubber or cellulose are composed of very large number of molecules. Polymers are constructed from relatively small molecular fragments known as **monomers** that are joined together.

Any synthetic material which can be easily moulded into any desired shape upon heating is called **plastic**. Plastic is also a polymer like the synthetic fibres. In some plastics, the arrangement of monomer is linear whereas in others it is crosslinked as shown in Fig. 6.6.

(*a*) Linear polymers



(*b*) Cross linked polymers

Fig. 6.6 *Arrangement in plastic*

6.8 HAZARDS OF POLYMERS

In 1940, when synthetic polymers were making a major impact in people's lives, some of the problems were not realized at the time. The problems which these materials have created as their uses multiplied and the world became more wary of 'chemicals'.

6.8.1 Release of Small Molecule

Many kinds of polymers contain small molecules — either unreacted monomers, or substances specifically added (plasticizers, uv absorbers, flame retardants, etc.) to modify their properties. Many of these smaller molecules are able to diffuse through the material and are released into any liquid or air in contact with the plastic and eventually into the aquatic environment. Those that are used for building materials can build up in closed environments and contribute to indoor air pollution.

6.8.2 Residual Monomer

The unpolymerized monomer that remains incorporated in a polymer after the polymerization reaction is completed is called **residual monomer**. Although there is little evidence that the small quantities that diffuse into the air or leach out into fluids pose a quantifiable health risk. People are understandably reluctant to tolerate these exposures.

6.8.3 Plasticizers

These substances are compounded into certain types of plastics to render them more flexibility by lowering the glass transition temperature. They accomplish this by taking up space between the polymer chains and acting as lubricants to enable the chains to slip more readily over each other. Many but not all are small enough to be diffusible and act as a potential source of health problems.

Polyvinyl chloride polymers are one of the most widely-plasticized types, and the odour often associated with flexible vinyl materials such as garden hoses, waterbeds, cheap shower curtains, raincoats and upholstery are testament to their ability to migrate into the environment.

6.8.4 Hazards to Animals

There are two general types of hazards that polymers can introduce into the aquatic environment. One of these relates to the release of small molecules that act as hormone disruptors. It is well established that small aquatic animals such as fish are being seriously affected by such substances in many rivers and estuarine systems, but details of the sources of these molecules have not been identified. One confounding factor is the release of sewage water containing human birth-control drugs (which have a feminizing effect on sexual development) into many waterways.

The other hazard relates to pieces of plastic waste that aquatic animals mistake for food or become entangled in.

6.8.5 Measures to Prevent Hazards Caused by Polymers

The environmental problems caused by polymers have become so alarming that people are advised to use the following methods to prevent health hazards.

1. Reduce your dependence on plastic bags and other items made of plastics.
2. Carry your own cloth bag or jute bag for shopping.
3. Never burn plastic bags and other plastic items in open place to avoid air pollution.
4. Keep separate garbage bins in your home for biodegradable (green bins) and non-biodegradable wastes (blue bins).
5. Use biodegradable plastics instead of non-biodegradable plastics.

6. Follow and practice the 4R principle of reduce, reuse, recycle and recover to minimise environmental pollution.

Some plastics cannot be melted – they burn or harden instead of melting. It is even more difficult to recycle these plastics as they can only be used in the same shape in which they were originally cast.

The above problems mean that the continued manufacture of non-biodegradable polymers is a cause for environmental concern.

Biodegradable polymers decompose naturally, so burying them is slightly less environmentally unsustainable as they will eventually break down.

They can be recycled, broken down into their original components and reused, but they still need to be collected, separated and cleaned.

A **6.3**

Quiz Time

Complete the following:

1. A single unit in a chain of polymer is called __________.
2. __________ is the product formed by the combination of large number of monomers.
3. PVC stands for __________.
4. 4-R principle stands for __________ __________ __________ __________.
5. The unpolymerized monomer that remains incorporated in a polymer after polymerization reaction is completed is called __________.

Green Chemistry in Bhutan

Green chemistry involves the design of chemical products and processes that reduce or eliminate the use of hazardous substances. It is a well known fact that conservation of environment constitutes an important part of Bhutan's spatial planning and strategic framework, and has always enjoyed a high priority. Bhutan has always made efforts to protect its environment and natural resources. A few noteworthy measures under the fold of green chemistry to check pollution are mentioned below:

- **Plastics:** For the prevention of pollution. Bhutan has come up with a new way of paving its rods. As part of effort to curb the use of fossil fuels and deal with the growing amount of plastic waste, the country plans to mix used plastic bottles and other waste plastic with bitumen to blacktop the roads.
- **Fertilizers and pesticides:** It is a well known fact that self-sufficiency in food has been achieved in Bhutan by using fertilizers and pesticides and exploring improved methods of farming, good quality of seeds irrigation, etc. But overuse of fertilizers and pesticides has resulted in the deterioration of soil, water and air.
- **Dry cleaning of clothes:** A commonly used dry cleaning solvent is tetrachloroethene. It pollutes water and is also carcinogenic. This has been replaced by other detergents which contain liquid carbon dioxide.
- **Synthesis of chemicals:** Conscious efforts are also being made to produce substances of daily use by chemical reaction which neither employ toxic chemicals nor released the same in atmosphere.

All countries including Bhutan are giving importance to the study of biotechnology which has a major role in checking pollution. Green chemistry directly or indirectly relates to us. We must realize that we do not have solutions for every problem but we can concentrate on issues, which we feel strongly about and can do something about. It is the duty of every one in the world to keep the environment free from pollution. We should take care to put into practice whether we preach. Remember, green chemistry or environment protection begins with us.

SELF EVALUATION

1. Recycling is an increasingly important process in modern industry. Plastics, however, are more difficult to recycle than glass, paper and metals.

   One way of recycling plastics is to decompose them and to re-use the decomposition products to make plastics once more. This decomposition process is called 'pyrolysis'.

   (*i*) Explain why plastics are difficult to recycle.

   (*ii*) Which feature of the pyrolysis process ensures that the plastic is decomposed and not burnt?
2. The large scale use of synthetic polymers created acute environmental problems. What types of polymers can be used to prevent the disbalance of the environment?
3. Write the merits and demerits of plastics in our life.

RECAPITULATION

1. The branch of chemistry that deals with processes and products that reduce the use and generation of hazardous substances is called green chemistry.
2. Microorganisms are tiny organisms that cannot be seen with naked eyes.
3. Ammonia formed is converted first into nitrites and then into nitrates by the action of *Nitrosomonas* and *Nitrobacter* bacteria respectively. The process is called nitrification. These nitrates are again absorbed by plants and the cycle is repeated.
4. Rhizobium is a nitrogen fixing bacteria found in the roots of leguminous plants.
5. The process of converting complex proteins into ammonium form of nitrogen is called ammonification.
6. The soil contains denitrifying bacteria called *Pseudomonas* which convert nitrate form of nitrogen into free nitrogen which goes back into the atmosphere. The process is called denitrification.
7. The root nodules of certain leguminous plants such as peas, beans, etc. contain nitrogen fixing bacteria called rhizobium.
8. Some non-leguminous plants like Alnus and Ginkgo also fix atmospheric nitrogen.
9. The fixation of nitrogen by bacteria and algae is called biological fixation.
10. When the plants and animals die, the putrefying bacteria and fungi present in the soil decompose the protein of dead plants and animals into ammonia. This process is called ammonification.
11. The plants are eaten by animals and the proteins present in them are hydrolysed in the animal body to amino acids.
12. Chile saltpeter, potassium salt, rock phosphate are some naturally occurring inorganic fertilizers.
13. Urine, decomposing vegetable matter, animal matter are some naturally occurring organic fertilizers.
14. Ammonium sulphate, ammonium sulphate nitrate, calcium ammonium nitrate, basic calcium nitrate, calcium cyanamide and urea are the nitrogenous fertilizers.
15. The carbon cycle is defined as a series of naturally occurring processes where carbon is exchanged between organisms and the environment.
16. The plant uses the carbon dioxide in order to grow by the process of photosynthesis.
17. Photosynthesis is a part of the carbon cycle.

18. The cycle by which plants photosynthesise and use atmospheric carbon dioxide to produce carbohydrates, which are in turn metabolized by animals to decompose into products that return carbon dioxide to the atmosphere is called carbon cycle.
19. Global warming is the increase in the Earth's average surface temperature due to greenhouse gases, such as carbon dioxide emissions from burning fossil fuels or from deforestation, which traps heat that would otherwise escape from the Earth.
20. Greenhouse gases keep heat close to the Earth's surface making it livable for humans and animals.
21. With the start of industry in the 1700's, humans began emitting more fossil fuels from coal, oil, and gas to run cars, trucks, and factories. Save on gas, but help prevent global warming.
22. There is more carbon dioxide in the atmosphere today than at any point in the last 800,000 years.
23. Consequences of global warming include drought, severe hurricanes, massive fires and the melting of polar caps.
24. Heat waves caused by global warming present greater risk of heat-related illness and death, most frequently among patients of diabetes who are elderly or very young.
25. Monomers are the small repeating unit of a polymer usually obtained from low molecular mass.
26. Polymers are macromolecules of relatively high molecular mass, whose structures are composed of a large number of simple repeating units.
27. The process of formation of polymers from monomers is known as polymerisation.
28. Polymers which occur in nature are called natural polymers.
29. Synthetic polymers are man-made polymers.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. The root nodules of leguminous plants contain bacteria.
2. The plant absorbs nitrate compounds from the soil and water and converts it into plant proteins.
3. Oxygen is a greenhouse gas.
4. Biodegradable polymers decompose naturally.
5. Polymers are low molecular mass substances.

II. Fill in the Blanks

Fill in the blanks by using suitable word(s):

1. The process of converting complex protein to ammonium is called ____________.
2. Plants absorb ____________ compounds from the soil.
3. Nitrogen is the main constituent of ____________.
4. Photosynthesis and ____________ are essential and opposite to one another.
5. Global warming has been associated with increased depletion of ____________.
6. Polymers are constructed from relatively small molecular fragments known as ____________ that are joined together.

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. A branch of chemistry that deals with the processes and products that reduce the use and generation of hazardous substances is called
   - (*a*) green chemistry
   - (*b*) agricultural chemistry
   - (*c*) environmental chemistry
   - (*d*) inorganic chemistry
2. Root nodules are found in
   - (*a*) leguminous plants
   - (*b*) all plants
   - (*c*) trees
   - (*d*) none of the plant
3. Carbon dioxide (CO_2) in the air enters a plant via the process of
   - (*a*) respiration
   - (*b*) ingestion (eating)
   - (*c*) photosynthesis
   - (*d*) decomposition
4. Respiration is the process by which ________________ is consumed and ________________ is produced.
   - (*a*) carbon dioxide, oxygen
   - (*b*) oxygen, carbon dioxide
   - (*c*) carbon dioxide, energy
   - (*d*) energy, carbon dioxide
5. The word 'polymer' is meant for material made from ________________.
   - (*a*) single entity
   - (*b*) two entities
   - (*c*) multiple entities
   - (*d*) any entity
6. Polymers are ________________ in nature.
   - (*a*) organic
   - (*b*) inorganic
   - (*c*) both (*a*) and (*b*)
   - (*d*) none of these
7. One of characteristic properties of polymer material ________________.
   - (*a*) high temperature stability
   - (*b*) high elongation
   - (*c*) high mechanical strength
   - (*d*) low hardness

IV. Matching

Match the items of Column I with the corresponding items of Column II:

Column I	Column II
1. Carbon dioxide	(*a*) Best fertilizer for corn crops
2. Ammonium sulphate	(*b*) Constitutes of monomer
3. Respiration	(*c*) Nitrogenous fertilizer
4. Polymer	(*d*) Greenhouse gas
5. Urea	(*e*) Reverse of photosynthesis

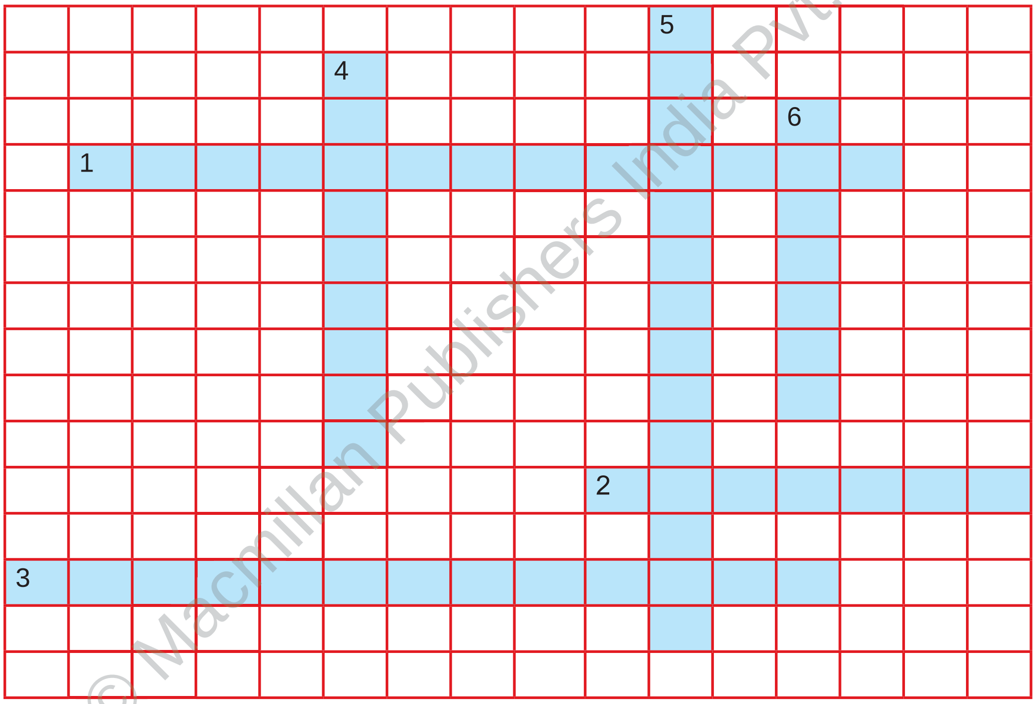
V. Answer the Following Questions

1. Write a short note on Nitrogen Fixation.
2. Explain the 'Nitrogen Cycle' with flow diagram.
3. What is a root nodule?
4. Explain the merits and demerits of fertilizers.
5. Write a short note on urea.

6. Explain the following:
   (*i*) Photosynthesis
   (*ii*) Respiration
   (*iii*) Weathering
7. Explain global warming and what are its causes.
8. Explain the role of carbon cycle in atmosphere.
9. List any three activities which would lead to an increase in the level of CO_2 in air.
10. Explain the greenhouse effect.
11. What do you mean by residual monomer?
12. What are the hazards of plastic? How can they be controlled? Give measures.
13. Classify different types of polymers.

VI. Puz e

Read the given clues carefully and write the appropriate word inside the boxes to solve the puz e.



Across:

1. A process in which ammonium is converted into nitrate.
2. It is a greenhouse gas.
3. Monomer of polyvinyl chloride.

Down:

4. Nitrogen fixing bacteria in leguminous plants.
5. A process by which green plants mak e their own food.
6. Simplest unit of polymers.

[Hints: 1. Nitrification 2. Methane 3. Vinyl chloride 4. Rhizobium 5. Photosynthesis 6. Monomer]

Web Links:

For more information on the following topics, visit the websites listed below:

Nitrogen Cycle

http://www.ducksters.com/science/ecosystems/nitrogencycle.php

Fertilizer

https://en.wik pedia.org/wik /Fertilizer

Carbon Cycle

https://eschooltoday.com/ecosystems/the-carbon-cycle.html

Global Warming

http://www.nationalgeographic.com/environment/global-warming/global-warming-overview/

Polymer

http://www.sciencedirect.com/science/journal/00323861

CHAPTER 7

ORGANIC CHEMISTRY: HYDROCARBONS

LEARNING OBJECTIVES

After going through this chapter, the students would be able to:

- Explain that alkanes are saturated hydrocarbons, alkenes and alkynes are unsaturated hydrocarbons.
- Apply the general formula for alkanes (C_nH_{2n+2}), alkenes (C_nH_{2n}) and alkynes (C_nH_{2n-2}).
- Apply the IUPAC rules to the nomenclature of alkanes, alkenes and alkynes.
- Draw and name structural isomers for simple alkanes, alkenes and alkynes.
- State the products of burning hydrocarbons and write balanced chemical equations for these reactions.
- Identify unknown hydrocarbons as being saturated or unsaturated by tests such as burning, adding bromine water or acidified potassium manganate (VII) acidified potassium permanganate.
- Describe that petroleum is a mixture consisting mainly of alkane hydrocarbons.
- Describe the process of fractional distillation of crude oil.
- State that cracking involves the breaking of C–C bonds in alkanes and cracking can be used to obtain more useful alkanes and alkenes.
- Describe the formation of addition polymers from the products of crude oil by cracking and polymerization.
- State some uses of addition polymers, e.g. polyethene and polyvinyl chloride (PVC).

7.1 INTRODUCTION

The word organic means "pertaining to life". It was believed that substances such as sugar, starch, proteins, acetic acid, urea, etc., could be obtained from living organisms, i.e., animals and plants. Such substances were called '**Organic substances**' and the science dealing with them were known as **'Organic chemistry'**.

7.2 MODERN DEFINITION OF ORGANIC COMPOUNDS

Since carbon is an essential constituent of all organic compounds, organic chemistry is defined as **chemistry of carbon compounds**. However, there are compounds of carbon such as carbon dioxide, carbon monoxide, carbonates, carbides, etc., which are studied in inorganic chemistry because of their properties. All organic compounds contain carbon and hydrogen, some organic compounds contain oxygen, nitrogen, sulphur and halogens along with carbon and hydrogen. Thus, we can define organic compounds as the compounds that are associated with life processes and contain one or more carbon atoms covalently linked to atoms of other elements. **Organic chemistry** is the branch of chemistry that deals with the scientific study of the structure, properties, and reactions of organic compounds and organic materials. Different species of organisms and different individuals of a species can be distinguished by the variation in organic molecules.

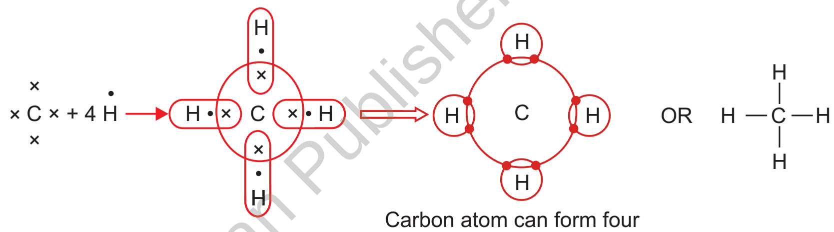
7.2.1 Reasons for Existence of Large Number of Organic Compounds

It is interesting to note that organic compounds are basically made up of carbon, hydrogen and oxygen with a few more elements but their number is far more than the inorganic compounds. At present, millions of organic compounds are known to us. This can be attributed to the unique character of carbon atom.

Following are the unique characteristics of carbon which are responsible for the existence of a large number of organic compounds.

1. **Tetra Covalency of Carbon:** Carbon has four valence shell electrons. It can neither lose nor gain electrons to attain an octet, thus it always forms covalent bonds by sharing electrons with other atoms. This characteristic of carbon atom by virtue of which it forms **four covalent** bonds is called **tetra covalency of carbon.**

   Since all the organic compounds contain carbon as an essential element, therefore, the nature of chemical bonding in organic compounds is always covalent.



Carbon atom can form four single covalent bonds e.g., CH_4 (Methane)

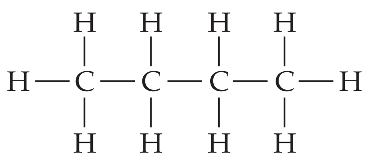
Fig. 7.1 *Tetra covalency of carbon*

2. **Catenation:** The tendency of an element to form chains of identical atoms is called catenation. Carbon has the unique property of combining with any number of carbon atoms to form straight chains, branched chains and rings of different sizes. This property is called **catenation**.

   Catenation i.e., self linking of carbon atoms leads to the formation of almost unlimited number of organic compounds.

   Examples:

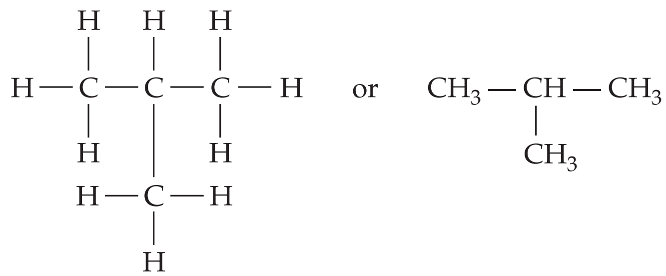
   (*i*) *Straight Chain Compound:*



or $CH_3—CH_2—CH_2—CH_3$

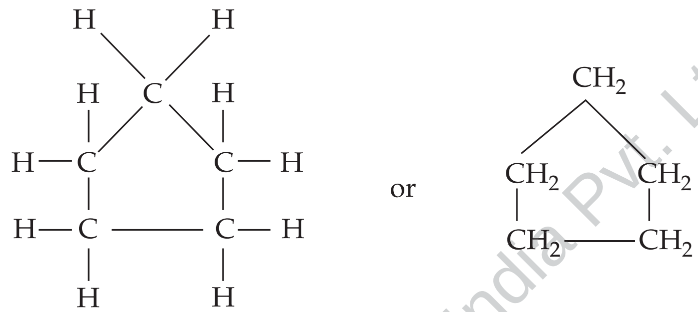
Straight chain of four carbon atom
(Normal butane)

(*ii*) *Branched Chain Compound:*



Branched chain of four carbon atoms
(2-methyl propane)

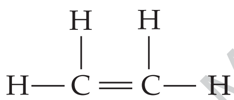
(*iii*) *Cyclic Compound:*



Closed chain (ring) of five carbon atoms (Cyclopentane)

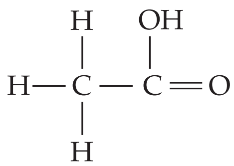
Carbon exhibits catenation due to tetra covalency of carbon and great strength of carbon—carbon bonds, i.e., high C—C bond energy. Due to stable C—C covalent bonds, compounds formed are stable. There are two conditions necessary to exhibit catenation. (*a*) The valency of element must be greater than or equal to two. (*b*) The element must form strong covalent bonds between its own atoms as compared to the bonds with atoms of other elements, especially oxygen.

3. **Formation of Multiple Bonds:** Due to its small size carbon atom can form multiple bonds, i.e., double and triple bonds with not only carbon but with atoms of other elements such as oxygen, nitrogen, etc.

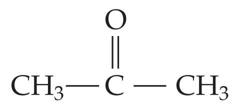


Ethene (C_2H_4)
Double bond

Ethyne (C_2H_2)
Triple bond

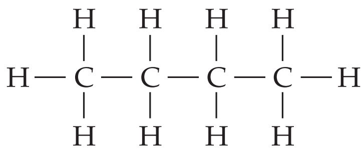


Acetic acid (CH_3COOH)



Acetone (CH_3COCH_3)

4. **Isomerism:** Another reason for the large number of organic compounds is the phenomenon of isomerism. Compounds having the same molecular formula, but different structural formula, are called **isomers** of one another, and the phenomenon is called **isomerism**. For example, in butane (C_4H_{10}), there are following two possible structures for the same molecular formula C_4H_{10}.

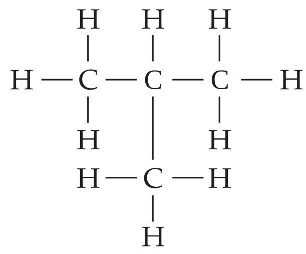
(*i*) 

or

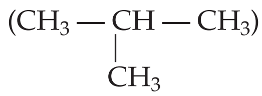
($CH_3 — CH_2 — CH_2 — CH_3$)

n-butane

Structure I

(*ii*) 

or



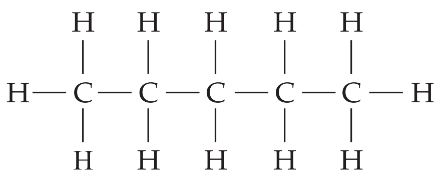
Isobutane

Structure II

Structure I contains a straight chain of four carbon atoms and represents the compound known as ***n*-butane**. Structure II contains a branched chain and represents **isobutane**. These compounds possess different chemical properties. Thus, with the same molecular formula C_4H_{10}, we have two organic compounds which differ in chemical structures and properties.

Other Examples of Isomers

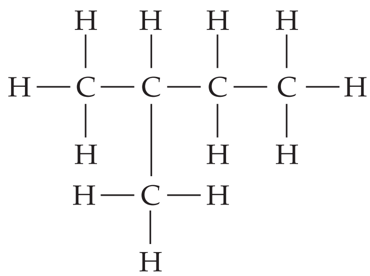
The isomers with molecular formula C_5H_{12} can have three structures:

1. 

n-pentane
Straight chain compound

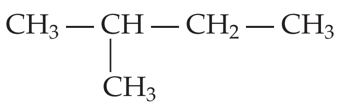
or

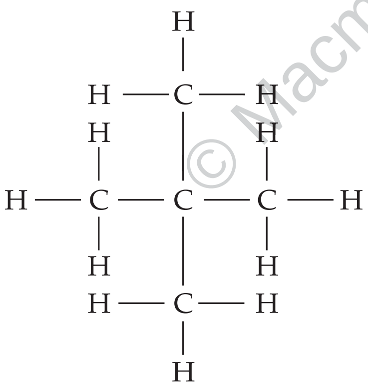
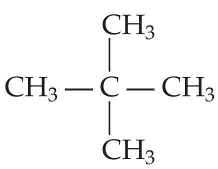
$CH_3 — CH_2 — CH_2 — CH_2 — CH_3$

2. 

Isopentane
One branched chain
or 2-methyl butane

or



3.  or 

Neopentane
Two branched chains
2, 2-dimethylpropane

PREFIX

n – denotes that alkane is unbranched.

iso– denotes that alkane contains $(CH_3)_2CH–$ and no other branches.

neo– denotes that alkane contains $(CH_3)_3C–$ and no other branches.

SELF EVALUATION

1. Write all the possible isomers having molecular formula C_6H_{14}.
2. Out of these, which one has multiple bond—C_2H_6, CH_3COOH, C_3H_8, and C_2H_5Cl?

7.3 SOURCES OF ORGANIC COMPOUNDS

1. **Plants:** Carbohydrates (glucose, sucrose, starch, cellulose), vitamins, alcohols, acetone, etc.
2. **Animals:** Proteins, fats, urea, etc.
3. **Fungi and Microorganism:** Alcohols, carboxylic acids, antibiotics (penicillin), vitamins, etc.
4. **Coal:** Destructive distillation of coal is the main source of organic aromatic compounds such as benzene, toluene, phenol, dyes, drugs, perfumes, etc.
5. **Natural Gas and Petroleum:** It is one of the major source of organic compounds. For example, gasoline, kerosene oil, vaseline, paraffin wax, etc.
6. **Wood:** Methyl alcohol, acetone, etc.
7. **Synthetic Source:** More than 90% of known organic compounds are synthetic i.e., prepared in laboratory.

7.4 IMPORTANCE OF ORGANIC COMPOUNDS

Organic compounds play an important role in our daily activities. There is hardly any walk of life where we do not need the organic compounds. Every industry is dependent on organic compounds. The following list clearly illustrates the importance of organic compounds:

1. **Food:** Carbohydrates (sugar, starch, etc.), proteins, fats, vitamins, enzymes, etc.
2. **Clothes:** Cotton, silk, wool, nylon, rayon, etc.
3. **Fuels**: Coal, wood, natural gas, petrol, etc.
4. **Medicines:** Penicillin, sulpha drugs, aspirin, iodoform, etc.
5. **Explosives:** Nitroglycerine, trinitrotoluene, nitro-cellulose, etc.
6. **Dyes:** Indigo, alizarin, malachite green, etc.
7. **Insecticide:** D.D.T., B.H.C., gammexane, etc.
8. **Household and Other Articles:** Soaps, cosmetics, perfumes, detergents, paper, rubber, plastics, resin, paints, etc.

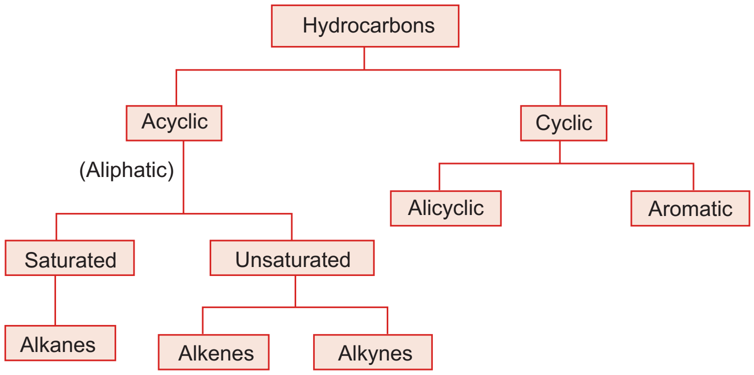
7.5 HYDROCARBONS

The simplest organic compounds containing carbon and hydrogen are called **hydrocarbons.** They are considered to be parent organic compounds while all other organic compounds are thought to have been derived from them by the replacement of one or more hydrogen atoms by other atoms or group of atoms.

The most important natural source of hydrocarbons is **petroleum (or crude oil)** which is obtained from underground oil deposits by drilling oil wells. The **natural gas** which occurs above petroleum deposits also contains hydrocarbons.

7.6 CLASSIFICATION OF HYDROCARBONS

Hydrocarbons can be broadly classified on the basis of carbon-carbon chain into the following types:



Acyclic or open chain hydrocarbons are of two types:

1. **Saturated Hydrocarbons–Alkanes:** The hydrocarbons in which all the carbon atoms are linked together by single covalent bonds are called **saturated hydrocarbons** or **alkanes**.

   For example:

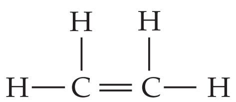
CH_4	$CH_3—CH_3$	$CH_3—CH_2—CH_3$
Methane	Ethane	Propane

   In all these compounds, all the four valencies of each carbon are satisfied by single covalent bonds.

   > The general formula of saturated hydrocarbons or alkanes is C_nH_{2n+2}, where n is the number of carbon atoms in one molecule of the alkane.

2. **Unsaturated Hydrocarbons:** The hydrocarbons which are having at least one carbon-carbon double or triple bonds are called **unsaturated hydrocarbons**. These are of two types:

   (*i*) **Alkenes:** These are unsaturated hydrocarbons having at least a carbon-carbon double bond.

   or C_2H_4 Ethene

   > The general formula for alkene is C_nH_{2n}, where n is the number of carbon atoms in one molecule of the alkene.

   (*ii*) **Alkynes:** These are unsaturated hydrocarbons having at least one carbon-carbon triple bond.

   $$H—C\equiv C—H$$

   or C_2H_2 Ethyne

   > The general formula of an alkyne is C_nH_{2n-2}, where n is the number of carbon atoms in one molecule of the alkyne.

7.7 REPRESENTATION OF ORGANIC COMPOUNDS

1. **Molecular Formula:** The formula of an organic compound which represents the kind of atoms and the number of each kind of atoms present in one molecule is called **molecular formula.**

   The molecular formula of an organic compound represents:

   (*i*) Kind of atoms present in its one molecule.

   (*ii*) The number of each kind of atoms present in one molecule.

   Example: Molecular formula of butane is C_4H_{10}. It represents:

   (*i*) One molecule of butane consisting of carbon and hydrogen atoms.

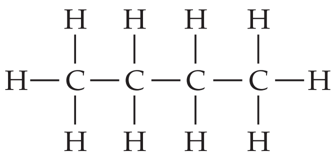
   (*ii*) One molecule of butane consisting of 4 carbon atoms and 10 hydrogen atoms.

2. **Graphic or Structural Formula:** The formula of an organic compound which represents the arrangement of various atoms in one molecule in space is called **graphic or structural formula.**

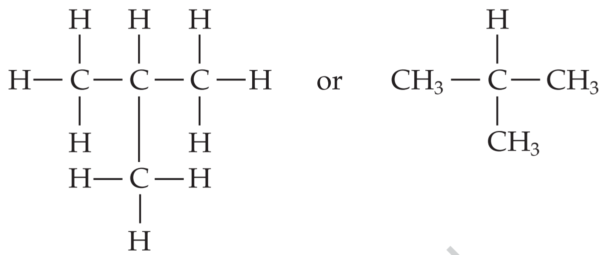
Organic compounds can have same molecular formula, but different structural formulae as shown below:

Molecular formula of butane is C_4H_{10}.

Graphic or structural formulae of butane are:



n-Butane



Iso-butane

3. **Condensed Formula:** A kind of structural formula which indicates the group of atoms joined together to each of the carbon atom in straight or branched carbon chain is called **condensed formula.**

Example: Condensed formula for C_4H_{10} whose graphic formulae are shown above can be shown as

$$CH_3 - CH_2 - CH_2 - CH_3$$

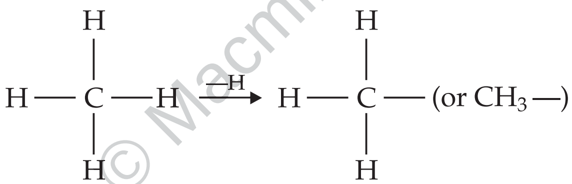
n-Butane

$$CH_3 - \underset{\displaystyle CH_3}{\underset{|}{CH}} - CH_3$$

Isobutane

7.8 ALKYL GROUP

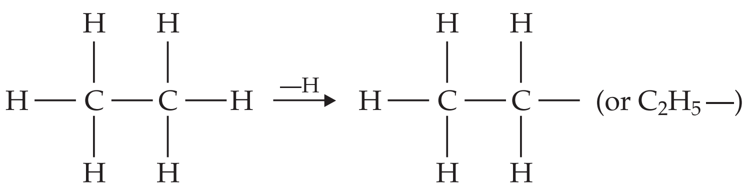
The group formed by the removal of one hydrogen atom from an alkane molecule is called an **alkyl group**. Examples of alkyl group are methyl group (CH_3—), ethyl group (C_2H_5—), etc. Methyl group is formed by the removal of one H atom from methane (CH_4).



Methane　　　Methyl group

The free line (–) shown on the carbon atom of an alkyl group means that one valency of carbon atom is free in alkyl group. The general formula of an alkyl group is C_nH_{2n+1} where n is the number of carbon atoms.

Ethyl group is formed by the removal of one H atom from ethane (C_2H_6).



Ethane　　　Ethyl group

The alkyl groups are usually denoted by the letter 'R'.

In the same way, other alkyl groups like propyl (C_3H_7), butyl (C_4H_9) etc., are formed.

7.9 FUNCTIONAL GROUP

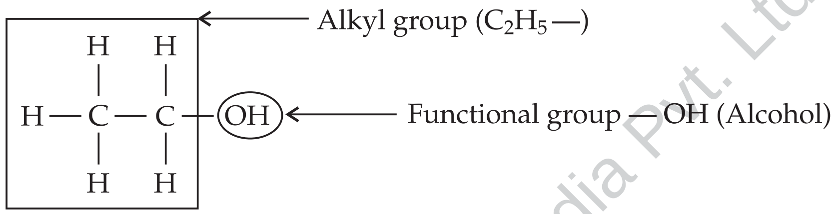
A **functional group** may be defined as an atom or group of atoms present in a molecule which largely determines its chemical properties.

As already stated, hydrocarbons are the parent organic compounds. All other compounds are considered to have been derived from them by replacing one or more of their hydrogen atoms by functional groups.

$$\underset{\text{Hydrocarbon}}{R-H} \xrightarrow[+G]{-H} \underset{\text{Organic compound}}{R-G}$$

In the organic compound R — G, R denotes the carbon hydrogen framework. In case of aliphatic compounds, R is an alkyl group. The group G is the functional group and decides the chemical properties of the compound. The alkyl group determines the physical properties of an organic compound such as melting point, boiling point, density, etc.

Example:



Ethanol

Organic compounds having the same functional group behave similarly and form a family. Thus, on the basis of functional groups present, organic compounds can be divided into families. This makes the study of organic compounds much easier. There exists a homologous series of compounds containing a particular type of functional group. Some important functional groups and the corresponding families are given in Table 7.1.

$$-\overset{|}{C}=\overset{|}{C}- \text{ and } -C\equiv C-$$

are also regarded as functional groups.

Table 7.1: Important Functional Groups and their Families

Functional group	Name	Example
— OH	Alcohol	CH_3-CH_2-OH Ethanol
— X	Halo	CH_3-CH_2-Br Ethyl bromide
— CHO	Aldehyde	CH_3-CHO Acetaldehye
$-\overset{\vert}{C}=O$	Ketone	CH_3COCH_3 Acetone
$-\overset{\overset{OH}{\vert}}{C}=O$	Carboxylic	CH_3COOH Acetic acid

7.10 HOMOLOGOUS SERIES

A **homologous series** is a group of organic compounds having similar structures and similar chemical properties in which the successive compounds differ by CH_2 group. The various organic compounds of a homologous series are called **homologues**. It is clear that the two adjacent homologues differ by 1 carbon atom and two hydrogen atoms (or CH_2 group).

7.10.1 Examples of Homologous Series

1.

Homologous series of alkanes General formula: C_nH_{2n+2} (where n is an integer)	
Molecular formula	**Name**
CH_4	Methane
C_2H_6	Ethane
C_3H_8	Propane
C_4H_{10}	Butane
C_5H_{12}	Pentane
C_6H_{14}	Hexane

2.

Homologous series of alkenes General formula: C_nH_{2n} (where n is an integer)	
Molecular formula	**Name**
C_2H_4	Ethene
C_3H_6	Propene
C_4H_8	Butene
C_5H_{10}	Pentene
C_6H_{12}	Hexene
C_7H_{14}	Heptene

3.

Homologous series of alkynes General formula: C_nH_{2n-2} (where n is an integer)	
Molecular formula	**Name**
C_2H_2	Ethyne
C_3H_4	Propyne
C_4H_6	Butyne
C_5H_8	Pentyne
C_6H_{10}	Hexyne
C_7H_{12}	Heptyne

4.

Homologous series of alcohol	
Molecular formula	**Name**
HCH_2OH	Methanol
C_2H_5OH	Ethanol
C_3H_7OH	Propanol
C_4H_9OH	Butanol
$C_5H_{11}OH$	Pentanol
$C_6H_{13}OH$	Hexanol

7.10.2 Characteristics of a Homologous Series

1. All the members of a homologous series can be represented by the same general formula. For example, all the members of the alkane series can be represented by the general formula C_nH_{2n+2}.
2. Any two adjacent homologues differ by CH_2 group in their molecular formulae. For example, first two homologues of the alkane series CH_4 and C_2H_6 differ by CH_2 group.
3. The difference in the molecular masses of any two adjacent homologues is 14 amu. For example, molecular mass of CH_4 $(12 + 4 \times 1)$ is 16 and that of next homologue C_2H_6 $(2 \times 12 + 6 \times 1)$ is 30. The difference in molecular masses of ethane and methane is 30–16 = 14 (i.e., $CH_2 = 12 + 2 \times 1$).
4. All the compounds of a homologous series show similar chemical properties, as the functional group of the homologues is similar. For example, members of homologous series of alcohols CH_3OH, C_2H_5OH, $CH_3CH_2CH_2OH$ show similar chemical properties.
5. The members of a homologous series show a gradual change in their physical properties with increase in molecular mass.
6. All the members of a homologous series can be prepared by the general method of preparation.

7.10.3 Importance of Homologous Series

1. Homologous series has simplified the study of organic compounds.
2. Since the physical properties show a regular gradation and chemical properties are similar, the behaviour of higher members of the series can be predicted, if the properties of first few members are known.
3. It helps to predict the properties of even those members of the series that are yet to be prepared.

7.11 NOMENCLATURE OF ORGANIC COMPOUNDS

The term nomenclature means the system of naming organic compounds. There are two systems of nomenclature which are commonly used:

(*i*) Trivial or common system, and (*ii*) IUPAC system.

Trivial or Common System: In early days of organic chemistry, each new compound was given an individual name. Such a name was based on the source, some property or some other trivial reason. For example, **urea** got its name since the compound was first obtained from the urine of mammals. Similarly, **methyl alcohol** was called **wood spirit** since it could be obtained by destructive distillation of wood. **Formic acid** was so named as it was obtained by distillation of red ants (Latin, *formica* = ants). **Acetic acid** got its name from *acetum* (Latin, *acetum* = vinegar). A group of sedatives was named **barbiturates** after the name of woman *Barbara*. These names of organic compounds are called trivial or common names.

IUPAC System: With the rapid growth of organic chemistry, the number of organic compounds increased fantastically. It became impossible to give common names to such a large number of compounds.

In 1957, the **International Union of Pure and Applied Chemistry** (IUPAC pronounced as eye-you pack) evolved a scheme for giving systematic names to organic compounds on the basis of their structures. IUPAC is the apex body in chemistry to set rules for naming organic compounds from their structures. This system is known as **IUPAC system** and the given name is known as **IUPAC name.**

The IUPAC name of any organic compound essentially consists of three parts, i.e.,

1. Word root 2. Suffix 3. Prefix

IUPAC Name = Prefix + Word root + Suffix

1. **Word Root:** It is the basic unit of the name and *represents the number of carbon atoms* present in the *parent chain* of the molecule. **Parent chain is the longest possible continuous chain of carbon atoms.** For compounds containing parent chains up to four carbon atoms ($C_1 - C_4$) special word roots are used while for compounds containing parent chains having more than four carbon atoms Greek number roots are used as given in Table 7.2.

Table 7.2: Word Roots for Different Carbon Chains

Chain length	Word root	Chain length	Word root
C_1	Meth–	C_6	Hex–
C_2	Eth–	C_7	Hept–
C_3	Prop–	C_8	Oct–
C_4	But–	C_9	Non–
C_5	Pent–	C_{10}	Dec–

2. **Suffix:** The suffix comes after the word root. It is of two types:

(*i*) ***Primary Suffix:*** It denotes the nature of carbon-carbon bond. In other words, it indicates whether the carbon chain is saturated or unsaturated. The three basic primary suffixes are given in Table 7.3.

Table 7.3: Basic Primary Suffixes

Nature of carbon chain or bond	Primary suffix	General name and formula	
Saturated (C — C)	–ane	Alkane	C_nH_{2n+2}
Unsaturated, having one double bond (C ═ C)	–ene	Alkene	C_nH_{2n}
Unsaturated, having one triple bond (C ≡ C)	–yne	Alkyne	C_nH_{2n-2}

C_nH_{2n+2}	—	CH_4 Methane	C_2H_6 Ethane	C_3H_8 Propane
C_nH_{2n}	—	C_2H_4 Ethene	C_3H_6 Propene	C_4H_8 Butene
C_nH_{2n-2}	—	C_2H_2 Ethyne	C_3H_4 Propyne	C_4H_6 Butyne

(*ii*) ***Secondary Suffix:*** A secondary suffix is then added to the primary suffix to indicate the nature of functional group present in the organic compounds.

3. **Prefix:** The prefix comes before the 'word root' and denotes the branched chains which are considered as substituents. Thus, it denotes the place of substituent alkyl group or functional group.

In naming a hydrocarbon, the following simple rules are followed:

(*a*) ***Selection of carbon chain:*** The longest chain of carbon atoms in the structure of the compound is selected. This determines the word root by giving the number to each carbon.

(*b*) ***Determine the nature of the bond*** between the carbon atoms and add the suffix to the word root. For example,

IUPAC name for $\mathbf{CH_3CH_2CH_2CH_2CH_2CH_3}$

(*i*) Number the carbon atoms in the longest chain.

$$\overset{1}{CH_3} - \overset{2}{CH_2} - \overset{3}{CH_2} - \overset{4}{CH_2} - \overset{5}{CH_2} - \overset{6}{CH_3}$$

The longest chain contains six carbon atoms, so the word root is **Hex**.

(*ii*) All the carbon-carbon bonds are single bonds. Hence, the suffix is **ane.** No prefix is there.

IUPAC name = word root + suffix

Hex + ane = **Hexane**

(*c*) ***If there is a branched chain***, it is considered as substituent.

In naming the hydrocarbons containing side chain, the carbon atom of the selected longest chain are numbered in such a way that the substituents (alkyl groups) get the lowest possible number. For example,

IUPAC name for $$\begin{array}{l} CH_3 - CH - CH_2 - CH_3 \\ \qquad\quad\;\; | \\ \qquad\quad CH_3 \end{array}$$

The hydrocarbon has five carbon atoms. There are 4 carbon atoms in the longest carbon chain, so this compound is a derivative of butane i.e., word root is **but**. There is also one substituent or alkyl group on one of the carbon atom. Therefore, we have to number the carbon chain in such a way that this CH_3 group gets lowest (smallest) number. There are two ways to number i.e., either from left or from right. By doing numbering from left, the CH_3 group gets the lowest number (2) while from right the number is (3).

$$\overset{1}{CH_3}-\underset{\displaystyle CH_3}{\overset{2}{CH}}-\overset{3}{CH_2}-\overset{4}{CH_3} \qquad\qquad \overset{4}{CH_3}-\underset{\displaystyle CH_3}{\overset{3}{CH}}-\overset{2}{CH_2}-\overset{1}{CH_3}$$

2-methylbutane (Correct name) — 3-methylbutane (Wrong name)

Thus, word root is **but**, and suffix is **ane** due to the presence of carbon-carbon single bond, and methyl group at carbon atom 2 is the substituent. Hence, the IUPAC name is **2-methylbutane.**

(*d*) ***If the primary suffix is–ene or –yne,*** the number representing the position of double or triple bond is written before the word root. For example,

IUPAC name for $CH_3-CH_2-CH=CH_2$

The numbering is done in such a way that the double bond gets the lowest number.

$$\overset{4}{CH_3}-\overset{3}{CH_2}-\overset{2}{CH}=\overset{1}{CH_2}$$

Word root is **but** due to four carbon, and suffix is **–ene** due to the presence of double bond, at carbon-1.

Hence, name is **1-butene** or **but-1-ene**.

Similarly, IUPAC name for $CH_3-C\equiv C-CH_3$

$$\overset{1}{CH_3}-\overset{2}{C}\equiv\overset{3}{C}-\overset{4}{CH_3}$$

word root—but, suffix—yne (triple bond)

Hence, IUPAC name is **2-butyne** or **but-2-yne**.

The longest chain or parent chain selected may not be straight but it must be continuous. The other alkyl groups are substituent groups.

$$CH_3-CH_2-\underset{\displaystyle CH_2-CH_2-CH_3}{CH}-CH_2-CH_3$$

Correct chain (C–6)

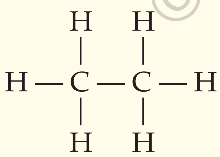
$$CH_3-CH_2-\underset{\displaystyle CH_2-CH_2-CH_3}{CH}-CH_2-CH_3$$

Wrong chain (C–5)

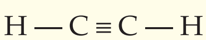
SELF EVALUATION

I. A to H are the structural formulae of some organic compounds:

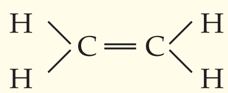
(A)



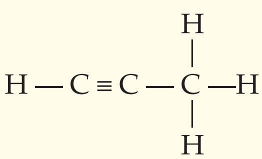
(B)



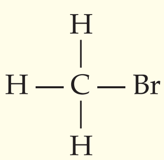
(C)



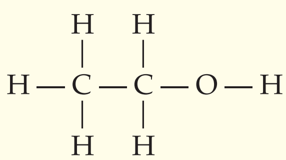
(D)



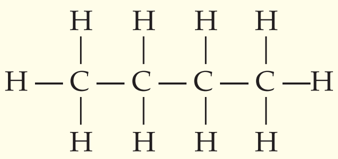
(E)



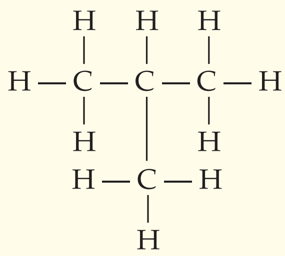
(F)



(G)



(H)



Give the letters which represent

(*i*) two alkanes

(*ii*) two compounds which are not hydrocarbons

(*iii*) B and D are members of a homologous series. Give a reason why this statement is correct.

II. Give the name and formulae of two members of each of the following.

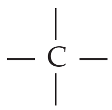
(*i*) Saturated hydrocarbon (alkane)

(*ii*) Unsaturated hydrocarbon (alkene and alkyne)

7.12 ALKANES

Alkanes are the simplest organic compounds made of carbon and hydrogen only. They have the general formula C_nH_{2n+2}, where n = 1, 2, 3, etc.

The carbon atoms in their molecules are bonded to each other by single covalent bond. Each carbon is again bonded to enough hydrogen atoms to give maximum covalence of 4. Since the carbon skeleton of alkane is fully saturated with hydrogens, they are also called **saturated hydrocarbons.**



Alkane

Alkanes contain strong C—C and C—H covalent bonds, so these are chemically inert. Hence, they are also called as **paraffins** (Latin, parum affinis = little affinity).

7.13 ISOMERISM IN ALKANES

Alkanes show **chain isomerism**. Such structural isomers, which differ in the arrangement of carbon chain are called chain isomers. They are also called nuclear isomers.

The first three hydrocarbons of the series (methane, ethane, propane) do not exhibit isomerism. The next alkane, *butane* (C_4H_{10}) exists in two chain isomeric forms.

(*i*) $CH_3 — CH_2 — CH_2 — CH_3$

n-butane

$$\begin{array}{c} \quad CH_3 \\ \quad | \\ CH_3 — CH — CH_3 \end{array}$$

(*ii*) 2-methyl propane or Isobutane

Pentane (C_5H_{12}) exists in three isomeric forms.

(*i*) $CH_3 - CH_2 - CH_2 - CH_2 - CH_3$

n-pentane

$$\begin{array}{l} \phantom{CH_3 - {}}CH_3 \\ \phantom{CH_3 - {}}\,| \\ CH_3 - CH - CH_2 - CH_3 \end{array}$$

(*ii*) 2-methyl butane

$$\begin{array}{l} \phantom{CH_3 - {}}CH_3 \\ \phantom{CH_3 - {}}\,| \\ CH_3 - C - CH_3 \\ \phantom{CH_3 - {}}\,| \\ \phantom{CH_3 - {}}CH_3 \end{array}$$

(*iii*) 2, 2-dimethyl propane

Hexane (C_6H_{14}) exists in five isomeric forms:

(*i*) $CH_3 - CH_2 - CH_2 - CH_2 - CH_2 - CH_3$

n-hexane

$$\begin{array}{l} \phantom{CH_3 - {}}CH_3 \\ \phantom{CH_3 - {}}\,| \\ CH_3 - CH - CH_2 - CH_2 - CH_3 \end{array}$$

(*ii*) 2-methyl pentane

$$\begin{array}{l} \phantom{CH_3 - CH_2 - {}}CH_3 \\ \phantom{CH_3 - CH_2 - {}}\,| \\ CH_3 - CH_2 - CH - CH_2 - CH_3 \end{array}$$

(*iii*) 3-methyl pentane

$$\begin{array}{l} \phantom{CH_3 - {}}CH_3 \\ \phantom{CH_3 - {}}\,| \\ CH_3 - C - CH_2 - CH_3 \\ \phantom{CH_3 - {}}\,| \\ \phantom{CH_3 - {}}CH_3 \end{array}$$

(*iv*) 2, 2-dimethyl butane

$$\begin{array}{l} \phantom{CH_3 - {}}CH_3 \quad CH_3 \\ \phantom{CH_3 - {}}\,| \qquad\;\; | \\ CH_3 - CH - CH - CH_3 \end{array}$$

(*v*) 2, 3-dimethyl butane

7.14 METHANE (CH_4)

It is the first member of alkane chain series. The structural formula of methane is

$$\begin{array}{c} H \\ | \\ H - C - H \\ | \\ H \end{array}$$

7.14.1 Preparation of Methane

Laboratory Method: *From sodium ethanoate (sodium acetate).*

Reactants: Sodium ethanoate (sodium acetate) and soda lime (a mixture of NaOH and CaO).

Procedure: When a mixture of sodium ethanoate and soda lime is heated in a hard glass tube, methane gas is evolved.

Reaction:

$$\underset{\text{Sodium ethanoate}}{CH_3 - \boxed{COONa + NaO} - H} \xrightarrow[\text{Heat}]{CaO} \underset{\text{Methane}}{CH_4} + Na_2CO_3$$

The above reaction is called **decarboxylation**, since one carbon atom is lost from the molecule of the organic compound.

7.14.2 Physical Properties of Methane

1. **Colour, Smell and Taste:** Methane is colourless, odourless and tasteless gas.
2. **Solubility in Water:** Methane is a non-polar covalent compound and hence, is insoluble in water. It is soluble in organic solvents.
3. It is lighter than air.
4. **Melting and Boiling Point:** Melting point and boiling point increase with increase in the carbon number.

7.14.3 Chemical Properties of Methane

Alkanes being saturated hydrocarbons are chemically less reactive. However, they undergo substitution reactions, oxidation and combustion. Methane and ethane being homologues show similar type of reactions.

1. **Combustion:** All hydrocarbons are combustible and burn to give carbon dioxide and water vapours in air.

   Methane burns with a blue flame in air.

$$\underset{\text{Methane}}{CH_4} + \underset{\text{Air (Excess)}}{2O_2} \longrightarrow CO_2 + 2H_2O + \text{Heat}$$

   In a very limited supply of air, methane gives carbon black.

$$\underset{\text{Methane}}{CH_4} + \underset{\text{Air (very limited)}}{O_2} \longrightarrow \underset{\text{Carbon black}}{C} + 2H_2O$$

2. **Oxidation of Methane:** When methane is mixed with oxygen in the ratio of 9 : 1 and is passed through hot copper tube it gets oxidized to methanol or methyl alcohol.

$$\underset{\text{Methane}}{2CH_4} + \underset{\text{Oxygen}}{O_2} \xrightarrow[200\,°C]{Cu} \underset{\text{Methanol or methyl alcohol}}{2CH_3OH}$$

OR

$$\underset{\text{Methane}}{CH_4} + \underset{\text{Oxygen}}{O_2} \xrightarrow[\text{Heat}]{\text{Molybdenum oxide}} \underset{\text{Formaldehyde or methanal}}{HCHO} + H_2O$$

   Methanol on further oxidation with acidified potassium dichromate yield aldehydes and carboxylic acids.

$$\underset{\text{Methanol}}{CH_3OH} \xrightarrow[\text{acidic} - H_2O]{K_2Cr_2O_7\ +[O]} \underset{\text{Methanal or formaldehyde}}{HCHO} \xrightarrow[\text{acidic} - H_2O]{K_2Cr_2O_7\ +[O]} \underset{\text{Methanoic acid or formic acid}}{HCOOH}$$

3. **Pyrolysis:** The thermal decomposition of an organic compound at high temperatures is called pyrolysis. It takes place in the absence of air. Pyrolysis of hydrocarbons is called cracking. It means breaking of C—H bonds and forming unsaturated hydrocarbons.

Cracking is the process of breaking up of a higher hydrocarbon into lower hydrocarbons by the action of heat and catalyst. Common catalysts used in cracking are alumina (Al_2O_3) or silica (SiO_2).

$$\underset{\text{Methane}}{2CH_4} \xrightarrow[Al_2O_3]{1500°C} \underset{\text{Ethyne}}{HC \equiv CH} + 3H_2$$

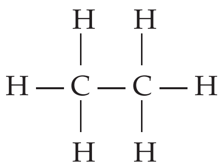
7.14.4 Uses of Methane

1. As a domestic fuel in the form of natural gas.
2. In the manufacture of methanol and hydrogen.
3. In the manufacture of carbon black which is used in printing inks, shoe polishes, gramophone records and rubber tyres.
4. In the manufacture of halogen derivatives such as CH_2Cl_2, $CHCl_3$, CCl_4, etc., which are used as solvents in laboratory and industry.

   $CHCl_3$ (chloroform) is used as a local anaesthetic. Chloromethane is used as a refrigerant. CCl_4 is used in fire extinguishers under the name of **Pyrene.**

7.15 ETHANE (C_2H_6)

It is the second member of the alkane chain series. The structural formula of ethane is



7.15.1 Preparation of Ethane

Laboratory Method: *From ethyl iodide.*

Reactants: Ethyl iodide and nascent hydrogen.

Procedure: Reduction with nascent hydrogen using zinc-copper couple and aqueous ethyl alcohol.

Reaction:

$$\underset{\text{Ethyl iodide}}{C_2H_5I} + 2[H] \xrightarrow[C_2H_5OH]{\text{Zn-Cu couple}} \underset{\text{Ethane}}{C_2H_6} + HI$$

7.15.2 Physical Properties of Ethane

1. **Colour, Smell and Taste:** It is a colourless, odourless, tasteless and non-poisonous gas.
2. **Solubility in Water:** It is sparingly soluble in water. It is soluble in acetone, ether, alcohol, etc.
3. The vapours are heavier than air.

7.15.3 Chemical Properties of Ethane

1. **Combustion:** Ethane is highly combustible. It burns to give carbon dioxide, water vapour and heat. Methane burns with a blue flame in air.

$$\underset{\text{Ethane}}{2C_2H_6} + \underset{\text{Air (Excess)}}{7O_2} \longrightarrow 4CO_2 + 6H_2O + \text{Heat}$$

2. **Oxidation of Ethane:** Ethane gets oxidized to alcohol and aldehyde.

$$\underset{\text{Ethane}}{2C_2H_6} + O_2 \xrightarrow[200\,°C]{Cu} \underset{\text{Ethanol or ethyl alcohol}}{2C_2H_5OH}$$

$$\underset{\text{Ethane}}{C_2H_6} + O_2 \xrightarrow[Heat]{MoO} \underset{\text{Acetaldehyde or ethanal}}{CH_3CHO} + H_2O$$

Ethanol on further oxidation with acidified potassium dichromate yield aldehydes and carboxylic acids.

$$\underset{\text{Ethanol or ethyl alcohol}}{C_2H_5OH} \xrightarrow[acidic - H_2O]{K_2Cr_2O_7 + [O]} \underset{\text{Ethanal or acetaldehyde}}{CH_3CHO} \xrightarrow[acidic]{K_2Cr_2O_7 + [O]} \underset{\text{Ethanoic acid or acetic acid}}{CH_3COOH}$$

3. **Pyrolysis of Ethane:**

$$\underset{\text{Ethane}}{CH_3 - CH_3} \xrightarrow[Al_2O_3]{500\,°C} \underset{\text{Ethene}}{H_2C = CH_2} + H_2$$

7.15.4 Uses of Ethane

1. Ethane is used as a fuel. It has high calorific value than methane. Liquefied ethane is also used as a fuel.
2. In the preparation of ethanol, acetaldehyde and acetic acid which find use in paints, varnishes adhesive, plastic, etc.
3. For making hexachloroethane which is artificial camphor.
4. It is a ripening agent for fruits.
5. It is used in producing welding gas and is a primary ingredient in mustard gas.

7.16 UNSATURATED HYDROCARBONS

The open chain hydrocarbons which contain lesser number of hydrogen atoms than the corresponding alkanes containing same number of carbon atoms are called **unsaturated hydrocarbons.** These are characterised by the presence of double or triple bonds.

7.17 ALKENES

Alkenes are unsaturated hydrocarbons that contain a carbon-carbon double bond ($-\overset{|}{C}=\overset{|}{C}-$) in their molecule.

They have **general formula C_nH_{2n}** (n = number of carbon atoms). Alkenes contain two hydrogen atoms less than alkanes and are designated as **unsaturated hydrocarbons.** As a class, alkenes are commonly known as **Olefins** (Latin, *Oleum* = Oil, *ficare* = to make) because the lower members form oily products on treatment with chlorine or bromine. Being highly reactive, alkenes rarely occur free in nature.

No alkene with one carbon atom is possible as there must be a double bond between two carbon atoms.

7.18 ISOMERISM IN ALKENES

Alkenes with four or more carbon atoms can form isomers.

Alkenes show different isomerism.

1. **Chain Isomerism:** Chain isomers are molecules with the same molecular formula but different arrangements of two carbon skeletons. For example,

   (*i*) $CH_3 - CH_2 - CH = CH_2$

   1-Butene

   (*ii*) $\begin{array}{l} CH_3 - \underset{\displaystyle CH_3}{\underset{|}{C}} = CH_2 \end{array}$

   2-methyl propene

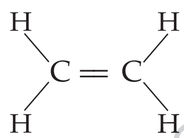
2. **Position Isomerism:** Position isomers are molecules with the same molecular formula but different movements of bonds (double, triple) or functional groups. For example,

   (*i*) $CH_3 - CH_2 - CH = CH_2$ (*ii*) $CH_3 - CH = CH - CH_3$

   1-Butene 2-Butene

7.19 ETHENE (ETHYLENE) C_2H_4

The simplest member of the alkene family is ethane (C_2H_5). The structural formula of ethene is



7.19.1 Preparation of Ethene

Laboratory Method: *By dehydration of ethyl alcohol.*

Reactants: Ethyl alcohol (Ethanol) and Conc. H_2SO_4 in the ratio of 1 : 2.

Procedure: Ethene is prepared by heating a mixture of ethyl alcohol and concentrated sulphuric acid at 170 °C, as a result ethyl alcohol decomposes into ethene.

Reaction:

$$\underset{\text{Ethyl alcohol or ethanol}}{CH_3 - CH_2OH} \xrightarrow[170°C]{\text{Conc. } H_2SO_4} \underset{\text{Ethene}}{CH_2 = CH_2} + H_2O$$

7.19.2 Physical Properties of Ethene

1. **State:** Colourless gas, neutral to litmus, faint sweet odour.
2. **Solubility:** Slightly soluble in water, highly soluble in organic solvents.
3. **Density:** Slightly lighter than air (vapour density of ethene is 14).

7.19.3 Chemical Properties of Ethene

Alkenes are reactive compounds due to the presence of double bond or unsaturation. Hence, alkenes undergo addition reaction.

1. **Addition Reaction:** The reactions in which molecules of the attacking reagent add across the double or triple bond of an unsaturated compound to yield saturated compound.

$$\underset{\text{(Unsaturated)}}{\underset{\text{Alkene}}{>C=C<}} + \underset{\text{molecule}}{\underset{\text{Attacking}}{XY}} \longrightarrow \underset{\text{(Saturated)}}{\underset{\text{Addition product}}{>\underset{X}{\overset{|}{C}}-\underset{Y}{\overset{|}{C}}<}}$$

(*i*) *Addition of Hydrogen (Hydrogenation):* Ethene reacts with hydrogen gas when heated in presence of a catalyst like nickel, to give ethane. The addition of hydrogen across the double bond is called hydrogenation.

$$\underset{\text{Ethene}}{CH_2 = CH_2} + H_2 \xrightarrow[300°C]{Ni} \underset{\text{Ethane (saturated)}}{CH_3 - CH_3}$$

Finely divided palladium or platinum may also be used as catalyst.

(*ii*) *Addition of Bromine:* Ethene reacts with cold and pure bromine water. The bromine looses its original red brown colour and gives a colourless liquid. The double bond breaks and bromine atoms get attached to each carbon atom.

$$\underset{\text{Ethene}}{CH_2 = CH_2} + Br_2 \longrightarrow \underset{\text{1, 2-dibromoethane}}{\underset{Br}{\overset{|}{CH_2}} - \underset{Br}{\overset{|}{CH_2}}}$$

2. **Oxidation Reactions:**

(*i*) *Combustion:* Ethylene burns in air with a luminous smoky flame to form carbon dioxide and water. The reaction is highly exothermic.

$$\underset{\text{Ethene}}{CH_2 = CH_2} + 3O_2 \longrightarrow 2CO_2 + 2H_2O + \text{Heat}$$

(*ii*) *Oxidation with Alkaline Potassium Permanganate:* With cold and dilute aqueous or alkaline solution of potassium permanganate **(Baeyer's reagent)**, it gets oxidised to form ethylene glycol.

$$\underset{\text{Ethene}}{CH_2{=}CH_2} + H_2O + O \xrightarrow[KMnO_4]{25° - 30°C} \underset{\text{(ethylene glycol)}}{\underset{\text{1, 2-ethanediol}}{\overset{OH}{\overset{|}{CH_2}} - \overset{OH}{\overset{|}{CH_2}}}}$$

The purple colour of potassium permanganate gets discharged in this reaction. Therefore, this reaction is used as a test for unsaturation (Baeyer's test).

7.19.4 Uses of Ethene

1. It is used for the manufacture of polythene – a plastic material.
2. For the artificial ripening of fruits.
3. As a general anaesthetic.
4. As a starting material for a large number of useful compounds, for example:

   Ethylene glycol—as antifreeze agent

   Oxyethylene flame—cutting and welding metals

   Ethanol—for cosmetics and toilet products

   Ethylene oxide—for making detergents
5. For making poisonous mustard gas (war gas).

7.20 ALKYNES

Alkynes are unsaturated hydrocarbons which contain at least one carbon-carbon triple bond ($-C{\equiv}C-$) in their molecules. The general formula of alkynes is C_nH_{2n-2}.

Alkynes are highly unsaturated than the alkenes. Alkynes contain four hydrogen atoms less than the corresponding alkanes and are characterised by the presence of triple bond in the molecule. The first and the most important member of this series is acetylene (ethyne), $HC{\equiv}CH$. Hence, alkynes are also called **acetylenes**.

7.21 ISOMERISM IN ALKYNES

Alkynes with four or more than four carbon atoms can form isomers.

Alkynes show different isomerism.

1. **Chain Isomerism:** Example: C_5H_8

   (*i*) $CH_3-CH_2-CH_2-C\equiv CH$

   1-Pentyne

   (*ii*)
   $$\begin{array}{c} \quad\quad\;\; H \\ \quad\quad\;\; | \\ CH_3-C-C\equiv CH \\ \quad\quad\;\; | \\ \quad\quad\;\; CH_3 \end{array}$$

   3-methyl but-1-yne

2. **Position Isomerism:** Example C_4H_6

   (*i*) $CH_3-CH_2-C\equiv CH$

   1-Butyne

   (*ii*) $CH_3-C\equiv C-CH_3$

   2-Butyne

7.22 ETHYNE (ACETYLENE) C_2H_2

It is the first member of alkyne series. Source of alkynes is natural gas and petroleum. Traces of ethyne is present in coal gas and in gas obtained by the decomposition of certain organic compounds. The structural formula of ethyne is

$$H-C\equiv C-H$$

7.22.1 Preparation of Ethyne

Laboratory Method: *From calcium carbide.*

Reactants: Calcium carbide and water.

Procedure: Ethyne is prepared when calcium carbide reacts with water. Ethyne is collected by downward displacement of water.

Reaction:

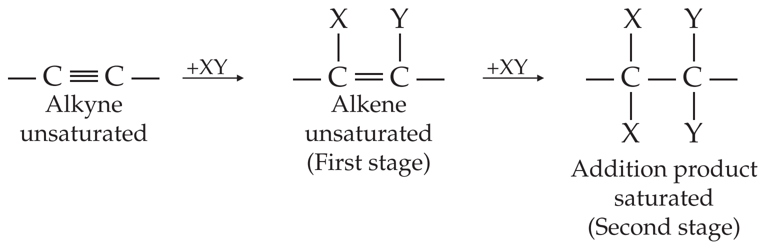
$$\underset{\text{Calcium carbide}}{CaC_2} + 2H_2O \longrightarrow \underset{\text{Ethyne (Acetylene)}}{C_2H_2} + Ca(OH)_2$$

7.22.2 Physical Properties of Ethyne

1. **State:** Colourless gas, faint garlic odour.
2. **Solubility:** Slightly soluble in water, more soluble in organic solvents.
3. **Density:** Slightly lighter than air
4. **Nature:** Slightly acidic.

7.22.3 Chemical Properties of Ethyne

1. **Addition Reactions:** Since alkynes contain a — C ≡ C — triple bond, they undergo addition reactions at triple bond in two stages



Following are the examples:

(i) *Addition of Hydrogen (Catalytic Hydrogenation):* Addition of hydrogen takes place in presence of nickel, platinum or palladium catalyst. In first stage, two hydrogen atoms are added to ethyne to form ethene. In second stage, two hydrogen atoms are added to ethene to give ethane (saturated).

First stage

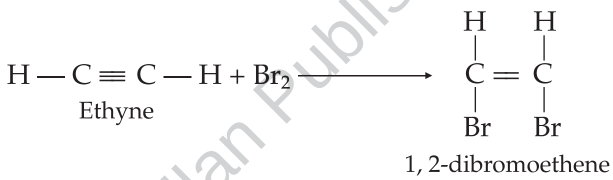
$$\underset{\text{Ethyne}}{HC \equiv CH} + H_2 \xrightarrow[300°C]{Ni} \underset{\text{Ethene}}{H_2C = CH_2}$$

Second stage

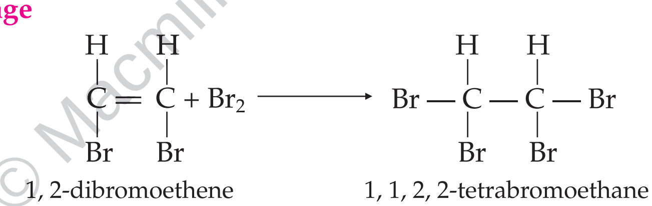
$$\underset{\text{Ethene}}{H_2C = CH_2} + H_2 \xrightarrow[300°C]{Ni} \underset{\text{Ethane}}{H_3C - CH_3}$$

(ii) *Addition of Bromine:* Ethyne reacts with bromine to give 1,2-dibromoethene which in turn reacts with bromine to produce 1, 1, 2, 2-tetra-bromoethane.

First stage



Second stage

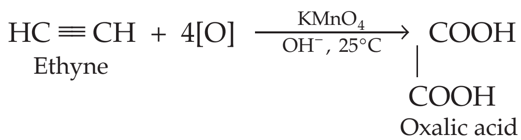


2. **Combustibility:** Acetylene burns in an atmosphere of pure oxygen to produce extremely high temperature (3000ºC) and a luminous but smoky flame.

$$\underset{\text{Ethyne}}{2C_2H_2} + 5O_2 \longrightarrow 4CO_2 + 2H_2O + \text{Heat}$$

This reaction is used in **oxyacetylene torch** for welding and cutting of metals.

3. **With Alkaline $KMnO_4$:** Acetylene (ethyne) decolourises the purple colour of alkaline $KMnO_4$ solution and gets oxidized to oxalic acid.



7.22.4 Uses of Acetylene

1. As an illuminant.
2. For the production of oxyacetylene flame. The temperature of the flame is above 3000ºC. It is used for cutting and welding metals.
3. For artificial ripening of fruits.
4. As a general anaesthetic under the name *Naracylene.*
5. As a starting material for the manufacture of a large variety of useful substances. For example, industrial solvents such as Westron ($C_2H_2Cl_4$) and Westrosal (C_2HCl_3).

Table 7.4: Comparison of Saturated and Unsaturated Hydrocarbons

S.No.	Saturated Organic Compounds	Unsaturated Organic Compounds
1.	All the four valencies of each carbon atom are satisfied by forming single covalent bonds with carbon and with hydrogen atoms.	1. The valencies of at least two carbon atoms are not fully satisfied by the hydrogen atoms.
2.	Carbon atoms are joined only by a single covalent bonds. $-\overset{\vert}{\underset{\vert}{C}}-\overset{\vert}{\underset{\vert}{C}}-$	2. Carbon atoms are joined by double covalent bonds $-\overset{\vert}{C}=\overset{\vert}{C}-$ or by triple covalent bonds. $-C\equiv C-$
3.	They are less reactive.	3. They are more reactive.
4.	They generally undergo substitution reactions.	4. They generally undergo addition reactions.

Table 7.5: Chemical Tests to Distinguish between Ethane, Ethene and Ethyne

S. No.	Chemical tests	Ethane	Ethene	Ethyne
1.	Combustion	Burns with a non-luminous flame	Burns with a luminous and slightly sooty flame.	Burns with a luminous and sooty flame.
2.	With bromine or bromine water	No change	Orange colour of bromine disappears.	Orange colour of bromine disappears.
3.	With alkaline $KMnO_4$ (Baeyer's test)	No change	Purple colour of $KMnO_4$ is discharged.	Purple colour of $KMnO_4$ is discharged.

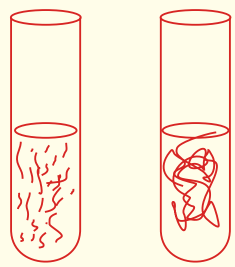
A luminous flame is obtained during combustion of unsaturated hydrocarbons due to greater proportion of carbon in unsaturated hydrocarbons (ethene and ethyne) as compared to alkanes.

A 7.1

Find which one is unsaturated hydrocarbon.

Procedure: Take two test tubes A and B. Put 3–4 drops of hydrocarbon sample in each test tube. Now add one drop of 2% $KMnO_4$ solution in each test tube and shake vigorously using small corks to stop the ends of the test tubes. Record your observations for each test tube.

Conclusion: The purple colour of the $KMnO_4$ changes to brown colour. This indicates that a reaction has taken place.



A B

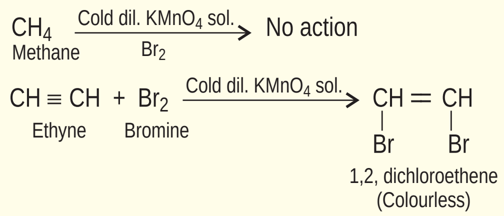
Fig. 7.2

A 7.2

How will you prove methane is a saturated hydrocarbon, while ethyne is an unsaturated hydrocarbon?

Procedure: Take two test tubes. Put Sample 1 in test tube A and Sample 2 in test tube B. Add the reagents provided. Record your observation and write the conclusion.

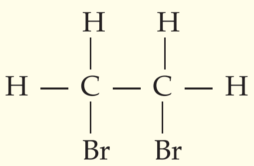
Conclusion: Methane is a saturated hydrocarbon while ethyne is an unsaturated hydrocarbon. Methane being a saturated hydrocarbon does not, decolourise Br_2 in cold dil. alkaline solution of $KMnO_4$ while ethyne decolourises.



On the basis of above experiment we can say methane is a saturated compound while ethyne or acetylene is an unsaturated hydrocarbon.

SELF EVALUATION

1. Bromine is used to produce dibromoethane which is used as an additive in petrol. It is made by adding bromine to ethene.



   (*i*) What is the molecular formula of ethene?

   (*ii*) What is the molecular formula of dibromoethene?

2. Given the molecule CH≡CH, the symbol, '≡' represents

   (*i*) one pair of shared electrons

   (*ii*) two pairs of shared electrons

   (*iii*) three pairs of shared electrons

   (*iv*) four pairs of shared electrons

7.23 FOSSIL FUELS

Dead remains of plants and animals which get buried deep under the Earth's crust over millions of years ago are said to have formed **fossils**. Some fossil fuels are coal, petroleum and natural gas.

7.23.1 Coal

Coal is the world's most widely used fuel which was formed by the decaying of large number of plants buried under the earth millions of years ago due to natural calamities such as earthquake, volcanoes, etc. The main constituent of coal is carbon. Small amount of nitrogen and sulphur is also present in coal.

The concentration of carbon present in coal determines the quality and types of coal. Based on the carbon content, coal is classified as Lignite (about 38% of carbon), Bituminous (about 65% of carbon), and Anthracite (about 95% of carbon).

In Bhutan, major coal mine is at Deothang, Samdrup Jongkhar.

Coal can be used directly or in other forms such as coke, coal gas, etc. **Coke** is coal without volatile materials. Some of the important uses of coal are:

(*i*) It is used as a source of heat energy in industries.

(*ii*) It is used as a fuel in thermal power plants.

(*iii*) It is used in the preparation of synthetic petrol and synthetic natural gas.

7.23.2 Petroleum

The word 'petroleum' is derived from Latin words '*Petra*' means 'rock' and '*Oleum*' means oil. Petroleum is formed by the decomposition of dead remains of organisms buried deep inside the Earth's crust under high pressure and temperature, millions of years ago. It is dark viscous oily liquid. Thus, we can define petroleum as a dark, viscous oily liquid formed by the decomposition of dead remains of organisms deep inside the Earth's crust under high pressure and temperature conditions. Petroleum is mainly a mixture of hydrocarbons. It also contains other compounds of carbon and sulphur, oxygen and nitrogen in small quantity.

Petroleum is obtained from deep underground or under sea by drilling holes up to the oil bed or oil field and is called **crude oil.**

Crude oil obtained from the oil field has to undergo processing in order to obtain various usable products. This is called **processing of crude oil.** Crude oil obtained from the oil fields is taken to the **oil refiner.** In refineries, the crude oil is first washed with basic or acidic solution depending on the nature of the impurities. The washed crude oil is then subjected to fractional distillation and a number of fractions or products of petroleum are obtained (Fig. 7.3). Some of the fractions of petroleum are petrol, paraffin wax, diesel, petroleum gas, fuel oil, kerosene, lubricating oil and asphalt.

7.23.3 Petroleum Gas

Petroleum gas is a very good fuel. It is obtained as a byproduct during fractional distillation of petroleum. **Petroleum gas** is a mixture of butane, ethane and propane. Butane is the main constituent in petroleum gas. For domestic purposes petroleum gas is provided in sealed cylinders in the form of **Liquefied**

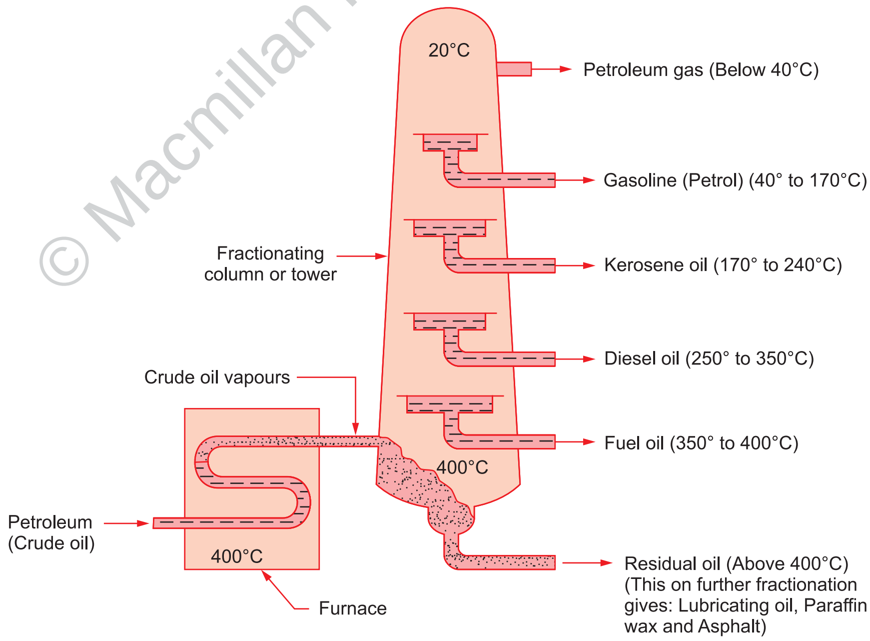


Fig. 7.3 *Fractional distillation of petroleum*

Petroleum Gas (LPG). Liquefied Petroleum Gas mainly consists of butane or isobutane with small quantities of ethane and propane. To detect leakage of the gas, foul smelling ethyl mercaptan (C_2H_5SH) is added to the gas.

Fractional Distillation of Crude Oil: If the difference in boiling points of two liquids is not much, simple distillation cannot be used to separate them. The technique of fractional distillation is used in such cases. In this technique, vapours of a liquid mixture are passed through a fractionating column before condensation. The fractionating column is fitted over the mouth of the round bottom flask. Vapours of the liquid with higher boiling point condense before the vapours of the liquid with lower boiling point. The vapours rising up in the fractionating column become richer in more volatile component. By the time the vapours reach to the top of the fractionating column, these are rich in the more volatile components. A fractionating column provides many surfaces for heat exchange between the ascending vapours and the descending condensed liquid. Some of the condensing liquid in the fractionating column obtains heat from the ascending vapours and revapouries. The vapours thus become rich in low boiling components. The vapours of low boiling components ascend to the top of the column and become pure in low boiling component and the pure liquid is collected in a receiver. After a series of successive distillations, the remaining liquids in the distillation get, enriched in high boiling components.

7.23.4 Natural Gas

Natural gas is another fossil fuel obtained as a by-product of mining. Natural gas occurs above crude oil in the oil field. It is a rich source of energy. It mainly consists of methane (95%) and a small quantity of ethane and propane. Natural gas is a clean fuel as it does not produce any polluting gases or particulate matter. Natural gas when compressed under high pressure is called **Compressed Natural Gas (CNG)** which is used as source of energy for automobiles.

In Bhutan, as of now petroleum products and CNG are imported from neighbouring countries. Some uses of natural gas are:

(*i*) It is used in the tyre industry as a source of carbon black.

(*ii*) It is used in fertiliser industries as a source of hydrogen.

(*iii*) It is used as a fuel in automobiles and industries.

7.23.5 Conditions of Combustion

What happens when a substance is burnt? When a substance is burnt, it combines with oxygen and in the process heat is produced. **Combustion** is burning of a fuel in presence of oxygen to produce heat. Thus combustion is an oxidation as well as an exothermic process. The substance which burns is called **combustible substance**.

Chemical Reactions:

$$\underset{\text{Methane}}{CH_4} + 2O_2 \longrightarrow CO_2 + 2H_2O$$

There are certain conditions which must be fulfilled for combustion of a substance. These conditions are:

(*i*) The substance to burn must be heated to its ignition temperature before it starts burning. The temperature at which a substance starts burning is called its **ignition temperature**.

(*ii*) There must be a supporter of combustion. Oxygen in air is the supporter of combustion. Combustible substances burn at a moderate rate in air, but at a faster rate in oxygen.

(*iii*) The substance to be burnt must be combustible. Wood, paper, coal, petrol, diesel, coke, LPG, kerosene, hydrogen, ethyl alcohol, natural gas, etc. are some examples of combustible substances.

7.23.6 Cracking

The process of decomposition of bigger hydrocarbon molecules into smaller hydrocarbon molecules by means of heat and catalyst is known as **cracking**, e.g., when decane is heated strongly in presence of catalyst, it cracks to give a mixture of low boiling *n*-pentane and pentene.

> **A 7.3**
>
> Make a list of ten materials that, you use in your daily life and list out, the products that they give on combustion. Are the products different when substances used are hydrocarbons?

$$\underset{\text{(Decane)}}{C_{10}H_{22}} \xrightarrow[\text{Cracking}]{\text{Heat catalyst}} \underset{(n\text{-pentane})}{C_5H_{12}} + \underset{\text{(pentene)}}{C_5H_{10}}$$

Thus, when bigger hydrocarbons (having higher boiling point) are subjected to cracking, a mixture of saturated and unsaturated smaller hydrocarbons (having low boiling point) is obtained which can be used as petrol. In fact, the petrol, obtained by cracking, has better properties than straight chain petrol. Thus, cracking is useful to increase the yield of petrol as well as to obtain good quality of petrol.

1. **Thermal Cracking:** When cracking is done by using heat only, it is called thermal cracking.
2. **Catalytic Cracking:** When cracking is carried out in the presence of both heat and catalyst, it is known as catalytic cracking. The quantity and yield of gasoline (petrol) is much better when it is obtained by catalytic cracking of heavy oils. The catalyst used for this purpose is generally aluminium silicate [Al_2SiO_5] or alumina (Al_2O_3).

7.24 POLYMERS

In Greek, '*poly*' means many and '*mer*' means unit. Polymers are consisting of large number of repeating structural units derived from simple molecules. The process of joining together a large number of simple small molecules (monomers) to make high molecular mass polymer through covalent bond is termed as **polymerisation**.

7.24.1 Classification of Polymers

1. **Addition Polymers:** A polymer formed by the direct repeated addition of monomer is called addition polymer. In these type of polymers, the monomers are unsaturated compounds and are generally derivatives of ethene. For example, polyethene, polypropylene and polyvinyl chloride.

> **Property of Addition Polymers**
>
> Addition polymer is commonly used as initiator because it decomposes under mild conditions to form *t*-butoxide radical.

2. **Condensation Polymers:** Condensation polymerisation involves a series of condensation reactions generally involving two different monomers. Each monomer normally contains two functional groups. During these reactions, there is a loss of small molecules usually water. For example, nylon, terylene, alkyd resins, bakelite, etc.




7.24.2 Polyethylene or Polythene Formation

It is obtained by polymerisation of ethylene under high pressure of 1000 to 2000 atmospheres at a temperature of 350 to 570 K in presence of trace of oxygen or a peroxide which initiates the polymerisation. The polymer obtained by this procedure of free radical addition and H-atom abstraction has highly branched structure and is called **low density polyethene.** Its mode of formation is depicted below:

$$\underset{\text{Ethene}}{n\mathrm{CH_2=CH_2}} \xrightarrow[\substack{\text{High pressure} \\ \text{1000–2000 atm}}]{\text{350–570 K}} \underset{\text{Polythene}}{-\!\!\left(\mathrm{CH_2-CH_2}\right)_n\!\!-}$$

Low density polyethene is chemically inert, tough but flexible and poor conductor of electricity.

Uses:

1. In the insulation of electric wires.
2. In the manufacture of squeeze bottles, toys and flexible pipes.

When polymerisation of ethylene is carried out in the presence of a catalyst at 330 to 350 K at atmospheric pressure, the polymer obtained has a linear structure and is called **high density polyethene.** It is also chemically inert but relatively tough and hard with high tensile strength.

Uses:

1. In the manufacture of house wares, bottles, pipes.
2. In the manufacture of containers (buckets, tubes, dustbins, etc.)

7.24.3 Polyvinyl Chloride (PVC)

PVC is a thermoplastic. It is the third-most widely produced synthetic plastic polymer, after polyethylene and polypropylene. PVC is a white brittle solid. It is insoluble in alcohol but slightly soluble in tetrahydrofuran. It is obtained by polymerisation of vinyl chloride under high temperature and pressure.

$$\underset{\text{Chloroethane}}{n\mathrm{CH_2=\underset{\mid \atop Cl}{CH}}} \xrightarrow{\text{Polymerisation}} \underset{\text{Polyvinyl chloride}}{-\!\!\left(\mathrm{CH_2-\underset{\mid \atop Cl}{CH}}\right)_n\!\!-}$$

Uses:

1. In the manufacture of raincoats, hand bags, curtain clothes, toys, etc.
2. In artificial flooring.
3. For making gramophone records.

SELF EVALUATION

1. What do you mean that a hydrocarbon can go through combustion? What are we really doing to the hydrocarbon?
2. How does the combustion of methane compare to the combustion of other hydrocarbons such as coal and oil?

RECAPITULATION

1. Organic chemistry is the study of hydrocarbons and their derivatives.
2. Organic compounds include hydrocarbons and their derivatives.
3. Hydrocarbons are the covalent compounds of carbon and hydrogen.
4. Catenation is the property by virtue of which atoms of the same element get linked together through covalent bonds so as to form long straight, branched or closed chains or rings. For example, a carbon atom shows catenation to maximum extent.
5. Carbon atom shows tetra covalency because it has four electrons in its valence shell.
6. An atom or group of atoms which decides the chemical properties of a compound is called functional group.
7. Structural formula gives us the relative arrangements of bonded atoms in a molecule.
8. Homologous series is a series of compounds having similar structural formulae, same functional group and hence similar chemical properties.
9. Saturated hydrocarbons are those compounds which contain $>C-C<$ single bonds in their molecules.
10. Unsaturated hydrocarbons are those compounds which contain $>C=C<$ (double bond) or $-C \equiv C-$ (triple bond) in their molecules.
11. Pyrolysis is the decomposition of higher molecular weight hydrocarbons into lower molecular weight hydrocarbons on heating to a high temperature.
12. The compounds having same molecular formula but different properties are called isomers and this phenomenon is called isomerism.
13. Dead remains of plants and animals which get buried deep under the Earth's crust over millions of years ago became fossil fuels. Some fossil fuels are coal, petroleum and natural gas.
14. The word 'petroleum' is derived from the Latin word '*Petra*' means rock and '*Oleum*' means oil. Petroleum can be defined as a dark viscous oily liquid formed by decomposition of dead remains of organisms deep inside the Earth's crust.
15. Petroleum gas is a mixture of butane, ethane and propane, butane being the main constituent.
16. Petroleum as it is obtained from the oil field is called crude oil.
17. Natural gas, a fossil fuel obtained as a by-product of mining, when gets compressed under high pressure, is called CNG (Compressed Natural Gas).
18. Combustion is burning of a fuel in presence of oxygen to produce heat.
19. Polymers are formed by the combination of large number of simple molecules called monomers.
20. Polymerisation is the process by which monomers are converted into polymers.
21. Polymers found in nature are called natural polymers.
22. Synthetic polymers are man-made polymers prepared in the laboratory.

EXERCISES

I. True or False Statements

State whether the following statements are true (T) or false (F):

1. Acetylene is more acidic than ethylene.
2. Organic compounds contain ionic linkages.
3. The large number of organic compounds is due to catenation.
4. $-\overset{|}{\underset{|}{C}}-\overset{|}{\underset{|}{C}}-$ bond is longer than $-\overset{|}{C}=\overset{|}{C}-$ bond.
5. Monomer of polyvinyl chloride is 1-chloroethene.

II. Fill in the Blanks

Fill in the blanks by using suitable word(s):

1. Chain isomerism arises due to different arrangements of ______________.
2. The organic compounds containing carbon and hydrogen atoms only are called ______________.
3. Petroleum gas obtained by ______________.
4. Compressed ______________ is a source of energy.
5. Polymer prepared by man is called ______________.

III. Multiple Choice Questions

Tick (✓) the only correct choice amongst the following:

1. On mixing a certain alkane with chlorine irradiating it with ultraviolet light, it forms only one monochloro alkane. The alkane could be
   (*a*) neopentane (*b*) propane (*c*) pentane (*d*) isopentane
2. Coal can be formed from
   (*a*) sunlight (*b*) steam (*c*) fossils (*d*) plants
3. Refining is
   (*a*) extracting petroleum gas (*b*) separation of various fractions of petroleum
   (*c*) heating of coal (*d*) sedimentation of fossil fuel
4. CNG is stored under
   (*a*) power generator (*b*) electric generator
   (*c*) solvent (*d*) none of these
5. Which of the following is not a natural polymer?
   (*a*) Cellulose (*b*) Starch (*c*) Polythene (*d*) Natural rubber

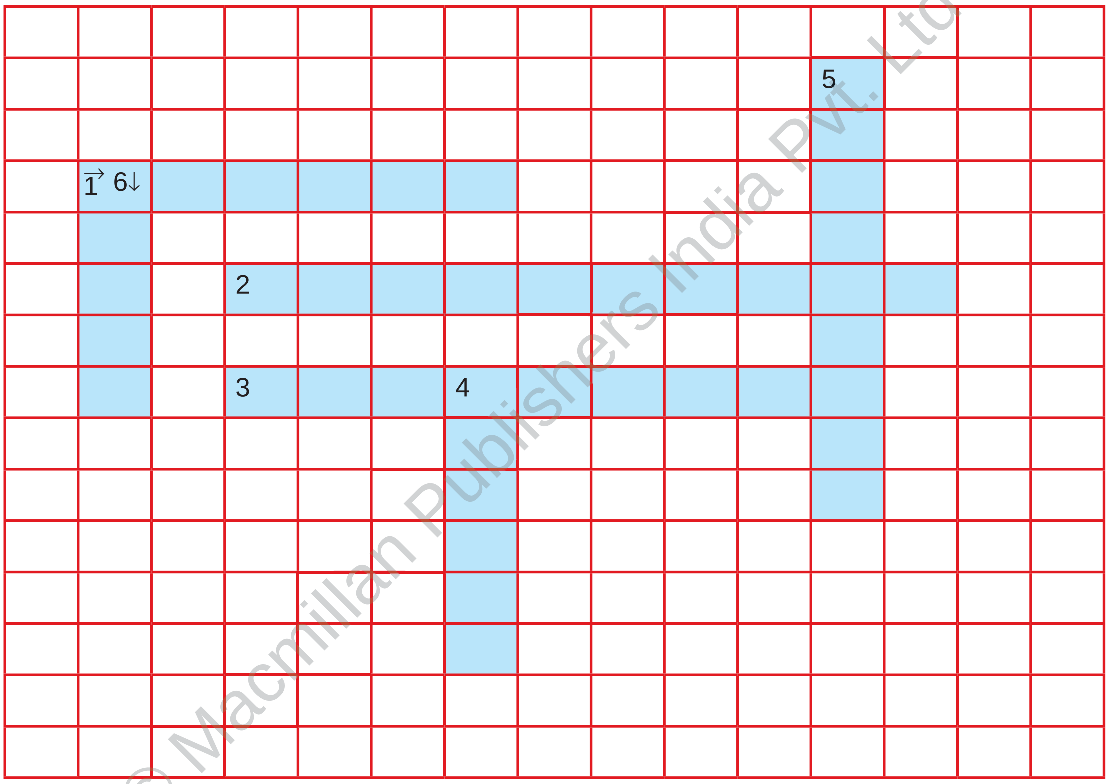
IV. Answer the Following Questions

1. Name the chain isomer of C_5H_{12} which has a tertiary hydrogen atom.
2. A compound is formed by the substitution of two chlorine atoms for two hydrogen atoms in propane. What is the number of structural isomers possible?
3. Arrange the following in the increasing order of C–C bond length: C_2H_6, C_2H_4, C_2H_2.

4. Name two reagents which can be used to distinguish between ethene and ethyne.
5. What happens when water is added to calcium carbide? Give chemical equation for the reaction.
6. Explain the term ‘Polymerisation’ with two examples.
7. Which of the acyclic hydrocarbon series is represented by C_nH_{2n}? Write the structural formulae and IUPAC names of all the hydrocarbons of this type having four carbons per molecule.
8. Write the preparation of polythene and polyvinyl chloride.
9. What is meant by chain isomerism? Discuss the isomerism of alkanes up to pentanes.
10. What are alkynes? Write different types of isomerisms in alkynes.

V. Puzzle

Read the given clues carefully and write the appropriate word inside the boxes to solve the puzzle.



Across:

1. A compound having same molecular formula but different structural formula is called.
2. Pyrolysis of hydrocarbon.
3. The commonly known alkane.

Down:

4. A six carbon alkane.
5. An important natural source of hydrocarbons obtained from underground oil deposits.
6. A set of rules for naming organic compounds from their structures.

[Hints: 1. Isomer 2. Catenation 3. Methane 4. Hexane 5. Petroleum 6. IUPAC.]

Web Links:

For more information on the following topics, visit the websites listed below:

Organic Compounds

https://www.britannica.com/science/organic-compound

Sources of Organic Compound

http://www.qsstudy.com/chemistry/sources-organic-compounds.html

Hydrocarbons

https://www.britannica.com/science/hydrocarbon

Functional groups and homologous series

http://ex ernal.webstorage.gr/images/Books- PDF/9780199151424.pdf

Nomenclature

http://www.chemguide.co.uk basicorg/conventions/names.html

Alkane, Alkene, Alkyne and their Isomers

http://www.introorganicchemistry.com/nomenclature.html

Fossil Fuels

https://www.studentenergy.org/topics/fossil-fuels

Glossary

Absolute scale of temperature: A scale of temperature based on absolute zero temperature.

Acid: A substance which dissolves in water, producing H^+ (*aq*) ions. Acid solutions turn litmus red and have a pH below 7. In their reactions acids act as proton donors.

Activation energy: The additional energy required by the molecules to attain threshold energy.

Activity series: A series of chemical elements arranged in the order of their electrode potential.

Addition polymer: A polymer formed by the direct repeated addition of monomers.

Alkali: A base that dissolves in water to give hydroxide ions.

Alkali metals: Group-1 elements Li, Na, K, Rb, Cs which are never found in elemental state in nature.

Alkali earth metals: The elements forming group-2 of the periodic table, i.e. Be, Mg, Ca, Sr, Ba and Ra.

Ammonification: When the plants and animals die, the putrefying bacteria and fungi present in the soil decompose the proteins of dead plants and animals into ammonia.

Analytical Chemistry: Analytical chemistry involves the qualitative and quantitative determination of the chemical compounds.

Anion: It is formed when an atom gains one or more electrons in its outermost shell and achieve negative charge.

Aqua regia: It is made by mixing one part nitric acid and three parts hydrochloric acid.

Atom: The smallest portion of an element that can take part in a chemical reaction.

Atomicity: It corresponds to the number of atoms in a given molecule.

Azidothymidine: It is used for helping AIDS victims.

Barometer: Instrument used for measuring atmospheric pressure.

Base: A compound releasing hydroxyl ions in aqueous solution, having pH greater than 7.

Basic oxide: They are formed by the combination of oxygen with highly electropositive elements.

Bimolecular reaction: It involves simultaneous collision between two species.

Binding energy: The amount of energy released when protons and neutrons bind themselves together to form the nucleus of an atom.

Biological fixation: The fixation of nitrogen by bacteria and algae is called biological fixation of nitrogen.

Boiling point: The temperature at which the vapour pressure of the liquid becomes equal to the atmospheric pressure.

Bond: It refers to the attractive forces that keeps ions, atoms or group of atoms bound together.

Bond Dissociation Enthalpy

Bond dissociation enthalpy: The change in enthalpy, when one mole of covalent bonds of a gaseous covalent compound is broken to form products in the gaseous phase.

Catalyst: A substance which changes the rate of a reaction without itself undergoing any permanent chemical change.

Cathode: In an electrolytic cell, the negative electrode from which current flows.

Cation: It is formed when an atom loses one or more electrons from its outermost shell and achieve positive charge.

Chain isomerism: Chain isomers are molecules with the same molecular formula but different arrangement of two carbon skeletons.

Charge density: The electric charge per unit volume of a body.

Chemical equation: A chemical equation is a shorthand representation of a chemical reaction using the symbols and formulae of substances involved in the chemical reaction.

Chemical reaction: A process in which one or more chemical elements or compounds form new compounds.

Closed system: A system which can exchange energy but there is no matter exchange with the surrounding.

Collision frequency: The total number of collisions taking place among the reacting species in a unit time.

Combination reaction: The reaction of two or more substances to produce a new substance.

Combustion: The oxidation reaction, in which carbon dioxide and water are formed along with the liberation of heat and light.

Compound: Any substance which consists of two or more different types of atoms in a fixed proportion. It is formed by chemical reaction.

Condensation: In the closed vessel, the molecule evaporating from the liquid surface are confined to limited space. These molecules may collide among themselves or with the molecules of air and in that process are pushed back to the surface of the liquid.

Condensed formula: A kind of structural formula which indicates the group of atoms joined together to each of the carbon atom in straight or branched carbon chain.

Coordinate bond: A bond formed between two atoms by sharing a pair of electrons, provided entirely by one of the combining atom.

Covalent bond: Covalent bond is that bond which involves mutual sharing of electrons between two atoms having contributed equal number of electrons.

Crude oil: Petroleum obtained from deep underground or under sea by drilling holes up to the oil bed or oil field.

Decomposition reaction: A chemical reaction in which a compound breaks down into two or more simpler compounds or elements.

Degenerate orbitals: Orbitals of equal energy.

Density: Density of a substance is its amount of mass per unit volume.

Diffusion: General transport of matter whereby molecules or ions mix through normal thermal agitation.

Dilute solution: Any solution with a small ratio of the quantities of solute to solvent.

Dilution: Process of decreasing concentration of solute by addition of solvent.

Dissociation: The process of separation of ions in water already as such in the solid state of the solute.

Dissociation energy: The amount of energy spend on dissociation of crystal when subjected to heat or any other source of energy.

Electrochemistry: The branch of chemistry concerned with electrolysis and other similar phenomena.

Electrochemical series: A series in which the elements are arranged in decreasing magnitude of their standard oxidation potentials.

Electrode potential: A potential exhibited by an electrode of any element in contact with a solution of its ions.

Electrode: An electroconducting body which when passed into an electrolyte exhibits a certain electrical potential w.r.t. the bulk of the electrolyte.

Electrolysis: The process of decomposing a substance usually in solution or as a melt, by the passage of an electric current.

Electron gain enthalpy: It is the enthalpy change taking place when an isolated gaseous atom accept an electron to form monovalent anion.

Electron: It is the ultimate indivisible negative charge or particle of negative electricity and forms an integral part of every atom.

Electronegativity: It is the tendency of an atom in a molecule to attract towards itself the shared pair of electrons.

Element: A substance which cannot be further divided by chemical methods.

Endothermic reactions: Those reactions which are accompanied by the absorption of heat.

Energy: Energy necessary to push against a force of 1N for 1m. Its unit is the joule.

Exothermic reactions: Those reactions which are accompanied by the evolution of heat.

Fat: An ester which is formed by the reaction of a fatty acid with any glycerol.

Fertilizers: These are the substances which are added to the soil in order to remove the deficiency of essential elements required for the proper growth of plants.

Fibres: Fibres are the thread forming solids which possess high tensile strength and high modulus.

Filter: A device used for separating solid particles from a liquid or gas.

Filtrate: The clear liquid obtained by filtration.

Fission: It is a process in which a large nuclei breaks up to give smaller nuclei.

Fluid: The substance which can flow.

Force of attraction: These are the intermolecular forces which act between the neighbouring particles (atoms, molecules or ions). They are weak forces.

Fossil: Dead remains of plants and animals which got buried deep under the Earth's crust over millions of years ago are said to have formed fossil.

Free radicals: Atoms or groups of atoms with unpaired electrons and therefore are highly reactive. They can be produced by high energy radiation.

Fuel: A substance burned to produce heat energy.

Functional group: A functional group may be defined as an atom or group of atoms present in a molecule which largely determines its chemical properties.

Fundamental properties: Basic rule which have to be applied in every law application is called fundamental properties.

Gangue: The clay and other silicate material present in ores.

Gas: The gaseous state is the most diffuse state of matter in which molecules have almost unrestricted motion.

Gel: A mixture formed between a solid and a liquid in which the solid forms a network that traps the liquid so that it cannot flow freely.

Green chemistry: The branch of chemistry that deals with the process and products that reduce or eliminate the use and generation of hazardous substances.

Ground state: Most stable lowest energy state of an atom, molecule or system.

Group: A vertical column of the periodic table containing elements with similar properties.

Half-life period: The time taken for the radioactivity in a sample of a radioisotope to fall to half of its original value.

Halogenation: A process resulting in the incorporation of halogens by addition.

Halogens: The elements fluorine, chlorine, bromine, iodine and astatine of group VII.

Heat of atomisation: The amount of heat required to dissociate one mole of an element into its gaseous atoms.

Heat of combustion: Heat change when one mole of a substance is completely burnt or oxidised in oxygen.

Heat of formation: The heat change that takes place when one mole of the substance is formed from its elements under given condition of temperature and pressure. It is usually represented by ΔH_f.

Heat of fusion: Heat of fusion is the heat change that takes place when one mole of a solid substance changes into its liquid state at its melting point.

Heat of hydration: The amount of heat change (i.e. the heat evolved or absorbed) when one mole of the anhydrous salt combines with the required number of moles of water so as to change into the hydrated salt.

Heat of neutralization: The heat change when one gram equivalent of an acid is neutralized by a base, the reaction being carried out in dilute aqueous solution.

Heat of solution: The heat change (i.e. amount of heat evolved or absorbed) when one mole of the substance is dissolved in such a large volume of the solvent that further addition of the solvent does not produce any more heat change.

Heat of sublimation: It is a process in which a solid on heating changes directly into its gaseous state below its melting point.

Heat of vaporisation: It is the heat change that takes place when one mole of a liquid changes into its gaseous state at its boiling point.

Heterogeneous reaction: A reaction which occurs between substances in different phases.

Homogeneous reaction: A reaction which occurs between substances in same phase.

Homologous series: A homologous series is a group of organic compounds having similar structures and similar chemical properties in which the successive compounds differ by $-CH_2$ group.

Humidity: It is mass of water vapour per unit mass of dry air.

Hydrogen bond: The attractive force between an H-atom and a highly electronegative atom (F, N and O).

Hydrogenation: A specific method of reduction in which hydrogen is added to a substance by the direct use of gaseous hydrogen.

Ionic crystal: A crystal composed of ions of two or more elements. The positive and negative ions are arranged in definite patterns and are held together by electrostatic attraction.

Ideal gas: It is a gas that follows Boyle's law, Charles's law and Avogadro's law strictly. Such gas is a hypothetical.

Inert gases: They are highly non-reactive and show no tendency to react with other elements.

Inner transition elements: *f*-block elements are commonly known as inner transition elements. Inner transition elements include lanthanoids and actinoids.

Inorganic chemistry: It is the study of the properties and behaviour of inorganic compounds.

Ionic bond: A type of chemical bond characterised by transfer of electrons from one atom to another.

Ionic radius: A value assigned to the radius of an ion in a crystalline solid, based on the assumption that the ions are spherical with a definite size.

Ionization enthalpy or ionization energy: The minimum energy required to remove the outermost electron from neutral gaseous atom to form ion.

Irreversible reaction: The reactions which occur only in one direction, under ordinary conditions of temperature and pressure.

Isolated system: A system which can exchange neither energy nor matter with the surrounding.

Isomers: These are two or more compounds that have the same chemical formula but a different arrangement of atoms.

Isomerism: The phenomenon due to which different compounds show same molecular formulae but differ in their physical or chemical properties.

Isotones: Atoms of different elements which contain the same number of neutrons.

Isotopes: Atoms with same atomic number but different mass number.

Joule: Unit of energy, symbol J.

Kolbe's method: A method of making alkanes by electrolysing a solution of a carboxylic acid salt.

Krypton: It is a noble gas having symbol "Kr". The element is practically inert and forms very few compounds like KrF_2.

Lamp black: A soft black carbon pigment obtained by incomplete combustion of natural gas or petroleum.

Lanthanoids: It is a series of elements in the periodic table from lanthanum to lutetium in which $4f$ orbitals are being filled.

Lime water: It is Calcium Hydroxide $Ca(OH)_2$.

Liquid paraffin: A highly refined white oil used medicinally as powder laxative.

Liquified petroleum gas: Mixture of C_3 and C_4 hydrocarbon obtained from petroleum refining operations.

Litmus: A natural colouring matter obtained from lichens after oxidation in the presence of NH_3.

Lone pair: A pair of electrons in a molecule which is not shared by two of the constituent atoms.

Law of conservation of mass: In all physical and chemical changes, the total mass of the reactants is equal to the mass of products.

Marsh gas: Methane formed by rotting vegetation in marshes.

Melting point: The temperature at which a solid changes into a liquid.

Mendeleev's periodic law: The physical and chemical properties of the elements are periodic functions of their atomic weights.

Metal: An element that is usually a hard crystalline solid, a good conductor of heat and electricity, and can lose electrons easily.

Metallic bond: A type of a chemical bond which holds together the atoms in a solid metal or alloy.

Metallic character: It is the tendency of atoms of an element to lose electrons and form positive ion.

Metalloids: Metalloids show the properties of metal as well as non-metal.

Methane: It is an alkane having molecular formula CH_4.

Methanol: It is a colourless liquid made by the dry distillation of wood and its molecular formula is CH_3OH.

Modern periodic law: The physical and chemical properties of elements are the periodic functions of their atomic numbers.

Molecular formula: The formula of an organic compound which represents the kind of atoms and the number of each kind of atoms present in one molecule.

Natural gas: Mixture of gaseous hydrocarbons coming out of earth, mainly methane.

Natural polymers: These polymers are found in nature, generally obtained from plants and animals.

Natural rubber: It may be considered as a linear polymer of isoprene and is also called polyisoprene.

Nitrification: A process in which ammonium is converted to nitrate by decomposition.

Nitrogen cycle: The circulation of nitrogen through the living and non-living components of the biosphere.

Nitrogen fixation: The process of converting atmospheric nitrogen into ammonium nitrates and nitrites which can easily be used by plants.

Noble gases: The elements He, Ne, Ar, Kr, Xe and Rn are the members of Group-18 of the periodic table. They are also known as inert gases as they are highly non-reactive and show no tendency to react with other elements.

Normal boiling point: It is the temperature at which the vapour pressure of a liquid is equal to 1 atm or 760 mm Hg.

Normal elements: These are the elements of sub-group 'A'.

Octet: It is a stable group of eight electrons in the outer shell of an atom.

Olefins: Alkenes are generally known as olefins.

Open system: In an open system, there is exchange of energy and matter between system and the surrounding.

Orbit: The path of an electron as it travels round the nucleus of an atom.

Organic chemistry: The branch of chemistry concerned with compounds of carbon.

Oxidation reaction: A reaction in which addition of oxygen or electronegative element takes place.

Oxides: These are the binary compounds formed between elements and oxygen.

Oxidising agent: It is a chemical species that removes an electron from another species.

Period: These are the horizontal rows in a periodic table. Modern periodic table contains seven periods.

Periodic law: The physical and chemical properties of the elements are periodic functions of their atomic weights.

Periodic table: A table in which the chemical elements are arranged in order of their increasing atomic numbers or atomic weights.

Periodicity: The repetition of the similar properties of the elements after a regular interval of time in the periodic table.

Permanent change: It is a chemical change in which one or more new substance are produced. This change cannot be reversed easily.

Pesticide: A chemical compound that kills or prevents the growth of pests which damage crops.

Petrochemicals: It refers to those chemicals which are obtained from petroleum and natural gas.

Petroleum: A complex mixture of hydrocarbons formed over a period of millions of years by decomposition of dead remains of plants and animals. Deep inside the earth's crust under high pressure and temperature. It is a dark viscous oily liquid.

Petroleum gas: It is a very good fuel and is a mixture of butane, ethane, and propane. Butane is a main constituent.

Photoelectric effect: The ejection of electrons from the metal surface when light falls on it.

Photosynthesis: The process by which green plants make their food from carbon dioxide and water in the presence of chlorophyll and sunlight.

Plasticizers: Substances which are added to a synthetic resin to make it flexible.

Polar covalent bonds: A covalent bond in which electrons are shared inequally and the bonded atoms acquire a partial positive and negative charge.

Polymers

Polymers: These are high molecular mass substances consisting of large number of repeating structural units called monomers.

Pyrolysis: It is the thermal decomposition of an organic compound at high temperatures.

Quicklime: A white amorphous powder obtained by heating limestone in vertical rotary kiln to 800–1000 °C. It is used for the manufacture of bleaching powder, sugar, glass, washing soda, etc.

Radon: It is a radioactive colourless atomic gas of group 18. It is used in radiotherapy. Its chemical symbol is Rn, atomic number is 86 and relative atomic mass is 222.

Rare earth metals: A group of metals (atomic number 57 to 71) most of which are extremely difficult to extract because of their very similar chemical properties.

Rate of chemical reaction: It is the change in the amount of the reactants or the change in the amount of the products in a given time interval.

Reactants: The substances which combine or react to give products are known as reactants.

Reagent: A substance which is added to a system to bring about a chemical reaction and can be used in testing for ion, radical or chemical substance.

Redox reaction: A chemical reaction in which oxidation and reduction process takes place simultaneously.

Reducing agent: A substance which provides hydrogen or any other electropositive element or removes oxygen or any other electronegative element.

Renewable energy sources: Sources of energy that do not use up the earth's finite mineral sources.

Respiration: The process of taking oxygen into cells, using it for energy release and then eliminating the waste products.

Reversible reaction: The chemical reactions which take place in both forward as well as in reverse direction.

Reversible: Reversible under strict thermodynamic sense is a special way of carrying out a process such that system is at all times in perfect equilibrium with its surroundings.

Rhizobium: A nitrogen fixing bacteria present in the root nodules of leguminous plants.

Rusting: Corrosion of iron due to the formation of hydrated iron (III) oxide ($Fe_2O_3 \cdot xH_2O$). This process occurs in the presence of water and oxygen.

Saturated hydrocarbons: Hydrocarbons which contain only carbon—carbon single bond. The IUPAC name for a homologous series of such compounds is alkane.

Screening or shielding effect: The effective nuclear charge experienced by a valence electron in an atom will be less than the actual charge on the nucleus because of shielding or screening of the valence electrons from the nucleus by the intervening core electrons.

Soap: These are used for cleaning purposes. They are sodium or potassium salts of long chain fatty acids, e.g. stearic, oleic, and palmitic acids.

Solid: It is a state of matter which has three-dimensional regularity of structure. They have definite volume and definite shape.

State function: These are those functions which depend only on the initial state and the final state of the system.

State Function

Stoichiometry

Stoichiometry: It deals with the calculation of masses (sometimes volume also) of two reactants and the products involved in a chemical reaction.

Straight chain hydrocarbons: The names of such compounds are based on their chain structure, and end with suffix 'ane' and carry a prefix indicating the number of carbon atoms present in the chain.

Structural formula: The formula of an organic compound which represents the arrangement of various atoms in one molecule in space.

Surroundings: The portion of the universe other than the system.

System: The part of the universe chosen for thermodynamic consideration.

Thermal cracking: A cracking done by using heat only.

Thermochemical equation: A balanced chemical equation not only indicating the quantities of the different reactants and products but also the amount of heat evolved or absorbed.

Transition elements: These are the elements of groups 3 to 12.

Unsaturated hydrocarbons: Hydrocarbons which contain double or triple bonds between carbon and carbon i.e.

$$-\overset{|}{C}=\overset{|}{C}-,\ -C\equiv C-$$

Uranium: A radioactive metallic element, which is used as a source of nuclear energy.

Valence electron: The electrons present in outermost orbit of an atom, which can take part in forming chemical bond.

Valency: It is the capacity of atoms of a given element tend to combine with or replace atoms of hydrogen.

van der Waals force: It is the attraction of intermolecular forces between the molecules. These are much weaker than chemical bonds.

Weathering: It is the process of breaking down or dissolving minerals and rocks on the earth's surface.

Xenon: It is a colourless, odourless gas of Group-18. It is used in photographic flash tubes.

Yeast: A microscopic fungus consisting of single oval cells that reproduce by budding, and capable of converting sugar into alcohol and carbon dioxide.

Specimen Question Paper

Chemistry | **Time:** 2 Hours
Class IX | **Total Marks:** 100

READ THE FOLLOWING DIRECTIONS CAREFULLY:

1. Do not write during the first fifteen minutes. This time is to be spent on reading the questions. After having read the question, you will be given two hours to answer all questions.
2. In this paper, there are two sections: A and B. Section A is compulsory. You are expected to attempt any five questions from Section B.
3. The intended marks for questions or parts of questions, are given in brackets [].
4. Read the directions to each question carefully and write all your answers in the answer sheet provided separately.

Section A (50 Marks)

Compulsory: Attempt **all** *questions.*

Question 1

a. *Each question in this section is provided with four possible options. Choose the most appropriate option.* [$1 \times 25 = 25$]

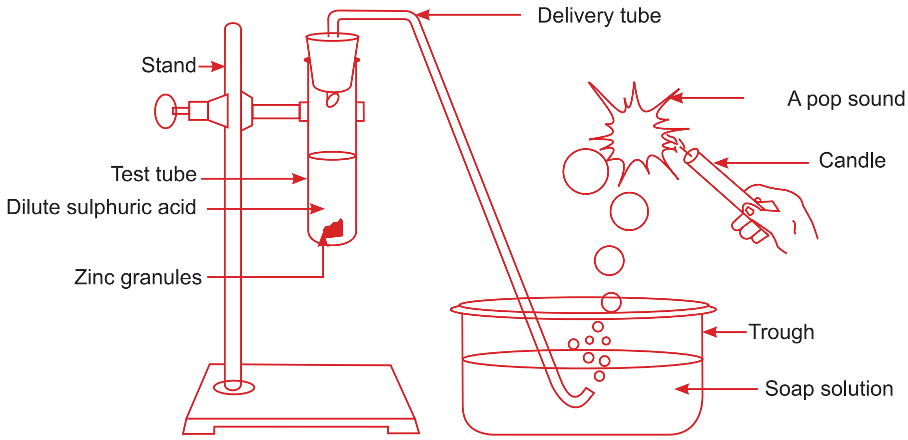
i. The correct order of increasing radius of the elements Si, Al, Na and P is
   A. Si, Al, P, Na B. P, Si, Al, Na
   C. Al, Si, P, Na D. Al, P, Si, Na

ii. Which of the following properties does not depend on periodicity?
   A. Atomic weight B. Atomic number
   C. Ionisation energy D. Electronegativity

iii. Which of the following is required for the formation of an ionic bond?
   A. An electron from the more electronegative element should be transferred to the less electronegative element.
   B. The total energy of the resulting molecule should be less than the total energy of the reactants.
   C. The lattice energy of the resultant molecule should be as low as possible.
   D. The ionic potentials of the reactant should be identical.

iv. The bonds present in N_2O_5 are
   A. Only ionic bond B. Covalent bond or Coordinate bond
   C. Only covalent bond D. Covalent bond and ionic bond

v. Name the gas evolved in the following experiment:



A. SO_2 B. CO_2 C. NH_3 D. H_2

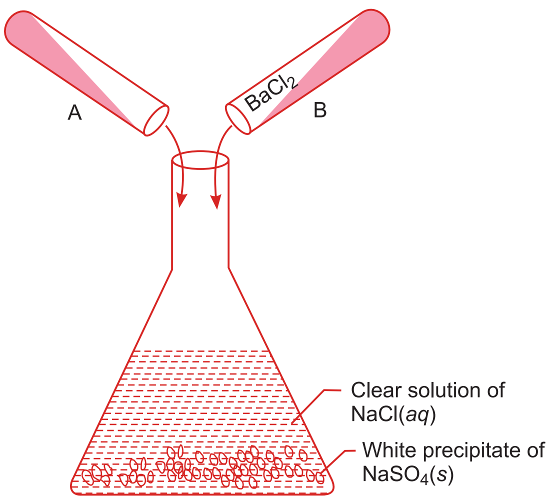
vi. Aqua regia is made by the mixing of

A. One part HNO_3 and three parts HCl
B. One part H_2SO_4 and three parts HCl
C. Two parts HNO_2 and three parts H_2SO_4
D. One part HCl and three parts HNO_3

vii. CO_2 reacts with carbon to form carbon monoxide. This is an example of

A. Displacement reaction
B. Isomerisation reaction
C. Addition reaction
D. Decomposition reaction

viii. Puran poured clear solution of test-tube A and equal quantity of solution of test-tube B in a conical flask. He shook it well and kept undisturbed for some time. After some time he observed that the white precipitate is settled down at the bottom and a clear solution is seen in the flask. Puran forgot to write the name of the solution of test-tube A. What it could be?



A. Barium sulphate
B. Sodium chloride
C. Barium chloride
D. Sodium sulphate

ix. Which of the following statements is correct for the activation energy of a reaction?

A. It increases with increase in temperature
B. When the activation energy is zero, the rate constant is temperature dependent
C. It decreases with decrease in temperature
D. It is nearly independent of temperature over a wide range.

x. Thermodynamics does not deal with

A. the feasibility of a chemical reaction
B. energy changes involved in a chemical reaction
C. the extent to which a chemical reaction proceeds
D. the rate at which a reaction occurs.

xi. Which statements among the following does not come under green chemistry?

A. Use of plastic cans for storage of substances.

B. Using H_2O_2 for bleaching purpose instead of chlorine based bleaching agents.

C. Using soap made from vegetable oils instead of using synthetic detergents.

D. Use of bicycle for travelling small distances instead of gasoline driven vehicles.

xii. What is true about ozone hole?

A. It is formed in troposphere.

B. It is vanishing of ozone layer completely around the earth.

C. It is releasing of ozone from stratosphere to troposphere.

D. It is thinning of ozone layer of stratosphere at some places on earth.

xiii. The formula of westrosol formula is

A. $CH_3 - CH_2 - CCl_3$ B. CCl_4

C. $CH_3 - \underset{\overset{|}{Cl}}{CH} - \underset{\overset{|}{Cl}}{CH_2}$ D. $CHCl = CCl_2$

xiv. Which of the following is a synthetic polymer?

A. Cellulose B. Polyvinyl chloride

C. Starch D. Natural rubber

xv. What happen when calcium carbide react with water and formed

A. CH_4 B. C_2H_6

C. C_2H_2 D. C_2H_4

xvi. In the flame test, which of the following is the colour of the flame by sodium salt?

A. Green B. Red

C. Violet D. Yellow

xvii. The radioactive noble gas is

A. Helium B. Radon

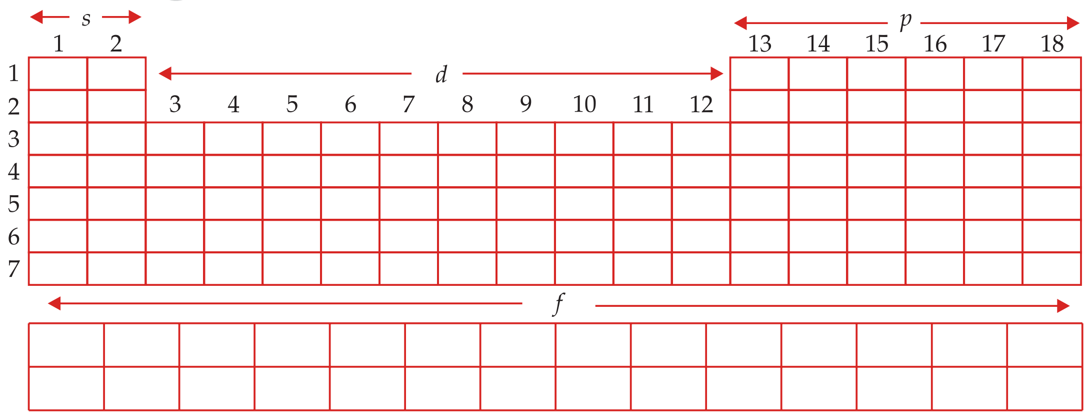
C. Krypton D. Argon

xviii. Molten NaCl conducts electricity due to the presence of

A. free electrons B. free ions

C. free molecules D. atoms of sodium and chloride

xix.



Elements of group 3 to 12 are known as

A. normal element B. transition element

C. inner transition element D. representative element.

xx. Rusting of iron is a

A. fast reaction B. slow reaction

C. redox reaction D. addition reaction

xxi. The energy that favours the dissociation of solute in water is known as

A. hydration energy B. lattice energy

C. affinity energy D. bond energy

xxii. Root nodules are found in

A. leguminous plant B. herbs

C. trees D. all plants

xxiii. One of the characteristic properties of a polymer material is

A. high temperature stability B. high elongation

C. high mechanical strength D. low hardness

xxiv. On mixing a certain alkane with chlorine indicating it with ultraviolet light, it forms only one monochloro alkane. The alkane could be

A. neopentane B. propane

C. pentane D. isopentane

xxv. Which of the following is not a natural polymer?

A. Cellulose B. Starch

C. Polythene D. Natural rubber

b. *Match each item under Column A with the most appropriate item in Column B. Rewrite the correct matching pairs in the answer sheet provided.* [5]

Column A	Column B
(i) Photoelectric effect	(a) Covalent compounds
(ii) Bad conductor of heat and electricity	(b) Decomposition reaction
(iii) Thermal stability of oxidation of metal	(c) Stoichiometry
(iv) Calculation of relative amounts of substances in a chemical reaction with evolution of heat	(d) Decreases from top to bottom in activity series
(v) Heating $CaCO_3$ in an example of	(e) Exothermic

c. *Fill in the blanks by writing suitable word(s).* [5]

i. The group 15 elements are commonly known as ________________.

ii. In ________________ compounds ions are arranged in a regular pattern to form a crystal lattice.

iii. Rusting of iron is a ________________ reaction.

iv. PVC is a widely used ________________.

v. Alkynes are ________________ hydrocarbons.

d. *State whether the following statements are 'True' or 'False' and correct the false statements:* [5]

i. Except lead, all metals are solid.

ii. Chemical reaction can indicate the concentration of the reactant and the product.

iii. Decomposition of calcium carbonate is a reversible reaction.

iv. Ginkgo fixes atmospheric nitrogen during fixation of nitrogen.

v. Monomer of polythene is ethene.

e. *Answer the following questions:* [5]

i. The horizontal rows in periodic table are called.
ii. A solid consists of an irregular shape.
iii. How much percentage of nitrogen found in ammonium sulphate fertilizer?
iv. The hydrocarbons contain a carbon—carbon double bond.
v. Alkanes are also known as.

f. *Write down one difference between the following pairs.* [5]

i. Metals and non-metals
ii. Oxidation reaction and reduction reaction
iii. Reversible reaction and irreversible chemical reaction
iv. Ethane and ethene
v. Chain isomerism and position isomerism

Section B (50 Marks)

*Attempt any **five** Questions.*

Question 2

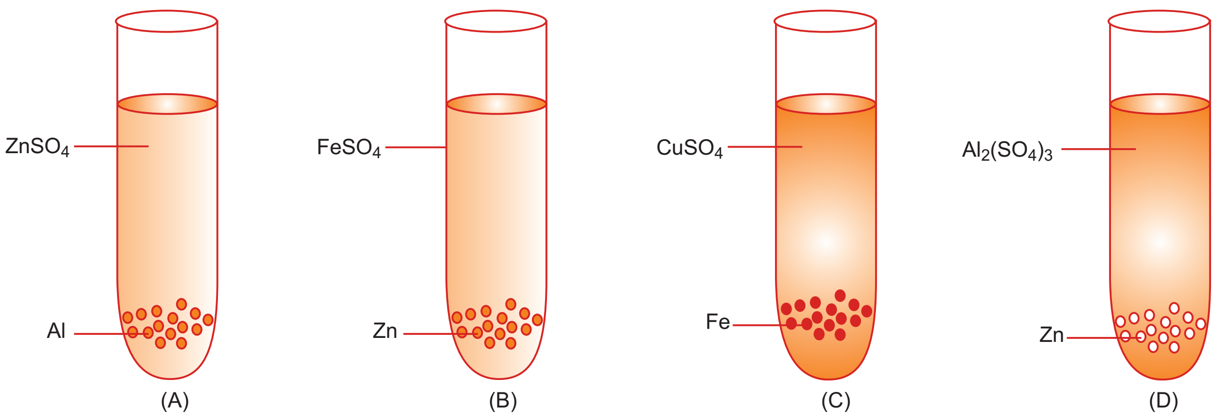
a. Atomic numbers of the elements of 3rd period of Modern Periodic table are listed below. Study the data carefully and answer the questions that follow:

Period 3 elements	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic Number	11	12	13	14	15	16	17	18

i. Which element has four valence electrons? [1]
ii. Which is the most reactive metal? [1]
iii. Which is the most reactive non-metal? [1]
iv. Which elements have valency equal to 2? Do they differ also? [1]
v. What is the valency of chlorine? [1]
vi. Name the element which forms amphoteric oxides. [1]
vii. Calculate the valency of phosphorus from the valence electrons. [1]

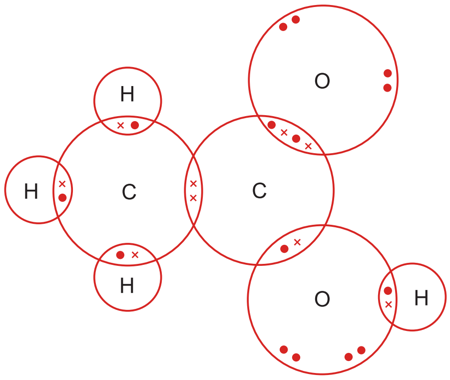
b. In the following four test-tubes, some metals are in contact with certain salt solutions.

i. After the experiment, the solution will become pale green in which test-tube? [1]
ii. What is a decomposition reaction? [2]



Question 3

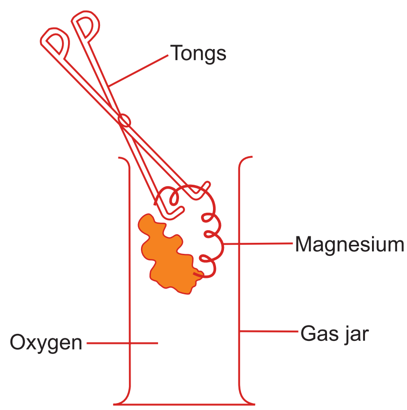
a. The diagram shows the arrangements of the outer electrons only in a molecule of ethanoic acid. Observe the diagram and answer the following questions.



i. Name the different elements present in this compound. [1]
ii. What is the total number of atoms present in this molecule? [1]
iii. Between which two atoms is there a double covalent bond? [1]
iv. How many covalent bonds does each carbon atom make? [1]
v. Would you expect, this compound to be a solid or a liquid at room temperature? Give a reason in support to your answer. [1]
vi. Ethanoic acid will dissolve in methylbenzene. Would you expect the solution to conduct electricity? Give a reason for this. [2]

b. Magnesium is a very reactive metal which will react vigorously with oxygen and even with nitrogen and carbon dioxide too.

The diagram given alongside shows the magnesium burning in oxygen. Now answer the following questions:



i. Copy out and balance the following equation for the reaction. [1]

$$.....\mathrm{Mg} + ... \longrightarrow .....\mathrm{MgO}$$

ii. Describe, what you see when the magnesium burns in oxygen. [2]

Question 4

a. Some types of chemical reactions are listed below:
decomposition; neutralisation; combustion; oxidation-reduction (redox).
Which types of reaction best describes the following changes?
i. hexane + oxygen $\longrightarrow$ carbon dioxide + water [1]
ii. calcium carbonate $\longrightarrow$ calcium oxide + carbon dioxide [1]
iii. magnesium + copper oxide $\longrightarrow$ magnesium oxide + copper [1]

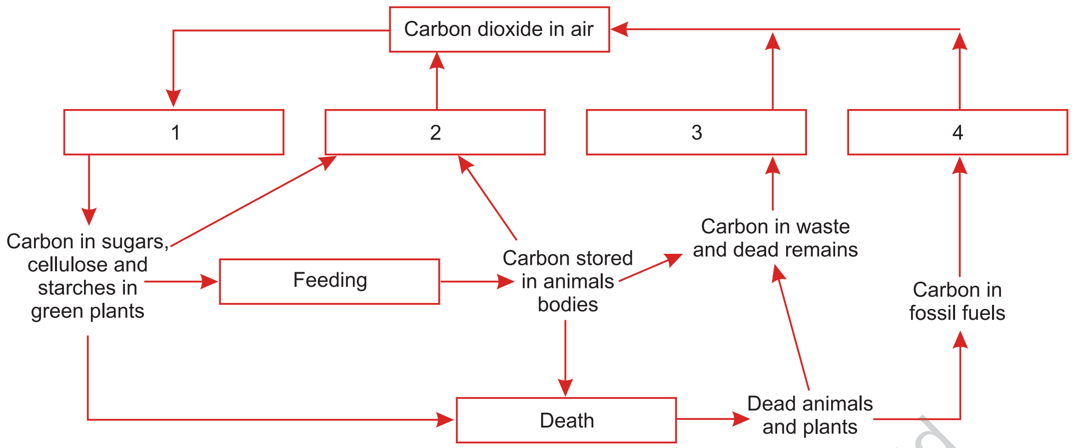
b. Copy and balance the following equations:
i. ....Cu + $O_2 \longrightarrow$ ....CuO [1]
ii. N_2 + ...$H_2 \rightleftharpoons$ ....NH_3 [1]

c. Write word and balanced chemical equations for the reactions between:
i. sodium and water, [1]
ii. zinc and copper sulphate solution. [1]

d. i. Show the formation of NaCl from sodium and chlorine atoms by the transfer of electron(s). [2]
ii. Why has sodium chloride a high melting point? [1]

Question 5

a. The diagram shows the parts of the carbon cycle. [4]



Using the best words from the following list:
combustion, decomposition, photosynthesis, respiration and transpiration
Write down the correct words in the above boxes.

b. What are fertilizers? Write two uses of fertilizers. [2]

c. Write the one difference between:
   i. Photosynthesis and Respiration [1]
   ii. Isolated system and Closed system [1]
   iii. Exothermic and Endothermic [1]

d. What is the meaning of negative (–) sign in the rate of reaction? [1]

Question 6

a. A chemical reaction occurs as a result of collision between two reacting molecules. Any factors which affects the collision rate will influence the reaction rate. Explain the factors affecting under the following factors. [1×5=5]
   i. Nature of reactants
   ii. Concentration of reactants
   iii. Effect of temperature
   iv. Effect of pressure
   v. Effect of catalyst

b. Define the following: [1×2=2]
   i. Heat of combustion
   ii. Heat of sublimation

c. Write the natural cause of global warming. [2]

d. Write the reaction for aerobic respiration. [1]

Question 7

a. Two alkenes A and B have 2 and 6 carbon atoms in their molecules, respectively. In what physical state will they occur at room temperature? [1]

b. A gas has only one carbon atom in its molecule. It burns with a blue flame and the resulting gas turns lime water milky. [1×3=3]
   i. Name the gas.
   ii. State the two properties based on above observation.
   iii. Write the chemical equation of the reactions that takes place when this gas burns in air.

c. Write the difference between alkane and alkyne. [1]

d. Write the preparation of the following: [1×2=2]
   i. Polythene   ii. Polyvinyl chloride

e. Write the short notes on the following: [1×2=2]
   i. Natural gas   ii. Petroleum gas

f. Complete the following equation: [1]

$$CH_4 + O_2 \longrightarrow \ldots\ldots\ldots\ldots\ldots\ldots + H_2O$$