# Class X Chemistry

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## Preface

Class X Chemistry is designed and written in strict accordance with the Science Curriculum Framework prepared by the Royal Education Council, Ministry of Education, Bhutan. With the implementation of various science textbooks for classes Ten and Twelve in 2017, the initiative of Royal Education Council has been to bring a major reform in the science curriculum of the country. The implementation of new science textbooks for classes IV to IX has started from academic session 2013 onward on the basis of different key stages. The science textbooks for different key stages are developed and written such that there is a spiral relationship in the flow of content from one key stage to another. Therefore, the authors hereby recommend that students acquire the scientific concept and skills of all key stages to maximize the learning in higher classes.

The salient features of this book includes:

- The intended learning objectives for each of the main topic in the chapter.
- Self-Evaluation questions as a follow up of each topic.
- Activities, both practical and theoretical, to explain abstract concepts.
- Summary for each chapter to make revision less time-consuming.
- Chapter-end exercise with all items of questions to maximize their competency in Chemistry.
- Specimen question paper.
- A glossary at the back of the book serves as a mini dictionary.
- List of references.

The incorporation of the above mentioned features will definitely make students adore and enjoy learning Chemistry. The inclusion of cartoon concepts to introduce new topics will make the subject interesting, interactive and easy to learn. It demonstrates that Chemistry too can be a fascinating subject, rather than a repelling one. With the successful nationwide implementation of transformative pedagogy in 2016, the teaching and learning process can be blended with cooperative learning structures to engage all students in learning.

Class X Chemistry is a student's most important tool to pursue their future in the science stream. Thus, this textbook would provide a greater opportunity to young learners to climb greater heights in Chemistry.

Every attempt has been made to make the textbook error – free. Any comments in the form of constructive feedbacks and suggestions from all individual users are welcome. We shall try our best to incorporate them in subsequent editions.

**- Authors**

# Contents

# Syllabus

## Strand: Material and their properties

### Chapter 1: Gas Laws

1 State Boyle's law, Avogadro's law and Charles' law.
2 Derive gas law equations.
3 Apply gas law equations to solve numerical problems.
4 Use the ideal gas equation pV= NkT and pV= nRT in numerical problems.
5 State Dalton's law.
6 Apply Dalton's law to calculate partial pressures of gases in a mixture of gases.

### Chapter 2: The mole concept and stoichiometry.

1 Explain that the quantity of one mole is set by defining one mole of carbon-12 atoms to have a mass of exactly 12 grams.
2 Define the term relative atomic mass ($A_r$), relative molecular mass ($M_r$) and relative formula mass (for ionic compounds).
3 Explain the concept of a mole as applied to electrons, atoms, molecules, ions, formulae and equations.
4 Explain Avogadro's constant as the number of particles per mole ($6.023 \times 10^{23}$ $mol^{-1}$).
5 Calculate empirical formula and molecular formula from composition by mass and percentage composition data.
6 Differentiate between empirical formula and molecular formula.
7 Calculate reacting masses from balanced chemical equations.
8 Compute reacting volumes of gases.
9 Evaluate the concentrations and volumes for reactions in solution.

### Chapter 3: Metallurgy

1 Name the ores of common metals e.g. bauxite ($Al_2O_3$) and haematite ($Fe_2O_3$).
2 Explain affect of the reactivity of a metal in determining the extraction process from its naturally occurring ores.
3 Explain that less reactive metal can be extracted by reduction with carbon or carbon monoxide (e.g. haematite).
4 Explain electrolysis.
5 Describe the purification and recycling of metal by electrolysis.
6 Describe the extraction of reactive metal by electrolysis, e.g. aluminium from its ore bauxite (aluminium oxide).
7 Outline the uses of common metals.

**Chapter 4: Halogens**

1 State the physical properties of the halogens (e.g. melting points and boiling points) and the changes in these properties as the order in group descends.

2 Describe the reactions of Group 17 elements $Cl_2$, $Br_2$ and $I_2$ with halide ions in aqueous solution ($Cl^-$, $Br^-$, $I^-$).

3 Describe the trends in reactivity of the reactions of Group 17 elements $Cl_2$, $Br_2$ and $I_2$ with halide ions in aqueous solution as order in group descends to predict the reactions of fluorine.

4 State the common uses of some of the halogens.

**Chapter 5: Transition elements**

1 Describe the similarities among transition elements and describe the characteristic properties of their compounds.

2 State some uses of transition elements.

**Chapter 6: Chemical energetics**

1 Explain that energy transfer is involved in making and breaking of chemical bonds in chemical reactions

2 Classify reactions as exothermic reaction and endothermic reaction.

**Chapter 7: Rates of reactions**

1 Explain the effect of temperature on the rate of enzyme-catalysed reactions and their dependence on pH.

2 State examples of enzymes being used in biotechnology.

3 Explain that the rate of many reactions depend on the frequency and energy of collisions between particles, using particle theory and explain the effect of temperature and concentration on the rates of chemical reactions.

**Chapter 8: Reversible reactions**

1 Explain common reactions and manufacturing processes as examples of equilibrium reactions.

2 State Le Chatelier's principle for equilibrium reactions

3 Explain (using Le Chatelier's principle) that changing concentration, pressure and temperature in an equilibrium reaction affects the position of the equilibrium.

4 Explain (using Le Chatelier's principle) that the yield of manufacturing processes depends on the reaction conditions employed, e.g. The Haber's Process.

**Chapter 9: Alcohols**

1 Explain the key functional group present in alcohols.
2 Name the first three alcohols, methanol, ethanol and propanol, in the homologous series of alcohols.
3 Describe the general properties of alcohols.
4 Describe industrial manufacture of ethanol by fermentation and by the reaction of ethene with steam.
5 Explain the principles of manufacture of alcohol in the distilleries.
6 Compare the economic and environmental advantages and disadvantages of production of alcohol.
7 State the uses of ethanol e.g. in alcoholic drinks, as a bio-fuel and as a solvent in methylated spirits.
8 Describe the social and health issues of drinking alcohol.

**Suggested practical work**

1 Carry out displacement reactions of the halogens in water solution of $Cl_2$, $Br_2$ or $I_2$ with their halides in water solution ($Cl^-$, $Br^-$, $I^-$)
2 Investigate the properties of transition metals such as Zn, Fe and Cu by using NaOH or $NH_4OH$ as testing reagents.
3 Identification of group 1 and 2 metal cations such as $Li^+$, $Na^+$, $K^+$, $Ca^{2+}$ by flame test.
4 Prepare tincture of iodine (2% w/v) as an antiseptic.
5 Investigate the reversible action of hydrated copper (II) sulphate crystals and anhydrous copper (II) sulphate.
6 Investigate the reaction of copper oxide and carbon.
7 Preparation of ethanol from glucose
8 Investigate the properties of alcohol with reference to alcohol as solute, its reaction with sodium and carboxylic acids.
9 To calculate the weight of solute dissolved in certain volume of solvent using PhET interactive simulations.
10 Construct a 3D molecular structure and shape of molecules of type Methane, Ethene, Ethyne, Water and ammonia using PhET interactive simulation.
11 Investigate the electrolysis of $CuSO_4$ and concentrated NaCl solution.

# 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 become 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 gas laws and mole concepts, metallurgy, halogen and transition elements, chemical energetics and rate of reactions, alcohols and modern chemistry and interrelationship of chemical science with other branches of science, and their attributes to people and environment .

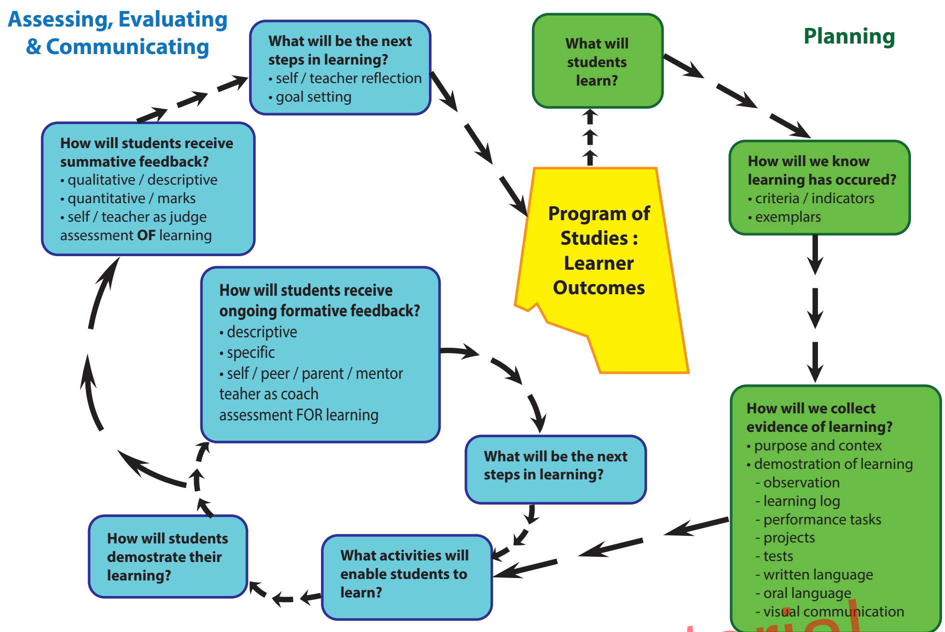
- **Working scientifically:** Basic understanding of the nature of science, and how science works. Demonstration of logical and abstract thinking and comprehension of complex situations. Explore how technological advances are related to the scientific ideas underpinning them. Compare, contrast, synthesize, 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.

**Assessing Student Learning in Classroom**



- 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:

### i. 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.

### ii. 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.

### iii. 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 & 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 & 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<br>T2= 2.5 | T1= 5<br>T2= 5 | T1= 2.5<br>T2= 2.5 | T1=30 | T2=50 |

## Assessment Techniques and Tools

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

### i. Observation Check list

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

a) Determine specific outcomes to observe and assess.

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

c) Ensure students know and understand what the criteria are.

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

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

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

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

### Sample Checklist

| Name | Topic: Gas Laws | | | | | | | | | Teacher's comments |
|---|---|---|---|---|---|---|---|---|---|---|
| | Scientific knowledge | | | Working scientifically | | | Scientific values | | | |
| | Describe Dalton's law of partial pressure of a gas and its application. | Explain Charles' law in relation to temperature and volume of gas at constant pressure. | Derive gas law equation. | Follows correct experimental procedures. | Handles equipment, apparatuses, and chemical safely. | Demonstrates ability to set up experiments. | Respects others ideas and views. | Shows curiosity to learn science. | Demonstrates concern for oneself and others. | |
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## ii. 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**

a) 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.

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

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

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

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

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

### iii. 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 science curriculum mandates students to carry out project work to help them to:

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

b) enhance deeper understanding of social and natural environment.

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

d) 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.

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

f) 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.

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

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.

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, depth of the fluid affects the fluid pressure. Develop a research question supported by a few questions to test your hypothesis. For example, how does the depth of fluid affect the pressure? Sub-questions may include, what is the fluid pressure at the same depth at different points? What is the fluid pressure as the depth increases?

### 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 pressure for the same fluid increases with the increase of depth!

### 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 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. | Problem is not new but meaningful. | Problem is stated but neither new nor meaningful. | Problem is not stated and | |

| Criteria | Scoring | | | | Total Score (28) |
|---|---|---|---|---|---|
| | 4 | 3 | 2 | 1 | |
| | Hypothesis is clearly stated in the “IF...THEN” format. | Hypothesis is clearly stated. | Hypothesis is not clearly stated. | 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. | |
| 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. | 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. | 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. | No appropriate tool used for analysis. Not enough explanation is made for acceptance and rejection of hypothesis. Conclusion is absent. | |
| | Reflection is stated clearly. | Reflection is stated. | Reflection is not clear. | Reflection is not stated. | |

| Criteria | Scoring | | | | Total Score (28) |
|---|---|---|---|---|---|
| | 4 | 3 | 2 | 1 | |
| 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 | |

## iv. 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 in which we live'.*Therefore, the purposes of doing practical in science classes are to –

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

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

c) 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.

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

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

f) 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.<br>All the procedures are followed systematically.<br>Full attention is given to relevant safety for oneself and others. | Purpose is clear purposeful.<br>All the procedures are followed but not done systematically.<br>Work is carried out with some attention to relevant safety procedures. | Purpose is inaccurate, general or extraneous.<br>A few procedures are skipped.<br>Safety procedures were frequently ignored | Purpose is vague or inaccurate.<br>Procedures are not followed<br>Safety procedures are ignored completely. | |

| Criteria | Scoring | | | | Total Score (16) |
|---|---|---|---|---|---|
| | 4 (Very good) | 3 (Good) | 2 (Fair) | 1 (Poor) | |
| Results & data representation | Representation of the data/results in tables and graphs with correct units of measurement.<br>Transformations in the results/data are evident.<br>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.<br>Transformations in some of the results/data are evident.<br>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.<br>Transformations in most of the results/ data are not evident.<br>Graphs and tables are scaled correctly, but without appropriate titles and labels. | Representation of the data/results in tables and graphs are not relevant.<br>Transformations in the results/data are not evident.<br>Some attempts are evident to produce graphs from the data/results. | |
| Analysis & discussion | All the tools used for analysis are appropriate.<br>A comprehensive discussion, containing a comparative analysis is evident.<br>The experimental findings are significant to the purpose of the experiment. | Most of the tools used for analysis are appropriate.<br>A comprehensive discussion, containing some comparative analysis is evident.<br>The experimental findings do not have strong significance to the purpose of the experiment. | Only a few tools are used for analysis.<br>A comprehensive discussion, containing a few comparative analysis is evident.<br>The experimental findings have weak significance to the purpose of the experiment. | No appropriate tools are used for analysis.<br>Comprehensive discussion is absent.<br>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.<br>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.<br>Limitations of experiment are identified. | Conclusions are not drawn from the findings and have no significance to objectives of the experiment.<br>Some limitations of experiment are identified. | No valid conclusions drawn from the findings.<br>Limitations of experiment are not identified. | |
| | | | | TOTAL SCORE | |

### Chapter-wise Weighting and Time allocation

| Chapters | Chapter title | Maximum time required (mins) | Weighting (%) |
|---|---|---|---|
| Chapter 1 | Gas Laws | 432 | 10% |
| Chapter 2 | The Mole concept and Stoichiometry | 691 | 16% |
| Chapter 3 | Metallurgy | 648 | 15% |
| Chapter 4 | Halogens | 389 | 9% |
| Chapter 5 | Transition Elements | 389 | 9% |
| Chapter 6 | Chemical Energetics | 389 | 9% |
| Chapter 7 | Reversible Reaction | 389 | 9% |
| Chapter 8 | Rates of Reaction | 389 | 9% |
| Chapter 9 | Alcohol | 605 | 14% |
| Total | | 4320 | 100% |

The total time required to complete the topics is 4320 minutes or 96 periods of 45 minutes in a period.

# Modern Periodic Table

| Atomic Number |
|---|
| Symbol |
| Name |
| Atomic Mass |

| 1 IA 1A | 2 IIA 2A | 3 IIIB 3B | 4 IVB 4B | 5 VB 5B | 6 VIB 6B | 7 VIIB 7B | 8 | 9 VIII 8 | 10 | 11 IB 1B | 12 IIB 2B | 13 IIIA 3A | 14 IVA 4A | 15 VA 5A | 16 VIA 6A | 17 VIIA 7A | 18 VIIIA 8A |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 H Hydrogen 1.008 | | | | | | | | | | | | | | | | | 2 He Helium 4.003 |
| 3 Li Lithium 6.941 | 4 Be Beryllium 9.012 | | | | | | | | | | | 5 B Boron 10.811 | 6 C Carbon 12.011 | 7 N Nitrogen 14.007 | 8 O Oxygen 15.999 | 9 F Fluorine 18.998 | 10 Ne Neon 20.180 |
| 11 Na Sodium 22.990 | 12 Mg Magnesium 24.305 | | | | | | | | | | | 13 Al Aluminum 26.982 | 14 Si Silicon 28.086 | 15 P Phosphorus 30.974 | 16 S Sulfur 32.066 | 17 Cl Chlorine 35.453 | 18 Ar Argon 39.948 |
| 19 K Potassium 39.098 | 20 Ca Calcium 40.078 | 21 Sc Scandium 44.956 | 22 Ti Titanium 47.867 | 23 V Vanadium 50.942 | 24 Cr Chromium 51.996 | 25 Mn Manganese 54.938 | 26 Fe Iron 55.845 | 27 Co Cobalt 58.933 | 28 Ni Nickel 58.693 | 29 Cu Copper 63.546 | 30 Zn Zinc 65.38 | 31 Ga Gallium 69.723 | 32 Ge Germanium 72.631 | 33 As Arsenic 74.922 | 34 Se Selenium 78.972 | 35 Br Bromine 79.904 | 36 Kr Krypton 84.798 |
| 37 Rb Rubidium 85.468 | 38 Sr Strontium 87.62 | 39 Y Yttrium 88.906 | 40 Zr Zirconium 91.224 | 41 Nb Niobium 92.906 | 42 Mo Molybdenum 95.95 | 43 Tc Technetium 98.907 | 44 Ru Ruthenium 101.07 | 45 Rh Rhodium 102.906 | 46 Pd Palladium 106.42 | 47 Ag Silver 107.868 | 48 Cd Cadmium 112.411 | 49 In Indium 114.818 | 50 Sn Tin 118.711 | 51 Sb Antimony 121.760 | 52 Te Tellurium 127.6 | 53 I Iodine 126.904 | 54 Xe Xenon 131.294 |
| 55 Cs Cesium 132.905 | 56 Ba Barium 137.328 | 57-71 | 72 Hf Hafnium 178.49 | 73 Ta Tantalum 180.948 | 74 W Tungsten 183.84 | 75 Re Rhenium 186.207 | 76 Os Osmium 190.23 | 77 Ir Iridium 192.217 | 78 Pt Platinum 195.085 | 79 Au Gold 196.967 | 80 Hg Mercury 200.592 | 81 Tl Thallium 204.383 | 82 Pb Lead 207.2 | 83 Bi Bismuth 208.980 | 84 Po Polonium [208.982] | 85 At Astatine 209.987 | 86 Rn Radon 222.018 |
| 87 Fr Francium 223.020 | 88 Ra Radium 226.025 | 89-103 | 104 Rf Rutherfordium [261] | 105 Db Dubnium [262] | 106 Sg Seaborgium [266] | 107 Bh Bohrium [264] | 108 Hs Hassium [269] | 109 Mt Meitnerium [268] | 110 Ds Darmstadtium [269] | 111 Rg Roentgenium [272] | 112 Cn Copernicium [277] | 113 Uut Ununtrium unknown | 114 Fl Flerovium [289] | 115 Uup Ununpentium unknown | 116 Lv Livermorium [298] | 117 Uus Ununseptium unknown | 118 Uuo Ununoctium unknown |

| | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| Lanthanide Series | 57 La Lanthanum 138.905 | 58 Ce Cerium 140.116 | 59 Pr Praseodymium 140.908 | 60 Nd Neodymium 144.242 | 61 Pm Promethium 144.913 | 62 Sm Samarium 150.36 | 63 Eu Europium 151.964 | 64 Gd Gadolinium 157.25 | 65 Tb Terbium 158.925 | 66 Dy Dysprosium 162.500 | 67 Ho Holmium 164.930 | 68 Er Erbium 167.259 | 69 Tm Thulium 168.934 | 70 Yb Ytterbium 173.055 | 71 Lu Lutetium 174.967 |
| Actinide Series | 89 Ac Actinium 227.028 | 90 Th Thorium 232.038 | 91 Pa Protactinium 231.036 | 92 U Uranium 238.029 | 93 Np Neptunium 237.048 | 94 Pu Plutonium 244.064 | 95 Am Americium 243.061 | 96 Cm Curium 247.070 | 97 Bk Berkelium 247.070 | 98 Cf Californium 251.080 | 99 Es Einsteinium [254] | 100 Fm Fermium 257.095 | 101 Md Mendelevium 258.1 | 102 No Nobelium 259.101 | 103 Lr Lawrencium [262] |

Alkali Metal · Alkaline Earth · Transition Metal · Basic Metal · Semimetal · Nonmetal · Halogen · Noble Gas · Lanthanide · Actinide

# 1 Gas Laws

## 1.1 Introduction

Many substances exist in gaseous state in our surrounding. The molecules in a gas are in constant random motion. Under similar conditions of temperature and pressure, gases exhibit certain physical properties different to matters in solid or liquid. The intermolecular distance among the gas molecules are comparatively large. Hence, the inter-molecular forces among the molecules are weak. The change in physical conditions like temperature or pressure results in change in the physical properties of gases, which is explained by gas laws.

## 1.2 Gas laws

**Learning objectives**

On completion of this topic, students should be able to:

» *state Boyle's law, Charles' law and Dalton's law of partial pressures.*

» *solve numerical problems based on the gas laws.*

» *verify Boyle's law and Charles' law.*

The physical properties of gases are described by three standard variables such as temperature (T), pressure (P) and volume (V).

**Temperature** is measured in degree Celsius and in solving numerical problems; it is converted into absolute temperature or Kelvin by adding 273 to degree Celsius.

$$0^{\circ}C = 273K$$

**Pressure** is measured in atmosphere or in height of column of mercury (cm Hg). The relationship between the different units of pressure is given by

$$1 \text{ atm} = 760 \text{ mm Hg} = 76 \text{ cm Hg} = 101.325 \text{ kilo pascal (kPa)}$$

The SI unit of pressure is pascal (Pa).

**Volume** is measured in litres (L), Millilitre (mL), cubic meter ($m^3$), cubic centimetre ($cm^3$) and cubic decimetre ($dm^3$).

$$1 \text{ litre} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$$

The SI unit of volume is $m^3$.

### 1.2.1 Boyle's law: Pressure-Volume relationship

Robert Boyle, an Irish scientist in 1660 performed a series of experiments in which he examined the effect of pressure on the volume of a given amount of a gas at a constant temperature. He found out that, when the pressure of a fixed quantity of a gas was doubled at constant temperature, the volume decreased to one-half and when the pressure was decreased to one-fourth, the volume increased by four times. Such behaviour of the gas was generalised in the form of Boyle's law which states that at a constant temperature the volume of a sample of gas is inversely proportional to its pressure. Mathematically,

$$P \propto \frac{1}{V} \Rightarrow PV = k \quad \text{(at constant temperature)} \qquad (1.1)$$

Where; P = pressure, V = volume, k = proportionality constant

The value of the constant of proportionality (k) depends upon the quantity of the gas at any constant temperature. Thus, if $V_1$ is the volume occupied by a given quantity of the gas at pressure $P_1$ and $V_2$ is its volume when the pressure changes to $P_2$ as shown in the Figure 1.1, then according to Boyle's law at constant temperature

$$P_1V_1 = k \qquad (1.2)$$

$$P_2V_2 = k \qquad (1.3)$$

$$P_1V_1 = P_2V_2 \qquad (1.4)$$

Thus, the equation (1.4) is called Boyle's law equation.

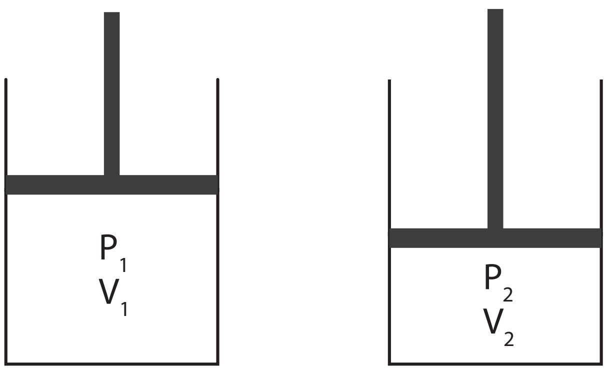


Figure 1.1 Illustration of Boyle's law

This means that if nothing else changes, the volume of a given mass of gas are inversely proportional to pressure. It is a linear relationship. If pressure on a gas doubles, its volume will decrease by half. Thus, Boyle's law states that the volume of a sample of gas is inversely proportional to its pressure, temperature remaining constant.

### Activity 1.1 Investigating Boyle's law .

#### Materials required

Syringe (60 mL) with cap and marshmallows.

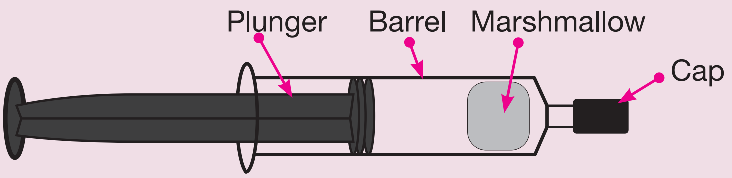


Figure 1.2 Syringe with marshmallows

#### Procedure I

1. Remove the cap from the syringe.
2. Hold the syringe in one hand and with the other pull the plunger out.
3. Carefully put two small marshmallows into the syringe. Put the plunger back and keep the air space as large as possible.
4. Seal the opening of the syringe by placing the cap so that no air can escape the syringe.
5. Slowly push the plunger in, and carefully observe the marshmallows. Repeat it two or three times.

#### Questions

1. What did you observe when the plunger is pushed in?
2. What happens to the volume of air inside the syringe?
3. Did it become more or less difficult to push the plunger in as the volume of the air in the syringe decreased?
4. Describe how the pressure changes as the volume of the air decreases using the marshmallow's example.
5. Compare the size of the marshmallows before and after pushing the plunger.
6. What can you conclude from the experiment?
7. Plot a graph of the volume against the pressure at constant temperature. Explain this inverse relationship between the pressure and volume of a gas

based on kinetic theory of gases.

**Procedure II**

1. Remove the cap from the syringe.
2. Compress the plunger in so that it just touches the marshmallows. Ensure that the plunger is not squeezing the marshmallows.
3. Place the cap back so that no air can escape the syringe.
4. Pull out the plunger and increase the volume. Carefully observe the marshmallows.

**Questions**

1. What did you observe when the plunger is pulled out?
2. What happens to the volume of air inside the syringe?
3. Describe how pressure changes as the volume of air increases using the marshmallows example.
4. What can you conclude from the experiment?

**Solved problems**

1. A sample of helium occupies a volume of 160 $cm^3$ at 10 atm and 25°C. What volume will it occupy if the pressure is decreased to 8 atm at constant temperature?

**Solution:**

$V_1 = 160\ cm^3$ $\quad$ $P_1 = 10$ atm

$V_2 = ?$ $\quad$ $P_2 = 8$ atm

Using Boyle's law equation,

$$P_1V_1 = P_2V_2$$

Rearranging to make $V_2$ the subject of the formula,

$$V_2 = \frac{P_1V_1}{P_2}$$

Substituting the values

$$V_2 = \frac{10\ atm \times 160\ cm^3}{8\ atm} = 200\ cm^3$$

Thus, the final volume occupied by helium is 200 $cm^3$

2. A balloon contains 7.2 L of helium. When the pressure is reduced to 1620 mm Hg, the balloon expands to occupy a volume of 25.1 L. What was the initial pressure exerted on the balloon?

**Solution:**

$V_1$ = 7.2 L    $P_1$ = ?

$V_2$ = 25.1L    $P_2$ = 1620 mm Hg

Using Boyle's law,

$P_1V_1 = P_2V_2$

Rearranging to make $P_1$ the subject of the formula:

$$P_1 = \frac{P_2V_2}{V_1}$$

$$P_1 = \frac{1620\,\text{mm of Hg} \times 25.1\text{L}}{7.2\text{L}}$$

$$= 5647.5\,\text{mm Hg}$$

Therefore, the initial pressure exerted on the balloon is 5647.5 mm Hg.

3. What will be the pressure required to reduce 600 mL of a dry gas at 750 mm pressure to 500 mL at the same temperature?

**Solution:**

$V_1$ = 600 mL    $V_2$ = 500 mL

$P_1$ = 750 mm    $P_2$ = ?

Using Boyle's law equation,

$P_1V_1 = P_2V_2$

Rearranging to make $P_2$ the subject of the formula:

$$P_2 = \frac{P_1V_1}{V_2}$$

$$P_2 = \frac{750\,\text{mm} \times 600\,\text{mL}}{500\,\text{mL}} = 900\,\text{mm Hg}$$

Therefore, the pressure required is 900 mm Hg.

4. Sulphur dioxide ($SO_2$), a gas that plays a central role in the formation of acid rain, is found in the exhaust of automobiles and power plants. Consider a 1.53 L sample of gaseous $SO_2$ at a pressure of 5.6 x $10^3$ Pa. If the pressure is changed to 1.5 x $10^4$ Pa at a constant temperature, what will be the new volume of the gas?

**Solution:**

PV = k,

Which is also written as;

$P_1V_1 = k = P_2V_2$ or $P_1V_1 = P_2V_2$

The given data are

$P_1 = 5.6 \times 10^3$ Pa $\quad$ $P_2 = 1.5 \times 10^4$ Pa

$V_1 = 1.53$ L $\quad$ $V_2 = ?$

The preceding for equation for $V_2$,

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(5.6 \times 10^3\,\text{Pa}) \times 1.53\,\text{L}}{1.5 \times 10^4\,\text{Pa}} = 0.57\,\text{L}$$

Therefore, the new volume of the gas is 0.57 L.

**Practice problems**

1. A cylinder containing carbon dioxide of volume 20 L at 2.0 atm was connected to another cylinder of certain volume at constant temperature. The final pressure of the gas in the cylinders was found to be 380 mm Hg. Calculate the volume of the second cylinder.
2. A gas of certain mass occupies a volume of 650 $cm^3$ under a pressure of 750 mm Hg. Calculate the pressure under which the volume of the gas will be reduced by 10 per cent of its original volume.
3. A gas tank holds 2785 L of propane, $C_3H_8$ at 830 mm Hg. What is the volume of the propane at standard pressure?
4. A balloon with a volume of 2.0 L is filled with a gas at 3 atm. If the pressure is reduced to 0.5 atm without a change in temperature, what would be the volume of a balloon?
5. 352 mL of chlorine under a pressure of 680 mm Hg are placed into a container under a pressure of 1210 mm Hg. The temperature remains constant at 296 K. What is the volume of a container?
6. The pressure on 40 mL of a gas is increased from 760 mm to 800 mm. Find the new volumes at the same temperature.

## 1.2.2 Charles' law: Volume-Temperature relationship

The air expands on heating, thereby decreasing its density. For this reason, balloons rise when inflated with warm air. The effect of temperature on the volumes of different gases at constant pressure was thoroughly studied by the French Physicist, J.A. Charles' who in 1787 observed that the expansion of the volumes of different gases was the same for equal rise in temperature. His work was extended by J.L. Gay-Lussac in 1802 who found that, a fixed mass of any gas expand or contract by $\frac{1}{273}$ of its volume at 0°C for every degree centigrade rise or fall in temperature respectively under constant pressure. Such behaviour of gases was generalised in a quantitative way by a law which is known as Charles' law. The law states that at a constant pressure, the volume of a given quantity of a gas increases or decreases by $\frac{1}{273}$ of its volume at 0°C for rise or fall in the temperature by 1°C.

If $V_0$ and V are the volumes of a given quantity of a gas at 0°C and t°C respectively at constant pressure, then from Charles' law

$$V = V_0 + \frac{1}{273} \times V_0 \times t$$

$$V = V_0\left(1 + \frac{t}{273}\right) \qquad (1.5)$$

$$V = V_0\left(\frac{273 + t}{273}\right) \qquad (1.6)$$

We may define now a new temperature scale such that any temperature 't' on this scale will be given by T = 273 + t.

$$V_t = V_0 \frac{T}{273}$$

For a given mass of a gas, $\frac{V_0}{273}$ is constant (k)

So, $V = kT$ (1.7)

This new temperature scale is known as the Absolute or Kelvin scale of temperature and is of fundamental importance in all sciences. In terms of this scale, equation (1.7) predicts that the volume of a definite quantity of a gas at constant pressure is directly proportional to the absolute temperature. Mathematically,

$$V \propto T$$

Or $$\frac{V}{T} = \text{constant} \qquad (1.8)$$

or in general, $$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad (1.9)$$

Equation (1.9) is known as Charles' law equation.

## Activity 1.2 Investigating Charles' law.

### Materials required

Beaker, ruled scale, thermometer, capillary tube and cooking oil.

### Procedure

1. Mount the capillary tube and the thermometer on a ruled scale as shown in the Figure 1.3.
2. Fill the beaker with oil and warm it.
3. Then carefully lower the ruled scale mounted with the capillary tube and thermometer into the beaker.

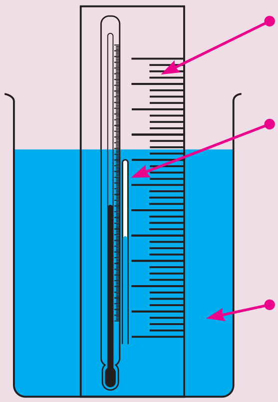


Figure 1.3

4. Note down the temperature of oil and the length of air space in the capillary tube.
5. Repeat step no. 4 at different intervals of time.
6. Fill in the data obtained in Table 1.1

Table 1.1

| Sl. No | Time interval | Temperature of oil | Length of air space in the capillary tube |
|---|---|---|---|
| 1 | | | |
| 2. | | | |

### Questions

1. What happens to the volume of air space with the decrease in temperature?
2. Why is it important to maintain the bottom of the air space in the capillary tube at the same depth below the surface of the oil bath?

3. Plot a graph for temperature versus air space using the data obtained in Table 1.1.


Be careful while handling the hot oil. You may need to wear gloves and laboratory spectacles for protection.

**Solved problems**

1. The temperature inside the refrigerator is about 4°C. A balloon is placed inside the refrigerator that initially has a temperature of 22°C and a volume of 0.5 litres. What will be the volume of the balloon when it is fully cooled by the refrigerator at constant pressure?

**Solution:**

$V_1 = 0.5\text{ L} = 500\text{ cm}^3$; $T_1 = 22 + 273 = 295\text{K}$

$V_2 = ?$ $T_2 = 4 + 273 = 277\text{K}$

Using Charles' law equation,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{500\,\text{cm}^3}{295\,\text{K}} = \frac{V_2}{277\,\text{K}}$$

$$V_2 = \frac{500\,\text{cm}^3 \times 277\,\text{K}}{295\,\text{K}} = 469.49\,\text{cm}^3 \text{ or } 0.46949\,\text{L}$$

The volume of the balloon when it is fully cooled by the refrigerator is $469.49\text{cm}^3$ or 0.46949 L.

2. To what temperature should 2.3 L of a gas at 25°C, be heated in order to expand its volume to 4.0 L under constant pressure?

**Solution:**

$T_1 = 25°\text{C} + 273 = 298\text{ K}$ $T_2 = ?$

$V_1 = 2.3\text{ L}$ $V_2 = 4.0\text{ L}$

Using the Charles' law equation,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{2.3}{298} = \frac{4.0}{T_2}$$

$$T_2 = \frac{4.0 \times 298}{2.3} = 518.26\,\text{K or } 245.26^\circ\text{C}$$

The gas should be heated to a temperature of 518.26 K or 245.26°C.

3. The volume of a given mass of a gas is 720 mL at 15°C. Assuming constant pressure, at what temperature will its volume be 960 mL?

**Solution:**

$V_1$ = 720 mL; $V_2$ = 960 mL

$T_1$ = 273 + 15 = 288 K $T_2$= ?

According to Charles' law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$T_2 = \frac{V_2 T_1}{V_1} = \frac{960 \times 288}{720} = 384\,\text{K or } 111^\circ\text{C}$$

The temperature necessary is 111°C.

**Practice problems**

1. Under what temperature will the volume of the gas at 0°C double itself if the pressure is kept constant?
2. A sample of a gas occupies 3000 $cm^3$ at 1°C. What volume will it occupy at -10°C at a constant pressure.
3. A sample of helium has a volume of 521 $dm^3$ at a pressure of 75 cm Hg and a temperature of 18°C. When the temperature is increased to 23°C, what is the volume of the helium?
4. A sample of oxygen occupies a volume of 1.6 L at 91°C. What will be the temperature when the volume of oxygen is reduced to 1.2 L?
5. A container contains 5 L of nitrogen gas at 25°C. What will be its volume if the temperature increases by 35°C keeping the pressure constant?
6. A sample of gas at 15°C and 1 atm has a volume of 2.58 L. What volume will this gas occupy at 38°C and 1 atm?

## 1.2.3 Avogadro's law

Avogadro in 1811 suggested the participation of molecules in the constitution of matter. He helped answer some of the basic questions about why substances react only in certain proportions. In particular, he helped solve the mystery of the reaction of gases with each other in small whole-number volume. He put forward a hypothesis, relating number of molecules of gases and their volumes under identical conditions of temperature and pressure. The hypothesis is generally known as Avogadro's law. It states that "equal volumes of all the gases under similar conditions of temperature and pressure contain equal number of molecules." He also proposed the relationship between the amount of gas expressed as the number of moles and volume occupied by it under the similar conditions of temperature and pressure. He found out that under the similar conditions of temperature and pressure, the volume (V) of the gas is directly proportional to its number of moles (n) expressed in gram moles.

$$V \propto n$$

$$V = kn \text{ (at constant pressure and temperature)} \qquad (1.10)$$

One mole of any gaseous substance at STP contains 6.023 x $10^{23}$ molecules. This number is known as the Avogadro's number or Avogadro's constant. It is represented by $N_A$. Avogadro's number may be expressed in terms of number of atoms, ions, molecules or electrons.

One mole of an ideal gas occupies 22.4 litres at STP (Standard conditions for temperature and pressure). This is often referred to as the molar volume of an ideal gas. Real gases may deviate from this value.

### 1.2.4 Gas equation (combining Boyle's and Charles' law)

If both the temperature and the pressure of a given mass of a gas are varied, the relationship between pressure, volume and the temperature is given by the combination of Boyle's law and Charles' law.

According to Boyle's law,

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

According to Charles' law,

$$V \propto T$$

Combining both the laws:

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

$$\text{Or } V = k\frac{1}{P} \times T$$

Thus, $\frac{PV}{T} = k$ If the volume of the gas changes from V to $V_1$, pressure from P to $P_1$ and temperature from T to $T_1$ then,

$$\frac{PV}{T} = \frac{P_1V_1}{T_1} = k \qquad (1.11)$$

This equation (1.11) is known as gas equation.

**Solved problems**

1. The given mass of a gas occupies a volume of 450 $cm^3$ at 14°C and 0.9 atm. What will be its volume at 28°C and 1.8 atm?

   **Solution :**

   P = 0.9 atm

   V = 450 $cm^3$

   T = 14 + 273 = 287 K

$P_1$ = 1.8 atm

$T_1$= 28 + 273 = 301 K

$V_1$= ?

Using gas equation,

$$\frac{PV}{T} = \frac{P_1V_1}{T_1}$$

$$V_1 = \frac{PVT_1}{TP_1} = \frac{0.9 \times 450 \times 301}{287 \times 1.8} = 235.97\,\text{cm}^3$$

The volume of the gas is 235.97 $cm^3$

2. 2.00 L of a gas is collected at 25.0°C and 745.0 mmHg. What is the volume at STP?

**Solution:**

$P_1$ = 745 mm Hg

$T_1$ = 25 + 273 = 298K

$V_1$ = 2 L

$P_2$ = 760 mm Hg

$V_2$ = ?

$T_2$ = 273K

Using gas equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{P_2T_1} = \frac{745 \times 2 \times 273}{760 \times 298} = 1.796\,\text{L}$$

The volume at STP is 1.796 L

**Practice problems**

1. A gas of certain mass occupies volume of 1.2 L at 37°C and 3.0 atm. At what temperature will the volume and pressure of this gas become one-third of their initial values?
2. A balloon filled with 0.5 L of certain gas at 23°C and 0.46 atm was placed inside a refrigerator. On opening the refrigerators after few hours, the volume and the pressure of a gas was found to have changed to 0.3 L and 0.7 atm respectively. Determine the temperature inside the refrigerator.

## 1.2.5 Ideal gas equation

The expressions of Boyle's law, Charles' law and Avogadro's law can be combined mathematically to give a general relation between pressure, volume, temperature and the number of moles of a gas.

According to Boyle's law:

$$V \propto \frac{1}{P} \text{ ( at constant T)}$$

According to Charle's law:

$$V \propto T \text{ ( at constant P)}$$

According to Avogadro's law:

$$V \propto n$$

On combining the three laws:

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

$V = R\frac{nT}{P}$ Where, R = molar gas constant = 0.0821 litre atm $K^{-1}$ $mol^{-1}$

$$PV = nRT \qquad (1.12)$$

$$\text{Or, } PV = \frac{w}{m}RT \qquad (1.13)$$

Where, w = mass of the gas and m = molar mass of the gas.

This equation (1.13) is known as the ideal gas equation.

**Solved problems**

1. At what temperature will 0.005 mol of a gas occupy 600 mL at a pressure of 750 mm Hg? (R = 0.0821 L atm / K mol)

**Solution:**

$P = 750 \text{ mm of Hg} = \frac{750}{760} \text{ atm} = 0.987 \text{ atm}$

V = 600 mL = 0.6 L ( since value of 'R' is in litres)

n = 0.005 mole

T = ?

Using ideal gas equation:

PV = nRT

$$T = \frac{0.987\,\text{atm} \times 0.6\,\text{L}}{0.005\,\text{mol} \times 0.0821\,\text{L atm/K mol}} = 1442.4\,\text{K}$$

The temperature of the gas should be 1442.4K.

2. A cylinder contains 5.0 g of neon at 256 mm Hg and at a temperature of 35°C. Calculate the volume of the gas?

**Solution:**

$P = 256 \text{ mmHg} = \frac{256}{760} \text{ atm} = 0.3368 \text{ atm}$

$V = ?$

$n = \frac{w}{m} = \frac{5.0}{20} = 0.25 \text{ mole}$

$R = 0.0821 \text{ L atm / K mol}$

$T = 35°C + 273 = 308 \text{ K}$

Using ideal gas equation,

$PV = nRT$

$$V = \frac{0.25 \times 0.0821 \times 308}{0.3368} = 18.76 \text{ L}$$

The volume of the gas is 18.76 L.

3. Suppose we have a sample of ammonia gas with a volume of 7.0 mL at a pressure of 1.68 atm. The gas is compressed to a volume of 2.7 mL at a constant temperature. Use the ideal gas law to calculate the final pressure.

**Solution:**

$V_1 = 7.0 \text{ mL};$ $V_2 = 2.7 \text{ mL}$

$P_1 = 1.68 \text{ atm},$ $P_2 = ?$

According to ideal gas equation, $PV = nRT$

Since 'n' and 'T' remain constant in this case, we can write $P_1V_1 = nRT$ and $P_2V_2 = nRT$.

Combining these two equations give

$P_1V_1 = nRT = P_2V_2$ or $P_1V_1 = P_2V_2$

On solving for final pressure, we have;

$$P_2 = \frac{P_1V_1}{V_2} = \frac{1.68 \times 7.0}{2.7} = 4.4 \text{ atm}$$

The final pressure is 4.4 atm.

4. A sample containing 0.35 mol of argon gas at a temperature of 13°C and a pressure of 568 torr is heated to 56°C and a pressure of 897 torr. Calculate the change in volume.

**Solution:**

Use ideal gas law to find the volume for each set of conditions.

State 1

$n_1 = 0.35$ mol

$P_1 = 568 \text{ torr} \times \frac{1\text{atm}}{760\text{ torr}} = 0.747$ atm

$T_1 = 13°C + 273 = 286$ K

State 2

$n_2 = 0.35$ mol

$P_2 = 897 \text{ torr} \times \frac{1\text{atm}}{760\text{ torr}} = 1.18$ atm

$T_2 = 56°C + 273 = 329$ K

Calculating Volume for state 1

$$V_1 = \frac{n_1RT_1}{P_1} = \frac{0.35 \times 0.0821 \times 286}{0.747} = 11\text{L}$$

Calculating volume for state 2

$$V_2 = \frac{n_2RT_2}{P_2} = \frac{0.35 \times 0.0821 \times 329}{1.18} = 8.01\text{L}$$

On going from state 1 to state 2, the volume changes from 11 L to 8.0 L. Thus, the change in volume, ΔV is given as

$\Delta V = V_2 - V_1 = 8.0 \text{ L} - 11 \text{ L} = -3 \text{ L}$

*Note: The change in volume is -3 L. The change in volume is negative as there is decrease in volume.*

5. Calculate the volume occupied by 2.34 grams of carbon dioxide gas at STP.

**Solution:**

As per the question, we have;

P = 1 atm (at STP)

$$n = \frac{\text{Weight in gram}}{\text{gram molecular mass}} = \frac{2.34\text{ g}}{44.0\text{ g mol}^{-1}} = 0.05\text{ mol}$$

$R = 0.0821 \text{L atm mol}^{-1}\text{K}^{-1}$

T = 273 K (at STP)

Using ideal gas equation, PV = nRT

$$V = \frac{nRT}{P} = \frac{0.05 \times 0.0821 \times 273}{1} = 1.12\text{L}$$

The volume occupied by the gas is 1.12 L.

**Practice problems**

1. Calculate the volume occupied by 4.2 g of nitrogen at STP.
2. A certain gas of mass 2.5 g at 25°C and 0.65 atm occupies a volume of 23.52 L. Determine the molecular mass of the gas.
3. A volume of 26.5 mL of nitrogen gas was collected in a tube at a temperature of 17°C and a pressure of 737 mm Hg. The next day the volume of the nitrogen was 27.1 mL with the barometer still reading 737 mm Hg. What was the temperature on the second day?
4. Calculate the mass of 15.0 L of $NH_3$ at 27° C and 900 mm Hg.
5. Determine the number of moles and the mass of the sample of argon occupying 37.8 L at STP.
6. A sample of hydrogen gas ($H_2$) has a volume of 8.56 L at a temperature of 0°C and a pressure of 1.5 atm. Calculate the moles of $H_2$ molecules present in this gas sample.
7. A sample of methane gas that has a volume of 3.8 L at 5°C is heated to 86°C at constant pressure. Calculate its new volume.
8. At what temperature will 0.654 moles of neon gas occupy 12.30 L at 1.95 atm?

## 1.2.6 Dalton's law of partial pressures

The English chemist, John Dalton investigated pressure exerted by the mixture of non- reacting gases and formulated the law which states that 'the total pressure exerted by a mixture of non-reacting gases in a vessel of known capacity at a constant temperature is equal to sum of the partial pressures of the constituent gases'.

$P_t = P_1 + P_2 + P_3 ...... + P_n$ (Where $P_t$ is the total pressure of the mixture and $P_1$, $P_2$, $P_3$... $P_n$ are the partial pressures of each component).

Partial pressure is the pressure that would be exerted by a gas if it alone occupied the same volume as the mixture at the same temperature.

### i. Partial Pressure

A container of fixed volume at constant temperature holds a mixture of **gas a** and **gas b** at a total pressure of 4 atm. The total pressure in the container is proportional to the number of gas particles.

More gas particles = greater pressure

Less gas particles = lower pressure

If each dot in Figure 1.4 represents 1 mole of gas particles, then there are 48 moles of gas particles in this container exerting a total pressure of 4 atm.

Imagine the container with no particles of **gas b**. Only particles of **gas a** are present in the same container at the same temperature. Now the container holds only 12 moles of gas particles instead of the 48 moles of gas particles it originally contained. Since pressure is proportional to the number of gas particles, the pressure exerted by **gas a** = 12 mol ÷ 48 mol x 4 atm = 1 atm

Imagine the container with no particles of **gas a**. Only particles of **gas b** are present in the same container at the same temperature. Now, the container holds only 36 moles of gas particles instead of the 48 moles of gas particles it originally contained. Since pressure is proportional to the number of gas particles, the pressure exerted by **gas b** = 36 mol ÷ 48 mol x 4 atm = 3 atm.

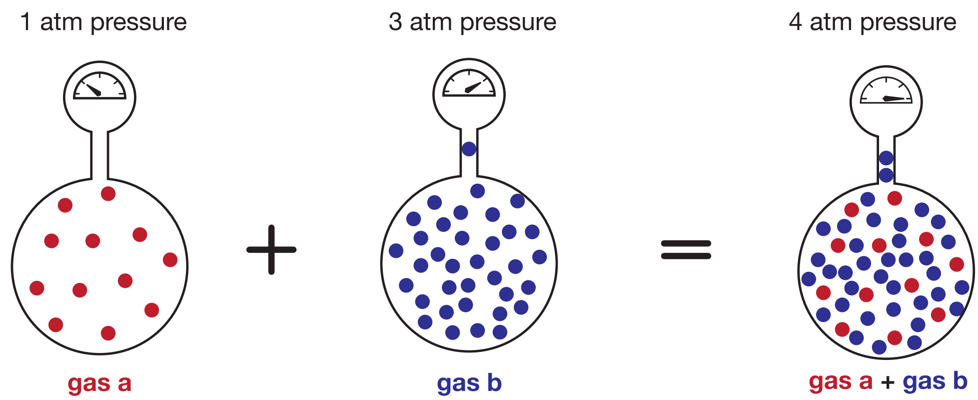


Figure 1.4 Total pressures in gas mixture

### ii. Total pressures

The total pressure in a gas mixture is the sum of the partial pressures of each individual gas.

$P_{total} = P_{gas\ a} + P_{gas\ b}$ (1.14)

If $P_1$, $P_2$, $P_3$...., $P_n$ are the partial pressures of the individual gases in a mixture then according to the Dalton's law of partial pressure the total pressure, P is given by

$P = P_1 + P_2 + P_3 ...$ (1.15)

It is assumed that the gases do not chemically interact under the conditions present in the experiment as shown in the illustration Figure 1.5. For example, suppose you have 1 litre of oxygen at a pressure of 159 mm of Hg and 1 litre of nitrogen at 593 mm of Hg. You now transfer one of the gases into the container occupied by the other. You will find that the total pressure is now 752 mm of Hg. Each gas is occupying the same volume of 1 litre, although they are mixed.

Each gas is therefore exerting its original pressure of 159 and 593 mm of Hg respectively. Within the single volume of 1 litre, the two pressures combine to produce a total of 752 mm of Hg.

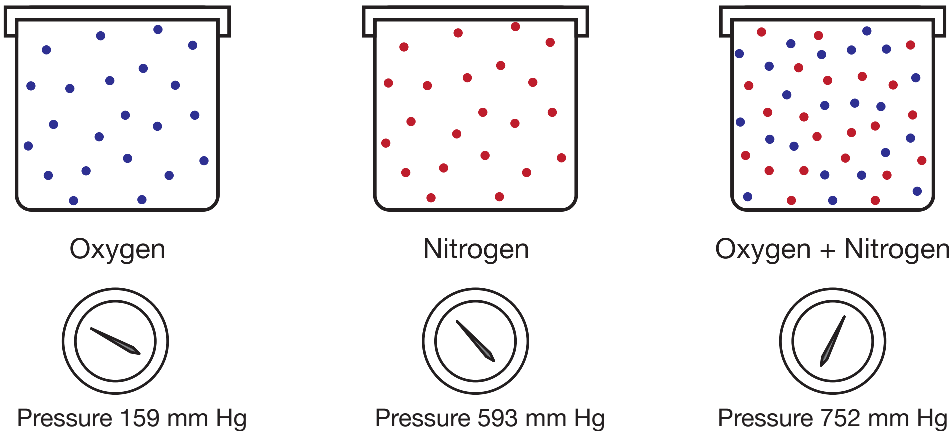


Figure 1.5 Dalton's law of partial pressure

By applying ideal gas equation, the partial pressure of each gas can be calculated as given below

$$P_1 = \frac{n_1 RT}{V}$$

$$P_2 = \frac{n_2 RT}{V}$$

$$P_3 = \frac{n_3 RT}{V}$$

Now, sum of the partial pressure = total pressure exerted by the mixture gases.

$$P_t = P_1 + P_2 + P_3$$

Or,
$$P_t = \frac{n_1 RT}{V} + \frac{n_2 RT}{V} + \frac{n_3 RT}{V}$$

$$= \frac{RT}{V}(n_1 + n_2 + n_3 + \ldots)$$

$$= n_t \frac{RT}{V} \qquad (1.16)$$

Where, $n_t = n_1 + n_2 \ldots + n_n$ and is the total number of moles of the gas mixture in volume 'V'. The equation (1.16) can be used for mixtures of gases as well as for pure gases.

## Solved problems

1. 10 g each of nitrogen and helium gas are placed together in a 10 L container at 25°C. Calculate the partial pressure of each gas and the total pressure of the gas mixture. (R = 0.0821 litre atm, 1 atm = 101.3 kPa)

**Solution:**

The number of moles (n) of each gas present:

| | Nitrogen ($N_2(g)$) | Helium (He(g)) |
|---|---|---|
| mass (g) | 10 g | 10 g |
| molar mass (g $mol^{-1}$) | 2 × 14 = 28 | 4 |
| n = mass ÷ molar mass | 10 ÷ 28 = 0.36 mol | 10 ÷ 4 = 2.5 mol |

Total number of moles of gases in the mixture ($n_t$) = 0.36 + 2.5 = 2.86 mol

Total pressure exerted by the gas mixture :

$$P_1 = n_t \frac{RT}{V}$$

$$= 2.86 \times \frac{0.0821 \times 298}{10}$$

$$= 6.99 \approx 7\,\text{atm}$$

$$\text{Partial pressure of nitrogen} = \frac{\text{number of moles of nitrogen}}{\text{total number of moles}} \times \text{total pressure}$$

$$= \frac{0.36}{2.86} \times 7 = 0.88\,\text{atm}$$

$$\text{Partial pressure of helium} = \frac{\text{number of moles of helium}}{\text{total number of moles}} \times \text{total pressure}$$

$$= \frac{2.5}{2.86} \times 7 = 6.11\text{atm}$$

2. At 15°C, 25 mL of neon at 1 atm and 75 mL of helium at 0.7 atm are both expanded into a 1 L sealed flask at the same temperature. Calculate the partial pressure of each gas and the total pressure of the gas mixture.

**Solution:**

Since the temperature and moles of each gas is constant, the pressure exerted by each gas is inversely proportional to its volume which is according to Boyle's law.

$$P_1V_1 = P_2V_2$$

$$P_2 = \frac{P_1V_1}{V_2}$$





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$P_2$ is taken as the partial pressure of the gases in the mixture.

For neon gas:

$V_1$ = 25 mL = 0.025 L. $V_2$ = 1.0 L. $P_1$ = 1.0 atm.

$$\text{Partial pressure of neon } (P_{neon}) = \frac{1.0 \times 0.025}{1.0} = 0.025 \text{ atm}$$

For helium gas:

$V_1$ = 75 mL = 0.075 L. $V_2$ = 1.0 L. $P_1$ = 0.7 atm.

$$\text{Partial pressure of helium } (P_{helium}) = \frac{0.7 \times 0.075}{1.0} = 0.0525 \text{ atm}$$

Total pressure exerted by the gas mixture = 0.025 + 0.0525 = 0.0775 atm

Thus, the total pressure exerted by the gas mixture is 0.0775 atm.

3. In a gaseous mixture at 25°C, the partial pressures of the components are:

Hydrogen = 150 mm

Carbondioxide = 300 mm

Ethane = 110 mm

Methane = 130 mm

What is the total pressure of the mixture and the volume per cent of each component?

**Solution:**

Total pressure of the mixture = sum of all the partial pressure of the components.

= 150 + 300 + 110 + 130 = 690 mm

$$\text{Volume per cent of } H_2 = \frac{150}{690} \times 100 = 21.74$$

$$\text{Volume per cent of } CO_2 = \frac{300}{690} \times 100 = 43.38$$

$$\text{Volume per cent of } C_2H_6 = \frac{110}{690} \times 100 = 15.28$$

$$\text{Volume per cent of } CH_4 = \frac{130}{690} \times 100 = 18.84$$

4. Mixtures of helium and oxygen can be used in scuba diving tanks to help prevent the bends. For a particular dive, 46 L He at 25°C and 1.0 atm and 12 L $O_2$ at 25°C and 1.0 atm were pumped into a tank with a volume of 5.0 L. Calculate the partial pressure of each gas and the total pressure in the tank at 25°C.

**Solution:**

Calculate the number of moles of each gas using the ideal gas law in the form, $n = \frac{PV}{RT}$

$$n_{helium} = \frac{1.0 \times 46}{0.0821 \times 298} = 1.9\,\text{mol}$$

$$n_{oxygen} = \frac{1.0 \times 12}{0.0821 \times 298} = 0.49\,\text{mol}$$

The tank containing the mixture has a volume of 5.0 L, and the temperature is 25°C. We can use these data and the ideal gas law to calculate the partial pressure of each gas using $P = \frac{nPV}{RT}$

$$P_{helium} = \frac{1.9 \times 0.0821 \times 298}{5.0} = 9.3\,\text{atm}$$

$$P_{oxygen} = \frac{1.0 \times 0.0821 \times 298}{5.0} = 2.4\,\text{atm}$$

The total pressure is the sum of the partial pressures:

$$P_{Total} = P_{helium} + P_{oxygen}$$

$$P_{Total} = 9.3 + 2.4 = 11.7\ \text{atm}$$

**Practice problems**

1. A cylinder contains 400g of oxygen and 600g of helium at a total pressure of 7.0 atm. Calculate partial pressures of the gases.
2. A flask contains 3.0 moles of nitrogen and 3.0 moles of neon. How many grams of argon must be pumped into the flask so as to make the partial pressure of the argon twice that of helium?
3. A 2.0 L container is pressurized with 0.25 atm of oxygen gas and 0.60 atm of nitrogen gas. What is the total pressure inside the container?
4. A mixture of 2 mol $H_2$ and 3 mol He exerts a total pressure of 3 atm. What is the partial pressure of He?

## Self Evaluation

1. Figure 1.6 shows the behaviour of a gas at constant temperature. Study the figure and answer the questions that follow.

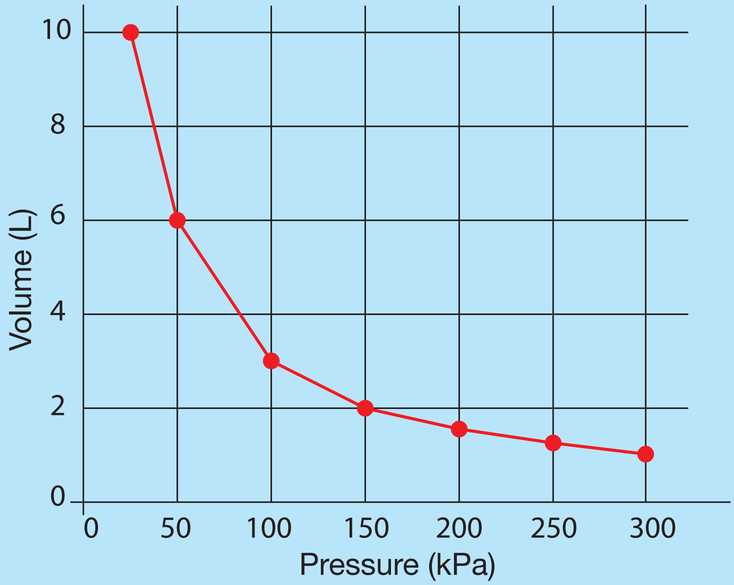



Figure 1.6

(a) Complete the Table 1.2

(b) What happens to the volume of a gas as the pressure increases?

(c) What happens to the pressure of a gas as the volume increases?

(d) What relation can you draw between pressure and volume of the gas at constant temperature?

**Table 1.2**

| Data for Gas sample at four pressures | | | |
|---|---|---|---|
| Sl.No | Pressure | Volume | Pressure x Volume |
| 1 | 25 kPa | ? | |
| 2 | ? | 6L | |
| 3 | 100 | ? | |
| 4 | ? | 2L | |
| 5 | 300 | ? | |

2. It is not advisable to place the cooking gas cylinder near the hot objects. Explain.

3. Carbon dioxide is usually formed when gasoline is burned. If 30 L of $CO_2$ is produced at a temperature of $1.00 \times 10^3$ °C and allowed to reach room temperature (25°C) without any pressure changes, what is the new volume of the carbon dioxide?
4. A tank contains 7.7 moles of gas at a pressure of 0.09 atm and a temperature of 56°C, what is the volume of the tank?

## Summary


1. Boyle's law states that the volume of a gas is inversely proportional to the pressure, if temperature is kept constant.
2. Charles' law states that the volume of a gas is directly proportional to its Kelvin temperature, if pressure is kept constant.
3. The pressure of a gas varies directly with the temperature, if the volume of the gas is kept constant.
4. Boyle's law and Charles' law can be combined into a single mathematical expression known as the combined gas law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
5. Avogadro's law states that equal volumes of different gases at the same temperature and pressure contain equal number of molecules or moles of molecules (n) that it contains.
6. The ideal gas law, PV = nRT, describes the relationship among the pressure, volume, number of moles, and temperature of an ideal gas.
7. The total pressure of the mixture is equal to the sum of the pressures that each gas would exert by itself in the same volume.


## Exercise

**I. Fill in the blanks with correct word(s).**

1. The volume of the gas would become zero at ______ temperature.
2. At constant temperature when the volume of the gas decreases, the pressure _____.
3. A sample of helium gas occupies 6 mL at a temperature of 250 K. At ____ K the gas expands to 9 mL.
4. The pressure of a gas mixture is equal to the_____ of the partial pressures of the constituent gases at a particular temperature.

5. The value of PV for 5.6 moles of a gas at 0°C is _____.

## II. Check whether the following statements are true or false and rewrite the false statements correctly.

1. The inflated balloon inside a car during a hot sunny day will experience an increase in pressure.
2. You drove continuously from Phuentsholing to Thimphu and observed that the pressure in your tyres increased. This is because of the increased friction between the road and the tyres that results in increase in temperature.
3. When the temperature of a sample gas increases from 100°C to 200°C, the average volume of its particles is doubled.
4. Charles' law states that when the pressure of a fixed mass of gas is held constant, the volume of the gas is directly proportional to its temperature.
5. The total pressure in a mixture of gases is equal to the partial pressure(s) of the gas that occupies the largest volume.

## III. Match the items of Column I with the corresponding items of Column II.

| Column I | Column II |
|---|---|
| 1. For a given mass of gas at constant temperature, the volume of the gas varies inversely with pressure. | a. Boyle's law. |
| 2. Law can be used to determine the total pressure of the mixture of gases. | b. Charles's law. |
| 3. The pressure of a gas is directly proportional to its Kelvin temperature if the volume is kept constant. | c. Graham's law. |
| 4. The volume of a fixed mass of gas is directly proportional to its Kelvin temperature, if the pressure is kept constant. | d. Dalton's law. |
| 5. The volume of gas is directly related to the number of moles at constant temperature and pressure. | e. Gay-Lussac's law. |
| | f. Ideal gas law. |
| | g. Avogadro's law. |

## IV. Choose the most appropriate response from the given options.

1. Gas pressure is caused by gas molecules
   - A heating up.
   - B hitting other gas molecules.
   - C hitting the walls of a container.
   - D reacting with other gas molecules.

2. Each of these flasks contains the same number of gas molecules. In which container is the pressure highest?


Figure 1.7

| | | | |
|---|---|---|---|
| A | Flask 1 | C | Flask 3 |
| B | Flask 2 | D | Flask 4 |

3. If the absolute temperature of a gas is doubled and the pressure of the gas is halved, how will the volume change?
   A It will increase by two times its original value.
   B It will decrease to one fourth of its original value.
   C It will stay the same as its original value.
   D It will increase by four times its original value.
4. At constant temperature, the product of pressure and volume of a given amount of gas is constant. This is according to:

| | | | |
|---|---|---|---|
| A | Charles' law | C | Boyle's law |
| B | Gay-Lussac's law | D | Avogadro's law |

5. 2 g each of the following gases at 60°C and 700 mmHg are taken. Which of them will have the highest volume?

| | | | |
|---|---|---|---|
| A | HBr | C | HF |
| B | HCl | D | HI |

6. At constant temperature in a given moles of an ideal gas
   A the ratio of pressure and volume always remains constant.
   B the product of pressure and volume remains constant.
   C volume always remains constant.
   D pressure always remains constant.
7. The food in a pressure cooker gets cooked well and faster than in a pot. This is because the food in the pressure cooker is cooked

| | | | |
|---|---|---|---|
| A | under high pressure | C | at constant volume. |
| B | at constant heat. | D | at constant pressure. |

8. If the pressure and absolute temperature of 4 litres of carbon dioxide are doubled, the volume of carbon dioxide would become.

| | | | |
|---|---|---|---|
| A | 2 litre | C | 6 litre |
| B | 4 litre | D | 8 litre |

## V. Write answers for the following questions

1. Table 1.3 shows the volume of a sample of gas at four temperatures.

**Table 1.3**

| Gas Samples at Four Temperatures | | |
|---|---|---|
| Trial | Temperature T | Volume V |
| A | 10.0°C | 100 $cm^3$ |
| B | 50.0°C | 114 $cm^3$ |
| C | 100.0°C | 132 $cm^3$ |
| D | 200.0°C | 167 $cm^3$ |

(a) What happened to the volume of gas when the temperature is changed after each trial?

(b) What kind of pattern or regularity do you observe in the data in Table 1.3?

(c) In trial B, C and D the temperature is doubled, what has happened to the volume?

(d) Based on the data in Table 1.3, plot a graph of 'temperature versus volume'.

(e) What kind of graph do you obtain? Explain the relationship between the two variables.

(f) What will happen to the volume of gas if the temperature is cooled below 0°C?

(g) Add two columns to the right hand side of the volume column in Table 1.3. Convert the temperatures to the Kelvin scale in one column and find the ratio of V/T in another column.

(h) Name and state the law associated with the above data.

2. A certain mass of gas occupies a volume of 2.5 L at 90 atm. What pressure would the gas exert if it were placed in a 10 L container at the same temperature?
3. A sample of gas at 1 atm had a volume of 1.2 L at 100°C. What would its volume be at 0°C and 700 mm Hg pressure?
4. 4.5 L of gas at 1.23 atm is expanded at constant temperature until the pressure is 0.74 atm. Calculate the final volume of a gas?
5. What volume is needed to store 0.050 moles of helium gas at 2 atm and 400 K?
6. 5.0 L of a gas is collected at 100 K and then allowed to expand to 20.0 L. What is the new temperature in order to maintain the same pressure (as required by Charles' law)?

7. What pressure will be exerted by 20.16 g hydrogen gas in a 7.5 L cylinder at 20°C?
8. A 50 L cylinder is filled with argon gas to a pressure of 100 atm at 30°C. How many moles of argon gas are in the cylinder?
9. To what temperature does a 250 mL cylinder containing 0.40 g helium gas need to be cooled in order for the pressure to be 2.49 atm?
10. A gas syringe contains 56.05 milliliters of a gas at 315.1 K. Determine the volume that the gas will occupy if the temperature is increased to 380.5 K.
11. A sample of gas at 1 atmosphere had a volume of 1.2 L at 100°C. What would its volume be at 0°C at the same pressure?
12. A balloon had a volume of 75 L at 25°C. How much the temperature should be raised in order for the balloon to have a volume of 100 L at the same pressure?
13. A container holds three gases: oxygen, carbon dioxide, and helium. The partial pressures of the three gases are 2 atm, 3 atm, and 4 atm, respectively. What is the total pressure inside the container?
14. A container with two gases, helium and argon, is 30 % by volume helium. Calculate the partial pressure of helium and argon if the total pressure inside the container is 4 atm.
15. If 60 L of nitrogen is collected over water at 40°C when the atmospheric pressure is 760 mm Hg, what is the partial pressure of the nitrogen?

## VI. Solve the cross word puzzle.

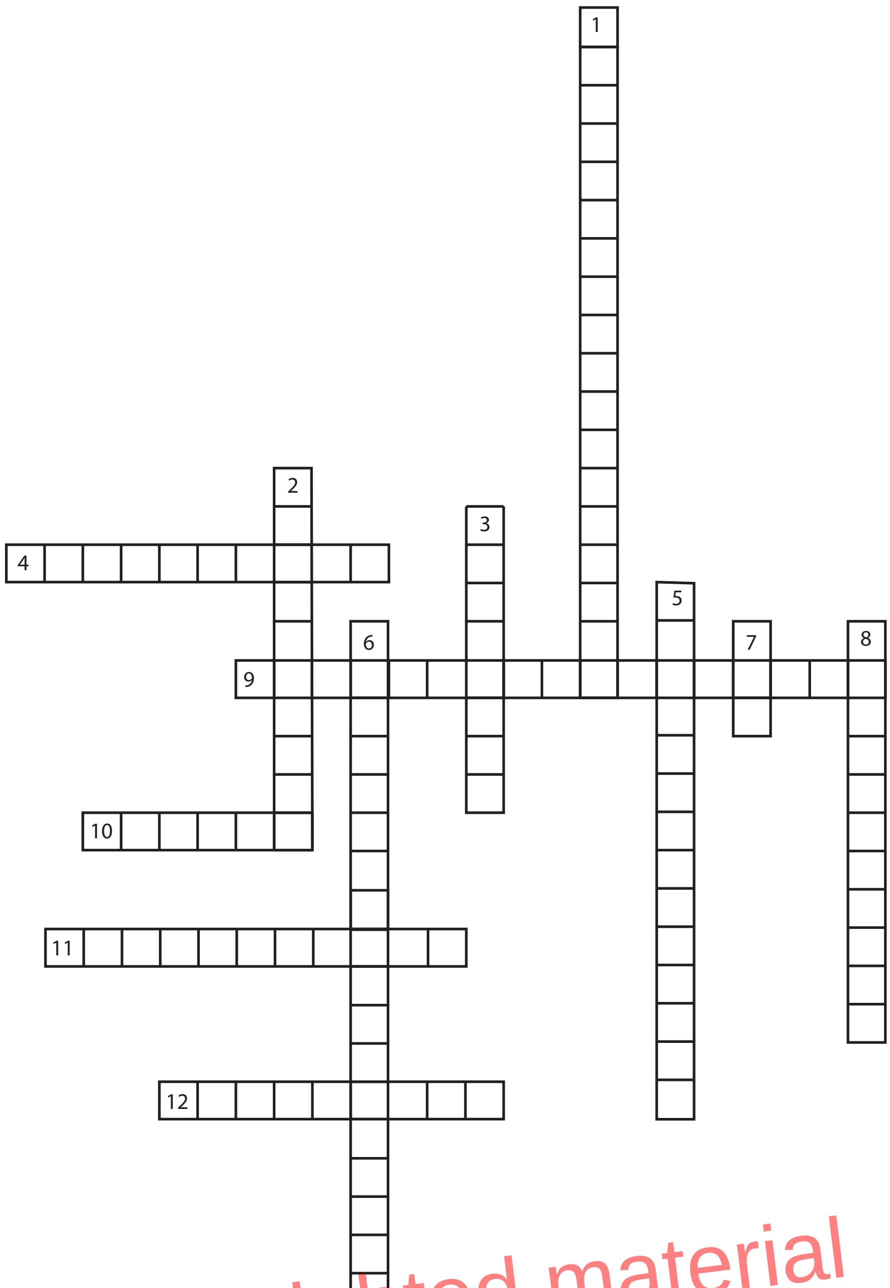
**Across**

4. At constant pressure, the volume of a gas is directly proportional to its Kelvin temperature.
9. The problem-solving method in chemistry that uses mathematical relationships to convert one quantity to another.
10. The SI unit for measuring pressure.
11. The temperature scale defined so that temperature of a substance is directly proportional to the average kinetic energy of particles such that zero on the scale corresponds to zero kinetic energy.
12. At constant temperature, the volume and pressure of a gas are inversely proportional.

**Down**

1. Statement that at the same temperature and pressure, equal volumes of gases contain equal numbers of particles.
2. The SI unit for measuring pressure that equals 1000 pascals.
3. A gas in which the particles undergo elastic collisions.

5. The combination of Boyle's law and Charles's law.
6. The pressure that supports a column of mercury 760 millimeters in height.
7. Abbreviation for standard temperature and pressure which is 0 degrees Celsius and 1 atmosphere.
8. The equation that expresses exactly how pressure, volume, temperature, and the number of particles of a gas are related.



# 2 The Mole Concept and Stoichiometry

## 2.1 Introduction

An equation of chemical reaction provides quantitative information relating the reactants and the products involved in it. So, it is possible to calculate the relation between the weights and the volumes of the substances taking part in a chemical change. A reaction always takes place through the interaction of integral number of molecules of the reactants, and the products formed are also in integral number of the molecules. There is a simple relationship among the gram-molecular weights or moles of the reactants and the products. The chemical calculations based on chemical equations are also calculated by using mole concept.

## 2.2 Relative atomic mass and Relative molecular mass, Avogadro's number and Mole

**Learning objectives**

On completion of this topic, students should be able to:

» *explain the terms such as relative atomic mass, relative molecular mass, Avogadro's number and mole.*

» *solve the numerical problems based on the concept of mole, relative atomic mass, relative molecular mass and Avogadro's number.*

### 2.2.1 Relative atomic mass (RAM or $A_r$)

Atoms are extremely small and cannot be seen or weighed directly but indirect methods in Physics have helped us to know the absolute mass of nearly all kinds of atoms. The mass of hydrogen atom is found to be $1.6735 \times 10^{-24}$ g while that of oxygen atom is $26.565 \times 10^{-24}$ g. These masses are so small that it is not convenient to use bigger units like kilograms or grams. It is appropriate to use the mass of some light atom as a unit and then relate mass of other atoms to it. The standard substance taken for this purpose is carbon-12 ($^{12}C$). One atom of carbon-12 is given the mass of 12.

$$\text{Atomic mass unit (a.m.u)} = \frac{1}{12} \times \text{mass of one atom of carbon-12}$$

The mass of an atom of the given element is compared with 1/12 the mass an atom of $^{12}C$ and the ratio thus obtained is known as relative atomic mass. The resulting masses of atoms are called relative atomic masses (RAM).

$$\text{RAM}(A_r) = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12}\text{mass of one atom of carbon - 12}}$$

## Relative atomic mass scale

**What is the Relative atomic mass scale ($^{12}$Carbon Scale)?**

- Masses of atoms are expressed relative to the mass of a carbon-12 atom.

↓

**Why is there a need for the scale?**

- Atoms are too small to be weighed directly.
- Inconvenient to express masses of atoms in terms of kg.

↓

**So what does it imply?**

- 1 atom of $^{12}C$ = 12
- $\frac{1}{12}$ the mass of a $^{12}C$ atom = 1

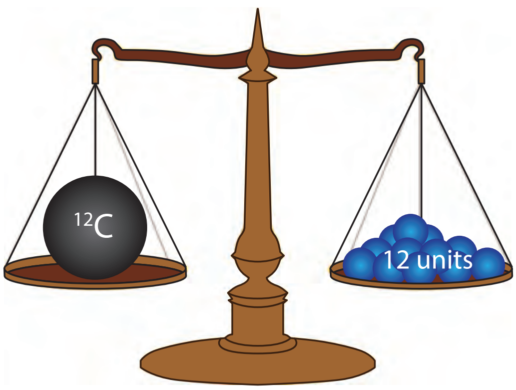



Figure 2.1

**Note**: Relative atomic masses have no units (dimensionless) since they are relative to the arbitrary standard (i.e. ratio).

However, in more advanced texts, relative masses are represented in terms of atomic mass units (a.m.u).

It is observed that most atomic masses are not whole numbers because majority of elements found in nature are a mixture of two or more isotopes of constant composition. The atomic weight of an element is the weight average of the atomic weights of its natural isotopes. For example, a sample of bromine prepared in the laboratory contains two isotopes: bromine – 79, ${}^{35}_{79}Br$ and bromine – 81, Br ${}^{35}_{81}Br$. A mass spectrometer can be used to find out the masses of these isotopes using the carbon-12 scale. Their relative isotopic masses are 78.919 and 80.917

respectively. The word ‘relative’ emphasizes that the masses are relative to an atom of carbon-12. The spectrometer will also show the proportions of each isotope.

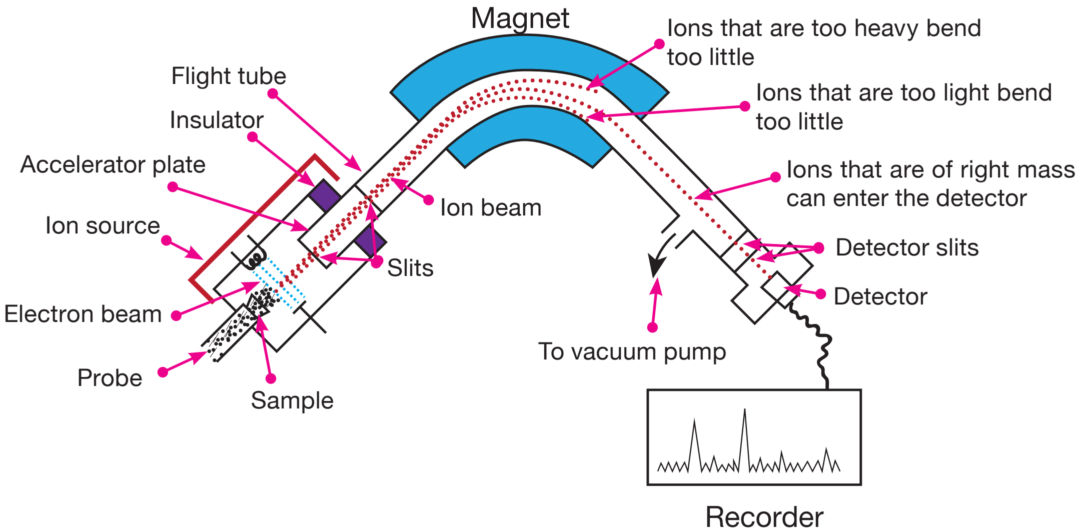



Figure 2.2 Mass spectrometer

For example, any naturally occurring sample of bromine contains approximately 50.52% of $^{79}_{35}Br$ and 49.48 of $^{81}_{35}Br$. Thus, the average mass of a bromine atom will be

$$\left(78.919 \times \frac{50.52}{100}\right) + \left(80.917 \times \frac{49.48}{100}\right) = 79.908$$

The value obtained is known as the relative atomic mass of bromine. Relative atomic masses are given the symbol $A_r$, and then show the element to which they refer in brackets. For example, $A_r$(Br) = 79.908.

The relative atomic mass or atomic weight of an element is the weighted average of the masses of the isotopes in the naturally occurring element relative to the mass of an atom of the carbon-12 isotope which is taken to be exactly 12. Relative atomic mass of an element is the number of times one atom of the element is heavier than $\frac{1}{12}$ times of the mass of an atom of carbon-12.

## 2.2.2 Gram atomic mass

There are several quantities that the chemists use to make their calculations easier. One of these is the gram atomic mass. A gram atomic mass of an element is that quantity of the element that has a mass in grams numerically equal to its atomic mass. The atomic weight of an element expressed in grams is gram atomic weight or one gram atom of that element. For example, the atomic mass of oxygen atom is 16 a.m.u, therefore its gram atomic mass is 16 g, and a 16 g sample of oxygen is equal to one atomic weight of oxygen.

The term can be applied either to the naturally occurring mixture of isotopes or to any particular isotope. A gram atomic mass is also called gram-atom, which is shorter and more convenient term.

**Table 2.1 Relative atomic mass and gram atomic mass**

| Element | Symbol | Relative atomic mass (At.wt.) | Gram atomic mass (gram atom) |
|---|---|---|---|
| Aluminium | Al | 26.98 a.m.u. | 27g |
| Carbon | C | 12.00 a.m.u. | 12g |
| Chlorine | Cl | 35.453 a.m.u. | 35.5g |
| Hydrogen | H | 1.008 a.m.u. | 1g |
| Iron | Fe | 55.847 a.m.u. | 56g |
| Nitrogen | N | 14.007 a.m.u. | 14g |
| Oxygen | O | 15.99 a.m.u. | 16g |

**Calculating relative atomic mass**

1. Naturally occurring silver is 51.84% silver-107 and 48.16% silver-109. Calculate the relative atomic mass of silver.

   Solution:

$$\begin{aligned}\text{RAM}(\text{Ag}) &= \left(\frac{51.84}{100}\times 107\right)+\left(\frac{48.16}{100}\times 109\right)\\ &= 55.469 + 52.494\\ &= 107.96\end{aligned}$$

   So, the RAM or $A_r(\text{Ag}) = 107.96$

2. Chlorine consists of two isotopes, 75% chlorine - 35 and 25% chlorine – 37. Calculate the relative atomic mass of chlorine.

   Solution:

$$\begin{aligned}\text{RAM}(\text{Cl}) &= \left(\frac{75}{100}\times 35\right)+\left(\frac{25}{100}\times 37\right)\\ &= 26.25 + 9.25\\ &= 35.5\end{aligned}$$

   So, the RAM or $A_r(\text{Cl}) = 35.5$

## 2.2.3 Relative molecular mass (RMM or $M_r$)

Since the molecules are very small their masses cannot be determined directly by weighing. The molecular masses are determined relative to the mass of $^{12}C$, which is taken as the standard substance. The absolute or actual molecular mass of a compound is the actual mass of one molecule of that compound. The mass of molecule of the given substance is obtained by adding together the relative atomic masses of all the atoms or ions present in a molecule or a compound.

$$\text{RMM}(M_r) = \frac{\text{Mass of one molecule of the substance}}{\text{Mass of } \frac{1}{12} \text{ atom of } ^{12}\text{C}}$$

Relative molecular mass of an element or compound is the number of times one molecule of the substance is heavier than $\frac{1}{12}$ the mass of an atom of carbon-12.

**Calculating relative molecular mass**

1. Calculate the relative molecular mass of a bromine molecule.

   **Solution:**

   Molecular formula of bromine molecules; $Br_2$

   $$\begin{aligned} \text{RMM of } Br_2 &= 2 \times A_r(\text{Br}) \\ &= 2 \times 79.908 \\ &= 159.816 \end{aligned}$$

It is not possible to discuss the atomic mass of a molecule or of a compound because more than one type of atom is involved. Instead, the relative formula mass (RFM) of the compound needs to be discussed. For example, sodium chloride, $Na^+ Cl^-$, contains sodium ions and chloride ions. Even though there are no sodium chloride molecules, the values $A_r$ (Na) = 23 and $A_r$ (Cl) = 35.5 are added to give $M_r$ (NaCl) = 58.5.

**$M_r$ = Relative formula mass (RFM) = Relative molecular mass (RMM)**

**= the sum of all the atomic masses for all the atoms or ions in a given formula**

If all the individual atomic masses of all the atoms in a formula are added together, then that is equal to relative formula mass or molecular mass. RFM is the sum of the relative atomic masses of all those elements shown in the formula of the substance.

**Calculating relative formula mass**

1. Calculate the relative formula mass for the following compounds and elements

   a. NaCl

   b. $N_2$

   c. $C_6H_{12}O_6$

   **Solutions:**

   a. RFM of NaCl = $1 \times 23 + 1 \times 35.5 = 58.5$

   b. RFM of $N_2 = 2 \times 14 = 28$

   c. RFM of $C_6H_{12}O_6 = 6 \times 12 + 12 \times 1 + 6 \times 12 = 180$

For substances that are molecular, the term gram molecular mass can be used in place of the term gram formula mass. Gram molecular mass (gram molecule) is the relative molecular mass of a substance expressed in grams.

Table 2.2 Relative molecular mass and gram molecular mass

| Substance | Relative molecular mass (Mol.wt.) | Gram molecular mass (gram mole) |
|---|---|---|
| Nitrogen | 28.014 | 28g |
| Oxygen | 31.998 | 32g |
| Chlorine | 70.906 | 71g |
| Carbon dioxide | 43.998 | 44g |
| Sulphur dioxide | 64.062 | 64g |
| Sulphuric acid | 98.076 | 98g |

## 2.2.4 Avogadro's number

The Italian physicist Amedeo Avogadro, first suggested the participation of molecules in the constitution of matter. According to him, the molecules and not the atoms were the smallest material particles that are capable of independent or free existence. He also held that the volumes of gaseous substances under identical conditions of temperature and pressure should be simply related to their number of molecules and not the atoms. So, Avogadro put forward a hypothesis, relating the volumes of gases under identical conditions of temperature and pressure with the number of molecules contained in them. The hypothesis is generally known as Avogadro's law.

The law states that equal volume of all gases under the same conditions of temperature and pressure contain the same number of molecules. The converse of Avogadro's law is also true. That is, if the samples of different gases at the same temperature and pressure contain the same number of molecules, then the volumes of all the samples must be equal. The coefficients in chemical equations tell the relative number of molecules, they also tell the relative volumes of gaseous substances, provided that these volumes are measured at the same temperature and pressure.

Example: If

| | | |
|---|---|---|
| 50 mL of $O_2$ at | STP contain | 'V' number of molecules, then |
| 50 mL of $CO_2$ at | STP also contain | 'V' number of molecules |
| 50 mL of $SO_2$ at | STP also contain | 'V' number of molecules |
| (Equal volume of gas) | (Same temperature, pressure) | (Equal number of molecules) |

**Activity 2.1 Worksheet**

**Instruction**

Study Table 2.3 and answer the questions that follow.

**Table 2.3**

| Elements | Hydrogen | Oxygen | Carbon dioxide |
|---|---|---|---|
| |  |  |  |
| 1. Volume | 22.4 L | 22.4 L | 22.4 L |
| 2. Pressure | 1 | 1 atm | 2 |
| 3. Temperature | 3 | 4 | 0°C |
| 4. $M_r$ | 5 | 6 | 7 |
| 5. Number of Molecules | 8 | 9 | 10 |

1. Fill up the blank spaces numbered 1 to 10.
2. Calculate the relative molecular mass for each type of gas.
3. How many molecules are present in the containers containing hydrogen, oxygen and carbon dioxide? Justify.

The specific number of molecules in one gram-mole of a substance is the molecular weight in grams and is equal to $6.023 \times 10^{23}$. For example, the molecular weight of oxygen is 32.0 and contains $6.023 \times 10^{23}$ oxygen molecules. The number $6.023 \times 10^{23}$ is called Avogadro's number ($N_A$) or Avogadro's constant. The units may be electrons, atoms, ions, or molecules, depending on the nature of the substance and the type of reaction.

**Solved Problems**

1. Calculate the number of molecules in 6.4 grams of sulphur dioxide at STP. (At.wt. S=32, O=16)

**Solution:**

Molecular weight = sum of atomic weights.

Molecular weight of $SO_2$ = 32 + (16)2 = 64

GMW of $SO_2$ is 64 grams.

If 64 grams of $SO_2$ contains = $6.023 \times 10^{23}$ molecules.

Then, 6.4 grams of $SO_2$ contains = $\frac{6.4}{64} \times 6.023 \times 10^{23}$

= $6.023 \times 10^{22}$ molecules.

2. How many grams are there in $1.8 \times 10^{23}$ molecules of sulphur dioxide at STP?

**Solution:**

If $6.023 \times 10^{23}$ molecules of sulphur dioxide contains = 64 grams.

Then, $1.8 \times 10^{23}$ molecules of sulphur dioxide contains

$$= \frac{1.8 \times 10^{23} \times 64}{6.023 \times 10^{23}} = 19.126\,g$$

## 2.2.5 Mole concept

Atoms and molecules are very small and even a little chemical sample contains an unimaginable number of them. Therefore, it is impossible to count the number of atoms or molecules in a sample. The mole, abbreviated as ‘mol’ is SI unit, which measures the number of particles in a specific substance. A mole is the quantity of any substance which contains as many elementary entities such as atoms, molecules or ions as there are in 12.00 grams of carbon-12. The number of particles is Avogadro's number, which is equal to $6.023 \times 10^{23}$. Thus, the weight of $6.023 \times 10^{23}$ molecules of any substance equals its gram molecular weight and they occupy 22.4 litres at STP. The volume 22.4 litres (22400 $cm^3$) is also called molar volume. For example, one mole of oxygen molecule means 32 grams of oxygen or $6.023 \times 10^{23}$ molecules of oxygen or 22.4 litres of oxygen at STP. Similarly, one mole of oxygen atoms means 16 grams of oxygen or $6.023 \times 10^{23}$ atoms of oxygen.

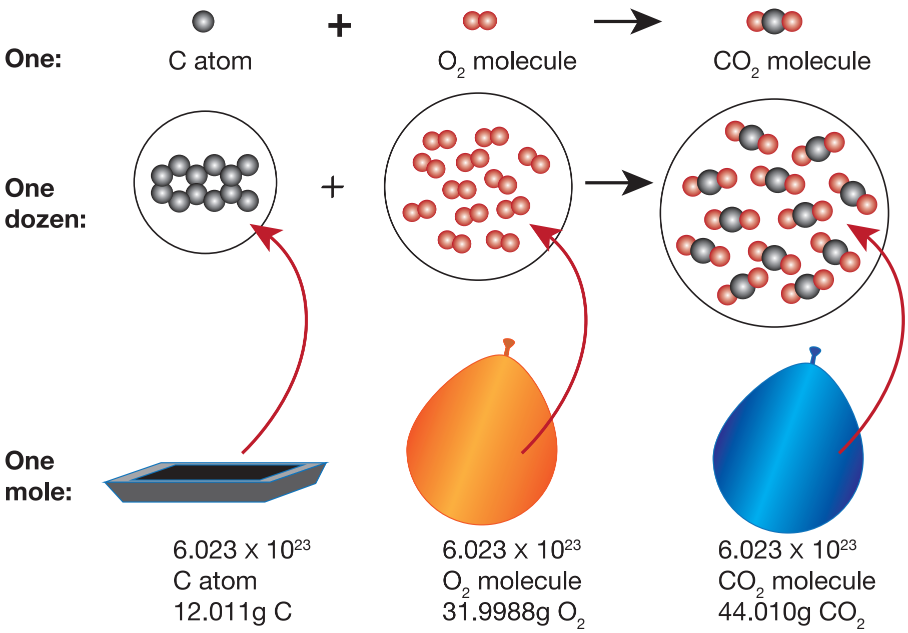



Figure 2.3 Mole of substances

Thus, mole is a unit for expressing number of atoms, molecules, ions, electrons etc. Hence, it can be conclude that

- One mole of any substances contains = $6.023 \times 10^{23}$ particles (the particles can be atoms, molecules, ions, electrons etc. Thus, we need to specify particles).

  For example,

  one mole of hydrogen atoms (H) contains = $6.023 \times 10^{23}$ atoms of hydrogen.

  one mole hydrogen molecules ($H_2$) contains = $6.023 \times 10^{23}$ molecules of hydrogen.

  one mole of water molecules ($H_2O$) contains = $6.023 \times 10^{23}$ molecules of water.

  one mole hydrogen ions ($H^+$) contains = $6.023 \times 10^{23}$ ions of hydrogen.

- One mole of an atom weighs one gram atomic weight of the atom.
- One mole of any substance (molecule) weighs one gram molecular weight of the substance.
- One mole of any gas weighs one gram molecular weight and occupies 22.4 litres (molar volume) at STP.

$$\text{One mole of atoms} = \frac{\text{Mass of element in grams}}{\text{Gram atomic mass of element}}$$

$$\text{One mole of substance (molecule)} = \frac{\text{Mass of substance in grams}}{\text{Gram molar mass}}$$

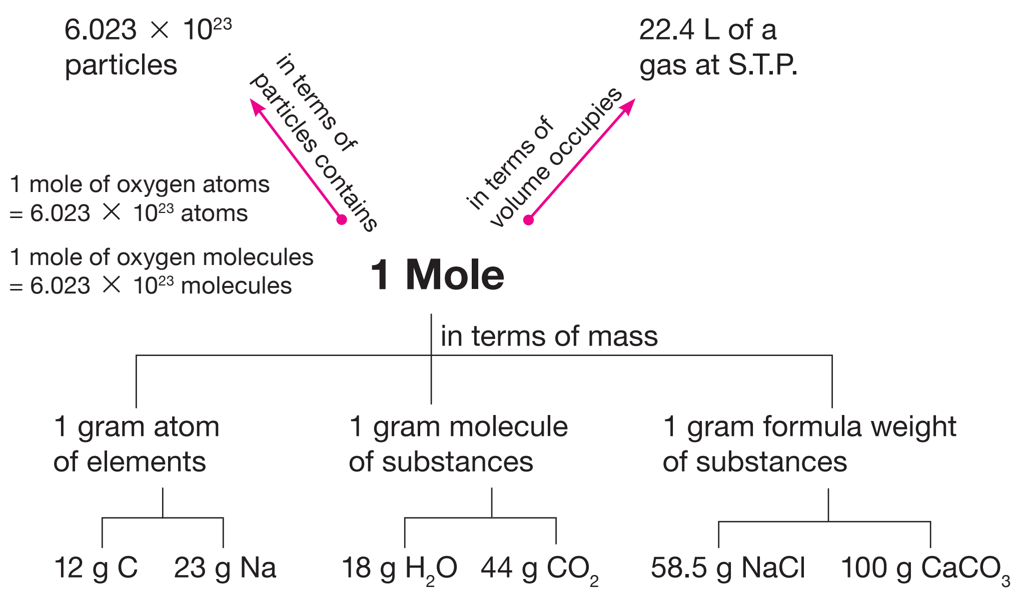



Figure 2.4 Mole concept

**Solved Problems**

1. If atomic mass of Ca atom is 40 g, find the number of atoms present in 1g of Ca.

   **Solution:**

   40g of Ca atoms contains = $6.023 \times 10^{23}$ atoms.

   Therefore, 1g of Ca atoms contains $= \frac{1}{40} \times 6.023 \times 10^{23}$

   $= 0.1505 \times 10^{23}$ atoms or $1.505 \times 10^{24}$ atoms





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2. Calculate the mass in grams of a single carbon (C) atom.

**Solution:**

The mass of one mole of carbon atom = 12g

One mole of carbon atom contains = $6.023 \times 10^{23}$ atoms.

$$\text{Mass of 1 C atom} = \frac{\text{Mass of a mole of atoms}}{6.023 \times 10^{23}}$$

$$\text{Mass of 1 C atom} = \frac{12\,\text{g}}{6.023 \times 10^{23}\,\text{atoms}}$$

Mass of 1 C atom = $1.994 \times 10^{-23}$ g

Therefore, the mass of a single carbon atom is $1.994 \times 10^{-23}$ g.

3. Calculate the number of moles of nitrogen in 7g of nitrogen (N=14).

**Solution:**

1 mole of any substance = 1 gram molecular weight of it.

Molecular weight of $N_2$ = $14 \times 2 = 28$

$\therefore$ gram molecular weight of nitrogen = 28 g

$\therefore$ 28 g of nitrogen = 1 mole of nitrogen

$\therefore$ 7 g of nitrogen = $\frac{1}{28} \times 7 = 0.25$ moles

**Alternative method.**

$$\text{Mole} = \frac{\text{weight in gram of a substance}}{\text{GMW (molecular weight)}}$$

$$\text{Mole} = \frac{7}{28} = 0.25 \text{ moles}$$

4. Calculate the mass of 50 cc of CO at STP. (C = 12, O = 16)

**Solution:**

1 mole of any substance = 1 gram molecular weight of it and occupies 22.4 litres at STP.

Gram molecular weight of carbon monoxide = 12 + 16 = 28 g

1 mole of CO = 1 g mol. wt and occupies 22400 cc at STP.

= 28 g of CO occupies 22400 cc at STP.

= ? g of CO will occupy 50 cc at STP = $\frac{28 \times 50}{22400}$ = 0.0625 g

5. In 3 moles of $O_2$: (a) How many $O_2$ molecules are there? (b) How many O atoms are there?

**Solution:**

a. 1 mole of $O_2$ molecule contain = $N_A$ = $6.023 \times 10^{23}$ molecules

3 moles of $O_2$ molecules contain = $\frac{6.023 \times 10^{23}}{1} \times 3$

= $1.8 \times 10^{24}$ molecules of $O_2$.

b. 1 mole of O atom contain = $N_A$ = $6.023 \times 10^{23}$ atoms

3 moles of O atoms contain = $\frac{6.023 \times 10^{23}}{1} \times 3 \times 2$ O-atom in $O_2$ atoms.

= $3.6 \times 10^{24}$ atoms of O.

6. How many atoms of oxygen are present in 300 g of $CaCO_3$?

**Solution:**

Molecular mass of $CaCO_3$ = 40 + 12 + (16 × 3) = 100 g.

1 mole of $CaCO_3$ contain = 3 moles of oxygen atoms

100 g of $CaCO_3$ contain = $3 \times 6.023 \times 10^{23}$ atoms

300 g of $CaCO_3$ contain = $\frac{3 \times 6.023 \times 10^{23}}{100} \times 300$

= $54.2 \times 10^{23}$ oxygen atoms.

### Self Evaluation

1. Use the information from the Table 2.4.

Table 2.4

| Element | Isotopes | Mass Number | Relative Abundance (%) |
|---|---|---|---|
| Nitrogen | $^{14}N$ | 14 | 99.63 |
| | $^{15}N$ | 15 | 0.37 |
| Oxygen | $^{16}O$ | 16 | 99.76 |
| | $^{17}O$ | 17 | 0.04 |

Calculate the

(a) relative atomic mass of oxygen and nitrogen atom respectively.

(b) relative formula mass of the dioxide formed by the two elements.

(c) number of atoms of oxygen present in one mole of dioxide formed by two elements.

2. For 196 g of pure $H_2SO_4$, calculate the

(a) number of moles of $H_2SO_4$.

(b) total number of $H_2SO_4$ molecules.
(c) total number of atoms present in it.
(d) number of atoms of each kind present in it.
(e) absolute mass of $H_2SO_4$ molecules.

3. How many moles of water are there in 1 L of water? Assume a density of 1.0 g $mL^{-1}$.
4. From 200 mg of $CO_2$, $10^{21}$ molecules are removed. How many molecules of $CO_2$ are left?
5. Calculate the number of water molecules contained in a drop of water weighing 0.06 g?
6. Find the number of aluminum ions present in 0.051 g of aluminium oxide, $Al_2O_3$.
7. Borax is the common name of sodium tetraborate, $Na_2B_4O_7$. In 20 g of borax,
   (a) how many moles of boron are present?
   (b) how many grams of boron are present?

## 2.3 Percentage composition, empirical formula and molecular formula

**Learning objectives**

On completion of this topic, students should be able to:

» *define the terms empirical formula and the molecular formula.*
» *explain the differences between empirical formula and molecular formula.*
» *solve numerical problems based on the percentage composition, empirical formula and molecular formula.*

### 2.3.1 Percentage composition

Percentage composition is the percentage by weight of each element present in the compound. The percentage of each element in the compound gives the mass of the element (in grams) present in 100 g of the compound. It compares the mass of one part of a substance to the mass of the whole. It is also the percentage by mass of atoms of an element present in one mole of the compound.

$$\text{Percentage composition} = \frac{\text{Total wt. of element in one molecule of the compound}}{\text{Gram molecular weight of the compound}} \times 100$$

### Determination of percentage composition of an element in a compound

**Solved Problems**

1. Calculate the % of copper, sulphur and oxygen in copper sulphate, $CuSO_4$. (Cu = 64, S = 32 and O = 16).

   **Solution:**

   The relative molecular mass of $CuSO_4 = 64 + 32 + (4 \times 16) = 160$

   Gram molecular weight of $CuSO_4$ = 160 g

   % composition of Cu in $CuSO_4 = \frac{64}{160} \times 100$ = 40% copper by mass in the compound

   % composition of S in $CuSO_4 = \frac{32}{160} \times 100$ = 20% sulphur by mass in the compound

   % composition of O in $CuSO_4 = \frac{64}{160} \times 100$ = 40% oxygen by mass in the compound

2. Calculate the percent by weight of sodium (Na) and chlorine (Cl) in sodium chloride. (Na = 23, Cl = 35.5).

   **Solution:**

   The relative molecular mass of NaCl = 23 + 35.5 = 58.5

   Gram molecular mass of NaCl = 58.5g

   % by weight of Na in NaCl = $\frac{\text{Mass of Na}}{\text{RMM}} \times 100 = \frac{23}{58.5} \times 100 = 39.32\%$

   % by weight of Cl in NaCl = $\frac{\text{Mass of Cl}}{\text{RMM}} \times 100 = \frac{35.5}{58.5} \times 100 = 60.68\%$

3. Find the percentage composition of Cu and $H_2O$ in the compound $CuSO_4.5H_2O$?

   **Solution:**

   The relative molecular mass of $CuSO_4.5H_2O$ = 250

   Gram molecular mass of $CuSO_4.5H_2O$ = 250 g

   Percentage composition of Cu in $CuSO_4.5H_2O = \frac{64}{250} = 26\%$

   Percentage composition of $H_2O$ in $CuSO_4.5H_2O = \frac{90}{250} = 36\%$

4. Calculate the percentage composition of water in hydrated magnesium sulphate $MgSO_4.7H_2O$.

   **Solution:**

   Molecular mass of $MgSO_4.7H_2O$= 24 + 32 + 16 × 4 + 7 (18) = 246 g

   Percentage composition of $H_2O$ in $MgSO_4.7H_2O = \frac{126}{246} \times 100 = 51.2\ \%$

## 2.3.2 Empirical formula

The empirical formula of a compound shows the atomic ratio of the elements present in a molecule of the compound. For example, the empirical formula of hydrogen peroxide ($H_2O_2$) is HO, because the simplest atomic ratio of hydrogen and oxygen in a molecule of it is 1:1. The empirical formula of a compound can be determined if the percentage composition of the elements in the compound is known.

Empirical formula is the simplest formula of a compound which gives the simple whole number ratio of various elements present in one molecule of the compound.

**Determination of empirical formula**

**i. From percentage composition**

A compound contains the following percentage composition 6.7% H, 40% C and 53.3% O. Determine its empirical formula.

Solution:

| Steps | Element | | |
|---|---|---|---|
| | H | C | O |
| 1. Write the percentage of the different elements. | 6.7 | 40 | 53.3 |
| 2. Divide the percentage of each element by the respective atomic masses to obtain the number of moles of atoms of the elements. | $\frac{6.7}{1} = 6.7$ | $\frac{40}{12} = 3.33$ | $\frac{53.3}{16} = 3.33$ |
| 3. Divide each ratio by the smallest number in order to obtain the simple whole number. | $\frac{6.7}{3.33} = 2$ | $\frac{3.33}{3.33} = 1$ | $\frac{3.33}{3.33} = 1$ |
| 4. Write the empirical formula by writing the symbols of elements with number of atoms as the subscript to the lower right of the symbol. | $CH_2O$ | | |

**ii. From mass**

A compound was analyzed and found to contain 13.5 g Ca, 10.8 g O, and 0.675 g H. Determine the empirical formula of the compound?

**Solution:**

| Steps | Element | | |
|---|---|---|---|
| | Ca | O | H |
| 1. Write the masses of each element, given in the problem. | 13.5 | 10.8 | 0.675 |
| 2. Divide the mass of each element by the respective atomic masses to obtain the number of moles. | $\frac{13.5}{40.1} = 0.337$ mol | $\frac{10.8}{16} = 0.675$ mol | $\frac{0.675}{1} = 0.675$ mol |
| 3. Divide each mole value by the smallest number of moles calculated. | $\frac{0.337}{0.337}$ = 1.000 | $\frac{0.675}{0.337}$ = 2.002 | $\frac{0.675}{0.337}$ = 2.002 |
| 4. Round off the ratio obtained in step 3 to the nearest whole number. This is called mole ratio of the elements. Note: If the number is too far to round, then multiply each solution by the same factor to get the lowest whole number multiple. | $\approx 1$ | $\approx 2$ | $\approx 2$ |
| 5. Write the empirical formula by writing the symbols of elements with mole ratio as the subscript to the lower right of the symbol. | $CaO_2H_2$ or $Ca(OH)_2$ | | |

## 2.3.3 Molecular formula

The molecular formula can be derived from the empirical formula when the molecular mass of the compound is known. Molecular formula gives the actual number of atoms ‘n’ of the different elements present in a molecule of the compound. The formula for determining the value of ‘n’ is

$$\text{Emperical fromula} \times n = \text{molecular formula, where "n" is an integer}$$

$$\therefore \text{Relative molecular mass} = n \times \text{molecular mass}$$

If the substance is a gas or a compound which can be completely volatilised, the molecular mass can be easily calculated from the relative density (or the vapour density).

$$\text{Relative molecular mass} = 2 \times \text{Vapour density}$$

## Determination of molecular formula

**Solved Problems**

1. Calculate the empirical and molecular formula of the compound having the following percentage composition; C = 26.59%, H = 2.22%, O = 71.19%. Its molecular weight is 90.

**Solution:**

**Step I**: Calculate the empirical weight of the compound from its empirical formula.

| Element | Percentage Composition | Atomic masses | Atomic ratio | Simplest ratio |
|---|---|---|---|---|
| C | 26.59 | 12 | $\frac{26.59}{12} = 2.22$ | $\frac{2.22}{2.22} = 1$ |
| H | 2.22 | 1 | $\frac{2.22}{1} = 2.22$ | $\frac{2.22}{2.22} = 1$ |
| O | 71.19 | 16 | $\frac{71.19}{16} = 4.45$ | $\frac{4.45}{2.22} = 2$ |

Hence, the empirical formula is $CHO_2$

The empirical formula weight of $CHO_2 = 12 + 1 + 16 \times 2 = 45$.

**Step II**: Divide its molecular weight by empirical weight which gives the value of (n)

Emperical formula × n = molecular formula, where "n" an integer

$(CHO_2)n = 90$

$(45)n = 90$

$n = \frac{90}{45} = 2$

**Step III**: Multiply the empirical formula by this number to get the molecular formula.

$(CHO_2) \times 2 = C_2H_2O_4$

The molecular formula is $C_2H_2O_4$

2. A compound with a molar mass of 34.0 g $mol^{-1}$ is known to contain 5.88% hydrogen and 94.12% oxygen. Find the molecular formula for this compound.

**Solution:**

**Step I**: Determine the empirical formula using the percentage composition of the compound.

| Element | H | O |
|---|---|---|
| % by mass | 5.88 | 94.12 |
| Molar mass / mol $g^{-1}$ (relative atomic mass in g) | 1.008 | 16.00 |
| Moles = mass ÷ molar mass | $\frac{5.88}{1.008} = 5.83$ | $\frac{94.12}{16.00} = 5.88$ |
| Divide throughout by the smallest number of moles calculated | $\frac{5.88}{5.83} = 1$ | $\frac{5.88}{5.88} = 1$ |
| Convert mole ratio to an empirical formula | $H_1O_1$ is HO | |

**Step II**: Determine the molecular formula using the empirical formula and molar mass of the compound

Empirical formula is HO

Molecular formula = n × empirical formula

i.e., Molecular formula = n (HO) = $H_nO_n$

The molar mass of the empirical formula HO = 1.008 + 16.00 = 17.008 g $mol^{-1}$

**Calculating for the value of "n"**

molar mass of compound = n × molar mass of empirical formula

34.0 = n × 17.008

$$n = \frac{34.0}{17.008} = 2$$

Now, substitute the value for 'n' into the molecular formula $H_nO_n$.

The molecular formula of the compound is $H_2O_2$.

**Alternative method**

Empirical formula is HO

Empirical weight of HO = 1+ 16 = 17

Molecular formula = n × empirical formula

$$\text{Where, } n = \frac{\text{molecular formula weight}}{\text{emperical formula weight}}$$

$$n = \frac{34}{17} = 2$$

Molecular formula = n × empirical formula

= 2 × HO

**Molecular formula = $H_2O_2$**

## 2.3.4 Differences between empirical formula and molecular formula

| Empirical Formula | Molecular Formula |
|---|---|
| The formula which gives the simple whole number ratio of the atoms of various elements present in one molecule of the compound is called empirical formula. | The formula which gives the actual number of atoms of various elements present in one molecule of the compound is called molecular formula. |
| The empirical formula expresses the elements by which it is formed. | This formula expresses the molecular structure of the compound or element. |
| It is determined by knowing the percentage of the elements of the compounds. | It can be determined by knowing the percentage of elements and the molecular mass of the compounds. |
| Sometimes the molecular formula and empirical formula of any compound become identical. e.g., HCl | Molecular formula of any compound sometimes becomes identical with empirical formula and sometimes it becomes multiple of the empirical formula.<br>e.g., $H_2O$ (empirical formula $H_2O$)<br>$C_6H_6$ (empirical formula CH) |
| Empirical formula of different compounds can be same.<br>e.g., empirical formula of benzene and ethyne is CH. | Molecular formula of any compound is always fixed except isomerism.<br>e.g., The molecular formula of benzene and ethyne are respectively $C_6H_6$ and $C_2H_2$. |
| Empirical formula is only for compound. | Molecular formula is for both compounds and elements. |

### Self Evaluation

1. Copper (II) carbonate in a crucible produced only 7.0 g of copper (II) oxide. What was the percentage yield of copper (II) oxide?
2. Find the percent composition of calcium (Ca), oxygen (O) and hydrogen (H)

in calcium hydroxide, $Ca(OH)_2$.

3. Bromine is used to make 1,2-dibromoethane. This is an additive in leaded petrol. 1,2-dibromoethane reacts with sodium hydroxide to form a compound that has the composition by mass: carbon-38.7%, hydrogen-9.7% and rest oxygen.
   (a) Calculate its empirical formula.
   (b) The relative molecular mass of the compound is 62. What is its molecular formula?
4. Calculate the percent composition of carbon in:
   (a) $CO_2$
   (b) $C_6H_{12}O_6$
5. A compound is found to contain 50.05 % sulphur and 49.95 % oxygen by weight. What is the empirical formula for this compound? The molecular weight for this compound is 64.07 g $mol^{-1}$. What is its molecular formula?
6. A compound is 19.3% Na, 26.9% S, and 53.8% O. Its formula mass is 238 g $mol^{-1}$. What is the molecular formula?

## 2.4 Calculations based on chemical reactions

**Learning objectives**

On completion of this topic, students should be able to:

» name the types of chemical reactions.

» solve numerical problems based on chemical equation.

» solve problems based on mass – mass and mass - volume relationship.

**Activity 2.2 Types of chemical reaction**

**Instruction**

Study the Table 2.6 and write down the missing information to complete the table.

**Table 2.6**

| Sl. No. | Reaction Type | Explanation | General Formula |
|---|---|---|---|
| 1 | Combination | Two or more elements or compounds (reactants) combine to form a single compound (products). | $A + B \rightarrow AB$<br>e.g. $S(s) + O_2(g) \rightarrow SO_2(g)$ |

| Sl. No. | Reaction Type | Explanation | General Formula |
|---|---|---|---|
| 2 | ? | The opposite of a combination reaction – a complex molecule breaks down to make simpler one. | $AB \rightarrow A + B$<br>For example: ? |
| 3 | Precipitation | Two solutions of soluble salts are mixed, resulting in an insoluble solid (precipitate) forming. | A + Soluble salt $\rightarrow$ Precipitate + soluble salt C<br>For example: ? |
| 4 | Combustion | ? | $A + O_2 \rightarrow H_2O + CO_2$<br>For example: ? |
| 5 | ? | The more reactive element displaces a less reactive element from its compound. | ?<br>e.g. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$ |

The chemical reactions in Table 2.6 can be represented with the help of chemical equation. A chemical equation gives molecular formulae of the reactants and the products in a chemical reaction. An equation conveys the following information. For example

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

i. *The molecular proportion of substances*: In this equation two molecules of solid potassium chlorate on heating in the presence of manganese dioxide give two molecules of solid potassium chloride and three molecules of oxygen gas.

ii. *The relative molecular masses of the substances*: 2 × 122.5g = 245g of potassium chlorate gives 2 × 74.5g = 149g of KCl and 3 × 32 = 96g of oxygen.

iii. *The volumes of gaseous substances*: 3 x 22.4 L = 67.2 L of oxygen at STP is evolved when 245g of potassium chlorate is heated.

## 2.4.1 Calculation based on chemical equations

### i. Mass – mass relationship

**Solved Problem**

1. The chief component of glass is silica ($SiO_2$). It can be dissolved by the hydrofluoric acid, HF, to form silicon tetra fluoride, $SiF_4$, a gas at room temperature according to the following reaction.

$SiO_2 + 4HF \longrightarrow SiF_4 + 2H_2O$

How many grams and how many moles of $SiF_4$ can be produced from 63.4g of HF?

**Solution:**

The balanced chemical equation for the reaction is

**Step I** $SiO_2 + 4HF \longrightarrow SiF_4 + 2H_2O$

**Step II** $4\,\text{moles HF} \longrightarrow 1\text{mole } SiF_4$

**Step III** $4(1+19) \longrightarrow 28.1 + 4 \times 19$

$80\text{g} \longrightarrow 104.1\text{g}$

**Step IV** $63.4\text{g} \rightarrow ?$

Molecular mass of HF = (1 + 19) = 20

$$20\text{g HF} = 1\text{mole}$$

$$63.4\text{g HF} = \frac{1}{20} \times 63.4 = 3.17\,\text{moles}$$

According to the balanced chemical equation,

4 moles of HF produce 1mole of $SiF_4$.

3.17 moles HF produces $= \frac{3.17}{4} = 0.793\,\text{moles}$

Molecular mass of $SIF_4$ = 104.1

1 mole of $SiF_4$ = 104.1g

0.793 mole of $SiF_4$ = $104.1 \times 0.793$ g

= 82.5g

2. Copper on reacting with conc.$H_2SO_4$ produces copper sulphate. If 1.28g of copper is to be converted to copper sulphate. Find the weight of the copper sulphate formed and the weight of the acid required at the same time according to the equation. (Cu = 64, S = 32, O = 16).

$Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2$

**Solution:**

The balanced chemical equation involved in the reaction is

a. $Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2$

64g [64 + 32 + 4 × 16=160g]

64g of Cu yields 160g of $CuSO_4$

$\therefore$ 1.28g of Cu will yield $\frac{160 \times 1.28}{64} = 3.2\,\text{g of } CuSO_4$

b. $Cu + 2H_2SO_4 \longrightarrow CuSO_4 + 2H_2O + SO_2$

64g 2[2 × 1 + 1 × 32 + 4 × 16]

$2 \times 98 = 196g$

196g of $H_2SO_4$ are required to react with 64g of Cu.

Hence, ?g of $H_2SO_4$ are required to react with 1.28g of Cu.

$$X = \frac{1.28 \times 196}{64} = 3.92g$$

$\therefore$ the weight of acid required is 3.92g

## ii. Mass – volume relationship

**Solved problem**

1. Calcium carbonate reacts with dilute HCl according to the equation.

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

Calculate the weight of $CaCl_2$ obtained from 10g of $CaCO_3$ and also the volume at STP of $CO_2$ obtained at the same time. (Ca = 40, C = 12, O = 16, Cl = 35.5)

Solution:

$$CaCO_3 \quad + \quad 2HCl \longrightarrow CaCl_2 \quad + \quad H_2O + CO_2$$

| | | |
|---|---|---|
| [40 + 12 + 48] | [40 + 71] | 1 mole |
| =100g | =111g | 22.4 litres at STP. |

100g of $CaCO_3$ gives 111g of $CaCl_2$

$$\therefore \text{10g of } CaCO_3 \text{ gives } \frac{111 \times 10}{100} = 11.1\text{g of } CaCl_2$$

Similarly, 100g of $CaCO_3$ liberate 22.4 litres of $CO_2$ at STP.

$$\therefore \text{10g of } CaCO_3 \text{ gives } \frac{22.4 \times 10}{100} = 2.24 \text{ litres of } CO_2$$

2. Calculate the volume of sulphur dioxide formed at STP in mL, by treating 4.8 g of copper with excess of hot concentrated sulphuric acid. (Cu = 64, H = 1, S = 32, O = 16).

**Solution:**

The equation obtained as per the problem is

$$Cu(s) + 2H_2SO_4(aq) \longrightarrow CuSO_4(aq) + 2H_2O(l) + SO_2(g)$$

64 g of Cu produces $\rightarrow$ 1 mole (of $SO_2$)

64 g $\rightarrow$ 22400 mL

Volume of $SO_2$ formed using 64 g of copper = 22400 mL

Volume of $SO_2$ formed using 4.8 g of copper = ?

$$X = \frac{4.8 \times 22400}{64} = 1680\,mL$$

Therefore, the volume of $SO_2$ produced at STP is 1680 mL.

### iii. Volume – volume relationship

**Solved Problem**

1. Propane gas burns in $Cl_2$ according to the equation

$$C_3H_8 + 4Cl_2 \longrightarrow 8HCl + 3C$$

What volume of chlorine will be used up when 40 litres of HCl gas is produced in the reaction?

**Solution:**

$$C_3H_8 \quad + \quad 4Cl_2 \quad \longrightarrow \quad 8HCl \quad + \quad 3C$$

1 Vol     4 Vol     8 Vol

As per the equation,

8 vol of HCl is produced when chlorine used up = 4 vol

$\therefore$ volume of Cl used up to produce 40 L of HCl = $\frac{4 \times 40}{8} = 20\,\text{litres}$

2. What is the volume of $O_2$ produced at STP when 61.25 g of $KClO_3$ is strongly heated?

(K = 39, Cl = 35.5, O = 16)

**Solution:**

The relative molecular mass of $KClO_3$= 39 + 35.5 + 3 × 16 = 122.5

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

$2(39 + 35.5 + 3 \times 16) \rightarrow 3 \times 22.4$ L

2(122.5) = 245

245 g of $KClO_3$ produces 3 x 22.4 litres of $O_2$

$\therefore$ volume of oxygen produced by 61.25g of $KClO_3$ = $\frac{61.25 \times 3 \times 22.4}{245} = 16.8\,L$

So, the volume of $O_2$ (measured at STP) produced = 16.8 L

## iv. Mass – number of particles relationship

**Solved Problem**

1. How many moles and how many atoms are contained in 10.0 g of nickel?

**Solution:**

Atomic mass of nickel = 58.69

The molar mass of nickel is 58.69 g $mol^{-1}$.

$$\text{Since, Mole} = \frac{\text{weight in gram of substance}}{\text{atomic weight}}$$

$$\therefore \text{the number of moles in Ni present} = \frac{10\,\text{g Ni} \times 1\,\text{mole Ni}}{58.69\,\text{g Ni}} = 0.170\,\text{mol Ni}$$

To determine the number of atoms, convert the moles of Ni to atoms using Avogadro's number:

Since, Number of atoms = Mole $\times$ Avogadro's number.

$$\therefore \text{the number of atoms present} = 0.170\,\text{mol Ni} \times \frac{6.023 \times 10^{23}\,\text{atoms Ni}}{1\,\text{mol Ni}}$$

$$= 1.02 \times 10^{23}\,\text{atoms}$$

2. A sample of gas contains 4.4 $\times$ $10^{24}$ carbon dioxide molecules. How many moles of carbon dioxide molecules are present in the sample?

**Solution:**

The number of moles of carbon dioxide molecules is

N(carbon dioxide molecules) = number of carbon dioxide molecules

= 4.4 $\times$ $10^{24}$ carbon dioxide molecules

1 mole of any substance contains Avogadro's number of molecules

$\therefore$ the number of moles of carbon dioxide molecules present in the sample

$$n = \frac{N}{N_A}, \text{ where } N_A = \text{Avogadro's number} = 6.023 \times 10^{23}$$

$$n(\text{carbon dioxide molecules}) = \frac{N(\text{carbon dioxide molecules})}{N_A}$$

$$n(\text{carbon dioxide molecules}) = \frac{N(\text{carbon dioxide molecules})}{6.023 \times 10^{23}}$$

$$= \frac{4.4 \times 10^{24}}{6.023 \times 10^{23}}$$

$$= 7.3\,\text{carbon dioxide molecules}$$

**Self Evaluation**

1. Carbon monoxide burns in oxygen to produce carbon dioxide according to the equation:

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

   How many molecules of oxygen would react with $5.0 \times 10^5$ molecules of CO?

2. Hydrogen and oxygen combine to form water: $2H_2 + O_2 \rightarrow 2H_2O$. A mixture of 22.4 L of hydrogen and 22.4 L of oxygen at 100°C is ignited.

   (a) Calculate the volume of steam produced.

   (b) What gas, if any will be present on cooling to room temperature?

3. Calculate the number of atoms of each element in 31.5 g of $HNO_3$(H = 1, N = 14, O = 16)

4. How much oxygen will contain the same number of atoms as the number of molecules in 73g of HCl? (H = 1, O = 16, Cl = 35.5)

5. From the equation $(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + 4H_2O + N_2$.

   Calculate:

   (a) the volume of nitrogen at STP, evolved when 63g of ammonium dichromate are heated.

   (b) the mass of $Cr_2O_3$ formed at the same time.

6. An acid of phosphorus has the following percentage composition: 2.47% H, 38.27% P, and 59.26% O. Find the empirical formula of the acid and its molecular formula, given that its relative molecular mass is 162.

   (H = 1, O = 16, P = 31)

7. Calculate the following based on the equation:

   $C + 2H_2SO_4 \rightarrow CO_2 + 2H_2O + 2SO_2$.

   (a) the mass of carbon oxidised by 49 g of $H_2SO_4$.

   (b) the volume of sulphur dioxide measured at STP, liberated at the same time.

8. When excess lead nitrate was added to a solution of sodium sulphate, 15.15g of lead sulphate was precipitated. What mass of sodium sulphate was present in the original solution?

   $Na_2SO_4 + Pb(NO_3)_2 \rightarrow PbSO_4 + NaNO_3$

## Summary

1. The study of quantitative relationships based on chemical formulas and equations is called stoichiometry.
2. The formula mass is the sum of the atomic masses of all the atoms represented by a formula. Either formula mass or molecular mass can be used to describe this sum in molecular substances. Only the term formula mass should be used for ionic substances.
3. The term relative atomic mass of an element is the number of times one atom of the element is heavier than one twelfth times of the mass of an atom of carbon-12.
4. The term relative molecular mass of an element represents how many times one molecule of the substance is heavier than $\frac{1}{12}$ of the mass of an atom of carbon-12.
5. A mole of substance contains Avogadro's number ($6.023 \times 10^{23}$) of units of that substance. Depending on the substance, a mole of a substance may have a mass equal to relative atomic mass, relative molecular mass or relative formula mass.
6. The mass of a sample of an element can be calculated from the number of moles present.
7. A mole of gaseous substance at STP occupies a volume of 22.4 litres.
8. The percentage composition of a compound is the percentage by mass of each of the elements in the compound.
9. The empirical formula of a compound can be determined from its percentage composition, the masses of each element in a sample of the compound or from the number of moles of each element that makes up the compound . In order to determine the molecular formula, the molar mass must be known.
10. A chemical equation is a condensed statement of facts about a chemical reaction. Reactants are substances that exist before a reaction takes place. Products are the substances that come into existence as a result of the reaction.
11. There are different types of chemical reaction such as synthesis, decomposition or analysis, displacement, precipitation, combustion, neutralization.
12. Calculations made from measurements enable chemists to determine the quantities of substances that react and are formed during chemical reactions.

13. The coefficients in a chemical equation represent the relative number of molecules and the relative number of moles of reacting substances. If the substances are gases at the same temperature and pressure, coefficients represent their relative volumes.
14. When the balanced equation for a chemical reaction is known and the mass of one substance taking part in the reaction is given, the masses of all the other substances taking part in the reaction can be determined.
15. If the balanced equation for a reaction is known, chemists can solve problems based on mass-mass relationship, mass-volume relationship, volume-volume relationship, and mass – number of particle relationship.

## Exercise

**I. State whether each of the statement is True or False.**

1. The standard molar volume of a gas is 22.4 L.
2. A given mass of a noble gas contains twice the number of atoms as molecules.
3. The molecular formula of a compound is a multiple of its empirical formula.
4. A mole of electrons means $6.023 \times 10^{23}$ electrons.
5. The vapour density of a gas is twice its relative molecular mass.

**II. Fill in the blanks by selecting the most appropriate word given in the bracket.**

(atoms, molecules, empirical, molecular, a, no, the same as, twice, half, one, two, HO, $H_2O$, $H_2O_2$, no)

1. At the same temperature and pressure, equal volume of gases contains the same number of _____.
2. The ____ formula of a compound is a simple multiple of its ______ formula.
3. There is ____ difference between the atomic and molecular masses of noble gas elements.
4. The relative molecular masses of hydrogen, nitrogen, oxygen and chlorine are ____ their relative atomic masses.
5. A compound of relative molecular mass 34 and empirical formula HO has the molecular formula_____.

**III. Choose the most appropriate response from the given options.**

1. What is the total number of moles of hydrogen gas contained in $9.03 \times 10^{23}$

| | | | |
|---|---|---|---|
| A | 1.50 moles | C | 6.02 moles |
| B | 2.00 moles | D | 9.03 moles |

2. A sample of 100 $cm^3$ of carbon monoxide was burnt in 100 $cm^3$ of oxygen. What was the composition of the gas remaining after the reaction? (All measurements were made at room temperature and pressure)

   A 100 $cm^3$ of carbon dioxide and 50 $cm^3$ of excess oxygen.

   B 100 $cm^3$ of carbon dioxide and 50 $cm^3$ of excess carbon monoxide.

   C 100 $cm^3$ of carbon dioxide only.

   D 200 $cm^3$ of carbon dioxide only.

3. The volume occupied by half a mole of a gas at STP is

   A 5.6 L C 22.4 L

   B 11.2 L D 44.8 L

4. The number of molecules in 35.5 g of chlorine gas is

   A $3.011 \times 10^{23}$

   B $6.023 \times 10^{23}$

   C $9.033 \times 10^{23}$

   D $1.204 \times 10^{23}$

5. The relative atomic mass of Ne is 20. Its relative molecular mass is

   A 10 C 30

   B 20 D 40

## IV. Write answers for following questions.

1. Define the term mole.

2. Calculate the relative molecular masses of the following substances:

   A $NH_3$

   B $C_8H_{18}$

   C $H_2SO_4$

   D $Cu(NO_3)_2$

3. Calculate the percentage by mass of nitrogen in the following fertilizers and nitrogen containing compounds. (Relative atomic mass H = 1, C = 12, N = 14, O = 16, P = 31)

   A $(NH_4)_2SO_4$

   B $(NH_4)_3PO_4$

   C $CO(NH_2)_2$

   D $CH_2(NH_2)COOH$

4. How much oxygen will contain the same number of atoms as the number of molecules in 73 g of HCl? ( H = 1, O = 16, Cl = 35.5)
5. Figure 2.5 shows two balloons containing oxygen gas and carbon dioxide gas respectively.

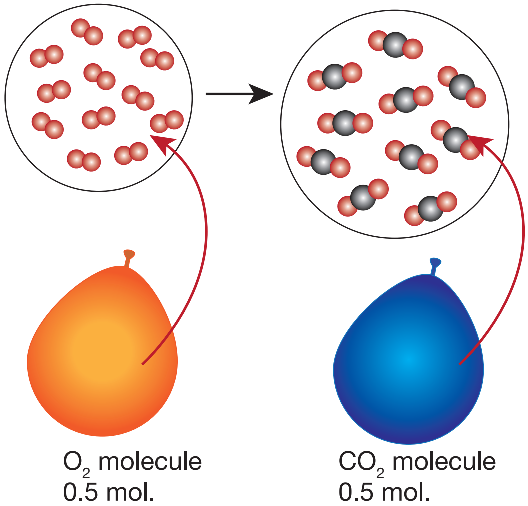



Figure 2.5

   (a) Calculate mass of oxygen gas in balloon A.

   (b) Calculate the volume of $CO_2$ gas in balloon B.

   (c) Compare and explain the number of gas molecules in balloon A and balloon B.

6. Some types of chemical reaction are listed below.

   Decomposition, neutralization, combustion, oxidation-reduction

   Which reaction type best describes the following changes.

   (a) Hexane + oxygen $\rightarrow$ carbon dioxide + water

   (b) Calcium carbonate $\rightarrow$ calcium oxide + carbon dioxide

   (c) Magnesium + copper oxide $\rightarrow$ magnesium oxide + copper

   (d) Hydrochloric acid + sodium hydroxide $\rightarrow$ sodium chloride + water

7. A crop of wheat uses up 20 kg of nitrogen per hectare of soil. What is the mass of the fertilizer calcium nitrate required to replace the nitrogen in a 10 hectare field. (N = 14, O = 16, Ca = 40)
8. 11.2 L sample of gas is determined to contain 0.5 moles of nitrogen. At the same temperature and pressure, how many moles of gas would there be in a 20 L sample?
9. How much of $KClO_3$ must be heated to produce as much $O_2$ as required to burn 24 g of carbon? (K = 39, Cl = 35.5, O = 16)

10. How many atoms of Na are present in 46 g of the metal? ($A_r$ of Na = 23)
11. How many $H_2O$ molecules are there in a snowflake weighing 1 mg?
12. Copper (II) oxide can be reduced to copper metal by heating it in a stream of hydrogen gas. Dry copper (II) oxide was placed in a weighed tube and the tube reweighed. The apparatus was then set up as shown in the Figure 2.6 below.

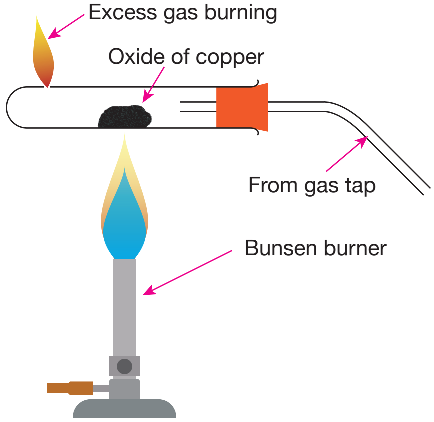



Figure 2.6

Hydrogen was passed through the tube for 15 seconds before the escaping gas was lit. The tube was heated for a few minutes. The apparatus was then allowed to cool with hydrogen still passing through. The tube was re-weighed. The process was repeated until there was no further change in mass.

(a) The results for the experiment are given in Table 2.7

**Table 2.7**

| | |
|---|---|
| Mass of empty tube | 46.12 g |
| Mass of tube + copper (II) oxide | 47.72 g |
| Mass of copper (II) oxide | ? |
| Mass of tube + copper | 47.40 g |
| Mass of copper produced | ?g |
| Mass of oxygen in the copper (II) oxide | ?g |

(b) Copy and complete the results in Table 2.7.

(c) How many moles of copper atoms are involved in the reaction? (Cu = 64)

(d) How many moles of oxygen atoms are involved in the reaction? (O = 16)

(e) From the results of the experiment how many moles of oxygen atoms will combine with one mole of copper atoms?

13. What is the volume of air required to completely burn 1 L of CO? Assume that all volumes are measured at the same temperature and pressure and also that air contains one-fifth by volume of $O_2$.
14. Determine the number of hydrogen atoms in 1.5 moles of water, $H_2O$, molecules.
15. What is the volume of oxygen required for the complete combustion of 100 L of ethane according to the following equation?

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O.$$

    Assume that all the volumes are measured at the same temperature and pressure.
16. The order of increasing relative molecular mass of the following gases is hydrogen, oxygen, carbon dioxide, sulphur dioxide, chlorine. Given 8 g of each gas at STP, which will contain the least number of molecules and which the most?
17. When vaporised, 64 g of methanol ($CH_3OH$) would occupy 44.8 L at STP. What is the vapour density of methanol?
18. 9.2 g sample of a compound contain 2.8 g of nitrogen and 6.4 g of oxygen. Find the empirical formula of the compound.
19. A metal M forms a volatile chloride containing 65.5% Cl. If the density of the chloride relative to hydrogen is 162.5, find the molecular formula of the chloride (M = 56, Cl = 35.5)
20. The empirical formula of a compound is $C_2H_5$. It has a vapour density of 29. Determine the relative molecular mass of the compound and its molecular formula.
21. 1.000 g sample of red phosphorus powder was burned in air and reacted with oxygen gas to give 2.291 g of a phosphorus oxide. Calculate the empirical formula and molecular formula of the phosphorus oxide given the molar mass is approximately 284 g $mol^{-1}$.

# 3 Metallurgy

## 3.1 Introduction

There are 118* elements at present in modern periodic table. These elements are divided into metals and non-metals. The majority of the elements are metals. In the ancient times only about eight metals namely gold, copper, silver, tin, iron, lead, mercury and arsenic were known to the people. Today, scientists have identified about ninety metals. These metals either exist in their free state or in the combined state in nature. Learning the art of obtaining metals from the naturally occurring minerals was a big step for civilization. Several physical and chemical methods are used to extract metals depending upon their position in metal activity series and nature of the ore. With the knowledge of science we have come a long way in learning the skill of extracting metals from their respective mineral ores. This chapter introduces the various processes involved in extraction of metals and their uses.

## 3.2 Metallurgy

**Learning objectives**

On completion of this topic, students should be able to:

» *define metallurgy.*

» *name the ores of some common metals.*

» *relate how the reactivity of a metal affects its extraction from the naturally occurring ores.*

» *outline the extraction process of metal with brief explanation.*

» *explain the uses of some common metals and their alloys.*

*On 30$^{th}$ December, 2015, IUPAC has announced the verification of the discoveries of four new chemical elements, thereby completing the 7$^{th}$ period of the periodic table. The element 113, 115, 117 and 118 were given the temporary working name and symbol as ununtrium (Uut), ununpentium (Uup), ununseptium (Uus) and ununoctium (Uuo) respectively.

The study of metal and the various processes involved in the extraction of metals from their respective ores are called metallurgy. The methods of extraction depend upon the chemical nature of the sources from which the metals are obtained.

## 3.2.1 Occurrence of metals

The surface of the earth's crust is made up of sand, silicates, metal compounds (minerals) and some free metals. Very few metals like gold, platinum, silver and mercury exist in the free or native state. The rest of the metals occur in the combined form as oxides, carbonates, sulphides, sulphates, silicates, halides, nitrates, phosphates, etc. Metals at the top of the activity series (K, Na, Ca, Mg and Al) are so reactive that they are never found in nature as free elements. Copper and silver are two metals which occur in free state as well as in the combined state as sulphide, oxide or halide. Various processes have to be employed in order to obtain different metals from their respective ores.

## 3.2.2 Some terminologies used in metallurgy.

### i. Minerals and ores

Metals and their compounds are found in earth as natural substance known as minerals. Minerals are naturally occurring chemical substances in which the metals and their compounds are found in the earth's crust. A metal may be present in the form of several minerals. Some minerals contain large amount of metal while other contains only traces of metal. Most of the minerals are associated with impurities but some contain less impurity while others contain more impurities which make the extraction of metal from them difficult. Thus, all the minerals are not used for the extraction of metals.

Those minerals from which the desired metals are extracted are called ores. An ore contains higher percentage of metals with lesser amount of impurities. For example, both haematite and iron pyrites are the minerals of iron. Haematite contains higher percentage of iron than iron pyrite. So haematite is used as an ore to extract  iron and iron pyrite due to its high sulphur content is used as raw material for the manufacture of sulphuric acid. All the ores are minerals, but all minerals are not necessarily ores. Table 3.1 shows examples of some types of ores.

Table 3.1 Different types of ores

| Oxides ores | Carbonate ores | Halide ores | Sulphide ores |
|---|---|---|---|
| Bauxite [$Al_2O_3.2H_2O$] | Calamine [$ZnCO_3$] | Cryolite [$Na_3AlF_6$] | Cinnabar [$HgS$] |
| Cuprite [$Cu_2O$] | Marble [$CaCO_3$] | Fluorspar [$CaF_2$] | Galena [$PbS$] |
| Haematite [$Fe_2O_3$] | Magnesite [$MgCO_3$] | Horn silver [$AgCl$] | Chalcocite [$Cu_2S$] |
| Zincite [$ZnO$] | Siderite [$FeCO_3$] | Rock salt [$NaCl$] | Zinc blende [$ZnS$] |
| Magnetite ($Fe_3O_4$) | Dolomite ($CaCO_3.MgCO_3$) | | |

### ii. Charge

The mixture of materials fed into a furnace, to extract a metal is called charge.

### iii. Gangue

The impurities like sand, rocky materials, earthy particles etc. associated with an ore are called gangue. They are also called matrix.

### iv. Flux

It is a chemical substance added to an ore during the extraction of metal that combines chemically with the gangue to form a fusible light mass.

### v. Slag

The product obtained by the combination of gangue with the flux is called slag.

$$\text{Gangue} + \text{Flux} \longrightarrow \text{Slag}$$

## 3.2.3 Processes involved in the extraction of metals.

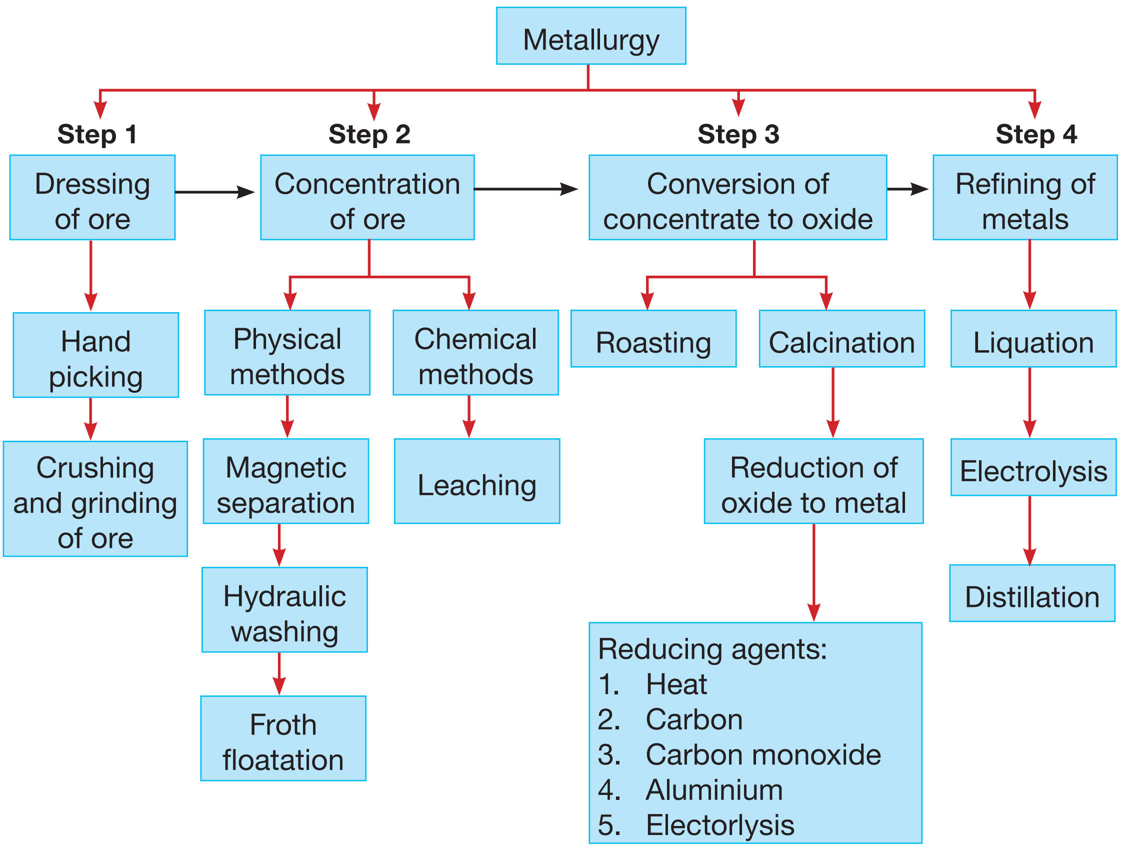



Figure 3.1 Overview of the metal extraction process.

### i. Dressing of ores (Step 1)

Generally, the natural ores contains some amount of gangue. The gangue is to be removed before the metal is extracted from an ore. The conversion of ores into their physical form from which the gangue can be easily removed is called dressing of ores. It is also called enrichment. The presence of gangue in the ores

- interfere with the process of extraction.
- make the metal impure.
- increase the cost of production of metal.

Hence, all efforts have to be made to remove the gangue from an ore.

The dressing of ore involves the following stages:

**a. Hand picking**

Large gangue especially the stony materials are removed manually from the ores by hand picking.

**b. Crushing and grinding of the ore**

Most of the ores found in the nature occur as huge lumps. They have to be broken into small pieces to make the extraction convenient. These large lumps are broken into smaller pieces with the help of jaw crushers or grinders. These pieces are then reduced to fine powder with the help of a stamp mill or ball mill. This process is called pulverization of the ore.

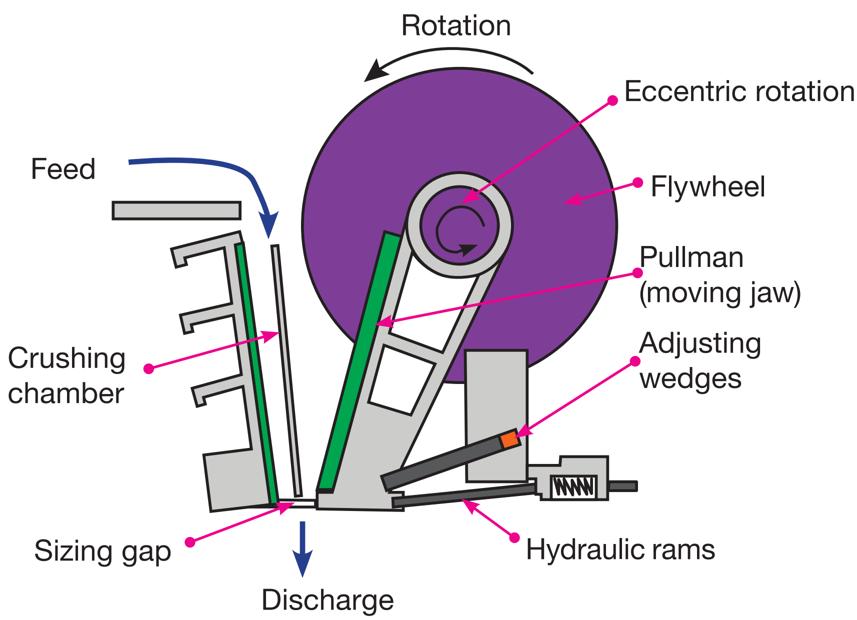



Figure 3.2 Jaw crusher.

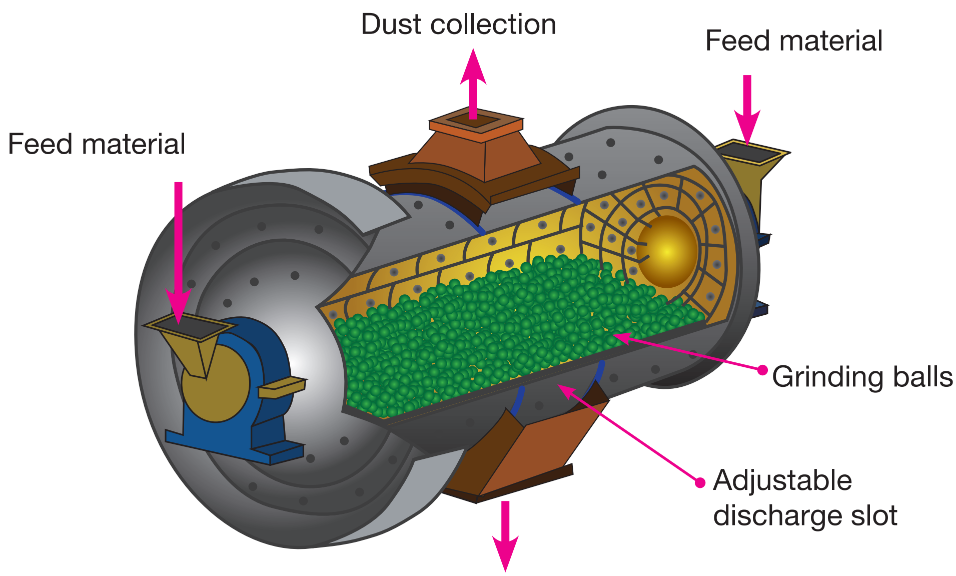


Figure 3.3 The ball mill.

## ii. Concentration of ores (Step 2)

The process of removing the gangue from the dressed ore is called concentration of ore. This is done by using the current of air or water depending on the nature of ore and the impurities present in it. The purified ore is called concentrated ore. Some important methods used in concentration of ores are:

### a. Froth floatation process.

This method is used generally for the concentration of sulphide ores. It is based on the preferential wetting properties of the ore with the frothing agent and water. When powdered ore is added to water containing pine oil (frothing agent), water wets the gangue particles and the pine oil wets sulphide ore particles.

Compressed air is passed through the mixture containing water, oil and crushed ore. Froth is formed in the oil layer and the ore particles embedded in the froth rise up with it. The froth is collected and pressed to remove the adhering oil from the ore. The gangue particles are collected at the bottom of the tank.

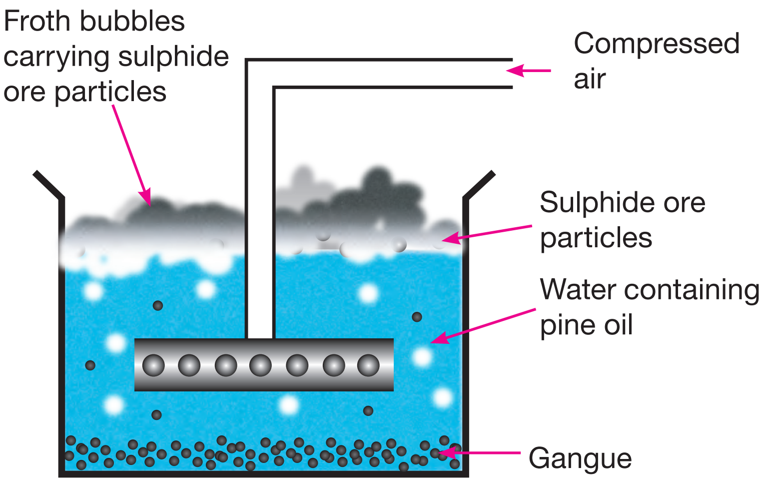



Figure 3.4 Froth floatation.

**b. Gravity separation or levitation.**

This process of concentrating ore is based on the difference in the specific gravities of the ore and gangue particles. The ore particles are generally heavier than the impurities, so the muddy particles can be removed by washing with water. When the powdered ore is washed with an upward stream of running water in a hydraulic classifier, the lighter gangue particles are washed away while heavier ore particles settle down and are collected from the base. This method is called levitation.

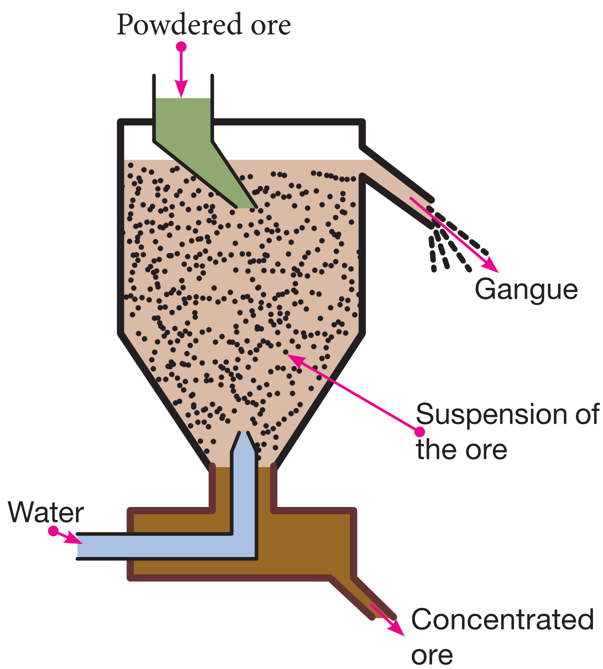



Figure 3.5 Hydraulic classifier.

The washing of powdered ore with water can also be carried out by using Wilfley washing table. It is a wooden table fitted with ridges and is placed in slanting position. The crushed ore is spread on the table and stream of water is passed over the vibrating table. The heavier ore particles collect in the grooves (in between the ridges) and the lighter impurities are washed away by the current of water.

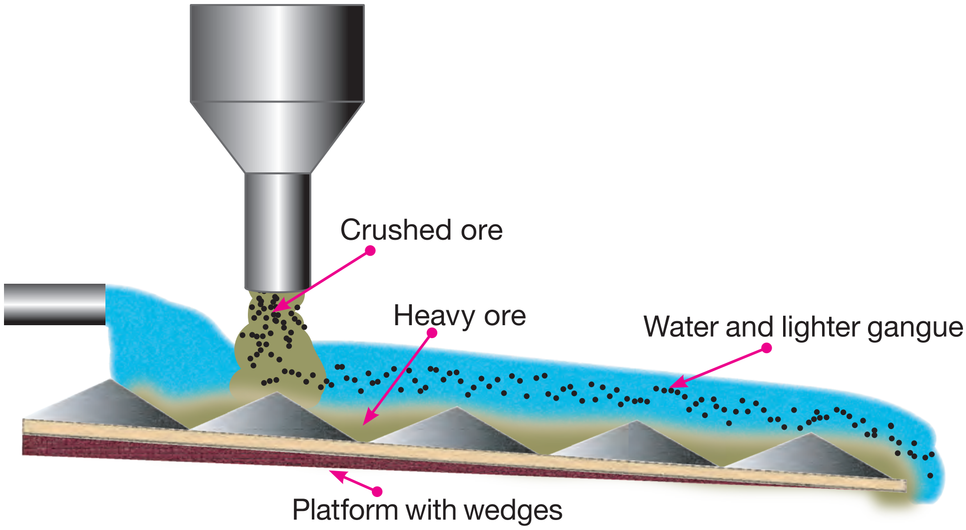



Figure 3.6 Wilfley washing table.

This method is suitable for concentration of heavy oxide and carbonate ores like cassiterite ($SnO_2$) and haematite ($Fe_2O_3$).

### c. Magnetic separation.

This method is based on the difference in magnetic properties of the ore and gangue. The magnetic separator consists of a rubber belt moving over two rollers, one magnetic and the other nonmagnetic.

The powdered ore is placed on a conveyer belt running on magnetic wheel. While falling from the conveyer belt at the other end having a magnetic roller, the magnetic particles are drawn in and the non-magnetic particles are thrown away. The magnetic and non-magnetic particles form separate heaps as shown in the Figure 3.7.

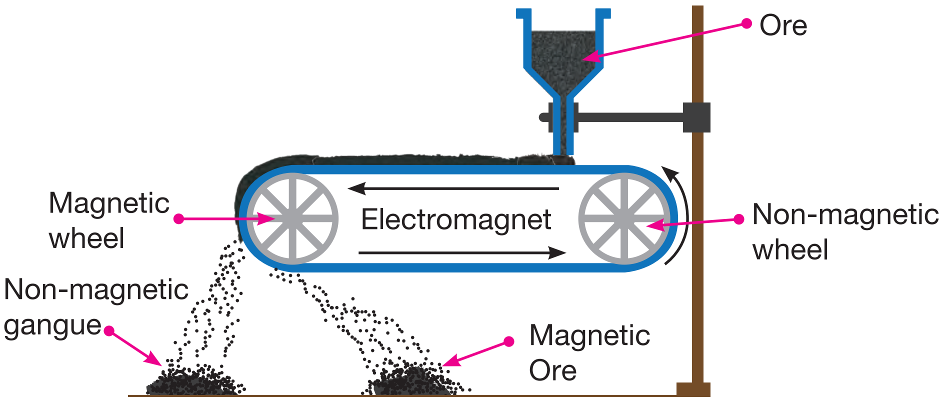



Figure 3.7 Magnetic separator

Magnetic separation is used for separating magnetite ($Fe_3O_4$) from its non-magnetic impurities and also for separating tin stone ($SnO_2$) from its magnetic impurities such as wolframite (tungstates of iron and manganese).

**d. Leaching.**

This is a chemical method of concentrating an ore. The desired metal compound present in the ore is leached, i.e., dissolved out, by a suitable reagent and the gangue is left behind. The desired compound is then obtained or precipitated from the solution by a suitable method. This method is used for concentration of bauxite ($Al_2O_3.2H_2O$).

### iii. Extraction of metal from the concentrated ore (Step 3).

Being electropositive, a metal loses electron(s) to combine with other elements and is thereby oxidized. Hence, a metal present in a combined state in an ore has to be reduced.

**a. Conversion of ores into metal oxides.**

The form in which a metal occurs in an ore may or may not be suitable for reduction. For example, the oxide of a metal is more suitable for reduction than its sulphide, hydroxide, hydrated oxide or carbonate form. To convert the ore into a suitable form for reduction, the ore is often calcined or roasted.

**b. Calcination.**

It is the process of heating the concentrated ore in a limited supply of air or in absence of air at a temperature below its fusion point. This is done in order to remove moisture, to convert carbonate ores into oxides or to remove any volatile impurities like sulphur and phosphorus. For example, carbonate ores get converted into oxide as shown in the equation below.

$$CuCO_3 \xrightarrow{\text{Heat}} CuO + CO_2 \uparrow$$

**c. Roasting.**

It is the process of heating the concentrated ore in excess supply of air at a temperature below its fusion point. This process is generally carried out to convert sulphide ores into their metallic oxides. Roasting can also remove moisture and volatile impurities like sulphur, arsenic and phorphorus. For example, conversion of sulphide ores into oxides is shown in the equation below:

$$2PbS + 3O_2 \xrightarrow{\text{Heat}} 2PbO + 2SO_2 \uparrow$$

## d. Reduction to metal.

The type of reducing agents and the process employed depends upon the nature of the metal in the ore. The calcined or roasted ore is reduced to metal by the following methods:

### (i) The carbon reduction process – Pyrometallurgy

This is the cheapest and by far the most common method of producing metals on a large scale. The oxides of metal below aluminium in the activity series can be reduced to their respective metals on heating with carbon (C). This process of heating metal oxide in the presence of carbon is also known as smelting. There is a great possibility of formation of carbon monoxide (CO) as a result of carbon reacting with the metal oxide or air. Carbon monoxide acts as a reducing agent and is a stronger reducing agent than carbon. For example, reduction of oxides ore to metal with carbon.

$$FeO + C \xrightarrow{\text{Heat}} Fe + CO \uparrow$$
$$Fe_2O_3 + 3C \xrightarrow{\text{Heat}} 2Fe + 3CO \uparrow$$

Reduction of oxides ore to metal by CO

$$Fe_2O_3 + 3CO \xrightarrow{\text{Heat}} 2Fe + 3CO_2 \uparrow$$
$$CuO + CO \xrightarrow{\text{Heat}} Cu + CO_2 \uparrow$$

These reactions are carried out at high temperature and as a result the metal formed by reduction melt down.

Although the majority of gangue is removed during the concentration of ore, some impurities are still left in it. Such impurities are also reduced during smelting and removed from the ore. In order to remove the remaining gangue, a substance known as flux is added. The impurities present in the ore combine with the flux to form slag. Slag being lighter, form the upper layer while molten metal form the lower layer.

**Zinc Oxide**

Zinc oxide is reduced only by carbon and is not reduced by hydrogen and carbon monoxide. It is due to the reactivity of zinc oxide which can be reduced by carbon monoxide and hydrogen only at very high temperature and pressure.

$$ZnO + C \xrightarrow{\text{Heat}} Zn + CO \uparrow$$

### Activity 3.1 Reduction of copper oxide (CuO) by carbon

#### Materials required

Test tube, cork or rubber bung, clamp and stand, Bunsen burner/spirit lamp, delivery tube, test tube, spatula, copper oxide, charcoal, beaker and lime water.

#### Procedure

1. Transfer one spatula of CuO into a clean glass test tube.
2. Carefully add one spatula of charcoal powder on top of copper oxide without any mixing.
3. Strongly heat the content in a test tube in a slanting position for five minutes on a spirit lamp or Bunsen flame.
4. Pass the evolving gas through a delivery tube fixed with cork or rubber bung into the lime water (as shown in Figure 3.8).
5. Allow it to cool and then observe the change that takes place at the junction of two layers.
6. Record your observation.

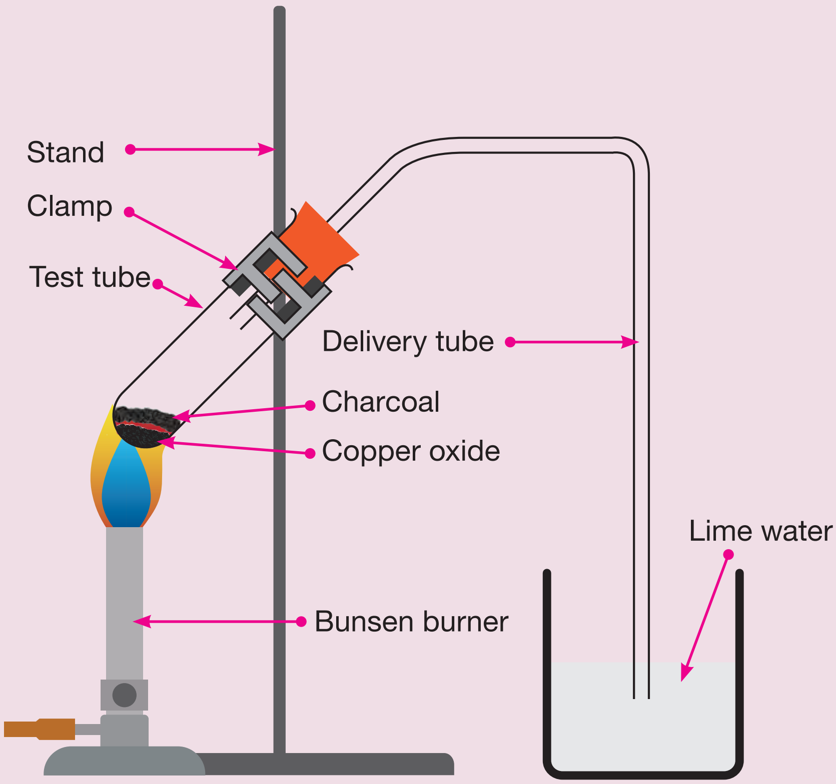



Figure 3.8 Reduction of copper oxide with carbon

**Question**

1. What is observed when CuO and charcoal is heated in a test tube?
2. Write the balanced chemical equation that takes place between charcoal and CuO.
3. What change if any is observed when the evolved gas is passed through the lime water? Why?

### (ii) Reduction with aluminium - Aluminothermy

Oxides of some metals like manganese, chromium, iron and tungsten which are less active than aluminium are reduced by heating them with aluminium powder. A mixture of metal oxide and aluminum powder, called thermite is ignited with a magnesium wire embedded in magnesium powder and barium oxide ($BaO_2$) mixture. Due to aluminum's strong affinity towards oxygen, it acts as reducing agent. This process is called aluminothermy.

$$\underset{\text{(thermite)}}{Cr_2O_3 + 2Al} \xrightarrow{\text{Heat}} Al_2O_3 + 2Cr + \text{Heat}$$

It is carried out in a clay crucible which can withstand very high temperature. The reaction is highly exothermic and heat produced is sufficient to melt the metal. This molten metal obtained can be used for welding purpose and the process is called thermite welding.

### (iii) Auto-reduction

In certain cases, no reducing agents are required. The ore is roasted and a portion of the ore reacts with the rest of the ore to give the metal.

$$\underset{\text{(unreacted ore)}}{Cu_2S} + \underset{\text{(roasted ore)}}{2Cu_2O} \xrightarrow{\text{Heat}} 6Cu + SO_2$$

### (iv) Electrolytic reduction

Metals are arranged in a vertical column in the decreasing order of their reactivity called metal activity series. The extraction process of highly reactive metals depends on the position in the activity series. They are generally extracted by electrolytic reduction.

| | | | |
|---|---|---|---|
| **Highly reactive** | Li | Lithium | Extraction by electrolysis |
| | K | Potassium | |
| | Ba | Barium | |
| | Na | Sodium | |
| | Ca | Calcium | Expensive |
| | Mg | Magnesium | |
| | Al | Aluminum | |
| | **C** | **Carbon** | |
| | Zn | Zinc | Extraction by reduction of metal oxide with C or CO |
| | Fe | Iron | |
| | Ni | Nickel | |
| | Sn | Tin | |
| | Pb | Lead | |
| | **H** | **Hydrogen** | Inexpensive |
| | Cu | Copper | |
| | Hg | Mercury | |
| | Ag | Silver | |
| | Au | Gold | |
| **Least reactive** | Pt | Platinum | |

Figure 3.9 Metal activity series.

Electrolytic reduction is used for extraction of highly active metals like potassium, sodium, calcium, magnesium, aluminum etc. These metals cannot be obtained by reduction of their oxide because the metal formed will immediately combine with the carbon due to high temperature to form the metal carbide. Therefore, electrolysis is carried out to extract these metals using suitable electrodes. In fact, electrolytic reduction can be used for any metal, but it is costlier than the carbon-reduction method.

## iv. Purification or Refining of metal (Step 4).

### a. Liquation.

This method of refining is used for metals with low melting point like lead and tin. The impure metal is placed on the sloping hearth of a furnace and heated gently. As the metal melts and flows down, the impurities called dross is left on the hearth of the furnace as shown in the Figure 3.10.

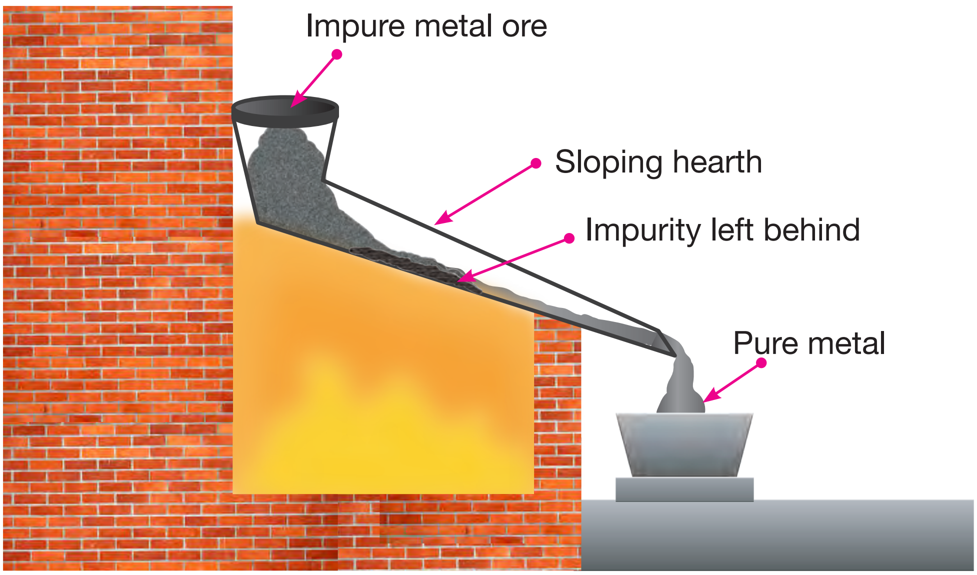



Figure 3.10 Purification of metal by liquation.

**b. Distillation method**

Metals like zinc, cadmium and mercury are purified by distillation. In this method, the impure metal is heated in a vessel. The metal forms vapours which condense separately in a receiver and the non-volatile impurities are left behind in the first vessel itself.

**c. Electrolytic refining**

The metals like copper, silver and gold are purified by electrolytic method. The block of the impure metal is made the anode and the thin sheet of pure metal is made as cathode. A suitable metal salt solution is chosen as the electrolytic solution. On passing the electric current, the metal ions from the electrolyte are deposited at the cathode in the form of pure metal. At the same time, an equivalent amount of the metal dissolves from the anode and passes into the electrolyte in the form of metal ions. The impurities either go into the solution or settle below the anode as anode mud as shown in the Figure 3.11.

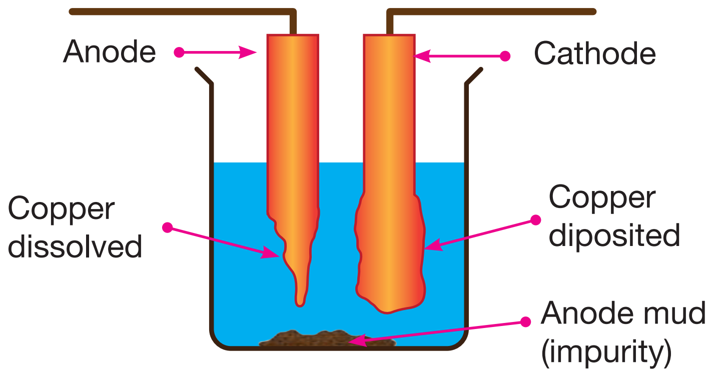



Figure 3.11 Electrolytic refining.

### d. Oxidative refining

This method is used where the impurities can be easily oxidized than the metal itself. When air is passed through the impure molten metal, the impurities like carbon, phosphorus, arsenic etc get oxidized to their volatile oxides. The pure metal is left behind. Impure iron known as the cast iron which contains carbon, phosphorus, silicon and manganese as impurities is refined by this method. When air is passed through cast iron, these impurities are oxidized to their oxides such as $CO_2$, $P_2O_5$, $SiO_2$, etc and are removed.

**Activity 3.2 Designing metal extraction process.**

**Instruction**

Design a schematic flow chart to show the entire process involved in extraction of the metal.

**Self Evaluation**

1. **The elements A, B, C, D, E, F, G, H, I, J, K and L represent metals in the decreasing order of reactivity. Which one of them is most likely to**
   (a) occur in the free state?
   (b) be highly reactive?
   (c) be least reactive?
   (d) occur as salt of fluorides and chlorides?
   (e) occur as oxides and sulphides?
2. **Give reason for the following:**
   (a) Metals in the middle of metal activity series cannot be obtained just

by heating their ores in air.

(b) In electrolytic refining impure metal is always assigned as an anode.

(c) Hydrogen is not a metal but it has been assigned a place in the reactivity series of metals.

(d) Metals at the top of activity series, like potassium, sodium, magnesium and aluminium cannot be obtained by the reduction of their oxides with carbon, carbon monoxide and hydrogen.

(e) Zinc oxide cannot be reduced by hydrogen and is reduced only by carbon.

**3. What is the difference between a mineral and an ore?**

**4. What are the methods used to concentrate an ore?**

**5. What is the purpose of the following processes?**

(a) Calcination

(b) Roasting

(c) Magnetic separation

(d) Liquation.

**6. When is the reducing agents used in the process of metal extraction? Write down the purpose of using it.**

## 3.3 Electrolysis

**Learning objectives**

On completion of this topic, students should be able to:

» *define electrolyte and non-electrolytes.*

» *explain electrolysis with simple illustration.*

» *differentiate cathodic reduction and anodic oxidation.*

» *investigate the electrolysis of copper sulphate solution.*

Some substances allow an electric current to pass through them and some substances do not. Those substances which allow an electric current to pass through them are called conductors and which do not allow are called non-conductors. When the electric current is passed through the substances in their molten state or in aqueous solution, the chemical compounds undergo decomposition. The process of decomposition of a chemical compound in its molten state or in aqueous solution by the passage of electricity through it, is termed as electrolysis.

Electrolysis finds its application largely in chemical industries and metal extraction process. It is used for electro refining, electroplating, electroforming and extraction of highly reactive metals like sodium from its ore in the molten state.

## 3.3.1 Types of conductors

### i. Metallic conductors

Metallic conductors conduct electricity without undergoing any chemical change. The conduction of electric current in them is due to presence of mobile free electrons. All metals are examples of metallic conductors except the carbon which in the form of graphite is a non-metallic conductor.

### ii. Electrolytic conductor

Electrolytic conductors are those substances which conduct electric current either in their molten state or in aqueous solution. The ions present in the electrolytes become free when they are in molten form or in aqueous solution. These free ions are responsible for the conduction of electric current. Some of the examples of electrolytes are NaCl, HCl, $CuSO_4$, NaOH etc.

There are substances which do not conduct electricity in aqueous solutions or in molten state. They are called non-electrolytes. Examples of non-electrolytes are pure water, alcohol, kerosene, urea, benzene, paraffin wax, carbon tetrachloride, sugar solution etc.

## 3.3.2 Electrolytic cell or Voltameter.

The vessel or an apparatus in which the electrolysis is carried out is known as Electrolytic cell. It is made up of insulating materials like glass. An electrolytic cell consists of the following components:

### i. Electrolyte.

An electrolyte is a substance which in aqueous solution or in molten state conducts electricity and is decomposed by the passing electric current. When electric current is passed through an electrolyte in the electrolytic cell, the electrolyte decomposes into ions. Ion is an electrically charged atom or the group of atoms. Ions carrying positive charge are called cations and ions carrying negative charge are called anions.

Electrolytes can be classified as strong electrolyte and weak electrolyte.

**a. Strong electrolyte.**

Electrolytes which are almost completely dissociated into ions in aqueous solution and allow more electric current to flow through them are known as strong electrolytes. Some of the examples are strong bases like NaOH and KOH, strong acids like HCl, $HNO_3$ and $H_2SO_4$ and salts prepared from strong acids and bases like NaCl, $KNO_3$, $Na_2SO_4$, etc.

**b. Weak electrolyte.**

Electrolytes which are dissociated into ions in aqueous solution only to a small extent and allow less electric current to flow through them are known as weak electrolytes. Some examples are weak bases like $NH_4OH$, $NH_3$ and its derivatives and weak acids like $H_2CO_3$, $H_3PO_4$, $CH_3COOH$, HCOOH etc and salts prepared from weak acids and bases like $NH_4NO_3$, $(NH_4)_2CO_3$ , etc.

## ii. Electrodes.

Electrodes are the graphite or metal plates or rods through which the electric current enters or leaves the electrolyte. The electrodes are dipped in the electrolyte during electrolysis. There are two types of electrodes namely;

a) Anode or positive pole is the electrode connected to the positive terminal of the battery. The current enters an electrolyte through anode.

b) Cathode or negative pole is the electrode connected to the negative terminal of the battery. The current leaves an electrolyte through cathode.

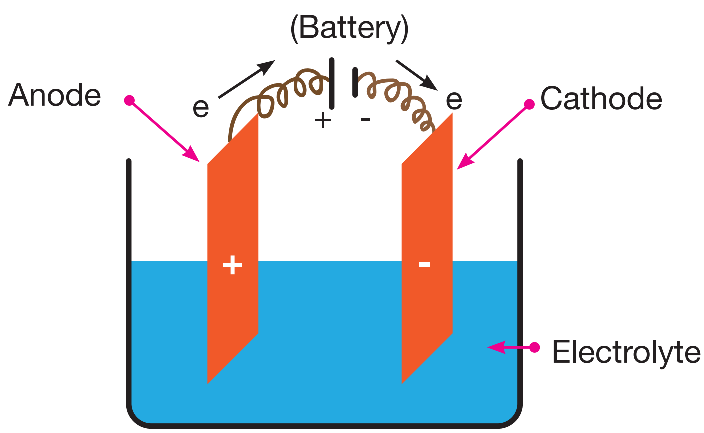



Figure 3.12 Electrolytic cell

### 3.3.3 Electron transfer process – oxidation and reduction.

A chemical reaction occurs by simultaneous loss of electron by an atom, ion or a molecule and gain of electron by the other atom, ion or a molecule. According to electronic concept the chemical reaction which involves loss of electron is called oxidation and the one which involves gain of electron is called reduction.

A chemical reaction which involves simultaneous oxidation and reduction reaction is known as redox reaction. For example, loss of two electrons by the magnesium atom is the oxidation reaction as shown in the equation.

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

And the gain of an electron by chlorine atom is the reduction reaction as shown in the equation.

$$Cl + e^- \longrightarrow Cl^-$$

The reaction between magnesium and chlorine to form $MgCl_2$ is an example of a redox reaction. It involves the transfer of electrons from magnesium to chlorine, in which magnesium is oxidized and chlorine is reduced.

$$Mg + Cl_2 \longrightarrow Mg^{2+} + 2Cl^- \longrightarrow MgCl_2$$

The substance which oxidizes other substance is called oxidizing agent or oxidant. The substance which reduces other substance is called reducing agent or reductant. The total number of electrons lost by the reductant is always equal to the number of electrons accepted by the oxidants.

In the reaction between Mg and $Cl_2$ to form $MgCl_2$

$$Mg \longrightarrow Mg^{2+} + 2e^- \quad \text{(Oxidation)}$$

Mg is oxidized to $Mg^{2+}$ ion and Mg acts as a reducing agent.

$$Cl_2 + 2e^- \longrightarrow 2Cl^- \quad \text{(Reduction)}$$

Cl is reduced to chloride ion and chlorine acts as an oxidizing agent.

### 3.3.4 Dissociation or ionization of the electrolyte

The dissociation of an electrolyte is explained by ionic theory, which was put forward by a Swedish Chemist Avante Arrhenius (1887). It states that when an electrolyte is dissolved in water or in fused state, it dissociates into charged particles called ions (cations and anions). These ions have tendency to reunite to form unionized electrolyte.

$$\underset{\text{(electrolyte)}}{M^{n+}A^{n-}} \longleftrightarrow \underset{\text{(cation)}}{M^{n+}} + \underset{\text{(anion)}}{A^{n-}}$$

Where, M is a metal or hydrogen and A is a non-metal or sometimes a radical like sulphate, hydroxyl, etc. When the ions enter the solution, they are surrounded by water molecules. This process is called hydration and the hydrated ions move freely in solution which is responsible for the conductivity of the solution. An electrolyte is electrically neutral. Thus, in an electrolyte, the total positive charge is always equal to the total negative charge.

### 3.3.5 Discharge of ions at the electrodes.

On passing the electric current through the electrolyte, cations being positively charged ion move towards the cathode, and the anions, being negatively charged move towards the anode. The cations take up electron(s) from the cathode and are discharged (i.e., lose their charge) there. The anions, in contrast, give up electron(s) to the anode and are discharged there.

At cathode $M^{n+} + ne^- \longrightarrow M$ Cathodic reaction (reduction half reaction)

At anode $A^{n-} \longrightarrow A + ne^-$ Anodic reaction (oxidation half reaction)

The number of electron(s) taken up at cathode must be equal to that given up at the anode. As the electrons are taken up by the cations at the cathode, the reaction occurring at this electrode is called cathodic reduction. And since electrons are given up by the anions at the anode, this reaction is called anodic oxidation.

Metal atoms are deposited at the cathode, but hydrogen atoms, if discharged, combine to form $H_2$ molecules. Similarly, gaseous elements like oxygen, chlorine or bromine are discharged at the anode as the molecules $O_2$, $Cl_2$ or $Br_2$ respectively.

**Preferential discharge of ions from aqueous solution at electrodes.**

The product of the electrolysis of an aqueous solution of an electrolyte may be different from those obtained from molten salt. For example, the electrolysis of molten NaCl produces Na at the cathode and on the other hand electrolysis of its aqueous solution produces $H_2$ gas. This is due to the simultaneous dissociation of water, though to very small extent, into $H_3O^+$or $H^+$ and $OH^-$ ions.

$$H_2O \leftrightarrow H^+ + OH^-$$

$$H_2O + H^+ \longrightarrow H_3O^+$$

So in the electrolytic cell, there are more than one ions competing for discharge at each electrode. For example,

- the metal ion ($M^{n+}$) of salt and the $H^+$ or $H_3O^+$ of water ion at the cathode.
- the anion of salt and the $OH^-$ of water at the anode.

Only one ion will be discharged in preference to the other at each electrode which depends on the following factors:

### i. Relative position of the ion in the electrochemical series.

In a electrochemical series, the cations and the anions are arranged in decreasing order of their reactivity. The elements from the top of the electrochemical series ionize most readily but get discharged at the electrode with greater difficulty. The elements from the bottom of the electrochemical series ionize less readily but get discharged at the electrode most easily. So, the ease of discharge of ions at the electrode in the series increases down the electrochemical series.

**At the Cathode**: If the metal 'M' is more reactive than the hydrogen which lie at the higher position than H in the electrochemical series, the metal will remain in the ionic form ($Mn^+$) and stay at the electrode. $H^+$ instead gets discharged and liberated as $H_2$.

$$2H^+ + 2e^- \longrightarrow 2H \longrightarrow H_2$$

$$2H_3O^+ + 2e^- \longrightarrow H_2 + 2H_2O$$

Thus, the cations lying above $H^+$ in the electrochemical series will not get discharged in preference to $H^+$. On the other hand, if M is less reactive than H (lying below H in the electrochemical series) then $M^{n+}$ ions get discharged in preference to $H^+$. The $H_3O^+$ (or $H^+$) remains in the ionic form in the solution.

$$M^{n+} + ne^- \longrightarrow M$$

So, the ions like $Cu^{2+}$ and $Ag^+$ will be discharged at cathode in preference to $H_3O^+$ (or $H^+$) ions. As the reactivity of cation decreases, the ease with which it accepts electrons increases.

| Cation | Ease of discharge increases ↓ | Anion |
|---|---|---|
| $K^+$ | | $SO_4^{2-}$ |
| $Na^+$ | | $NO_3^-$ |
| $Ca^{2+}$ | | **$OH^-$** |
| $Mg^{2+}$ | | $Cl^-$ |
| $Al^{3+}$ | | $Br^-$ |
| $Zn^{2+}$ | | $I^-$ |
| $Fe^{2+}$ | | |
| $Sn^{2+}$ | | |
| $Pb^{2+}$ | | |
| **$H^+$** | | |
| $Cu^{2+}$ | | |
| $Ag^+$ | | |

Figure 3.13 Ease of discharge of ions

**At the Anode**: At anode the competing ions are the anion of the salt and $OH^-$ ions of water for discharge. Like cations, the anions can also be arranged in order of ease of discharge. This arrangement is called the electronegative series of non-metallic species. The ease of ion that gets discharged at anode increases down the electronegative series. As the reactivity of anion decreases, the ease with which it loses electron increases.

Thus, $Cl^-$, $Br^-$ and $I^-$ are discharged in preference to $OH^-$ and the products formed at anode are $Cl_2$, $Br_2$ and $I_2$. The $OH^-$ ions are discharged in preference to $SO_4^{2-}$, $NO_3^-$ ions, and product is $O_2$.

### ii. Concentration of the ions in the electrolyte.

The products obtained at electrode is also dependent on the concentration of ions in the electrolyte. In many of the cases it is observed that greater the concentration of the ion, greater is its probability to get discharged at the electrode. For example, the electrolysis of very dilute solution of NaCl produces $H_2$ at the cathode as expected, but at the anode, $OH^-$ ions get discharged instead of $Cl^-$ ions, evolving $O_2$ gas. This is due to higher concentration of $OH^-$ ions than $Cl^-$ ions in the solution.

$$OH^- \longrightarrow OH + e^-$$
$$4OH \longrightarrow O_2 + 2H_2O$$

### iii. Nature of the electrode:

The nature of the electrode also influences the product formed at the electrodes. For example, when the concentrated solution of NaCl is electrolysed using platinum electrodes, $H^+$ ions get discharged at the cathode in preference to $Na^+$ ions. However, when mercury is used as cathode, $Na^+$ ions get discharged at the cathode and $Cl^-$ ions at the anode. Likewise, if active electrodes like, copper and zinc are used, they react with the electrolyte giving different products as compared to the usual product obtained using inert electrode like platinum and graphite.

## 3.3.6 Electrolysis of concentrated sodium chloride solution.

A concentrated solution of NaCl can be electrolysed in the laboratory. When electric current is passed through the solution, it results into four different ions. The two cations, $Na^+$ and $H^+$ move towards cathode and the two anions, $Cl^-$ and $OH^-$ move towards anode. The electrodes used here must be inert electrode like graphite or platinum, which do not react with the electrolytic solution.

$$NaCl \xrightarrow{\text{Electrical energy}} Na^+ + Cl^-$$
$$H_2O \xrightarrow{\text{Electrical energy}} H^+ + OH^-$$

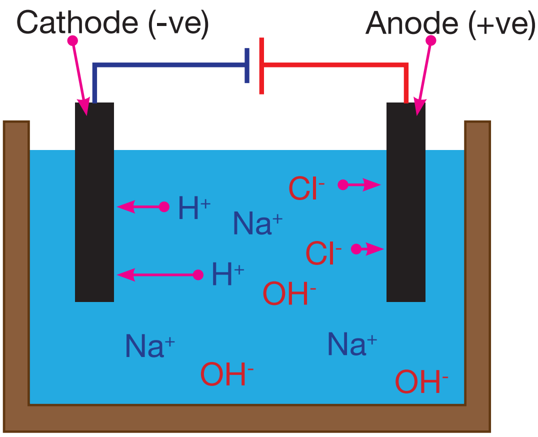



Figure 3.14 Electrolysis of NaCl solution.

### At cathode

$H^+$ ions get discharged in preference to $Na^+$ as it is less reactive than sodium and being lower in the electrochemical series. So, the hydrogen gas bubbles off at the cathode.

### At anode

$Cl^-$ ions are discharged more readily than $OH^-$ ions for chlorine being lower in the electronegative series. In anode, the pale green chlorine gas bubbles off.

| Cathodic reaction | Anodic reaction |
|---|---|
| $H^+$ ion accepts one electron to form an atom of hydrogen<br>$H^+ + e^- \longrightarrow H$<br>Two hydrogen atoms combine to form $H_2$<br>$H + H \longrightarrow H_2$<br>Overall reaction<br>$2H^+ + 2e^- \longrightarrow H_2$ | $Cl^-$ ion loses one electron to form an atom of chlorine.<br>$Cl^- \longrightarrow Cl + e^-$<br>Two chlorine atoms combine to form $Cl_2$<br>$Cl + Cl \longrightarrow Cl_2$<br>Overall reaction<br>$2Cl^- \longrightarrow Cl_2 + 2e^-$ |

**Activity 3.3 Investigating the electrolysis of $CuSO_4$ solution using copper cathode and graphite anode.**

**Materials required**

Water trough or large beaker, copper plates and graphite rod as electrodes, copper wires, batteries, spatula, glass rod, ammeter/torch bulb, $CuSO_4$ and water.

## Procedure

1. Dissolve a spatula of crystals of $CuSO_4$ (blue vitriol) in about 500 mL water.
2. Add about 5 - 6 mL of concentrated $H_2SO_4$ to make the solution clear.
3. Pour the solution in the electrolytic cell or large beaker up to $\frac{3}{4}$ of its volume.
4. Complete the circuit using copper plate as cathode and graphite rod as anode. Connect torch bulb or ammeter within the circuit to ensure that the circuit is complete as shown in the Figure 3.15.
5. Switch on an electric current using a 4.5 V, 6 V, 9 V or 12 V batteries.
6. Observe the changes.

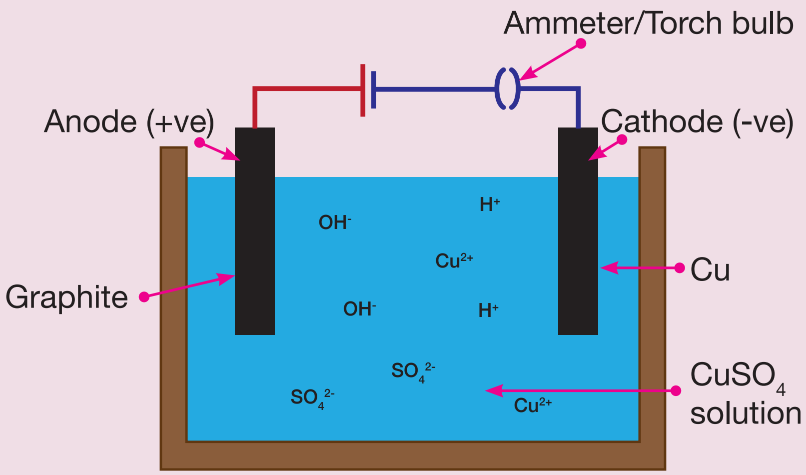



Figure 3.15 Electrolysis of $CuSO_4$ solution

## Questions

1. What happens to the colour of the solution? Explain your observation.
2. What is observed at the cathode?
3. Identify reduction electrode and oxidation electrode.
4. Write down the cathodic reaction and anodic reaction.
5. What is observed as the voltage of the batteries increase? Give reason.
6. Write the equations showing the dissociation of electrolyte.

### Self Evaluation

**1. Fill in the blanks in the following table:**

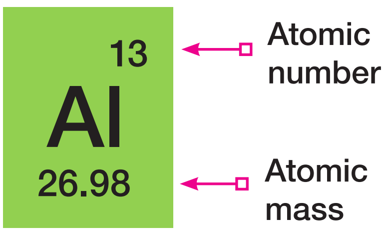
| Electrolyte | Cathode (graphite) | | Anode (graphite) | |
|---|---|---|---|---|
| | Reaction | Product | Reaction | Product |
| $CuSO_4$ (aq) | $2Cu^{2+} + 4e^- \rightarrow 2Cu$ | Cu | | |
| $AgBr_2$ (aq) | | | $Br^- \rightarrow 2Br + 2e^-$ | |

**2. What kind of a compounds act as an electrolyte? Explain with an example.**

**3. Classify the following substances into strong electrolytes, weak electrolyte and non-electrolyte.**

(a) Acetic acid
(b) Dilute HCl
(c) Ammonium hydroxide
(d) Benzene
(e) Distill water
(f) Potassium hydroxide
(g) Carbon tetrachloride
(h) Sodium acetate
(i) Sodium sulphate
(j) Carbonic acid

## 3.4 Aluminum




Aluminum is the most abundant metal found in the Earth's crust, mostly in the form of silicates. The word aluminum is derived from the Latin word '*alumen*' for alum which occurs as aluminum compound naturally used by the ancient people for dyeing textiles.

**Learning objectives**

On completion of this topic, students should be able to:

» *name the chief ores of an aluminum.*
» *outline the extraction process of aluminum.*
» *state the uses of aluminum.*
» *discuss the electro refining of aluminum by Hoopes method.*
» *explain electrolytic reduction of alumina by Hall-Heroult process.*

### 3.4.1 The chief ores of aluminum.

Bauxite: $Al_2O_3.2H_2O$ (Hydrated aluminium oxide), Cryolite: $Na_3AlF_6$ (Sodium aluminium fluoride), Corundum: $Al_2O_3$(Aluminium oxide), Feldspar: $K_2O.Al_2O_3.6SiO_2$(Potash feldspar).

### 3.4.2 Extraction of aluminum.

Davy, in 1808, first isolated the metal in an impure state. In 1825, the Danish chemist and physicist, Hans Christian Oersted, obtained the pure metal by reducing $AlCl_3$ by potassium.

$$AlCl_3 + 3K(s) \longrightarrow 3KCl(s) + Al(s)$$

This method was very expensive. In 1886, Charles Hall from U.S.A and Paul Heroult from France simultaneously reduced alumina ($Al_2O_3$) dissolved in cryolite ($Na_3AlF_6$) electrolytically on a large scale. Thus, the most important ore from which aluminum is extracted is bauxite ($Al_2O_3.2H_2O$) and it involves three basic steps as discussed below:

#### i. The purification of bauxite.

Bauxite contains some impurities, which need to be removed before subjecting it to electrolysis. It contains majority of alumina ($Al_2O_3$) and some water. The ore is generally contaminated with iron oxide, silica and titanium oxide. It is purified by Baeyer's process, in which bauxite is heated at 150°C along with caustic soda (NaOH solution) under pressure for several hours in a tank called digester. Aluminum oxide dissolves in alkali to form water soluble sodium meta aluminate ($NaAlO_2$) while ferric oxide and silica remain insoluble and settle down.

$$\underset{\text{(bauxite)}}{Al_2O_3.2H_2O} + 2NaOH \xrightarrow{150^\circ C} \underset{\text{(sodium meta aluminate)}}{2NaAlO_2} + 3H_2O$$

The solution is filtered to remove insoluble impurities. The filtrate containing sodium meta aluminate solution is agitated with freshly prepared aluminium hydroxide, $Al(OH)_3$ for about 36 hours when sodium meta aluminate hydrolyses to form precipitate of $Al(OH)_3$. The precipitate is filtered and strongly heated to obtain pure alumina ($Al_2O_3$).

$$NaAlO_2 + 2H_2O \longrightarrow Al(OH)_3\downarrow + NaOH$$

$$2Al(OH)_3 \xrightarrow{\text{Heat}} \underset{\text{(pure alumina)}}{Al_2O_3} + 3H_2O$$

#### ii. The electrolytic reduction of alumina – Hall-Heroult process

Alumina ($Al_2O_3$), being an ionic compound is very stable oxide and cannot be reduced by hydrogen, carbon and carbon monoxide. Pure $Al_2O_3$, dissolved in molten cryolite and fluorspar acts as an electrolyte and is reduced electrolytically using Hall-Heroult process. Since alumina has very high fusion temperature (about 2000°C), it is very difficult to electrolyse alumina alone. Molten electrolyte is covered with a layer of powdered coke to maintain the temperature and also to avoid molten metal from vaporizing at high temperature.

Functions of Cryolite ($Na_3AlF_6$) and Fluorspar ($CaF_2$).

**a. Cryolite.**

Cryolite is added to lower the fusion temperature of electrolytic bath. The mixture melts at 950°C instead of 2000°C thereby saving electrical energy.

**b. Fluospar along with cryolite.**

Pure alumina is a non-conductor and fluorspar acts as a solvent for the electrolytic mixture. Fluospar and cryolite enhance the conductivity of the mixture.

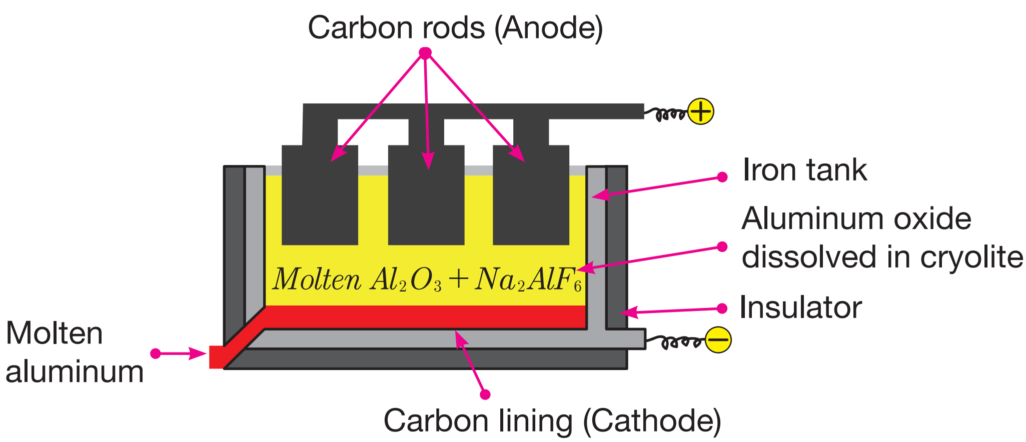



Figure 3.16 Electrolysis of fused alumina.

Carbon lining of an iron tank acts as cathode. The anode consists of a number of carbon rods attached to the copper clamps and dipped in fused alumina. The temperature of the electrolytic bath is maintained at 900°C-950°C. When an electric current is passed, aluminum is discharged at cathode. Aluminum being heavier than the electrolyte sinks to the bottom and is tapped out periodically from the outlet. Oxygen is liberated at the anode. Oxygen attacks carbon rods forming CO or $CO_2$, thus anodes have to be replaced from time to time.

**Electrode reaction:**

$$2Al_2O_3 \longleftrightarrow 2Al^{3+} + 2AlO_3^{3-}$$

$$\text{At cathode:} \quad Al^{3+} + 3e^- \longrightarrow Al$$

$$\text{At anode:} \quad 4AlO_3^{3-} \longrightarrow 2Al_2O_3 + 3O_2 + 12e^-$$

Thus, the overall chemical reaction taking place during electrolysis is

$$2Al_2O_3 \longleftrightarrow 4Al + 3O_2$$

Aluminum with 99.8% purity is obtained from this process.

### iii. The electro refining of aluminum – Hoopes method

Aluminum is further purified by Hoope's process. The electrolytic cell consists of three layers.

a) The top layer is of pure molten aluminum which acts as a cathode.

b) The middle layer consists of a mixture of the fluorides of Al, Ba and Na as ($Na_3AlF_6$) and $BaF_3$.

c) The lowest layer consists of molten impure aluminum which acts as an anode.

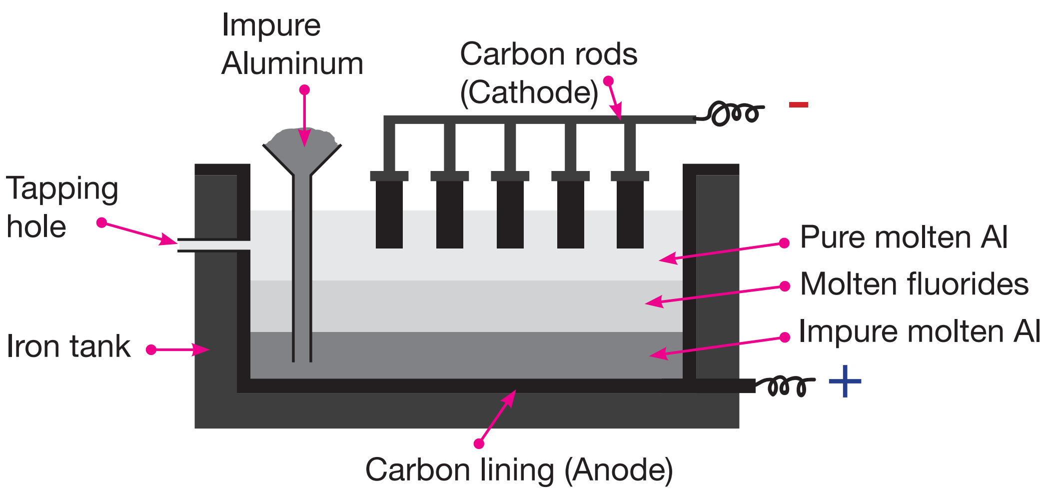



Figure 3.17 Purification of aluminum by Hoope's process.

The graphite rods immersed in pure aluminum and the impure molten aluminum along the carbon lining at the bottom of the iron tank act as electrode. On electrolysis, aluminum is deposited at cathode from the middle layer and an equivalent amount of aluminum is taken up by the middle layer from the bottom layer. Thus, aluminum is transferred from the bottom to the top layer through middle layer while impurities are left behind.

### 3.4.3 Uses of Aluminum

Some of the uses of Aluminum based on their special characteristic are given in Table 3.2.

Table 3.2 Uses of Aluminum

| Use | Characteristics |
|---|---|
| House hold uses: Making cooking utensils, household fitting, window frames, picture frames, electrical wires etc. | Aluminum being cheap and corrosion resistant.<br>It is good conductor of heat and electricity |
| Aircraft and other transport such as ships' superstructures, container vehicle bodies, trains bodies etc. | It is light, strong and resists corrosion. |
| It is used in making electric transmission cables and wires. | It is light, resists corrosion and good conductor of electricity |
| Packaging – used in making food containers including trays, foils, bottle caps, cans etc. | It resists corrosion and has no reaction with weak organic acids, it keeps food safe and clean from contamination. |
| Powdered aluminum mixed with linseed oil is used as paint for iron poles. | It being corrosion resistant. |
| A mixture of aluminum powder and ammonium nitrate is used as an explosive called **ammonal**. | The ammonium nitrate functions as an oxidizer and the aluminum as fuel. |
| Alloys of aluminum are used in making air ship, cheap balances, coins etc. | Aluminum is light but is weak, resists corrosion and alloyed to increase its strength. |
| Al is used in aluminothermy for the extraction of metals like Cr and Mn from their oxide which cannot be reduced by carbon. | Al has strong affinity for oxygen. |

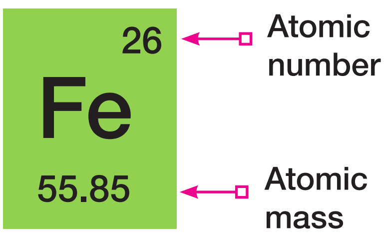
**Self Evaluation**

1. **Fill in the following blanks with appropriate word.**
   (a) Extraction of aluminum from alumina is done by _____.
   (b) The process of reduction of oxides by aluminum is known as ____.
   (c) During Hall-Heroult process, only $Al_2O_3$ is used up but ____ remains in the cell.
   (d) Aluminum is obtained from $Al_2O_3$ by _____ reduction.

(e) In the electrolytic refining of aluminum, the molten pure Al acts as ______.

2. **Give reason for the following.**
   (a) Cryolite is used in electrolytic reduction of alumina.
   (b) Aluminium is used in making utensils and food packaging.
   (c) Fluospar is used in Hall-Heroult process.
   (d) The positive anode is replaced from time to time during electrolytic reduction of alumina ($Al_2O_3$).
   (e) Aluminium is extracted by reduction of alumina rather than reduction of $AlCl_3$ by potassium.
3. Name the ore and give its chemical formula from which aluminium is extracted.
4. Name the processes used to purify aluminium ore.

## 3.5 Iron




Iron is the second most abundant metal in nature beside aluminum. Iron being reactive, it is rarely found in native state. It oxidizes readily in the presence of water and air and is found mainly in the form of oxides. Iron was known in Egypt before 3400 BC and in India from about 1000 BC. At around 1400 BC, Hittites of central Turkey discovered the method of extracting iron by reducing its ore with charcoal. Iron is present as a very important mineral in red blood cells of animals which helps in the transportation of oxygen.

**Learning objectives**

On completion of this topic, students should be able to:

» *name the important ores of an iron.*
» *explain the extraction process of iron.*
» *explain the different types of iron.*
» *state the uses of different forms of iron.*





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### 3.5.1 The chief ores of iron.

Haematite ($Fe_2O_3$), Siderite ($FeCO_3$), Magnetite ($Fe_3O_4$), Limonite ($Fe_2O_3.3H_2O$), Iron pyrites ($FeS_2$).

### 3.5.2 Extraction of cast iron or pig iron.

The cast iron is generally extracted from haematite and the extraction process involves the following steps:

#### i. Concentration of ores.

The ore is crushed with jaw crusher into small pieces so that it can be easily reduced by the carbon monoxide in the furnace and also they do not obstruct the flow of gas to the upper layers. The crushed ore is washed with the stream of water to remove impurities and then further concentrated by magnetic process. If pyrites ore is used, it is concentrated by froth floatation process.

#### ii. Calcination and roasting.

The concentrated ore is roasted and calcined with a little coal in shallow kiln (furnace) in excess air. This is done in order to remove moisture and to convert carbonate ores into oxides. It can also remove some of the volatile impurities like sulphur and phosphorus from the ore.

Some of the chemical reactions that can take place during calcinations are

$$2Fe_2O_3.3H_2O \xrightarrow{\text{Heat}} 2Fe_2O_3 + 3H_2O \uparrow$$
$$4FeCO_3(s) + O_2 \xrightarrow{\text{Heat}} 2Fe_2O_3(s) + 4CO_2 \uparrow$$

#### iii. Reduction of ore (smelting) in a blast furnace.

Smelting is carried out in a specially designed furnace known as a blast furnace. It is cylindrical in shape with about 25-30 m in height and 8-10 m in diameter at its widest part. It is made up of steel shell lined with fire-bricks. The upper two-third of the furnace flares out downwards to provide proper flow of the charge; the lower one-third narrows down as the materials melt and get collected.

**a. The blast furnace has the following important features.**

i) Double cup and cone

The double cup and cone arrangement is at the top of the furnace. When the charge is placed over it, the cone gets lowered and the charge gets in. Then the cone automatically gets lifted and the cup is closed. The function of this arrangement is to permit the entry of charge, help in even distribution of charges at the top and to prevent the exit of blast furnace gases during charging.

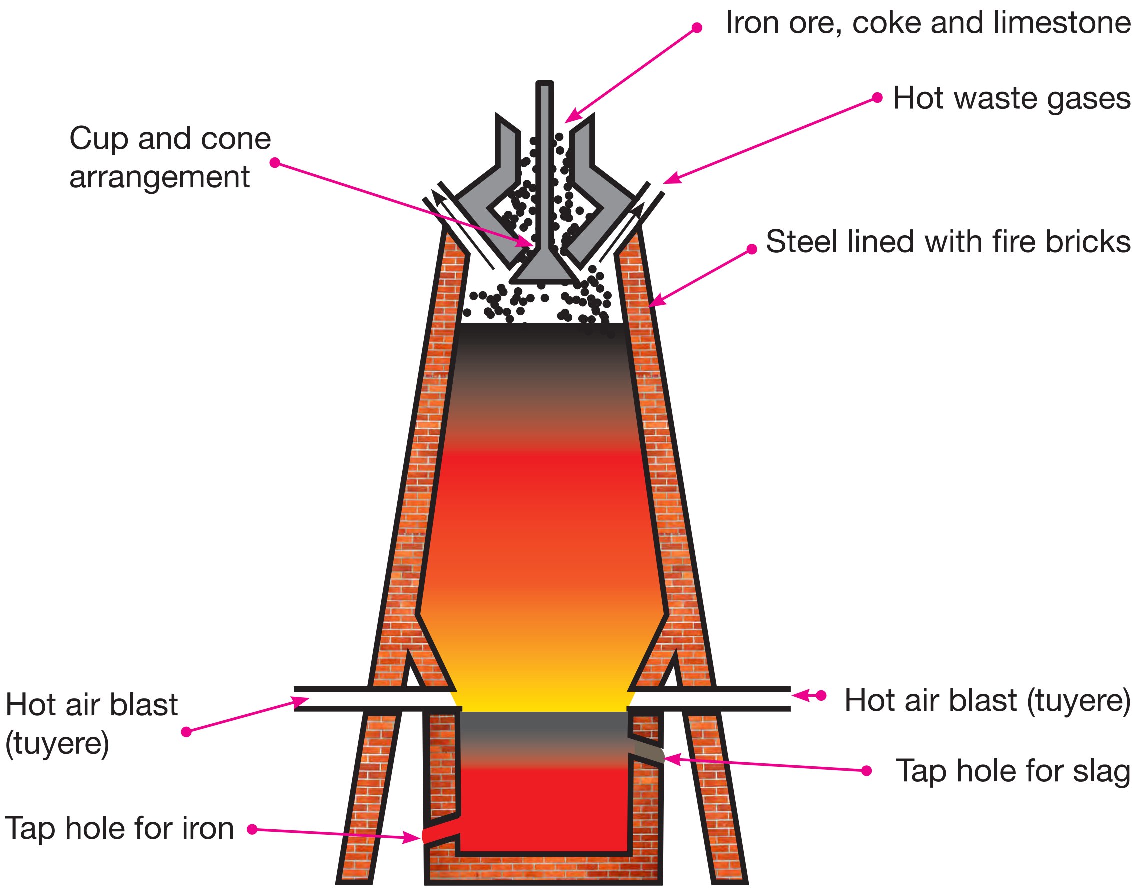



Figure 3.18 Blast furnace

ii) Tuyeres

These are water-jacketed iron pipes through which a hot compressed blast of air is introduced into the furnace to heat the charge.

iii) Tap holes

Just below the tuyeres are the two tap holes, the upper one for the discharge of molten slag and lower one for the discharge of molten iron. These are kept closed except when the molten materials are being discharged to prevent from coming in contact with air.

The charge of roasted ore, coke and lime stone is fed into the furnace by the double cup and cone arrangement. The coke acts as fuel and reducing agent for the ore. Haematite contains silica ($SiO_2$) as the chief acidic impurity, so in order to remove it; a basic flux like limestone ($CaCO_3$) is used. A hot air which supports combustion is passed through tuyere. The temperature inside the furnace is very high which can melt the iron and the slag. The molten metallic iron and the molten slag drip to the bottom of the furnace. The molten iron obtained from the blast furnace is called pig iron or cast iron containing carbon. The slag, being lighter floats on top of molten iron and are tapped separately through the respective tap holes. The waste gases coming out of the furnace contains CO, $CO_2$ and $N_2$.

**b. Chemical reactions in the blast furnace.**

The chemical reactions that occur in the blast furnace and temperature at various zones are shown in the Figure 3.19.

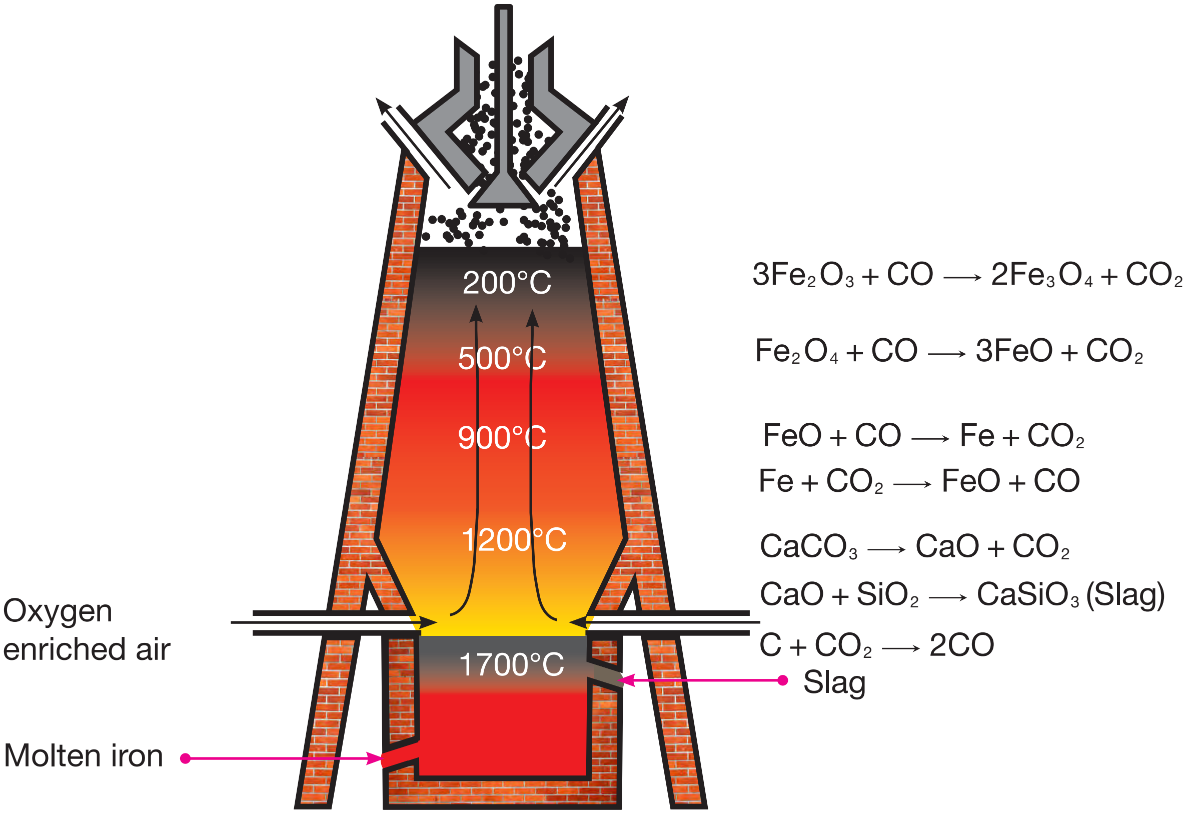



Figure 3.19 Blast furnace reactions at various regions

i) In the lower region, the coke is oxidized to carbon dioxide with the liberation of heat, which increases the temperature to 1400°C – 1750ºC.

$$C(s) + O_2(g) \xrightarrow{1700^\circ C} CO_2(g) \quad \Delta H = -405.9\,kJ$$

ii) In the middle region, the carbon dioxide passing upward through the layer of white–hot coke, is reduced to carbon monoxide. The limestone added as flux dissociate into calcium oxide and carbon dioxide.

Both the reactions being endothermic, the temperature in this region drops from 700ºC to 950ºC. Further, calcium oxide combines with silica (gangue) to produce calcium silicate, the fusible slag.

$$CO_2(g) + C(s) \xrightarrow{\text{Heat}} 2CO(g) \qquad \Delta H = +163.2\,kJ$$

$$CaCO_3(s) \xrightarrow{700^\circ C - 950^\circ C} CaO(s) + CO_2(g) \qquad \Delta H = +181.1\,kJ$$

$$CaO(s) + SiO_2(s) \xrightarrow{\text{Heat}} \underset{\text{(Slag)}}{CaSiO_3(l)}$$

iii) In the upper region, the ore is reduced to iron by carbon monoxide at temperature around 200°C to 550°C is given in the equation;

$$3Fe_2O_3 + CO \xrightarrow{220^\circ C} 2Fe_3O_4 + CO_2$$

$$Fe_2O_4 + CO \xrightarrow{400^\circ C} 3FeO + CO_2$$

$$FeO + CO \xrightarrow{550^\circ C} Fe + CO_2$$

**c. Products from the blast furnace and their uses.**

i) Pig iron or cast iron.

The molten iron obtained from the blast furnace contains 2.0% to 5.0% dissolved carbon and small amount of Si, Mn, P and S. It is named as pig iron because in an early method, the molten iron from the blast furnace used to be poured into moulds which resembled baby pigs.

Pig iron containing 1.0% to 3.0% silicon as an impurity is called cast iron because it can expand on solidification and take the shape of the mould. Pig iron is used in:

- manufacture of gutter covers, railings, drain pipes, radiators etc.
- cast iron is used in casting articles.
- making wrought iron
- making steel

ii) Slag.

The slag composes mainly of calcium silicate ($CaSiO_3$), calcium aluminates ($CaO.Al_2O_3$) and calcium alumina silicates ($CaO.Al_2O_3.SiO_2$). Slag is made into granules by running it into water when it solidifies into rock like materials. It is crushed and used

- in manufacturing of cement.
- along with tar in black topping of roads.
- as a ballast for rail tracks.
- for concrete aggregate and concrete sand.

iii) Fuel gases.

The mixture of waste gases which comes out of the outlet near the top of the furnace is called fuel gas or blast furnace gas. It consists of about 25% of CO, 10% of $CO_2$ and 6% of $N_2$. As it contains high percentage of combustible CO, it can be used for preheating the blast of air to be admitted through the tuyeres and also to convert poisonous CO into non-poisonous $CO_2$.

**Types of Iron**

Depending on the percentage of carbon along with other impurities, there are three types of Iron.

**a) Pig iron or cast iron**

The molten iron obtained from the blast furnace contains 2% to 5% dissolved carbon and small amount of Si, Mn, P, S etc. As it is the most impure form of iron, it is hard and brittle and cannot be welded. Since cast iron can expand on solidification, it can be used for casting articles of different shapes.

**b) Steel**

It contains about 0.1% to 1.5% of carbon along with other impurities like P, S, Si etc. It comes in between pig iron and wrought iron and its properties depend on the nature of other elements present in it. It has enormous application in steel industry in production of furniture and construction materials.

**c) Wrought iron**

It is the most pure form of iron with about 0.25% of carbon and traces of P and Si in the form of slag. It is soft, magnetic, malleable, and ductile and has high tensile strength. It can be heated, reheated and worked into various shapes. It is used in railings, outdoor stairs, handrails, fences and gates.

## 3.6 Alloy

An alloy is a homogeneous mixture of metals or a metal and a non-metal, which cannot be separated by physical means. The most important metallic component of an alloy (often representing 90 percent or more of the material) is called the main metal, the parent metal, or the base metal. The other components of an alloy (which are called alloying agents) can be either metals or non-metals and they are present in much smaller quantities. Alloys are produced to obtain certain desirable properties and to give greater strength or resistance to corrosion. An alloy is usually made by melting the components together and solidifying the mixture. Some common alloys of metal are described in Table 3.3.

**Table 3.3 Uses of common metals and their alloys.**

| Alloy | Composition | Properties | Use |
|---|---|---|---|
| Magnalium | Al = 68-95%.<br>Mg= 5-30%<br>Cu= 1-2 % | Strong, corrosion resistance, light weight, tough, low density and easy to weld. | For making cheap balances and other light scientific instruments and can be used on high speed lathes. |
| Duralumin | Al = 95%<br>Cu = 4%<br>Mn = 0.5%<br>Mg = 0.5% | Strong, hard, corrosion resistance, light weight and ductile. | For making automobile and air ships. |
| Aluminum bronze | Al = 90%<br>Cu = 10% | High strength, tarnish resistant and corrosion resistance. | For making picture frames, coins, trays etc. |
| Stainless Steel | Fe = 73%<br>Cr = 17 -19 %<br>Ni = 7 -9 % | Hard and resists corrosion. | Cutlery, cooking utensils, kitchen sinks, industrial equipment for food and drink processing and automobile parts. |
| Manganese steel | Fe = 83%<br>Mn = 13% | Very hard and resistance to wear and tear. | Used for making rock drills, rails or railway track, military helmets and crushing machines spring. |
| Tungsten Steel | Fe= 83%<br>W = 14%<br>Cr = 3% | Very hard even at high temperature. | Used for making cutting tools for high speed lathes. |
| Invar | Fe = 64%<br>Ni = 36% | Hard and has low coefficient of expansion. | Used in making measuring instruments and clock pendulums. |
| Alnico | Fe= 63%, Ni=20%<br>Al= 12%, Co =5% | Highly magnetic. | Used in making permanent magnets. |

| Alloy | Composition | Properties | Use |
|---|---|---|---|
| Brass | Zn = 20-40%<br>Cu = 60-80% | Malleable, ductile and resists corrosion. | Utensils, inexpensive jewelry, hose nozzles and condensers tubes. |
| Bronze | Cu=80%<br>Zn= 2%<br>Sn=18% | Strong, hard and can be polished. | Statues, parts of machines, coins and medals, tools and electrical hardware. |
| German silver | Cu=50%<br>Zn= 25%<br>Ni=25 % | Hard, silvery and takes up polish. | Making utensils and ornaments. |
| Gun metal | Cu=87%<br>Zn=3%<br>Sn=10% | Hard, brittle and easily cast. | Guns, casting and gears. |

## Self Evaluation

**1. State whether the following statements are true (T) or false (F)**

(a) The cast iron is generally extracted from siderite.

(b) Alnico is used in making coins, cheap balances and pendulum.

(c) Blast furnace gases have their application in cement industry.

(d) For any basic gangue, acidic flux has to be utilized.

(e) Duralumin, brass and magnalium form rust easily.

**2. Fill in the Blanks with the suitable word(s)**

(a) The iron pipes through which a hot compressed blast of air is introduced into the furnace are called_____.

(b) 2.0 to 4% of carbon is found in ______.

(c) Smelting is done in ______.

(d) The concentration of the haematite is done by______.

(e) ______is used in drain pipes, gutter covers, railings etc.

**3. Explain the function of each of the following component of the blast furnace.**

(a) Double cone and cup arrangement.

(b) Tuyeres

(c) Tap holes

(d) Steel body with fire bricks.

4. **Give the composition of the following alloys:**

(a) duralumin

(b) gun metal

(c) alnico

(d) tungsten steel

(e) manganese steel

## Summary

1. There are 118 elements at present in the universe which are divided into metals, non-metals, metalloids and noble gases.
2. Metallurgy is the process of extracting the pure metal from its ore.
3. Metals can be extracted from their ores by various methods depending on the reactivity of the metal concerned.
4. The process of decomposition of a chemical compound in its molten state or in aqueous solution by the passage of electricity through it is termed as electrolysis.
5. Electrolytes conduct electricity in their aqueous solution or in molten state where as non-electrolytes do not conducts electricity in their aqueous solution or in molten state.
6. Electrodes are the graphite or metal plates or rods through which the electric current enters or leaves the electrolyte
7. Electrolysis is used in chemical industries and electrometallurgy for extraction of reactive metals from their ores, electro refining, electroplating etc.
8. Aluminum being one of the reactive metals is extracted from its ore by electrolysis.
9. Aluminum is useful as construction metals, house hold utensils and food industry because it is light and resistant to corrosion.
10. The extraction of iron in the blast furnace is the most important example of chemical reduction.
11. The different forms of iron produced can be used for various purposes.
12. Alloy is a homogeneous mixture which is made up of at least two different chemical elements which cannot be separated by physical means, one of

which is a metal.

13. Various alloys are used as construction materials, manufacturing of automobiles, utensils, furniture etc.

# Exercise

## I. Fill in the blanks with correct word(s).

1. The iron pipes through which a hot compressed blast of air is introduced into the furnace are called_____.
2. 2 to 4% of carbon is found in ______.
3. Smelting is done in ______.
4. The concentration of the haematite is done by______.
5. ______is used in drain pipes, gutter covers, railings etc.

## II. Match the items of Column I with the corresponding items of Column II.

| Column I | Column II |
|---|---|
| 1. The substance added to get rid of the gangue | a. $CaO.Al_2O_3.SiO_2$ |
| 2. Removal of moisture from the $Al_2O_3.2H_2O$ is done by | b. Haematite |
| 3. Heating of ores in excess supply of air | c. Flux |
| 4. The formula for calcium alumina silicate | d. Roasting |
| 5. The chemical name of ore $Fe_3O_4$ is | e. Magnetite |
| | f. $CaO.Al_2O_3$ |
| | g. Calcination |

## III. State whether each of the statement is True or False.

1. The oxidation reaction takes place at cathode.
2. Pyrites ores are concentrated by froth floatation process.
3. The degree of dissociation indicates the strength of an electrolyte.
4. The ease of discharge of cation at cathode decreases down the electrochemical series.
5. Ammonal is an explosive mixture of aluminum powder and $NH_4NO_3$.

## IV. Choose the most appropriate response from the given options.

1. The extraction of aluminum by electrolytic reduction of alumina uses

A Hall-Herout process. C thermite process.

B Bayers process. D Hoopes process.

2. In the froth floatation process for concentration of the ores, the ore particles float because
   - A they are lighter than other materials.
   - B they are insoluble in water.
   - C they bear electrostatic charge.
   - D their surface is not easily wetted by water.
3. Which of the following species will be deposited at the cathode on electrolysis of an aqueous solution of potassium bromide?
   - A K
   - C $Br_2$
   - B $H_2$
   - D $O_2$
4. Heating an ore in the absence of air below its melting point is called
   - A leaching.
   - C smelting.
   - B roasting.
   - D calcination.
5. Which one of the following ores is best concentrated by gravity separation method?
   - A Magnetite.
   - C Galena .
   - B Alumina.
   - D Cassiterite.
6. The leaching process of concentrating ores is based upon the
   - A magnetic properties of gangue and ore.
   - B preferential wetting of ore particles by oil.
   - C difference in the specific gravity of ore and gangue particles.
   - D dissolution of desired metal from ore particles in suitable reagent.

## V. Write answers for following questions.

1. Aluminum is extracted from its chief ore, bauxite. The ore is first purified, in relation to this process, answer the following questions
   - (a) Name the process by which bauxite is purified.
   - (b) Write down the balanced equation for this process.
   - (c) Name the chemical used for dissolving aluminum oxide.
2. With the help of a diagram explain electrolytic refining of aluminum by Hoopes process.
3. Solid sodium chloride does not conduct electric current. Give reason.
4. Explain briefly the extraction process of iron from its chief ore.
5. Draw and describe the blast furnace. Write the reactions taking place in various zones inside it during the manufacture of cast iron.
6. Write short notes for the following:
   - (a) Smelting

(b) Aluminothermy.

7. Sonam is given two substances, one is a strong electrolyte and other is a weak electrolyte. What simple experiment would she conduct to distinguish between them?
8. Write the balanced chemical equation for the following reactions:
   (a) Aluminium powder is warmed with hot and concentrated caustic soda solution.
   (b) When carbonate ores are heated in a limited supply of air.
9. Write the names, composition and uses of the following alloys.
   (a) Stainless steel
   (b) Brass
   (c) Aluminum bronze
10. Answer the following questions with respect to electrolytic reduction of alumina ($Al_2O_3$) by Hall-Heroult process.
    (a) Name the following:
        (i) The electrolyte used.
        (ii) The material of which anode and cathode must be made of.
        (iii) The material of which electrolytic cell is made up. Give reason for the answer.
    (b) What are the purposes of using cryolite and fluorspar during the reduction process?
    (c) Why are anodes replaced periodically?
    (d) Write down the chemical reaction which takes place at the anode and cathode during electrolytic reduction of alumina.
11. What are the factors that affect electrolysis? Explain briefly.

# 4 Halogens

## 4.1 Introduction

The elements of group VIIA or 17 constitute a family known as the halogen. The term halogen has been derived from the Greek word ***halos*** meaning salt and ***genes*** meaning former. The family comprises the elements Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I) and Astatine (At). The elements of this family are highly reactive and thus, they readily combine with most of the elements to form various products. Most of the chemical and the physical properties, as exhibited by halogens are results of their smaller atomic size due to increasing nuclear charge across the period. The halogens and their compounds have wide variety of uses due to their oxidizing and reducing property.

## 4.2 Basic information of halogens

**Learning objectives**

On completion of this topic, students should be able to:

» *identify the elements that constitute group VIIA or 17.*
» *explain why halogens exhibit similar chemical and physical properties.*
» *compare the properties of halogens with other groups.*
» *identify the native source of halogen.*
» *explain how halogens achieve their stable state.*

### 4.2.1 Occurrence and source

The halogen atoms are highly reactive to occur in their elemental state in nature. Thus, all halogens exist as diatomic molecules. Astatine is highly radioactive halogen and is the rarest element found in nature.

The halogen atoms in combined state are found in the form of various minerals and salts. For example, fluorspar ($CaF_2$), cryolite ($Na_3AlF_6$) and fluorapatite ($Ca_5(PO_4)_3F$) are minerals that contain fluorine. Chlorine is considered as the twentieth most abundant elements and is found in nature as salts and salt water. Bromine is found in some salt brine and in the sea. Table 4.1 summarizes the basic information of group 17 elements.

**Table 4.1**

| Basic information | Halogens | | | | |
|---|---|---|---|---|---|
| | Fluorine | Chlorine | Bromine | Iodine | Astatine |
| Symbol | F | Cl | Br | I | At |
| Atomic number | 9 | 17 | 35 | 53 | 85 |
| Mass number | 18.9 amu | 35.4 amu | 79.9 amu | 126.9 amu | 210.0 amu |
| Isotopes (*atoms with same atomic number but with different mass number*) | F-18, F-19. | Cl-35, Cl-36, Cl-37, Cl-38. | Has variety of isotopes. The two most stable isotopes are Br-79, Br-81. | 37 known isotopes which undergo radioactive decay except I-127. | More than 30 short-lived isotopes of have been identified. At-210 has the longest half-life of 8 hrs 10 min. |
| Discovery | Joseph Henri Moissan in 1886. | Carl Wilhelm Scheele in 1774. | Antoine J. Balard in 1826. | Bernard Courtois in 1812. | D.R. Corson in 1940. |
| Source | Mineral Fluorite. | Salt. | Sea Water. | Sodium and potassium compounds. | Synthetically obtained by bombarding bismuth (Bi) with α-particles. |

**Do you know?**

Chlorine and iodine are named by their colour: *chloros* means 'yellowish green' and *ioeides* is 'violet' in Greek. Bromine is named by its smell *bromos* is Greek word for 'stink'.

## 4.2.2 Electron configuration

The atoms of an element are composed of three fundamental particles called electron, proton and neutron. The proton and neutron together constitute a nucleus and is located at the central core of an atom. The electrons revolve round the nucleus and exist in discrete energy levels known as first, second, third energy level and so on. Each energy level can accommodate only certain number of electrons as shown in the Figure 4.1. The placing of electrons in various energy levels is guided by the expression $2n^2$. In the expression, 'n' refers to shell number and it has value 1, 2, 3, 4 etc.

1st energy level (shell) → (n =1), thus the number of electron in it is $2(1)^2 = 2$

2nd energy level (shell) → (n = 2), thus the number of electron in it is $2(2)^2 = 8$

3rd energy level (shell) → (n= 3), thus the number of electron in it is $2(3)^2 = 18$

4th energy level (shell) → (n= 4), thus the number of electron in it is $2(4)^2 = 32$

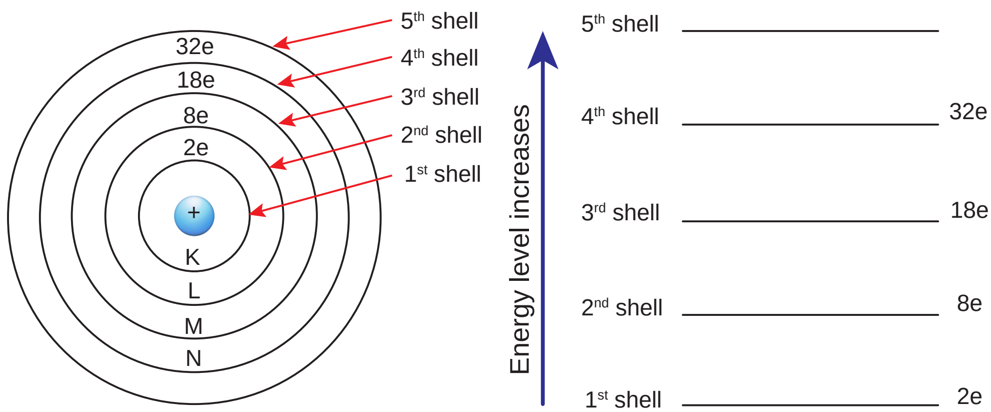
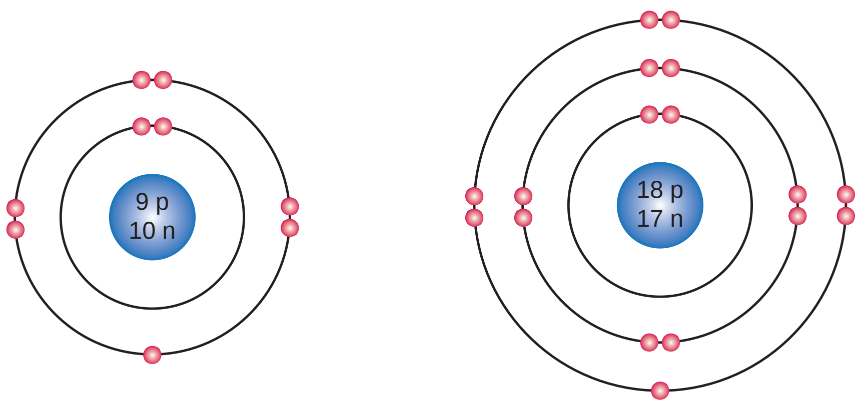



Figure 4.1 Electron holding capacity of shells

The distribution of electrons in various energy levels or shells of an atom is called the electron configuration. The Figure 4.2 shows the distribution of electrons in various energy levels in atoms of halogen.




(a) Electronic configuration of fluorine 2,7. (b) Electronic configuration of chlorine 2,8,7.

Figure 4.2 Orbital structures.

The neutral halogen atoms have seven electrons in their valence shell, a configuration in which one electron short of noble gas configuration. Thus, the halogen atoms require one more electron to achieve their noble gas configuration or octet state, which can be achieved by any of the following processes:

### i. Mutual sharing of electron between two similar halogen atoms

In this process, the two similar halogen atoms mutually share one electron each to achieve their octet state. This also shows that all halogen atoms exist as diatomic molecules ($F_2$, $Cl_2$, $Br_2$ and $I_2$) bonded by a single covalent bond. Generally, the diatomic molecules are non-polar in nature since there is no difference in the electronegativity value between the combining atoms. In such a bonding the shared pair of electrons is located centrally between two combining atoms as shown in the Figure 4.3.

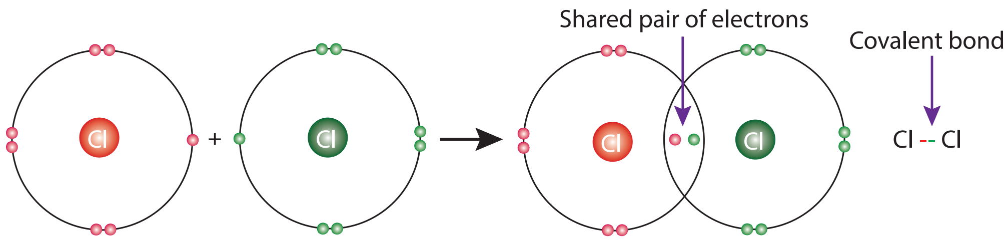



Figure 4.3 Formation of non polar covalent molecule.

### ii. Mutual sharing of electron between one halogen and one non-metal

When one electron each from halogen and non-metal (other than the halogens) is mutually shared, the type of compound formed is called polar covalent compound. It is so called since there is a development of partial polarity (charges) in combining atoms due to difference in the electronegativity value. In such bonding, the shared pair of electrons is attracted towards the more electronegative atom. Figure 4.4 illustrates the formation of polar covalent compound with respect to the formation of HCl.

*Cl being more electronegative attracts the shared pair of electron towards itself.*

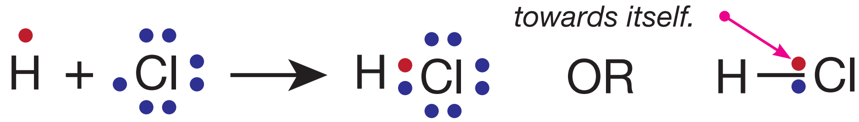



Figure 4.4 Formation of polar covalent compound.

### iii. Losing and gaining or transference of electron

In this case the electronegative halogen atoms combine with the electropositive metal atom to form ionic compounds. During this process the metallic atoms donates an electron(s) while the halogen atoms accepts it to achieve their octet state. The combining atoms of ionic compound are held by a strong chemical bond known as electrovalent or ionic bond. The ionic compounds like sodium chloride (NaCl), sodium bromide (NaBr) and magnesium chloride ($MgCl_2$) are formed by this process.

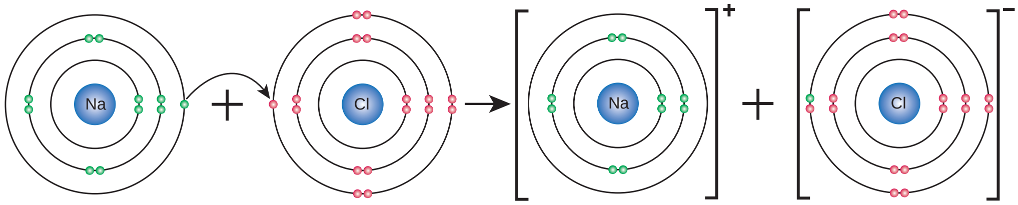



Figure 4.5 Formation of ionic compound.

**Activity 4.1 Placing electrons in their places**

Answer the following questions with respect to the atom with its atomic number 17.

1. Draw its atomic structure and place the electrons in their respective energy levels.
2. Write the electronic configuration of it.
3. What is the valence electron in it?
4. In which period and group is the given element located? Write answer without referring periodic table?
5. How can the given element achieve its noble gas configuration?
6. Write electronic configuration of atom in its ionic state.
7. Why are halogens highly reactive? Explain the reasons based on the electronic configuration.

### 4.2.3 Safety and storage of elemental halogens

The halogens are corrosive, poisonous and highly toxic although some of their compounds are used as table salt (NaCl) and toothpaste. They are highly reactive to combine readily and explosively with the metals. Thus, the halogen atoms are preferred to be stored in glass containers. However, fluorine reacts with a glass container and form silicon tetrafluoride ($SiF_4$) in the presence of small amount of water. Thus, the glass containers must be completely dry to avoid any further reactions. The halogens are powerful oxidizing agent. The container of chlorine may explode if it is exposed to heat due to oxidation. Similarly, bromine should not come in contact with ammonia, acetylene and other elements to avoid explosion during storage. In a laboratory, the halogens must be handled safely in a well ventilated fume hood to avoid leakage of poisonous vapors as its inhalation will cause severe damage to the respiratory system. As a safety measures in times of unavoidable accidents, the reducing agents like sodium sulphite must be available in a laboratory for destruction of excess halogen.

**Self Evaluation**

1. Explain why halogens are rarely found in native state?
2. Compare the conditions required for formation of covalent and ionic compound.
3. If an atom 'X' and 'Y' with seven electrons each in its valence shell undergoes chemical bonding, what will be the type of bond formed? Give an example to illustrate your answer.
4. Why do neutral atoms lose or gain electron?
5. Name the family in a periodic table that does not participate in chemical bonding under ordinary conditions. Give reason to support your answer.
6. When are compound said to be polar?

## 4.3 General properties

**Learning objectives**

On completion of this topic, students should be able to:

» *apply the concept of nuclear charge and effective nuclear charge to explain the trends in periodic properties of elements.*

» *explain the trends in general properties of halogens down the group.*

» *compare the trends in properties of halogen with other groups.*

» *investigate the displacement reactions of halogen.*

» *explain the oxidizing and reducing property of halogen.*

### 4.3.1 Nuclear charge and effective nuclear charge

The nuclear charge of an atom is equal to the number of proton in the nucleus. For example, the nuclear charge of chlorine atom is +17 as it has 17 protons in the nucleus. Thus, the nuclear charge of an atom is the total charge of all the protons in the nucleus.

Electrons are held in an atom or ion by the electrostatic attraction between the positively charged nucleus and the negatively charged electrons. In multielectron species, the electrons do not experience the full positive charge of the nucleus due to shielding by electrons which lie between the electron of interest and the nucleus. The amount of positive charge that actually acts on an electron of valence shell is called the effective nuclear charge. Thus, the effective nuclear charge is that portion of the total nuclear charge that a given electron in an atom experiences. Usually, the value of effective nuclear charge is lesser than the value of nuclear charge. For example, fluorine has nine electrons and nine protons. Its nuclear charge is +9. However, its effective nuclear charge is +7, because of the shielding due to two electrons. Effective nuclear charge of an atom can be approximated by:

Effective nuclear charge = Atomic number - Shielding electrons

$$Z_{eff} = Z - S$$

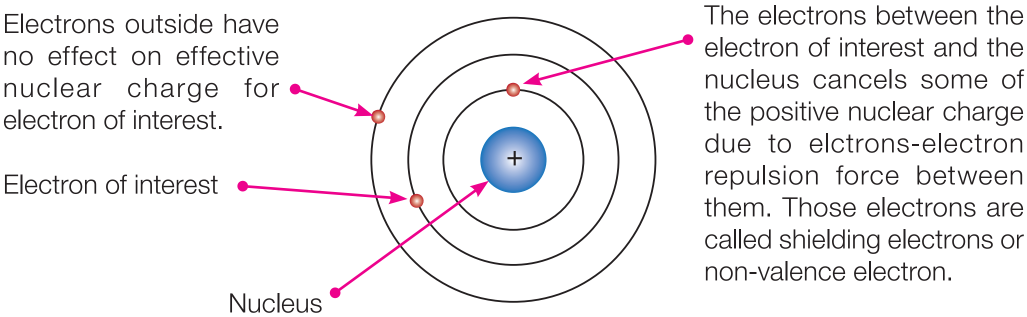



Figure 4.6 Effective nuclear charge diagram.

### Shielding effect

- The shielding effect describes the attraction between an electron and the nucleus in any atom with more than one electron shell. Shielding effect can be defined as a reduction in the nuclear charge on the electron cloud, due to a difference in the attraction forces of the electrons on the nucleus.
- The shielding effect explains why valence-shell electrons are more easily removed from the atom. The effect also explains atomic size. The more shielding, the further the valence shell can spread out and the bigger atoms will be.

Table 4.2 Calculated effective nuclear charge.

| Elements | Nuclear charge | Non-valence or shielding electrons (the electron other than valence electrons). | Effective nuclear charge ($Z_{eff} = Z - S$) |
|---|---|---|---|
| Carbon (C) | +6 | 2 | 6 – 2 = +4 |
| Chlorine (Cl) | +17 | 10 | 17 – 10 = +7 |
| Calcium (Ca) | +20 | 18 | 20 – 18 = +2 |

| Elements | Nuclear charge | Non-valence or shielding electrons (the electron other than valence electrons). | Effective nuclear charge ($Z_{eff} = Z - S$) |
|---|---|---|---|
| Sodium ion ($Na^+$) | +11 | 2 | 11-2 = +9 |

**Example problems on calculating the effective nuclear charge.**

1. Calculate the effective nuclear charge experienced by the electrons of neon atom in
   a. first or K shell
   b. valence shell

**Solution**

a. For first shell

$Z_{eff} = Z - S$

$Z_{eff} = 10 - 0 = +10$

*No. of shielding electron is 0, because there is no electron between the electron of interest (electron in 1st shell) and the nucleus.*

b. For valence shell

$Z_{eff} = Z - S$

$Z_{eff} = 10 - 2 = +8$

*No. of shielding electron is 2, because there are two electrons between the electron of interest (electron in 1st shell) and the nucleus.*

2. What charge is experienced by the electrons of sodium atom in
   a. first shell
   b. second shell and
   c. valence shell

**Solutions**

a. For first shell

$Z_{eff} = Z - S$

$Z_{eff} = 10 - 0 = +10$

*No. of shielding electron is 0, because there is no electron between the electron of interest (electron in 1st shell) and the nucleus.*

b. For second shell

$Z_{eff} = Z - S$

$Z_{eff} = 11 - 2 = +9$

*No. of shielding electron is 2, because there are two electrons between the electron of interest (electron in 2nd shell) and the nucleus.*

c. For valence shell

$Z_{eff} = Z - S$

$Z_{eff} = 11 - 10 = +1$

*No. of shielding electron for the third shell is 10, because there are 10 electrons (2e in 1st shell and 8e in valence shell) between the electron of interest (electron in valence shell) and the nucleus.*

**Activity 4.2 The trends in nuclear and effective nuclear charge of halogens.**

Fill in the empty boxes in Table 4.3.

Table 4.3

| Halogens | No. of proton | Nuclear charge | Shielding electron | Effective nuclear charge on valence electron |
|---|---|---|---|---|
| Fluorine | 9 | | | +7 |
| Chlorine | 17 | | 10 | |
| Bromine | 35 | +35 | | |
| Iodine | 53 | | 46 | |

Questions

1. Describe the trend in nuclear charge and effective nuclear charge of halogen down the group.
2. Predict the trend in nuclear and effective nuclear charge across the period. Illustrate your answer with the elements of 2nd period from Li to Ne.
3. Differentiate between nuclear charge and effective nuclear charge.
4. Define the term shielding effect.

## 4.3.2 Periodic properties of halogens

### i. Atomic size

The distance from the atomic nucleus to the outermost shell is considered as the atomic radius.

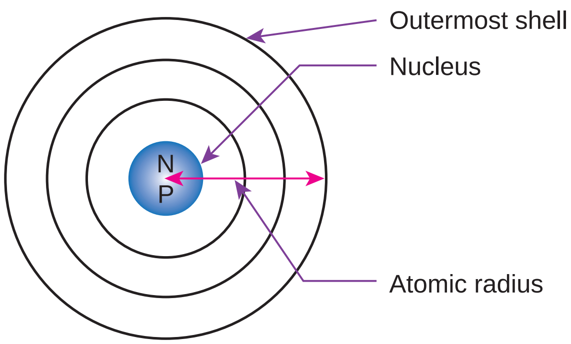



Figure 4.7 Atomic radius of element

Halogen family being located in extreme right corner of the periodic table has the smallest atomic radius in its respective periods. The decrease in atomic size across the period is due to increasing nuclear charge across the period.

Down the group, atomic radius increases due to increase in atomic number and hence the number of shell. In such cases, there will be more shielding among the electrons, thereby spreading the valence electrons shell further from the nucleus, which results in larger atomic radius.

**Comparing size of atom and ion (anion)**

The atoms of an element may lose or gain one or more electrons in order to achieve their octet state. The neutral atoms on losing or gaining electron(s) will develop a charge on it. Thus, the atoms with a net charge on it are referred as ions. The ions are of two types:

a) **Cations**: If a neutral atom loses one or more electrons, it acquires a net positive charge on it and is known as a cation. The metal atoms have tendency to lose electrons to form cation.

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

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

$$Al \rightarrow Al^{3+} + 3e^{-}$$

The ionic radius (i.e., the radius of cations) is smaller than their parent atom. This is because when an atom loses one or more electrons from it, the effective charge on the nucleus increases. Thus, this increased effective nuclear charge pulls the remaining electrons causing smaller cationic size.

b) **Anions**: If a neutral atom gains one or more electrons, it acquires a net negative charge on it and is known as an anion. Generally, non-metals accept electrons to form anion. In case of the halogens, they accept one electron to acquire a net negative charge of -1.

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

$$Cl + e^- \rightarrow Cl^-$$

The ionic radius (i.e., the radius of anions) is larger than their parent atom. Likewise, the ionic radius of halide is larger than their parent halogen atom. For example, the size of chloride ion ($Cl^-$) is larger than the chlorine atom (Cl). This is because the halide ions (i.e., anions) are formed on adding one electron to their parent atoms. With the addition of electron(s), the effective nuclear charge decreases resulting in great electronic repulsion which causes an expansion in electron cloud.

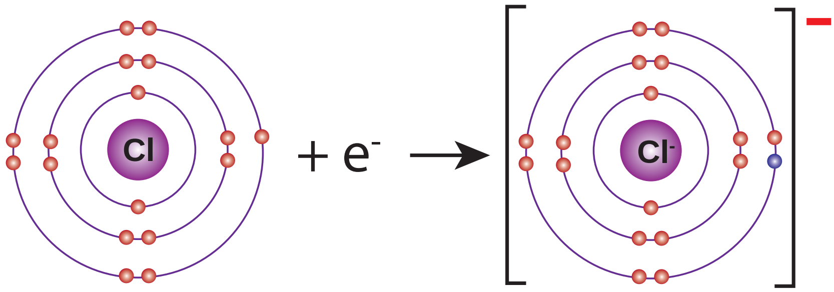



Figure 4.8 Size of chlorine atom and chloride ion.

## ii. Electronegativity

Electronegativity is the tendency of an atom in a molecule to attract the shared pair of electrons towards itself during covalent bonding. Table 4.4 describes the general trend in the electronegativity value down the group and across the period.

**Table 4.4**

| IA | IIA | IB | IIB | IIIA | IVA | VA | VIA | VIIA |
|---|---|---|---|---|---|---|---|---|
| H 2.1 | | | | | | | | |
| Li 1.0 | Be 1.5 | | | B 2.0 | C 2.5 | N 3.0 | O 3.5 | F 4.0 |
| Na 0.9 | Mg 1.2 | | | Al 1.5 | Si 1.8 | P 2.1 | S 2.5 | Cl 3.0 |
| K 0.8 | Ca 1.0 | | | | | | | Br 2.8 |
| | | | | | | | | I 2.5 |

Increasing electronegativity

Increasing electronegativity

The decrease in value of electronegativity down the group is attributed to the increase in size of atom on moving down the group. As the atomic size increases, the valence electrons are set farther apart which results to decrease in the effective nuclear charge. Fluorine has the highest electronegativity value among the elements of periodic table.

When atoms combine to form a molecule, the bonding between them may be ionic or covalent (polar or non-polar) depending on the difference in the electronegativity.

### Table 4.5 Determining the nature of bond

| Electronegativity difference (ΔEN) between combining atoms | Nature of bond | Example |
|---|---|---|
| Less than 0.5 | Non polar covalent. | $Cl_2$, $CS_2$ |
| 0.5 – 1.6 | Polar covalent. | $NH_3$, $H_2O$ |
| 1.6 - 2.0 | Ionic bond if bonding involves metals. | NaBr |
| | Polar covalent if bonding involves only non metal. | HF |
| Greater than 2.0 | Ionic bond. | NaCl |

### iii. Ionization energy or ionization enthalpy

Ionization energy is defined as the minimum amount of energy required by a neutral atom to remove electron(s) from the valence shell to form cation. It is expressed in unit of electron volts (eV). This process is accompanied by the absorption of heat and hence, it is an endothermic process.

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

The ionization energy increases on moving across the period due to increase in nuclear charge. The increasing nuclear charge across the period increases the attraction between nucleus and valence electrons. Thus, halogen atoms have the highest ionization energy in their respective period and have great difficulty in removing the electron from a valence shell.

On moving down the group, ionization energy decreases. This is because the electron to be removed from the valence shell is increasingly distant from the nucleus, as a result of the atoms getting bigger. The attraction of the nucleus for the electron becomes less, and it becomes easier to pull it away. Electrons in the inner energy levels also produce a shielding effect. These inner electrons reduce the attraction of the nucleus for the outer electrons. The shielding effect will increase as the number of inner energy levels increases. Within the halogen group, fluorine has the highest ionization energy.

### iv. Electron affinity

Electron affinity is the ability of an atom to accept an electron readily. Thus, the term electron affinity may be defined as the amount of energy released when a neutral gaseous atom accepts an electron to form anion. Since, the energy is released on changing from atom to anion, the process is exothermic.

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

The electron affinity increases on moving across the period due to decrease in atomic size or increase in nuclear charge. Thus, the halogens have relatively higher electron affinity values in their respective period. Halogens can readily accept one electron from alkali metals during chemical bonding to form halide salts.

On moving down the group, electron affinity value decreases because size of the atom increases. However, fluorine has a lesser value of electron affinity than the chlorine. A lesser value of electron affinity for fluorine is due to the small size of fluorine atom. In a compact L-shell of fluorine, the added electron brings about electron-electron repulsions, which gives rise to low values of electron affinity.

## 4.3.3 Physical properties

### i. Physical state

At room temperature (25°C), fluorine and chlorine are gases, bromine is a volatile liquid and iodine is a volatile solid. The existence of halogen atoms in all three states at room temperature is the result of increasing strength of the **van der Waal**'s forces of attraction between the molecules due to increase in atomic size.

### ii. Colour and solubility

Generally, the solute can best dissolve in a solvent when both solute and solvent molecules exhibit similar polarities. The phrase “like dissolves like” is one such rule of thumb to determine the solubility of substances. The molecules of halogen are non-polar in nature while the water molecule is polar in nature. When two molecules of different polarities are mixed, they do not mix and hence, the halogens have low solubility in water.

Chlorine, bromine and iodine are soluble in water to some extent. However, it is difficult to determine the solubility of fluorine as it reacts with water violently. Iodine is sparingly soluble in water. However, iodine forms aqueous solution in presence of iodide ions. The iodine reacts with iodide ions to form triiodide ions which are responsible for its brown colour.

$$I_2(aq) + I^{-}(aq) \rightarrow I_3^{-}(aq)$$

Table 4.6 Colours and solubility of halogens.

| Halogen | Colour and solubility | | |
|---|---|---|---|
| | In pure form | In non-polar solvents | In water |
| Fluorine | Pale yellow gas | (Reacts with solvents) | (Reacts with water) |
| Chlorine | Pale green gas | Pale green solution | Pale green solution. |
| Bromine | Dark red liquid | Orange solution | Yellow or Orange-red depending on concentration solution. |
| Iodine | Grey solid | Purple solution | (Slightly soluble) but forms a brown solution if excess KI is present |

### iii. Density

The density of a substance is the measure of mass to its volume. On moving down the group, the mass as well as the volume of atom increases. However, the increase in mass is more than the increase in volume. Hence, the density increases as you go down the groups in the periodic table.

On moving across the period, the mass increases but the volume decreases due to increase in the nuclear charge. Hence, the density increases on moving across the period.

### iv. Melting point and boiling point

When bonds between the combining atoms are strong, the more energy is required to break that bond. As such there is no distinguishable trend in the periodic table. However, the general conclusion is that the melting point and the boiling point of metallic atoms are higher than those of non-metallic atoms.

**Do you know?**

Group 1 and group 17 are on opposite side of periodic table and show opposite trends in their reactivity and melting points.

**Activity 4.3 Trends in boiling point and melting point of halogen family.**

The graph in Figure 4.9 shows a melting point and boiling point of halogen family at room temperature. Based on the information answer the questions as follow:

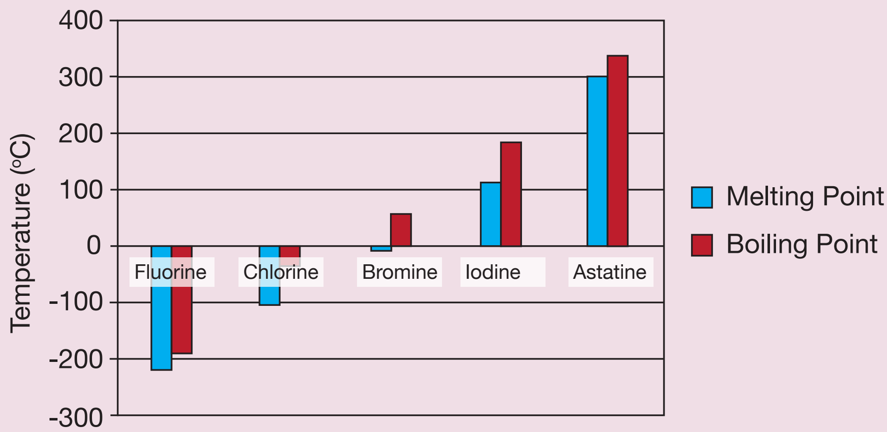



Figure 4.9

**Questions**

1. Describe the trend in melting point and boiling point of the halogens down the group.
2. Explain the reasons for the trend observed in question 1.

### v. Oxidation state

An oxidation state, also called oxidation number is the number assigned to the atoms in a chemical combination. This number represents the number of electrons that an atom can gain, lose, or share when chemically bonding with an atom of another element.

All halogens uniformly exhibit an oxidation state -1. Except for fluorine, other halogens also exhibit positive oxidation states of +1, +3, +5 and +7. Fluorine is the most electronegative of all the halogens, and hence, it has -1 oxidation state only.

## 4.3.4 Chemical properties

### i. Combustibility

Halogens are neither combustible nor a supporter of combustion.

### ii. As oxidizing agents

Halogens in its molecular state are powerful oxidizing agents. The oxidizing power of it is because of their high affinity for electron. They accept electron and get reduced to form halide ion. The oxidizing power decreases down the group from fluorine to astatine.

$$F_2(g) + 2e^- \rightarrow 2F^-(aq)$$

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

$$Br_2(g) + 2e^- \rightarrow 2Br^-(aq)$$

$$I_2(g) + 2e^- \rightarrow 2I^-(aq)$$

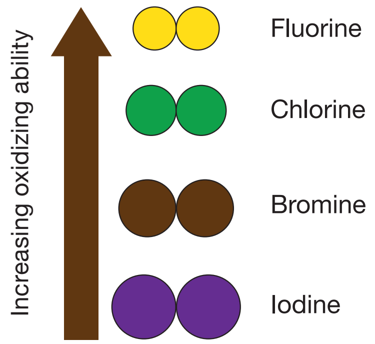



Figure 4.10 Trend in oxidizing power of halogen

### iii. As reducing agents

The halogens in their ionic form are powerful reducing agent. The reducing power of halides is its ability to lose electron and get oxidized to form halogen molecules. Halides are the binary compounds, of which one part is a halogen atom and the other part is an element or radical that are less electronegative than the halogen.

$$2Na + Cl_2 \rightarrow 2NaCl$$

In the reaction, the product NaCl is binary compound which undergoes dissociation as;

$$NaCl \rightarrow Na^+ + Cl^-$$

Here, $Cl^-$ is halide (chloride ion) and it acts as a reducing agent by losing electron. During the process, $Cl^-$ ion is converted to neutral chlorine atom. Thus, chloride ion is said to be oxidised.

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

Similarly, all halides acts as a reducing agent.

$$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$$

$$2I^-(aq) \rightarrow I_2(aq) + 2e^-$$

As the number of shells in an ion increases, the electrons in the valence shell are less strongly held due to increased shielding effects. Thus, those valence electrons are lost more easily and the halide ions are readily oxidized. Among halides, the iodide ions are the most powerful reducing agent followed by bromide ions. Fluoride ions have no significant reducing properties. Thus, the reducing power of the halides increases down the group.

In the reaction equation between chlorine and bromide, chlorine undergoes the reduction by gaining electron while bromide ion undergoes oxidation by losing electron. In other word, chlorine is said to be reduced while bromide ion is oxidized. The kind of reaction in which oxidation and reduction occurs simultaneously is called redox reaction.

$$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$$

Oxidation

Reduction

### iv. As bleaching agent

The term bleaching is defined as the process of removing the colours from coloured organic matter by using chemical agents or upon exposure to sunlight. The chemicals used in the process are called bleaching agent. The bleaching action of chlorine is due to oxidation and hence, the change is permanent.

**Mechanism of bleaching actions:**

1. Moisture of coloured organic matter combines with chlorine to form hydrochloric acid (HCl) and Hypochlorous acid (HClO).

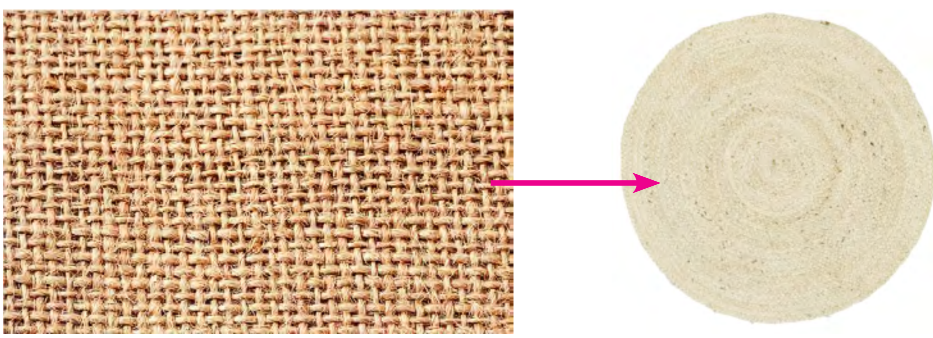
$$H_2O + Cl_2 \rightarrow HCl + HClO$$

2. HClO being unstable decomposes to produce nascent oxygen [O].

$$HClO \rightarrow HCl + [O]$$

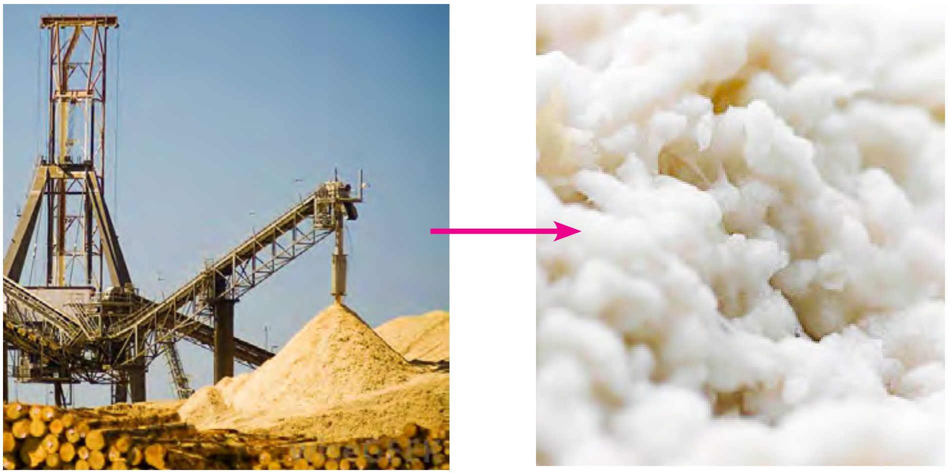
3. The nascent oxygen oxidizes the colouring matter to a colourless compound.

$$\text{Colouring matter} + [O] \rightarrow \text{Colourless compound}$$



Raw jute product

Bleached jute product



Raw wood pulp Bleached wood pulp

Figure 4.11 Bleached products

### v. Displacement reaction.

The displacement reaction of halogens with halide ions provide a clear illustration of the trends in oxidizing power of the halogens and the reducing power of the halides in aqueous solution. The more reactive halogens displace the ions of the less reactive halogens from its compound or in other word the halogen at top will displace any halide ion below it. For example, when chlorine is bubbled through a solution of potassium bromide, the solution changes from colourless to orange due to production of bromine:

$$2KBr(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + Br_2(aq) \quad \text{(Molecular equation)}$$

$$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq) \quad \text{(Ionic equation)}$$

**Reaction mechanism**: Since chlorine atom has strong affinity for electron than bromine atom, it takes electron from a bromide ion (i.e., potassium bromide) to form a chloride ion. On the other hand, bromide ion loses an electron to form bromine atom. Similarly, chlorine and bromine displaces iodide ions from its solution and the colour changes from colourless to dark orange or brown owing to the formation of iodine.

$$Cl_2(aq) + 2I^-(aq) \rightarrow Cl^-(aq) + I_2(aq)$$

$$Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$$

Smaller the size of halogen atom, the stronger is its oxidizing power. Thus, the reactivity of elements in halogen family decreases down the group due to increase in atomic size.

## Activity 4.4 Investigating the displacement reaction.

### Materials required

1. Solutions of chlorine, bromine and iodine.
2. Solution of potassium chloride, potassium bromide and potassium iodide.
3. Test tubes, test tube rack, test tube brush and dropper or Pasteur pipettes.

### Procedure

1. Place about 2 $cm^3$ each of KCl, KBr and KI solution in three separate test tubes.
2. To each of these solution add chlorine water drop by drop and record colour changes observed if any or otherwise record it as "No reaction" if there is no colour changes after adding up to nine drops. (Shake each test tube gently to mix the content after every drop).
3. Repeat procedure 1 and 2, but this time replace chlorine water with bromine water in procedure 2.
4. Once again repeat procedure 1 and 2, but this time replaces chlorine water with iodine water in procedure 2.

### Safety precautions

Halogens are highly toxic and corrosive to skin. Do not inhale any vapors of halogen. Wear safety goggles and PVC glove to handle halogens. Keep lab with good ventilation.

### Results

Retain all the content of test tubes in a rack and record your observation in Table 4.7.

**Table 4.7 Observation table**

| Halogen solution | Halides solution | | | |
|---|---|---|---|---|
| | KCl solution | KBr solution | KI solution | Conclusion |
| Chlorine water | | | | |
| Bromine water | | | | |
| Iodine water | | | | |

### Questions

1. Write down the names of halogen in order of their increasing reactivity.
2. How many boxes in the Table 4.9 are filled with ‘No reaction’? Explain why is the reaction not possible in each combination?
3. What results might you get if fluorine water and potassium fluoride were in the experiment too?
4. Write the balanced chemical equation for the reaction taking place between
   (a) chlorine and potassium bromide
   (b) bromine and potassium iodide.

### vi. Reactions with alkali metals.

The halogen atoms and the alkali metals are both unstable. Thus, the halogen atoms react readily with the elements of alkali metals. During the course of reaction, one electron lost by the alkali metals is gained by the halogen atoms to form metal halides.

$$M(s) + X(g) \rightarrow M^{+}X^{-}(s)$$
$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$
$$2K + Br_2(g) \rightarrow 2KBr(s)$$

### vii. Reaction with water

In the reactivity series of halogen, fluorine and chlorine have larger reduction potentials, and can oxidize water.

a) Fluorine reacts with water vapor to form oxygen and ozone along with hydrofluoric acid.

$$2F_2(g) + 2H_2O(g) \rightarrow 4HF(g) + O_2(g)$$
$$3F_2(g) + 3H_2O(g) \rightarrow 6HF(g) + O_3(g)$$

b) Chlorine dissolves in water to small extent. Thus, chlorine reacts with water to from a solution which contains two acids (i.e., hydrochloric acid and hypochlorous acid). This solution is called chlorine water.

$$Cl_2(g) + H_2O(l) \rightarrow HClO(aq) + HCl(aq)$$

Oxidation

Reduction

If chlorine water is tested with blue litmus paper, it will first turn red due to acidity of hydrochloric acid, but then is rapidly decolorized by the action of the hypochlorous acid, which acts as a bleaching agent. This solution has disinfectant properties and is used in swimming pools, water supplies, household cleaners, etc., in a variety of concentration. Iodine and bromine do not react with water because they have smaller reduction potentials than oxygen.

### viii. Reaction with hydrogen

All halogens react directly with hydrogen to form hydrogen halide. For example, a mixture of moist hydrogen and chlorine directly combine in the presence of diffused sunlight to form hydrogen chloride. The reaction is explosive in direct sunlight. In the dark, no reaction occurs, so activation of the reaction by light energy is required.

$$H_2 + Cl_2 \xrightarrow{\text{diffused sunlight}} 2HCl$$

**Self Evaluation**

1. Explain with reasons for the following statements.
   (a) The size of chloride ion is larger than the chlorine atom.
   (b) Electronegativity decrease as we move down the halogen group.
   (c) Iodine is denser than the fluorine.
   (d) Chlorine is used as a bleaching agent in paper industry.
   (e) Mixture of hydrogen and chlorine should not be exposed to direct sunlight.
2. Differentiate between
   (a) oxidation and reduction
   (b) oxidizing agent and reducing agent
3. Is the reaction given below feasible? Justify.
   $Br_2(aq) + 2NaCl(aq) \rightarrow Cl_2(aq) + 2NaBr(aq)$
4. Identify the following reaction as oxidation and reduction.
   (a) $F_2(g) + 2e^- \rightarrow 2F^-(aq)$
   (b) $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$
   (c) $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$
   (d) $2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$

## 4.4 Uses of halogens

**Learning objectives**

On completion of this topic, students should be able to:

» *state the uses of halogens.*

» *prepare tincture of iodine.*

» *explain the ill effects of chlorofluorocarbons (CFCs).*

### 4.4.1 Fluorine

The fluorine atom being highly reactive can reacts with almost all the elements including some of the noble gases under extraordinary conditions. Even the water will burst into flame if fluorine is bubbled through it. Although elemental fluorine is highly toxic, its fluoride form has wide range of uses as discussed in following points:

**Do you know?**

Davy first identified fluorine as an element, but was poisoned while trying unsuccessfully to decompose hydrogen fluoride. Two other chemists were also later poisoned in similar attempts, and one of them died as a result.

a) It helps to prevent the tooth decay and hence, is added in toothpaste in the form of fluoride salt (eg., sodium fluoride).

b) It is used to manufacture the tough non-stick plastic called Teflon, to coat cooking pans.

c) Chlorofluorocarbon (CFCs) is used as refrigerants and propellants. However, excessive use of CFCs has negative impact to the environment by depleting the ozone layer in the atmosphere.

d) It is used as an etching agent in a glass industry.

**Tooth decay and fluorine**

Our teeth have a hard outer layer called enamel which is mainly calcium carbonate, $CaCO_3$, and hydroxyapatite, $[Ca_3(PO_4)_2]_3.Ca(OH)_2$. Formation of lactic acid ($C_3H_6O_3$), is the main cause of tooth decay. It is formed when bacteria in saliva feed on sugars present in the sticky plaque on tooth surfaces. An increase in $H^+$ concentration causes the minerals in tooth enamel to decay faster.

Most often fluoride ions are added to the water supply in many cities. Likewise, most of the toothpaste contains fluoride ions which replaces hydroxide ions in hydroxyapatite to form fluoroapatite, $[Ca_3(PO_4)_2]_3.CaF_2$. This replacement makes the enamel more resistant to decay. Fluoride ions alone won't prevent tooth decay. Brushing and flossing teeth after every meal keeps away plaque from building up on tooth enamel.

## 4.4.2 Chlorine

Chlorine is a highly poisonous gas with a sharp smell. However, it can combine with other elements to form compounds safer for human consumption. The most notable compound of chlorine used daily is table salt (NaCl). Some of the other uses are shown in Figure 4.12.

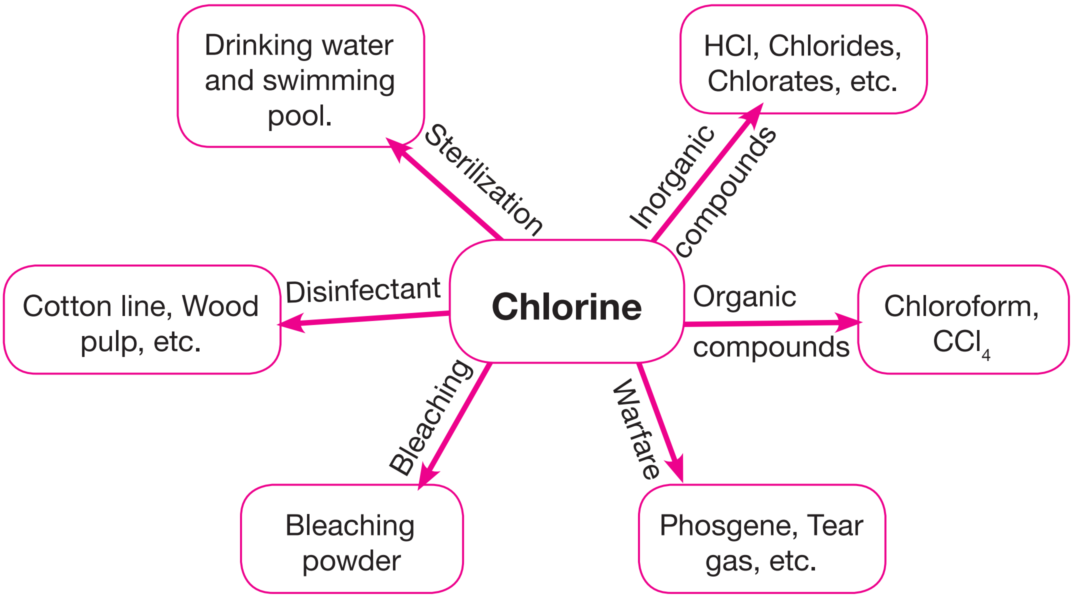



Figure 4.12 Uses of chlorine

**Sterilizing swimming pool and Chlorine**

Sterilization of swimming pool has two main purposes; prevent the growth of disease causing bacteria and foul causing algae. Thus, chlorine compounds are used to sterilize pool water. Liquid chlorine contains sodium hypochlorite, NaClO. Dry chlorine is calcium hypochlorite, $Ca(ClO)_2$. When hypochlorite ions dissolve in water, hydrolysis occurs and weak hypochlorous acid, HClO, is produced.

$$ClO^-(aq) + H_2O(l) \rightarrow HOCl(aq) + OH^-(aq)$$

The amount of undissociated hypochlorous acid in the pool water depends on the pH. If the pH is too high, the hydrolysis reaction will shift toward the reactants and reduce the concentration of HClO. If the pH is too low, too much acid will form. A high concentration of acid can cause eye irritation, damage plaster, and corrode the metal piping and filters in the pool.

If the pH of the pool water is too high, solid sodium hydrogen sulphate can be used to react with the $OH^-$ ions.

$$NaHSO_4(s) + OH^-(aq) \rightarrow Na^+(aq) + SO_4^{2-}(aq) + H_2O(l)$$

If the pH is too low, sodium carbonate can be used to neutralize some of the acid.

$$Na_2CO_3(s) + 2H^+(aq) \rightarrow 2Na(aq) + H_2O(l) + CO_2(aq)$$

## 4.4.3 Bromine

Bromine readily transform into gas since it has a boiling point of 58.8°C which is much lower than the water. Like any other halogens, its vapors are highly irritating to the eyes and throat. Some of the most significant uses are as follow.

a) It is used to purify the swimming pools.

b) It is used to manufacture organic pesticides and fungicides.

c) It is used as flame retardants for plastic products.

d) Both bromine and iodine are used in car head light.

## 4.4.4 Iodine

The iodine in iodide form is best known for the prevention of goiter. The thyroid gland produces a growth-regulating hormone that contains iodine, and lack of it may result to a goiter. The iodine required by our body is supplied by the table salts (NaCl). Thus, sodium iodide or potassium iodide is added to table salts to meet the requirement. Some of the other uses are as follow.

a) It is used in preparation of tincture of iodine.

b) It is also used as a food supplement in animal feed.

c) It is used to make dyes and for development of photography in photographic film.

### Activity 4.5 To prepare tincture of iodine

**Theory**

The tincture of iodine usually contains 2 to 7% of elemental iodine along with potassium iodide or sodium iodide dissolved in a mixture of ethanol and water. It is a weak solution of iodine and is often used as an antiseptic during pre-operative skin preparation.

**Materials required**

1. Spatula
2. Glass rod
3. Measuring cylinder (500 mL)
4. Beaker
5. 10 g of iodine crystals
6. 12 g of sodium iodide
7. 250 ml of ethanol
8. Distilled water


Figure 4.13 Tincture of iodine and its application

**Procedure**

1. Dissolve 10 g of iodine and 12 g of sodium iodide in 250 mL of ethanol.
2. Add enough distilled water to make solution up to 500 mL.

**Identification**

1. Tincture of iodine is dark red-brown liquid and has characteristic odour.

## 4.4.5 Astatine

Astatine is the rarest of all the elements. This is because astatine isotopes are radioactive with short half-lives. It is used to treat a condition known as hyperthyroidism, a disease related to a highly active thyroid gland.

### Self Evaluation

1. Write chemical equations to show sterilization of water.
2. What causes the tooth decay?
3. Why are fluoride salts used in the tooth paste?
4. What is tincture of iodine? Why is it used during pre-operative skin preparation?
5. State one physical property to identify tincture of iodine.

## Summary

1. The elements of group VIIA or 17 (i.e., fluorine, chlorine, bromine, iodine and astatine) constitute the family called halogen.
2. Halogens have seven electrons in the valence shell (i.e., one electron short of noble gas configuration).
3. Halogen atoms combine with the metallic atoms to form ionic compounds and with the non-metallic atoms to form covalent compounds.
4. The size of halide ion is larger than their parent halogen atom.
5. Fluorine has the highest electronegativity value in a periodic table.
6. Halogen atoms have the highest ionization energy in their respective period.
7. Ionization energy decreases down the group due to increase in atomic size.
8. The melting and boiling points of halogen atoms increases steadily down the group.
9. The reactivity of halogens decreases down the group. A more reactive halogen can displace a less reactive halogen from its salts.
10. Oxidation is a process involving the loss of electrons while reduction is a process involving the gain of electrons.
11. Halogens in its molecular state acts as an oxidizing agent while its ionic state acts as a reducing agent.
12. Oxidizing property of chlorine is widely used as a bleaching agent to remove the colours from the coloured organic matters.
13. Tincture of iodine is used as an antiseptic during pre-operative skin preparation.
14. Chlorine is used as disinfectants to sanitize the drinking water and the swimming pools.

# Exercise

## I. Fill in the blanks with correct word(s).

1. The atomicity of halogen is ______.
2. The halogen used in non sticky cooking pan is _______.
3. The more reactive halogens will _____ the less reactive halides from its compounds in solution.
4. The halogens form ______ compound with non-metals.
5. Those substances which are capable of undergoing oxidation are the _____ agents.

## II. State whether the following statements are True or False.

1. The family name given to the element belonging to 3rd period of 17th group is alkaline earth metal.
2. Chemical combination of sodium and bromine results into ionic bonding.
3. Fluorine in fluoride form is used to prevent the tooth decay.
4. Halogen atoms are the most reactive non-metals in their respective period.
5. Halogens require two electrons to acquire its noble gas configuration.
6. Atoms with stronger effective nuclear charge have greater electron affinity.

## III. Match the items of Column I with the corresponding items of Column II.

| Column I | Column II |
|---|---|
| 1. Fluorine | a. Radioactive. |
| 2. Chlorine | b. Sterilizing wound. |
| 3. Bromine | c. Prevent tooth decay. |
| 4. Iodine | d. Sterilizing water. |
| 5. Astatine | e. Making pesticides and plastics. |

## IV. Choose the most appropriate response from the given options.

1. Which type of bond is formed when electrons are transferred from one atom to another?

   A Covalent bond. C Ionic bond.
   B Hydrogen bond. D Metallic bond.

2. Which is the order of increasing reactivity of the halogens?

   A $F < Cl < I < Br$. C $I < Cl < Br < F$.
   B $F < Cl < Br < I$. D $I < Br < Cl < F$

3. Which of the following is correct electron configuration for chloride ion?
   - A 2, 8, 7
   - B 2, 8, 6, 1
   - C 2, 8, 8
   - D 2, 8, 7, 1
4. Which of the following compounds has high polarity?
   - A H–Br
   - B H–F
   - C H–Cl
   - D H–I
5. What happens to the trend in atomic radius and metallic character of halogen family?
   - A The atomic radius and the metallic character both increase.
   - B The atomic radius increases and the metallic character decreases.
   - C The atomic radius decreases and the metallic character increases.
   - D The atomic radius and the metallic character both decrease.
6. Which property of the halogens increases from fluorine to iodine?
   - A Effective nuclear charge
   - B Melting point
   - C Electronegativity
   - D Chemical reactivity
7. In the displacement reactions between halogens and halides, the halogen acts as an oxidizing agent. This means that the halogen
   - A oxidizes the halide ion to the halogen.
   - B gains the electron.
   - C is reduced to form the halide ion.
   - D loses the electron.
8. Which of the following pair would react more readily?
   - A K + Br
   - B K + I
   - C K + F
   - D K + Cl

## V. Explain the following statements.

1. The atomic radius of a chloride ion is larger than the chlorine atom.
2. The elements in halogen family have similar properties.
3. The halogens become less reactive down the group.
4. These reactions does not occur:

   (a) $I_2(aq) + 2Cl^-(aq) \rightarrow 2I^-(aq) + Cl_2(g)$

   (b) $I_2(aq) + 2Br^-(aq) \rightarrow 2I^-(aq) + Br_2(g)$

## VI. Write answers for the following questions.

1. Write down the electron configuration of
   (a) fluorine.
   (b) phosphorus.
   (c) argon.
2. Answer the following questions in reference to the unknown element 'X' with its electronic configuration 2, 8, 7.
   (a) To which period and group does the element 'X' belong? Why?
   (b) What is the valency of the element 'X'? Explain.
   (c) Is the element 'X' electropositive or electronegative? Why?
3. Name the appropriate halogen in each case.
   (a) The rare and radioactive.
   (b) The only non-metal liquid at room temperature.
   (c) Used as bleaching agent.
   (d) Sublimes at room temperature.
4. Label (a), (b), (c) and (d) with an appropriate terms given in the bracket (reducing agent, formed by oxidation, oxidizing agent, formed by reduction).

$$\underset{(a)}{Cl_2} + \underset{(b)}{2Br^-} \longrightarrow \underset{(c)}{Br_2} + \underset{(d)}{2Cl^-}$$

5. State all the possible advantages and disadvantages of chlorination of drinking water?
6. How do the reactivity of alkali metals and halogens vary down the group?

## VII. Solve the cross word puzzle.

**Across**

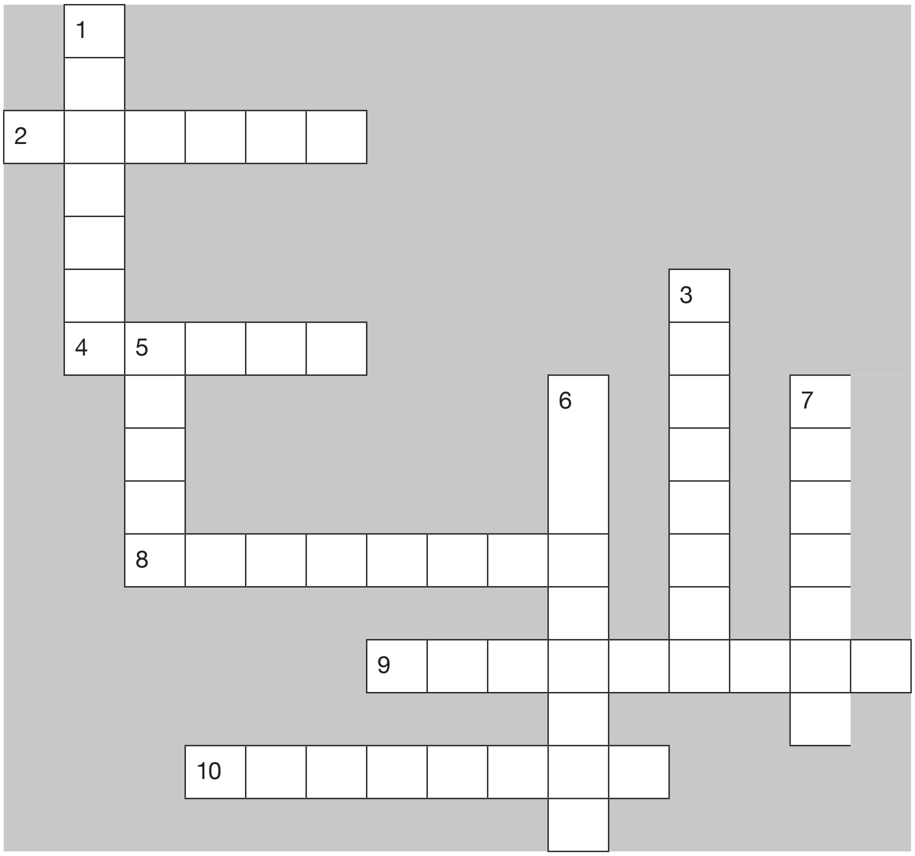
2. The halogen used as a food supplement in the animals feed.
4. The number of electrons present in the valence shell of halogen ions.
8. This halogen combines with an alkali metal to form table salt.
9. The process of being oxidized.
10. The most reactive non-metal.

**Down**

1. This halogen is a liquid at room temperature.
3. The types of bond existed in halogen molecules.
5. The type of compounds formed when halogen combine with metals.

6. The electrons around the nucleus other than valence electron are called ___ electron.

7. An organic compound used in preparation of the tincture of iodine







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# 5 Transition Elements

## 5.1 Introduction

The transition elements, most commonly refers to as a d-block element includes the elements of group 3 to 12 in a modern periodic table. The d-block elements are called transition elements as their properties are intermediate between the elements of s-block, which typically form ionic compounds, and the element of p-block, which largely form covalent compounds. The d-block elements, either in their atomic state or in any of their common oxidation state, the last electron enters the d-orbital which leaves them incompletely filled. This incompletely filled d-orbital in most transition elements is accounted for exhibiting the typical metallic properties and, have offered wide range of applications in biological systems and modern technologies.

## 5.2 Electron configuration and position in periodic table

**Learning objectives**

On completion of this topic, students should be able to:

» *classify the transition elements into four series.*

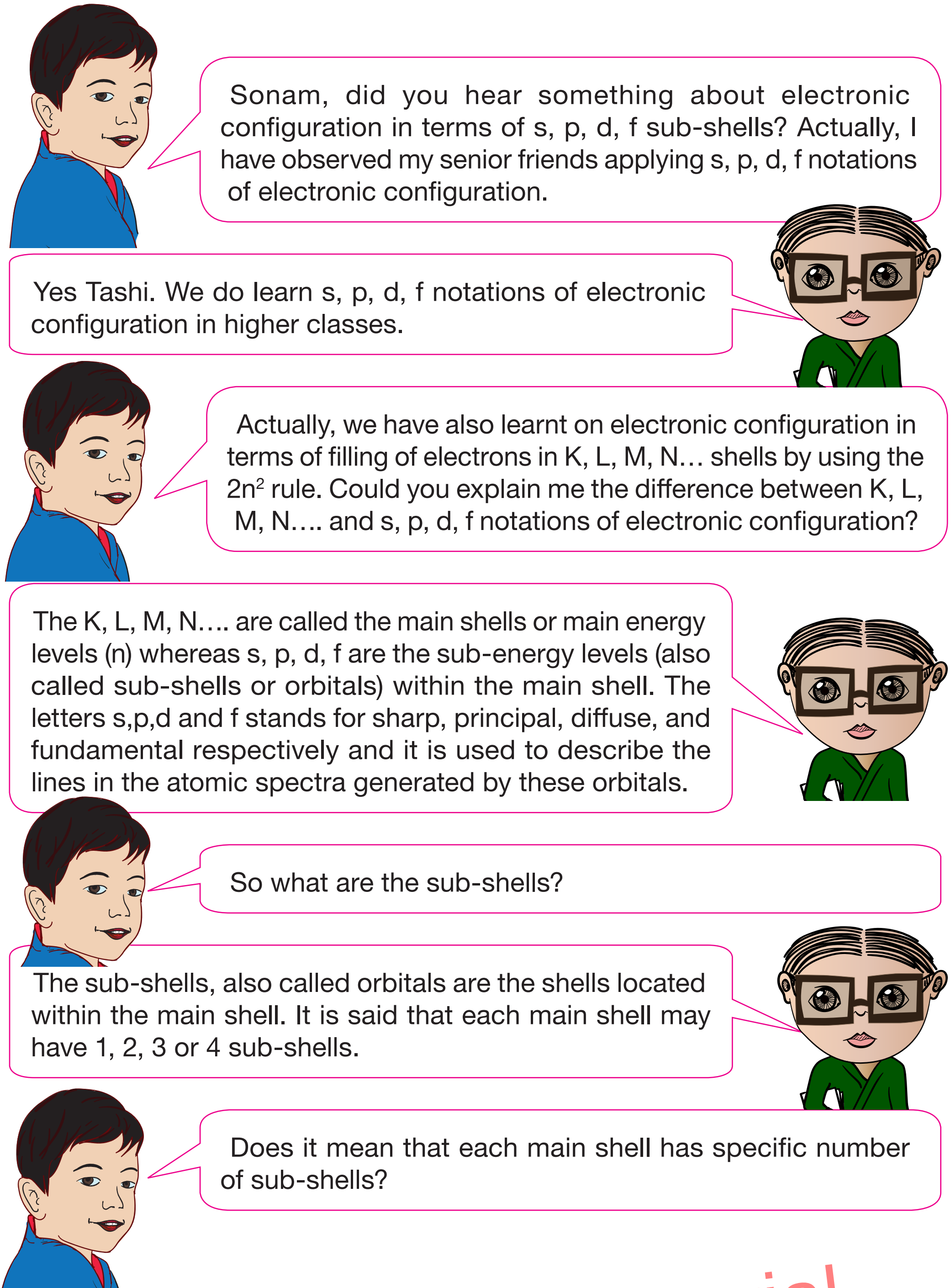
» *identify the position of transition elements in a periodic table.*

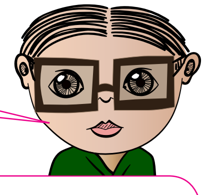
» *apply Aufbau's principle in writing electronic configuration in s, p, d, f notations.*

### 5.2.1 Electron configuration in s, p, d, f orbital notation

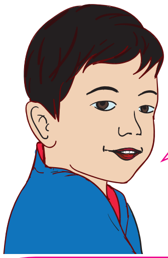
The concept of s, p, d, f notations of electronic configuration is required for the detail study of atomic structure. We know that the electrons in an atom are revolving round the nucleus in a fixed path called shell. In s, p, d, f notations of electronic configuration, we would come to know that each shell has sub-shells in which electrons are located.

The following cartoon illustration introduces the concept of electron configuration in terms of s, p, d, f notations.

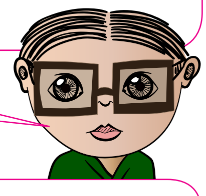




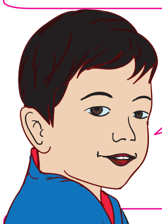
Yes, as you said.  Look at the Figure 5.1.



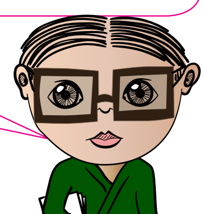
What I understood from the Figure 5.1 is that; K shell has one sub-shell (s -orbital), L shell has two sub-shells (s and p-orbitals), M shell has three sub-shells (s, p and d-orbitals) and N shell has four sub-shells(s, p, d and f-orbitals). Am I right?



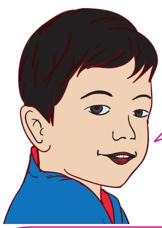
Yes you are right.



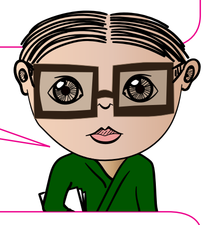
What is the electron holding capacity of each sub-shell? Is there any other direct way to calculate the electron holding capacity of each sub shell?



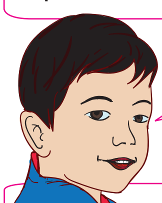
To find the electron holding capacity for s, p, d and f orbitals, we have other principles which are dealt in higher classes. For now you can refer Figure 5.1 and Table 5.1.



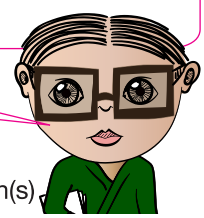
In terms of K, L, M, N… notations of electronic configuration, Na (Z=11) has its electronic configuration as 2, 8, 1. Sonam, could you show me how to write the electronic configuration in terms of s, p, d and f notations for Na?



Yes, definitely.  The electron configuration for Na is $1s^2 2s^2 2p^6 3s^1$. You could refer Table 5.2.



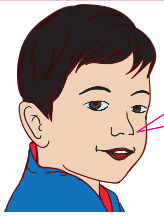
What does each numeral, letter and superscript represent in the configuration of this type?



Look at the illustration given below with respect to $1s^2$.

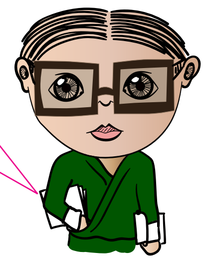
Number of electron(s)

Main shell or energy level → $1s^2$ ← Orbital or sub-shell



Is there any sequence or rule while filling the orbitals with electrons?

Oh! Yes, I nearly forgot. There are few principles which govern the filling of orbitals with electrons. One such principle is Aufbau's Principle. According to this principle, the electrons are filled to the orbitals one by one in order of increasing energies. This means that the electrons starts filling the orbitals of the lowest energy level first and continues to fill up the orbitals of higher energy levels. The order of filling of orbitals as per Aufbau's principle is: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, etc. You could refer Figure 5.2.



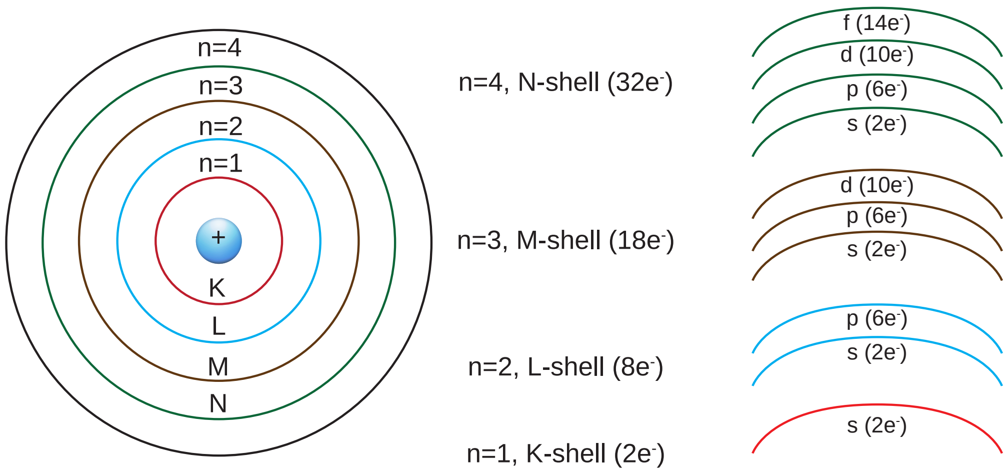



Figure 5.1 Electron holding capacity.

**Table 5.1**

| Orbital | s | p | d | f |
|---|---|---|---|---|
| Electron holding capacity of each orbital | 2 | 6 | 10 | 14 |

### Table 5.2 Electron configuration of sodium atom.

| Main shell (energy level(n) | K (n=1) | L (n=2) | | M (n=3) | | | N (n=4) | | | |
|---|---|---|---|---|---|---|---|---|---|---|
| Orbital(s) in each main shell. | s | s | p | s | p | d | s | p | d | f |
| Electronic configuration in terms of s, p, d, and f orbital notation. | $1s^2$ | $2s^2$ | $2p^6$ | $3s^1$ | | | | | | |

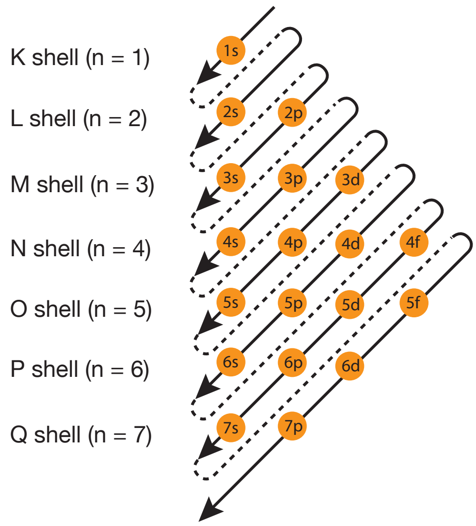



Figure 5.2 Order of filling of orbitals (Aufbau's Principle).

In the electronic configuration of copper (i.e., $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$), 4s orbital is filled before filling of the 3d orbital. This would mean that the energy level of 3d orbital is higher than the energy level of 4s orbital.

## 5.2.2 Position in a periodic table

In a modern periodic table, the elements are classified into four blocks, i.e., s, p, d and f block as shown in the Figure 5.3. The classification is based upon the name of orbital where the last electron enters, which is discussed as follow:

### i. s-Block elements

Those atoms of the elements in which the last electron enters the s-orbital are called s-block elements. Thus, the elements of group 1 and 2, including hydrogen and helium in which s-orbitals are progressively filled in are s-block elements.

| | |
|---|---|
| $_{1}H$ | $1s^1$ |
| $_{2}He$ | $1s^2$ |
| $_{4}Be$ | $1s^2\ 2s^2$ |
| $_{11}Na$ | $1s^2\ 2s^2\ 2p^6\ 3s^1$ |

### ii. p-Block elements

Those atoms of the elements in which the last electron enters the p-orbitals of their outermost shell are called p-block elements. It constitutes the elements from group 13 to 17 in which p-orbitals are progressively filled.

| | |
|---|---|
| $_{5}B$ | $1s^2\ 2s^2\ 2p^1$ |
| $_{7}N$ | $1s^2\ 2s^2\ 2p^3$ |
| $_{9}F$ | $1s^2\ 2s^2\ 2p^5$ |
| $_{10}Ne$ | $1s^2\ 2s^2\ 2p^6$ |

### iii. d-Block elements

Those elements between s-block and p-block, i.e., from group 3 to 12 comprise the d-block elements. In these elements, the last electron enters the d-orbital, i.e., the second last shell. Usually, the d-orbitals of this block are incompletely filled. For example, the electronic configuration of $_{26}Fe$ is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$. The last electron, during filling of the orbital is entering d-orbital of second last shell and it remains incompletely filled.

The d-block elements are classified into four series as discussed below:

a) 3d-series (1$^{st}$ transition series): This series constitute ten elements, beginning with Sc-21 and ending at Zn-30. The elements of this series lie in 4$^{th}$ period.

b) 4d-series (2$^{nd}$ transition series): This series constitute ten elements, beginning with Y-39 and ending at Cd-48. The elements of this series lie in 5$^{th}$ period.

c) 5d-series (3$^{rd}$ transition series): This series also have ten elements, beginning with La-57, Hf-72 to Hg-80. The elements of this series lie in the 6$^{th}$ period.

d) 6d-series (4$^{th}$ transition series): This series forms a part of the 7$^{th}$ period and contains elements from actinium, Ac – 89 and those beyond rutherfordium, Rf -104.

According to IUPAC, transition elements are those which have incompletely filled d-orbitals in their ground state (atomic state) or in any of their oxidation states. Taking this definition into account, not all of the d-block elements are transition element. For example, zinc and scandium are not transition metals although they are the members of d-block elements. This is because $Zn^{2+}$ has its d-orbital completely filled, while $Sc^{3+}$ has no electron in d-orbital.

In case of copper, it forms two ions (i.e., $Cu^{+}$ ion and $Cu^{2+}$ ion). In $Cu^{+}$ ion, it has outer configuration of $3d^{10}$ which has completely filled d-orbital, while $Cu^{2+}$ has $3d^{9}$ in outer configuration. Thus, copper is regarded as transition element since $Cu^{2+}$ ion having incompletely filled d-orbital is the most common form.

### iv. f-Block elements

This block elements lie at the bottom of the table, which includes the elements in Lanthanoids and Actinoids series. These elements are commonly known as inner transition elements or rare earth elements. Those atoms of the elements in which the last electron enters the f-orbitals of their outermost shell are called f-block elements.

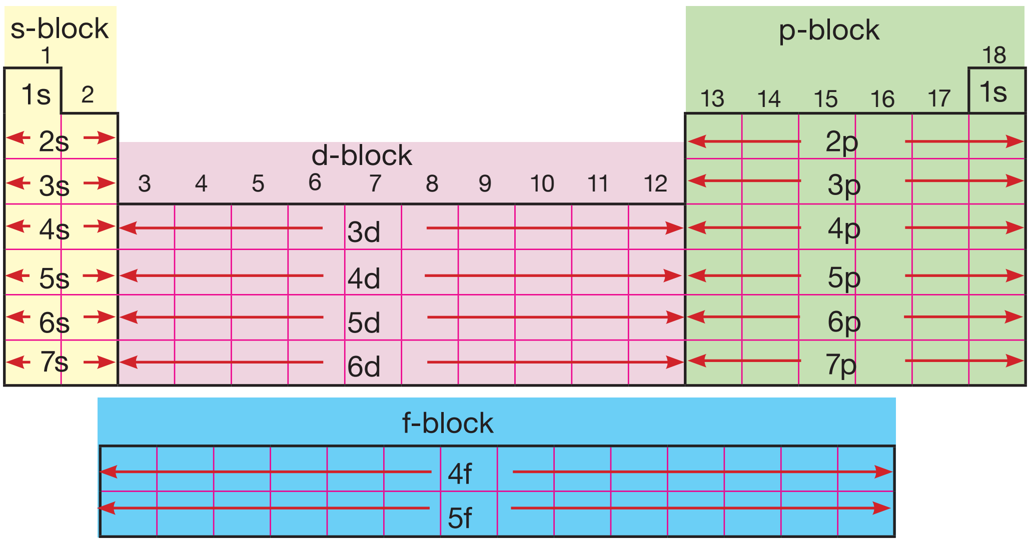



Figure 5.3 Periodic table showing different blocks

**Self Evaluation**

1. Write electron configuration for the following elements in s, p, d, f notations by using Aufbau's principle.

   a) Copper b) Zinc c) Iron d) Manganese e) Sulphur f) Calcium

2. Define Aufau's principle.
3. Which orbital is of higher energy level in following case:
   (a) 2s and 2p
   (b) 3p and 3s
   (c) 4s and 3d
   (d) 5s and 4d

## 5.3 Characteristics of transition elements

**Learning objectives**

On completion of this topic, students should be able to:

» *discuss the characteristics of transition elements.*

» *explain the reasons for observed trends among the transition elements.*

» *explain why d-block elements exhibit variable valencies.*

» *discuss the similarities and the differences between d-block and s-block elements.*

### i. Metallic character

Since the valence electron in transition elements is two, they are good conductor of heat and electricity. Normally they are hard, ductile, and malleable and have a metallic lustre. They also form alloys with other metals. The transition elements can form both covalent and metallic bonding. The presence of unfilled d-orbital favours covalent bonding.

### ii. Melting and boiling point

The transition metals being heavy metals exhibit very high melting point and the boiling point. Transition elements typically melt above 1000°C except Zn, Cd and Hg as they have completely filled d-orbital.

### iii. Colour

All transition metals of first row except zinc (Zn) form coloured ions. Since transition metals have incompletely filled d-orbitals, it is possible to promote the electrons from lower energy level to higher energy level. This process is accompanied by the emission of radiation from which the compounds absorb a particular colour. However, some elements other than Zn also appear colorless depending on their oxidation state. For example, $Sc^{3+}$, $Ti^{4+}$ and $Cu^{+}$ have completely filled d-orbitals and hence they appear colourless.

Many naturally occurring substances like minerals and gem stones are coloured due to the presence of transition metal ions or compounds. For example, blue aquamarine is due to iron compounds, green emeralds due to iron and titanium ions, red and blue sapphires are due to traces of iron, titanium, chromium and copper ions and red rubies due to chromium compounds.

### iv. Ionization potential

The ionization potential of transition elements is intermediate between the elements of s-block and p-block. Thus, the elements of d-block are less electronegative than the s-block but are more electronegative than the p-block. As a result, these elements do not form ionic compounds as readily as the s-block elements. The tendency to form ionic compound decreases as the size of the atom increases.

### v. Atomic volume and densities

The atomic volumes of transition metals are smaller than the metals of group 1 and 2 due to increased nuclear charge. The decrease in atomic volume increases the density. Thus, the densities of transition metals are higher than the metals of groups 1 and 2.

### vi. Low reactivity

Most of the transition elements react with mineral acids, liberating hydrogen gas. The noble metals like platinum (Pt) and gold (Au) have low reactivity due to their high melting point, boiling point and ionization potential.

### vii. Magnetic properties

Most transition elements are paramagnetic in nature. The paramagnetic character of the transition metals increases on moving across the period as the number of unpaired electron increases from one to five. The middle elements are found to possess the maximum paramagnetic property. The magnetic property decreases with the decrease in the number of unpaired electrons. The transition metals which contain paired electrons depict diamagnetic behavior.

**Table 5.3 Types of magnetic substances**

| Magnetic substance | Paramagnetic substances | Ferromagnetic substances | Diamagnetic substances |
|---|---|---|---|
| Definition | Those substances which acquires magnetic property in the presence of magnetic field but loses their property when the magnetic field is removed. The substance shows paramagnetism when it contains one or more unpaired electrons. | Those paramagnetic substances which retain their magnetic property even upon removing the magnetic field. | Those substances which are repelled by The magnetic field. A substance shows diamagnetism when it contains only paired electrons. |

| Magnetic substance | Paramagnetic substances | Ferromagnetic substances | Diamagnetic substances |
|---|---|---|---|
| Example | Platinum, chromium, manganese. | Iron, nickel, cobalt. | Zinc, cadmium, copper. |

### viii. Variable oxidation state (valency).

Most of the 3d-block elements exhibit variable oxidation state. This is because the core or kernel left after their atom loses valence electrons is unstable, and tends to lose one or more electrons further. For example, iron exhibit two oxidation states.

$_{26}Fe$ = 2, 8, 14, 2

$_{24}Fe^{2+}$ = 2, 8, 14 (core or kernel)

The core being unstable loses one or more electron to give $Fe^{3+}$ ion

$_{23}Fe^{3+}$ = 2, 8, 13

Table 5.4 Oxidation state of 3d-block elements

| 3d series transition element | Sc | Ti | V | Cr | Mn | Fe | Co | Ni | Cu | Zn |
|---|---|---|---|---|---|---|---|---|---|---|
| Electronic configuration | $s^2d^1$ | $s^2d^2$ | $s^2d^3$ | $s^1d^5$ | $s^2d^5$ | $s^2d^6$ | $s^2d^7$ | $s^2d^8$ | $s^1d^{10}$ | $s^2d^{10}$ |
| Oxidation states | | | | +1 | | | | | +1 | |
| | | +2 | +2 | +2 | +2 | +2 | +2 | +2 | +2 | +2 |
| | +3 | +3 | +3 | +3 | +3 | +3 | +3 | +3 | +3 | |
| | | +4 | +4 | +4 | +4 | +4 | +4 | +4 | | |
| | | | +5 | +5 | +5 | +5 | +5 | | | |
| | | | | +6 | +6 | +6 | | | | |
| | | | | | +7 | | | | | |

### ix. Complex ion formation

The transition metals have high tendency to form complex ions (coordination compound) with the ligands.

**Complex ion** (coordination complex or coordination compound): It is an ion comprising of one or more ligands attached to a central metal ion by means of a coordinate bond.

**Ligand**: It is a species which can use its lone pair of electrons to form a coordinate bond with a transition metal. Ligands can be a neutral molecule such as $NH_3$, $H_2O$ or negatively charged ions such as $Cl^-$, $OH^-$, $CN^-$ etc.

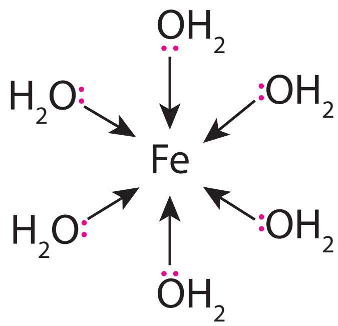


Figure 5.4 Complex ion

The formation of coordination compound is favoured if the ions of transition elements have following two features:

a) High charge density (i.e., smaller cationic size).

b) Vacant d-orbital to accept lone pairs of electron from the ligands.

$[Fe(H_2O)_6]^{2+}$ is an example of complex ion. In this complex, $Fe^{2+}$ is central metal ion and six molecules of water are the ligands. Each water molecule donates a lone pair of electrons from its oxygen atom to $Fe^{2+}$ ion to form a coordinate bond. The central ion is assigned with the number known as the coordination number depending upon the number of lone pair of electrons accepted from the ligands. In the $[Fe(H_2O)_6]^{2+}$, the coordination number is 6 because the central ion has accepted six lone pairs of electrons as shown in the Figure 5.4.

**Coordination number**: It is the number of coordinate bonds formed by the central metal ion by accepting the lone pairs of electron from the ligands during the formation of complex ion.

Some of the other examples of complex ions are $[Fe(CN)_6]^{4-}$, $[Cu(NH_3)_4]^{2+}$, $[Ni(CN)_4]^{2-}$, $[CoCl_4]^{2-}$, $[Cu(NH_3)_4(H_2O)_2]^{2+}$.

### x. Catalytic properties

Catalysts are the substances that alter the rate of chemical reaction; they generally speed up the chemical reactions. Some of the commonly used catalyst in chemical reactions are Pt, Ni, Fe, Cr, $MnO_2$, $V_2O_5$, etc. The catalytic property of transition elements is due to their vacant d-orbital and their ability to adsorb and activate reacting substances. The following processes illustrate the use of transition element as catalyst.

### a) Haber process

The process is used to prepare ammonia ($NH_3$) through a reaction between nitrogen and hydrogen.

$$N_2(g) + 3H_2(g) \xrightleftharpoons{Fe} 2NH_3(g)$$

In this process, iron (Fe) as a catalyst helps to increases the rate of reaction and lowers the temperature at which the reaction takes place.

### b) Contact process

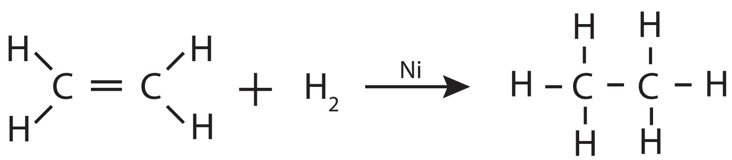
In this process, sulphur dioxide ($SO_2$) is converted into sulphur trioxide ($SO_3$) in manufacturing of sulphuric acid ($H_2SO_4$).

$$2SO_2(g) + O_2(g) \xrightleftharpoons{V_2O_5} 2SO_3(g)$$

Here, vanadium (V) oxide ($V_2O_5$) is used as a catalyst.

### c) Hydrogenation of alkenes

The hydrogenation reaction involves the addition of hydrogen to the reactant molecules. In this reaction, unsaturated hydrocarbon (alkene) is converted to saturated hydrocarbon (alkane) on addition of hydrogen across the C = C double bond.



(ethene) (ethane)

Nickel lowers the temperature and the pressure needed to complete the reaction.

### Self Evaluation

1. Iron exhibits variable valencies forming $Fe^{2+}$ and $Fe^{3+}$ ions. Write down the name and formula of each compound that iron can form with chlorine.
2. Vanadium (V) reacts with oxygen gas to form vanadium (v) oxide. Write a balanced equation for the reaction.
3. Describe the trends in melting point and boiling point among the transition element?
4. Define the following terms.
   (a) Ligands. (b) Coordination number.

## 5.4 d-Block elements of group 11 and the uses of transition elements

**Learning objectives**

On completion of this topic, students should be able to:

» *discuss the similiarities among the elements of group 11 or IB.*

» *recognize the variation of properties among the transition elements.*

» *state the uses of transition elements and their compounds*

The elements of group 11 (i.e., Cu, Ag and Au) in a periodic table are also called a coinage metals.

### 5.4.1 Similarities among copper, silver and gold.

Cu, Ag and Au belong to the same sub-group and show similar properties which are discussed in the Table 5.5.

Table 5.5 Similarities among Cu, Ag and Au

| Properties | Copper | Sliver | Gold |
|---|---|---|---|
| Physical properties | All three metals are hard, malleable and ductile. They have metallic lustre and take high polish. They have high density, melting points and low atomic volume. | | |
| Occurrence | All these metals are found in their native state in nature, owing to their inert character. | | |
| Electropositive character | All these elements are weakly electropositive. | | |
| Electronic configuration | The electronic configuration of these metals show that the ultimate shell has 1 electron ($1s^1$ electron) while the penultimate shell contains 18 ($s^2$ $p^6$ $d^{10}$) electrons which contribute to their inert nature. The electronic configurations of Cu, Ag and Au atoms are given below in the Table 5.6. | | |
| Variable valency | +1, +2 | +1, +2 | +1, +3 |
| Displacement reactions | Cu, Ag and Au atoms are not able to displace hydrogen from acids, as they are less electropositive, but metals like Zn, Fe and Mg can displace these metals from their salt solution. | | |
| Formation of complex salts. | $[Cu(NH_3)_4SO_4]$ (Tetrammine copper (II) Sulphate) | $[KAg(CN)_2]$ (Potassium argentocyanide) | $[KAu(CN)_2]$ (Potassium aurocyanide) |

### Table 5.6 Electronic configuration

| Element | At. No. | Complete Electron configuration | Valence-shell configuration |
|---|---|---|---|
| Cu | 29 | 2,8,18,1 or [Ar] $3d^{10}\ 4s^{1}$ | $3d^{10}\ 4s^{1}$ |
| Ag | 47 | 2,8,18,18,1 or [Kr] $4d^{10}\ 5s^{1}$ | $4d^{10}\ 5s^{1}$ |
| Au | 79 | 2,8,18,32,18,1 or [Xe ] $4f^{14}\ 5d^{10}\ 6s^{1}$ | $4f^{14}\ 5d^{10}\ 6s^{1}$ |

## Activity 5.1 Identification of copper ion in copper compounds by flame test

### Material required

$CuSO_4$ (any copper salts or compounds), watch glass, conc. HCl, platinum/nichrome wire and Bunsen burner.

### Procedure

1. Take 5 g sample of copper sulphate in a watch glass.
2. Moisten the sample with conc. HCl.
3. Clean platinum or nichrome wire by repeatedly dipping it into conc. HCl acid and holding it in a Bunsen burner, till the wire does not produce any colour in the flame.
4. With the help of this clean platinum or nichrome wire, pick very small amount of moist sample and place it in a flame.
5. Observe and record the result.
6. Repeat steps 1 to 5 for any other copper compounds.

### Safety Precaution

Conc. HCl is highly corrosive.

### Question

1. Why platinum or nichrome wire is used for the flame test?
2. Discuss some of the limitations in identifying ions using flame tests technique.
3. Why copper exhibits colour with flame?

## 5.4.2 Similarities of group 11 elements with other transition elements.

### i. Similarities of Cu with Zn

Although Cu (group 11) and Zn (group 12) are metals of different group, they exhibit some similar characteristics as discussed below.

a) Both the metals exhibit bivalency.

b) Sulphides, carbonates and phosphate of both the metals are insoluble in water, whereas their chloride and sulphates are soluble in water.

### ii. Similarities of Ag with Cd

a) Both the metals are white in colour.

b) Their oxides are coloured and get reduced when heated with carbon.

c) Both the metals form complex compounds, e.g. $[Ag(CN)_2]^-$, $[Cd(CN)_4]^{2-}$, $[Ag(NH_3)_2]^+$, $[Cd(NH_3)_4]^{2+}$.

### iii. Similarities of Au with Pt

a) These metals are noble metals hence, air, water, alkalis or acids have no action on them at ordinary conditions. Thus, they are used as raw materials for preparing the jewelries. However, they are attacked by fused cyanides and nitrates.

b) They occur in their native state in nature.

c) Both are heavy metals with high melting point and show metallic lustre. They are malleable, ductile and good conductor of heat and electricity.

d) Both metals dissolve in aqua regia, i.e., mixture of concentrated $HNO_3$ and HCl in the volume ratio of 1:3.

e) These elements show variable valency as given below:

Au +1, +3.

Pt +2, +4.

**Activity 5.2 Action of alkalis on some compounds of transition element.**

**Materials required**

Iron(II) Sulphate solution ($FeSO_4$), iron (III) sulphate solution $Fe_2(SO_4)_3$, copper (II) sulphate, zinc sulphate, sodium hydroxide, test tubes and test tube stand

and a dropper.

**Procedure**

1. Take small amount of iron (II) sulpahte solution in a test tube.
2. Then slowly add equal amount (5 mL) of sodium hydroxide to it.
3. Observe and record the result.
4. Repeat the same process for $CuSO_4$ solution and $ZnSO_4$ solution.

**Safety Precaution**

- Concentrated sodium hydroxide is corrosive in nature.
- Do not mix the dropper.

**Table 5.7 Observation Table**

| Transition metal compound. | Transition metal ion. | Colour of the precipitate formed when mixed with NaOH. | Name of the precipitate formed when mixed with NaOH. |
|---|---|---|---|
| Iron (II) Sulphate | $Fe^{2+}$ | | |
| Iron (III) Sulphate | $Fe^{3+}$ | | |
| Copper (II) Sulphate | $Cu^{2+}$ | | |
| Zinc Sulphate | $Zn^{2+}$ | | |

**Question**

1. Explain each of the observation and result with a balanced chemical equation.

## 5.4.3 Reaction involving transition elements

### i. Iron (Fe)

Iron is moderately reactive metal with steam and acids thereby displacing hydrogen from steam or acid to form hydrogen gas.

$$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2\uparrow$$

With the solution of iron (II) salt, a grey gelatinous precipitate of iron (II) hydroxide is formed on adding an alkali:

$$FeCl_2(aq) + 2NaOH(aq) \longrightarrow \underset{\text{(grey gelatinous precipitate)}}{Fe(OH)_2\downarrow} + 2NaCl(aq)$$

The precipitate is not affected by adding excess alkali. Also, the same precipitate is formed if ammonia solution is used instead of sodium hydroxide.

With solutions of iron (III) salts, a red-brown gelatinous precipitate of iron (III) hydroxide is formed when an alkali is added:

$$FeCl_3(aq) + 3NaOH(aq) \longrightarrow \underset{\text{(red gelatinous precipitate)}}{Fe(OH)_3\downarrow} + 3NaCl(aq)$$

### ii. Zinc (Zn)

Zinc is moderately reactive metals that will displace hydrogen from steam or dilute acids to form hydrogen.

$$Zn(s) + 2H_2O(g) \longrightarrow ZnO(s) + H_2\uparrow$$

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

When zinc (II) carbonate is heated, it decomposes to give zinc oxide with the evolution of $CO_2$.

$$\underset{\text{(white)}}{ZnCO_3(s)} \longrightarrow \underset{\text{(white)}}{ZnO(s)} + CO_2\uparrow$$

Solutions of zinc (II) salts produce a white precipitate of zinc (II) hydroxide when sodium hydroxide solution is added.

$$ZnSO_4(aq) + 2NaOH(aq) \longrightarrow \underset{\text{(white precipitate)}}{Zn(OH)_2\downarrow} + Na_2SO_4(aq)$$

Zinc hydroxide, is an amphoteric hydroxide and on addition of excess NaOH solution, $Zn(OH)_2$ dissolves with it to form sodium zincate, $Na_2Zn(OH)_4$ which is colourless.

$$\underset{\text{(white precipitate)}}{Zn(OH)_2\downarrow} + 2NaOH(aq) \longrightarrow \underset{\text{(colourless solution)}}{Na_2Zn(OH)_4(aq)}$$

**Amphoteric Compounds**

They are the molecules or ions that can act as an acid as well as a base. The word is derived from Greek word *amphoteroi* meaning 'both'. Many metals like Cu, Zn, Sn, Pb, Al and Be, form amphoteric oxides and hydroxides.

## 5.4.4 Uses of transition elements

Most of the transition metals and their compounds exhibit catalytic properties and have varied application in chemical industries. The noble metals like gold and silver are inert to air and water, and are used for making jewelries. Some of the transition metals are useful in our day to day life as shown in the Figure 5.5.

(a) Mercury: in thermometer.

(b) Tungsten: in filament.

(c) Titanium: in hip replacement.

Figure 5.5 Use of transition elements.

**Table 5.8 Uses of some transition metals and their compound.**

| Sl. No. | Transition metal and their compounds | Uses |
|---|---|---|
| 1 | $CuCl_2$ | Used in manufacturing $Cl_2$ from HCl by Deacon process. |
| 2 | Cu | Used as an electrodes during electrolysis. They are good conductors of heat and electricity and used for making calorimeter and electrical wires. |
| 3 | Fe | Iron is used as a catalyst in the Haber-Bosch process for manufacturing $NH_3$. |
| 4 | $FeCl_3$ | Used in the production of $CCl_4$ from $CS_2$ and $Cl_2$. |
| 5 | $MnO_2$ | Used as a catalyst in the laboratory preparation of oxygen from $KClO_3$ (or from decomposition of $KClO_3$). |
| 6 | Ni | Used in production of $H_2$ from $NH_3$ and production of $H_2O_2$. Used in polymerization of alkynes. Used as a catalyst in hydrogenation of vegetable fats and oil (vanaspati/ghee production) |

| Sl. No. | Transition metal and their compounds | Uses |
|---|---|---|
| 7 | Pd | Used in hydrogenation of phenol (aliphatic compound) to cyclohexanone (aromatic compound). |
| 8 | Pt / $V_2O_5$ | Used to oxidize $SO_2$ to $SO_3$ in the Contact process for manufacturing $H_2SO_4$. |
| 9 | Pt / Rh | Used as a catalyst in Ostwald process during preparation of $HNO_3$ to oxidize $NH_3$ to NO. |
| 10 | Zn | i. Used in galvanization of iron.<br>ii. It is employed as cathode plate in dry cell.<br>iii. An alloy of Zn with Cu is used in kitchen utensils, jewelries and ornaments, machine parts, statues etc.<br>iv. Granulated Zn is used in laboratory preparation of $H_2$. |

**Self Evaluation**

1. Fill in the blanks by using the word(s) given the bracket.

   (catalysts, coloured, conductors, densities, less, melting)

   The transition elements have high ________, high ________ points and are good ________. They are ________ reactive than the alkali metals, and often form ________ compounds. The transition elements and their compounds are useful ________ in the chemical industry.

2. Compare the following pairs of element with respect to their valency and electropositive character.

   (a) Ag and Au

   (b) Au and Pt

3. Name the transition metal(s) used in

   (a) Haber-Bosch process.

   (b) barometer and thermometer.

   (c) galvanization of iron.

   (d) making electrodes during electrolysis of acidified water.

   (e) Contact process.

## Summary

1. Transition elements are the elements of d-block series in periodic table.
2. Transition elements are classified into four series as 3d, 4d, 5d and 6d series corresponding to the filling of 3d, 4d, 5d and 6d orbitals.

3. The transition elements are located in-between the s-block and the p-block elements in a periodic table.
4. The properties of transition elements are intermediate between the highly electropositive elements of s-block and highly electronegative elements of p-block.
5. The transition elements are less reactive than the metals of groups 1 and 2.
6. During the filling up of electrons, the electrons are filled to the orbitals one by one in order of increasing energies.
7. Cations of d-block metals are small, have a high charge density and vacant orbital of low energy. Thus, they form complex ions readily.
8. Transition elements have certain distinctive properties from other elements.
   - the metals and their compounds often make good catalysts.
   - they have high melting and boiling point
   - the formation of coloured compounds.
   - the formation of complex compounds
   - they exhibit magnetic property
   - their variable valency
9. Paramagnetic substances are those substances which have tendency to acquire magnetic property in the presence of magnetic field.
10. Most of the transition elements are sufficiently electropositive and have low reactivity but there are some transition metals which are acted upon by acids, alkalis and water.
11. Many transition metals and their compounds are employed in chemical industry for the production of other compounds.
12. Some of the oxide and hydroxide of transition metals act as amphoteric compounds.

# Exercise

**I. State whether each of the statement is True or False.**

1. Sc and Zn are not transition metals.
2. Complex ions are often coloured.
3. Number of dative bonds to central metal ion is called oxidation number.
4. The location of transition elements is in between s-block and f-block series.
5. In p-block elements, the electron enters the d-orbitals.

## II. Match the items of Column I with the corresponding items of Column II.

| Column I | Column II |
|---|---|
| 1. Building materials, tools and vehicles. | a. Copper |
| 2. Water pipes and electric cables. | b. Iron |
| 3. In coins and catalyst in the manufacture of margarine. | c. Zinc |
| 4. Jewelry and plating of teeth. | d. Nickel |
| 5. Artificial hip joints, pipes in nuclear power stations. | e. Titanium |
| 6. Galvanization of steel. | f. Gold |
| | g. Silver |
| | h. Mercury |

## III. Choose the most appropriate response from the given options.

1. Which of the following ions are expected to be paramagnetic?

   A $_{29}Cu^{+}$ C $_{26}Fe^{2+}$

   B $_{30}Zn^{2+}$ D $_{21}Sc^{3+}$

2. In a periodic table, the element in a second row of transition elements beneath cobalt is

   A Zinc (Zn) C Palladium (Pd)

   B Nickel (Ni) D Copper (Cu)

3. The coordination number of complex ion, $[Fe(H_2O)_6]^{2+}$ is

   A 5 C 6

   B 4 D 3

4. The electronic configuration of $_{24}Cr$ is

   A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$

   B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$

   C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^5$

   D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$

5. Which of the following are not the characteristics of the transition elements and their compounds?

   A Exhibit more than one oxidation state

   B They form coordination compounds.

   C Many of their compounds are colored.

   D Most of the transition metals exhibit diamagnetic properties.

## IV. Give reasons for the following.

1. Transition metals have tendency to form complex compounds with the ligands.

2. The compounds containing transition metals are usually coloured.
3. Mercury is used as thermometric fluid.
4. Chromium is paramagnetic in nature.
5. 4s orbital is filled before 3d orbital according to Aufbau's principle.

## V. Write answers for following questions.

1. The chemical properties of zinc are similar to those of iron.
   (a) Write the chemical equation for the reaction of Zn and Fe with dilute hydrochloric acid.
   (b) Explain with equation the action of alkali on the solution of any salt of Zn and Fe.
   (c) What happens to the precipitate of both zinc hydroxide and iron hydroxide when excess of alkali is added?
2. Define the term galvanization. State its purposes.
3. Potassium and copper are two elements belonging to same period in the periodic table.
   (a) Write one property which is common to both the elements.
   (b) Describe two differences in their chemical properties.
4. What is the role of platinum in Ostwald process?
5. Discuss the following properties of transition elements.
   (a) Atomic volume and densities.
   (b) Ionization potential.
6. Write down the electronic configuration in s, p, d, f notation for the following atoms.

   A V
   B Cr
   C Co
   D Cu
   E Zn

# 6 Chemical Energetics

## 6.1 Introduction

Molecules have energy, and when they react, changes in the total energy of the reacting molecules take place. During chemical reaction, some bonds in the molecules are broken and some new bonds are formed. Energy is needed to break a chemical bond and on the other hand energy is released when a bond is formed. Hence, all the chemical reactions are accompanied by energy change. These energy changes appear in the form of evolution or absorption of heat, light, electricity etc. The energy evolved during a chemical reaction is in the form of heat and is absorbed in the form of thermal, electrical or photo energy. The amount of energy evolved or absorbed during chemical reaction is always same for the same quantities of reacting substances.

The branch of chemistry which deals with the energy changes during chemical reactions is called Chemical energetics, while the branch of chemistry that deals with the quantities of heat released or absorbed during chemical reactions is called Thermochemistry.

## 6.2 Energy change in chemical reactions

**Learning objectives**

On completion of this topic, students should be able to:

» *state the law of conservation of energy.*

» *explain endothermic and exothermic reactions with energy diagram.*

» *differentiate internal energy and enthalpy.*

» *mention some of the applications of energy transfer in everyday life.*

» *explain types of heat of reaction.*

### 6.2.1 Law of conservation of energy

The energy may be released or absorbed in a chemical reaction, but the total energy of a reacting system and the surrounding remains constant. The transfer of energy between system and the surrounding is guided by the law of conservation of energy by Mayer and Helmholtz (1840). According to them, the law states that the 'energy can neither be created nor destroyed, but it can be converted from one form into another.

The amount of energy that leaves a system is same as that dissipates into the surrounding, which implies that there is no lost in the energy (i.e. the total energy of the universe is always constant). However, during the course of physical and chemical change, one form of energy may be converted into another form. For example, water in a Chhukha dam possesses potential energy. When this water is released, the potential energy gets converted to kinetic energy. As the force of water rotates the turbine, kinetic energy is converted to electrical energy. This proves that whenever, one form of energy disappears an equal amount of energy in some other form appears, proving the law of conservation of energy.

### 6.2.2 Internal energy (E)

The chemical reactions take place due to interaction between various atoms and molecules of the substance undergoing a change. It is assumed that, these atoms and molecules are associated with some energy of their own which is called their internal energy. It is denoted by E.

The energy stored in a substance by virtue of its molecules is called its internal energy.

The internal energy is the sum of all the energies such as vibrational, rotational and translational kinetic energy and potential energy from intermolecular forces.

$$E = E_v + E_r + E_t + E_p$$

#### Change in internal energy (ΔE)

The internal energy is different for different substance and therefore, in a chemical reaction the internal energy of a reactant is different from that of the products. For example, during a chemical reaction whereby reactant is converted to product at a constant temperature and volume, the change in internal energy of the reactant ($E_R$) and the internal energy of the product ($E_P$), ΔE is given by:

$$\Delta E = E_P - E_R$$

Sign of ΔE depends on the fact whether $E_R$ is greater or smaller than $E_P$. Both positive and negative value of internal energy change is possible. For example:

a) When $E_R < E_P, \Delta E = (E_P - E_R)$ will have a positive value. In this case energy will be absorbed in the reaction, i.e., the reaction will be endothermic. Thus, for all endothermic reactions that take place at constant volume and temperature, ΔE will have a positive value.

b) When $E_R > E_P, \Delta E = (E_P - E_R)$ will have a negative value. In this case energy will be released in a reaction, i.e., the reaction will be exothermic. Thus, for all exothermic reactions that take place at constant volume and temperature, ΔE will have a negative value.

### 6.2.3 Enthalpy or Heat content (H)

When the chemical reactions are performed in an open vessel at a constant pressure (P=1 atmosphere), it is convenient to introduce a new function called enthalpy or heat change which is denoted by H.

The energy contained in a chemical bond that can be converted into heat is known as enthalpy. Enthalpy is a Greek word meaning 'to warm'. Enthalpy is related to internal energy (E) and is given by:

$$H = E + PV$$

Where E is internal energy, P is the pressure and V is the volume of the system.

#### Change in enthalpy (ΔH)

Enthalpy cannot be measured directly, however we can measure the enthalpy change in a thermochemical reaction, upon converting reactant into a product at constant pressure and temperature. If $H_R$ and $H_P$ are the enthalpy of reactants and products respectively, then the change in enthalpy, ΔH is given as:

$$\Delta H = H_P - H_R$$

The change in the heat content, in a chemical reaction that is carried out at constant pressure is referred to as change in enthalpy. Thus, enthalpy change is the amount of heat evolved or absorbed in a reaction carried out at constant pressure and temperature. It is assumed that the work done is only of pressure-volume type.

Sign of ΔH depends on the fact whether $H_R$ is greater or smaller than $H_P$. Both positive and negative value of change in heat content is possible. For example:

c) When $H_R < H_P, \Delta H = (H_P - H_R)$ will have a positive value (Figure 6.2). In this case heat energy will be absorbed in a reaction, i.e. the reaction will be endothermic. Thus, for all endothermic reactions that take place at constant pressure and temperature, ΔH will have a positive value. Positive value of ΔH means that the heat content of the reaction will be raised.

d) When $H_R > H_P, \Delta H = (H_P - H_R)$ will have a negative value (Figure 6.4). In this case heat energy will be released in a reaction, i.e., the reaction will be exothermic. Thus, for all exothermic reactions that take place at constant pressure and temperature, ΔH will have a negative value and have a tendency to proceed spontaneously. It also means that the heat content of the reaction will be lowered.

## 6.2.4 Heat (enthalpy) of reaction

The heat of reaction is defined as the amount of heat released or absorbed in a chemical reaction when numbers of moles of reactants completely react to form the products. It is also called enthalpy of reaction, and is denoted by ΔH. The quantities of heat shown with the balanced chemical equations are all heats of the corresponding reactions. For example, the thermochemical equation below shows that the heat of reaction involving the combination of 1 mole each of carbon and oxygen gas is -94.00 k.cal.

$$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -94.00\,k.cal$$

The heat of reaction is generally referred to as the difference between the total heat content of the reactant and the total heat content of the product. This relationship between the heat of reaction and the difference between the heat content of the products and reactants can be mathematically expressed as:

Sum of the change in enthalpy = (sum of the enthalpies of product) – (sum of the enthalpies of reactant).

$$\Sigma\Delta H = \Sigma H_P - \Sigma H_R$$

Every chemical reaction is represented by an equation and has a heat of reaction.

The reaction may involve the combustion of substance, neutralization of an acid by a base, dissolution of a salt, formation of a compound, etc. Thus, the heat (enthalpy) of reactions are of following types:

### i. Heat or Enthalpy of combustion

Combustion or burning is a chemical process in which a substance reacts rapidly with oxygen and gives off heat. The amount of heat released when 1 mole of a substance undergoes complete combustion in the presence of excess of $O_2$ or air at a given temperature is called heat of combustion of that substance. For example, the heat of combustion of propane in air is -530.60 k.cal as shown in the equation below:

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g) \quad \Delta H = -530.60\,k.cal$$

It should be noted that the combustion reactions are exothermic and hence, the heat of combustion has negative value.

### ii. Heat or Enthalpy of neutralization

Neutralization is a chemical reaction between an acid and a base. Heat is evolved when in aqueous solution an acid is neutralized by a base, or vice versa. The amount of heat evolved when 1 mole of an acid is neutralized by 1 mole of a base in dilute solutions is called the heat of neutralization. For example, heat of neutralization of HCl by NaOH in dilute solution at 25°C, is -13.7 k.cal.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) \quad \Delta H = -13.7\,k.cal$$

### iii. Heat of solution

When a solute dissolves in a solvent, heat is either absorbed or released. For example, when potassium iodide (KI) dissolves in water with absorption of heat, the solution becomes colder than the water taken. On the other hand, when lithium chloride (LiCl) dissolves in water with evolution of heat, the solution becomes hotter than the water taken.

The amount of heat absorbed or evolved when 1 mole of solute dissolves in solvent that further addition of solvent to the solution produces no further change in heat content is called the heat of solution. When dilute solutions are made in water, it is indicated by using the symbol 'aq' for aqueous. Thus, from the equation for the dissolution of salt in water,

$$NaCl(s) + (aq) \longrightarrow NaCl(aq) \quad \Delta H = +1.2\,k.cal$$

the heat of solution (ΔH) of NaCl is 1.2 k.cal. This would mean that 1.2 k.cal of heat is absorbed when 1 mole of NaCl (= 58.5 g) is dissolved in excess of water.

Now consider the following reaction:

$$HCl(g) + 50H_2O(l) \longrightarrow HCl.H_2O(aq) \quad \Delta H = -17.51k.cal$$

The heat of solution (ΔH) of HCl in water is -17.51 k.cal. This would mean that 17.51 k.cal of heat is released when 1 mole of HCl (= 36.5 g) is dissolved in excess of water.

### iv. Heat of formation and stability

The amount of heat absorbed or evolved when 1 mole of substance is formed from its element at STP is called the heat of formation. Here the standard state (i.e., STP) of any substance is taken at 25°C and at 1 atmospheric pressure. Thus, it is also known as standard heat of formation. The standard heat of formation is represented by $\Delta H°f$. For example, the standard heat of formation of $H_2O$ (l) is given by the equation

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l) \quad \Delta H°f = -68.3\,k.cal$$

is equal to -68.3 k.cal. It is the amount of heat evolved in the formation of 1 mole of $H_2O$(l) from its element namely $H_2$(g) and $O_2$(g).





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In a similar manner, the standard heat of formation of HI (g) as given by the equation:

$$\frac{1}{2}H_2(g) + \frac{1}{2}I_2(s) \longrightarrow HI(g) \quad \Delta H^\circ_f = +6.2\,k.cal$$

is equal to +6.2 k.cal, since it is the amount of heat required in the formation of 1 mole of HI(g) from its element namely $H_2(g)$ and $I_2(s)$ . Note that the standard heat of formation of HI(g) is not equal to +12.4 k.cal as shown by the equation as +12.4 k.cal is the amount of energy required for the formation of 2 moles of HI(*g*) .

$$H_2(g) + I_2(s) \longrightarrow 2HI(g) \quad \Delta H^\circ_f = +12.4\,k.cal$$

Consequently, heat of formation (ΔH°f) of 1 mole of HI(g) = $+\frac{12.4}{2}$ k.cal = +6.2 k.cal

The negative value of the heat of formation indicates that the energy is released when the compound is formed from its elements. The higher is the value of negative heat of formation, greater is the amount of energy required for the decomposition of the compound, and more stable the compound is. Thus, carbondioxide with ΔH°f = -94.00 k.cal is more stable than water (liquid) whose ΔH°f = -68.3 k.cal.

When a compound is formed with positive value of heat of formation, energy is absorbed. The positive value of ΔH°f, indicates that the enthalpy of a compound formed is greater than the sum of enthalpies of the elements from which the compound is formed. Such a compound is relatively less stable than its elements. For example, the heat of formation of HI, is +6.2 k.cal. It decomposes to some extent even at room temperature producing violet vapour of iodine. They have positive heat of formation and are not stable. Some other examples of such compounds are hydrogen peroxide ($H_2O_2$), acetylene ($C_2H_2$), hydrogen sulphide ($H_2S$), etc.

## 6.2.5 Thermochemical reactions

Every thermochemical reaction is accompanied by thermochemical equation which indicates the heat lost or gained during a change. A chemical equation that includes the quantity of heat released or absorbed during a reaction is called thermochemical equations. Based on evolution and absorption of the heat during a reaction, thermochemical reactions are of two types:

### i. Endothermic reactions

A reaction in which the energy is supplied to the reactants from a surrounding to obtain products is called endothermic reaction. It can be represented as:

$$\text{Reactant} + \text{Energy absorbed by reactants} \longrightarrow \text{Product}$$





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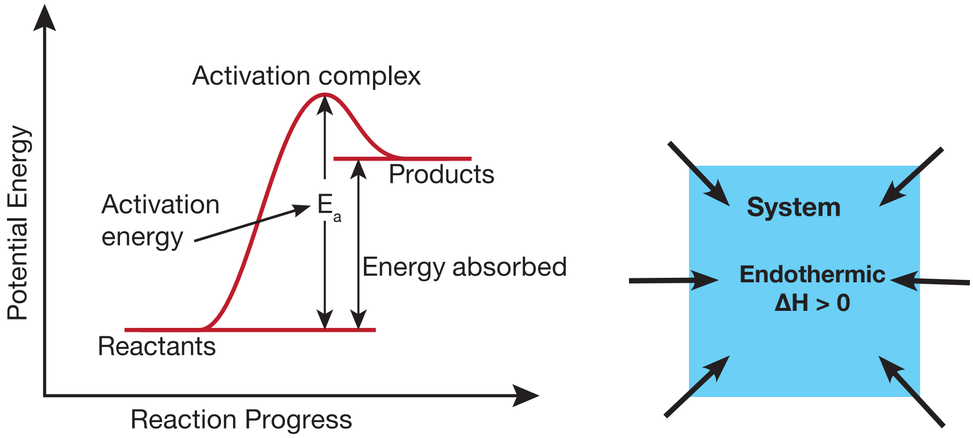


Figure 6.1 Energy diagram for endothermic reaction

If an endothermic reaction is carried out at constant volume and constant temperature, then $\Delta E = (E_P - E_R)$ will have a positive value.

And also, if an endothermic reaction is carried out at constant pressure and constant temperature, then the sum of enthalpies of reactants ($H_R$) is less than products ($H_P$) and hence the enthalpy change, ΔH is given by

$$\Delta H = \Sigma H_P - \Sigma H_R$$

will have a positive value. Thus, for any endothermic reaction, ΔH is positive (Figure 6.2). The positive value of ΔH indicates that endothermic reactions are accompanied by the absorption of heat energy in which the heat content of the reaction is raised.

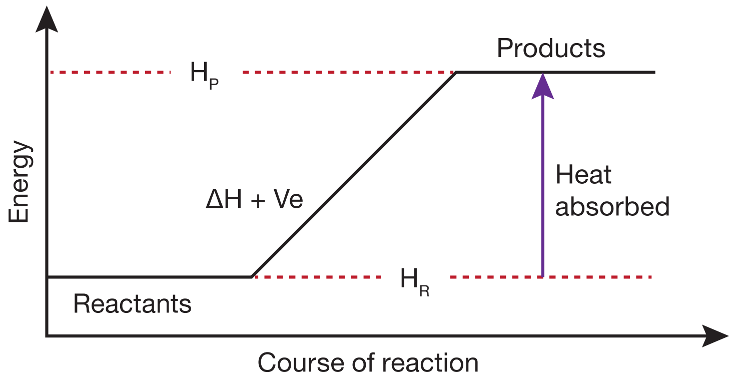


Figure 6.2. ΔH in endothermic reaction.

Some of the thermochemical equations for endothermic reactions are discussed below:

**a. Reaction between nitrogen and oxygen**

It is one of the reactions that take place when fuel is burnt in car engines. The equation for this reaction is:

$$N_2(g) + O_2(g) \longrightarrow 2NO(g) \quad \Delta H = +43.3\,k.cal$$

Here the bonding in the reactants is stronger than the product. Therefore, energy is consumed in by the reactants.

**b. Decomposition of HgO**

$$HgO(g) \longrightarrow 2Hg(l) + O_2(g) \quad \Delta H = +43.2\,k.cal$$

Dissociation of molecules into atoms, removal of electron to form ions (ionization), melting of solids, fusion, vaporization of liquids, sublimations, synthesis of protein in living bodies, etc. are some of the other examples of endothermic change.

## ii. Exothermic reactions

A reaction in which the energy is evolved or released to the surrounding along with the products is called exothermic reaction. It can be represented as:

$$\text{Reactant} \longrightarrow \text{Product} + \text{Energy released}$$

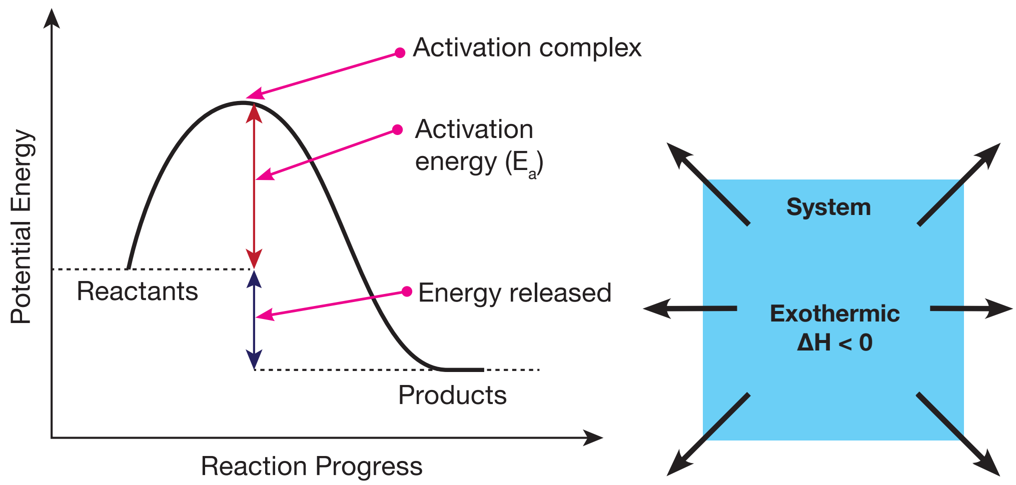



Figure 6.3. Energy diagrams for exothermic reaction

If an exothermic reaction is carried out at constant volume and constant temperature, then $\Delta E = (E_P - E_R)$ will have a negative value.

And also, if an endothermic reaction is carried out at constant pressure and constant temperature, then the sum of enthalpies of reactants ($H_R$) is greater than products ($H_P$) and hence the enthalpy change, $\Delta H$ is given by

$$\Delta H = \Sigma H_P - \Sigma H_R$$

will have a negative value. Thus, for any exothermic reaction, ΔH is negative (Figure 6.4). The negative value of ΔH indicates that exothermic reactions are accompanied by the release of heat energy in which the heat content of the reaction is lowered.

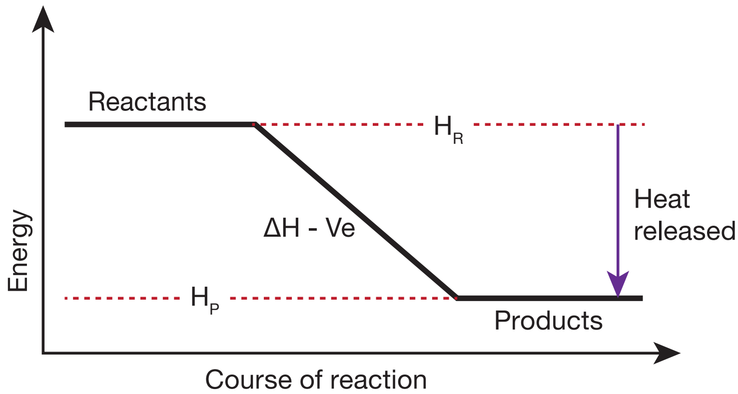



Figure 6.4. ΔH in exothermic reactions

The thermochemical equations for some exothermic reactions are discussed below:

**a. Burning of some of the substances like Zn, C and S in air:**

These processes result in the evolution of heat energy as the reactions proceed.

$$2Zn(s) + O_2(g) \longrightarrow 2ZnO(s) \quad \Delta H = -166.5\,k.cal$$
$$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -94.4\,k.cal$$
$$S(s) + O_2(g) \longrightarrow SO_2(g) \quad \Delta H = -71.1\,k.cal$$

The other examples of exothermic reactions are hydration process, formation of MX (M is metal and X is halogen), formation of anion by gaining electron (reduction), etc.

**b. Burning of methane ($CH_4$)**

Methane is the hydrocarbon molecule which is commonly used as fuel. When it is burnt, it reacts with oxygen and produce carbondioxide and water vapour.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

During this reaction, bonds are first broken and the new bonds are formed. In $CH_4$ the carbon atom is covalently bonded with hydrogen atoms. Similarly, the oxygen atoms in oxygen molecule are held together by covalent bond. During the chemical reaction all the bonds are needed to be broken. The energy is required to break these bonds by pulling the atoms apart. Thus, the energy has to be absorbed from the surrounding and hence, the process is endothermic.

The new bonds are formed between carbon and oxygen to form $CO_2$ and, hydrogen and oxygen to form $H_2O$. In this case the energy is released to the surrounding and hence, the process is exothermic.

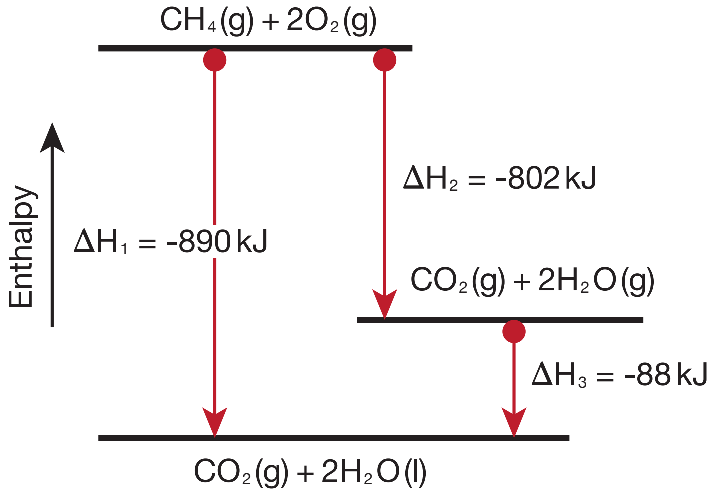



Figure 6.5. Combustion of methane gas

When methane reacts with oxygen, the total energy released is greater than the total energy absorbed. Thus, the overall reaction is exothermic. The energy is released as heat.

### Activity 6.1 Investigating endothermic and exothermic reaction.

#### Materials required

#### Apparatus

Test tubes, beakers, measuring cylinder, spatula, test tube stand, scissor/sharp knife, glass rod and thermometer (-10°C to 110°C).

#### Chemicals

NaOH pellets, $NaHCO_3$ solution, NaOH solution, HCl, water, magnesium ribbon, citric acid, zinc powder and sulphuric acid.

#### Procedure

| Sl. No. | Procedure | Observation | |
|---|---|---|---|
| | | Initial Temperature (°C) | Final Temperature (°C) |
| 1 | i. Measure 5 mL water and pour it in a clean dry test tube. Place it on a stand or in a beaker. Measure the tempera-ture of water (Initial temperature) | | |

| Sl. No. | Procedure | Observation | |
|---|---|---|---|
| | | Initial Temperature (°C) | Final Temperature (°C) |
| | ii. To it add about 3 pellets of NaOH. Stir well with a glass rod. Measure temperature of the solution. (Final temperature) | | |
| 2 | i. Take 5 mL NaOH solution in a clean and dry test tube and place it on a test tube stand or in a beaker. Record the temperature of NaOH solution (Initial temperature)<br>ii. Measure 5 mL of HCl and carefully add this to NaOH solution. Stir well with the glass rod and record the temperature (Final temperature) | | |
| 3 | Follow the same procedure and investigate the temperature change with the following:<br>i. 5 mL $H_2SO_4$ and 2 cm piece of Mg ribbon.<br>ii. 5 mL $NaHCO_3$ solution and 4 spatula of citric acid.<br>**Or**<br>Mix one spatula of citric acid and one spatula of sodium hydrogen carbonate in a dry test tube or watch glass, and add this mixture to 2 mL water in another test tube.<br>iii. 5 mL $CuSO_4$ solution and 1 spatula of Zinc powder.<br>**Note**: Need to wait as you record the temperature until the reaction is complete. | | |

### Safety Precaution

- Concentrated sodium hydroxide and acids (HCl and $H_2SO_4$) are corrosive in nature.
- Use scissor/sharp knife to cut magnesium ribbon.

**Questions**

1. Indentify endothermic and exothermic reactions from the investigation.
2. Which two substances can be used in a cold pack?
3. Golfers need a hand warmer to keep their hands warm on a cold day. Which chemicals could be used in these warmers?

## 6.2.6 Applications of energy change

The energy changes accompanying chemical reaction have numerous practical applications in day to day life. Some of them are discussed below:

1. Methane is the simplest hydrocarbon molecule used as fuel gas. When a mixture of natural gas, $CH_4$ and air is burnt in the kitchen, it produces heat which is used for cooking.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g) + \text{Heat energy}$$

2. When coal, a fossil fuel is burnt in air, it gives heat which is used as fuel.

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

3. Burning of candle (wax) gives heat and light energy.

$$C_{15}H_{32}(s) + 23O_2(g) \longrightarrow 15CO_2(g) + 16H_2O(g) + \text{Heat energy} + \text{Light energy}$$

4. During photosynthesis, the formation of glucose ($C_6H_{12}O_6$) takes place with the absorption of light energy.

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

5. When electric current is passed through acidified water, it is decomposed into its constituent elements, namely $H_2$ and $O_2$.

$$2H_2O(aq) \xrightarrow{\text{Electric energy}} 2H_2(g) + O_2(g)$$

6. The combustion of diesel and petrol ($C_8H_{18}$) give mechanical energy which is used as motor fuel.

$$C_8H_{18}(l) + 12.5\,O_2(g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(g) + \text{Heat}$$

7. The chemicals in a car battery undergo reaction to produce electrical energy, which is used in running the engine.

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s) + \text{Electrical energy}$$

8. Energy change also plays vital role in some important biochemical reactions. There are many chemical reactions which take place in the living bodies, both in plants and animals. Such reactions are called biochemical reactions. Heat is invariably produced in all biochemical reactions. Some of the examples are:

   a) The slow burning down of food (respiration) comprising of carbohydrates and fats, results in the formation $CO_2$ and $H_2O$. During this process, the heat is produced which maintains the normal body temperature and also provide energy to the body in carrying out various activities.

   b) The fermentation reaction which occurs with the help of bacteria and enzyme also produce heat.

**Self Evaluation**

1. Identify whether following reactions are endothermic or exothermic reaction.

   (a) $C(s) + \frac{1}{2}O_2(g) \longrightarrow CO(g)$ $\Delta H = -26.4$ k.cal

   (b) $H_2(g) + I_2(s) \longrightarrow 2HI(g)$ $\Delta H = +12.4$ k.cal

   (c) $H_2O(g) + C(s) \longrightarrow CO(g) + H_2(g)$ $\Delta H = +31.4$ k.cal

   (d) $CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g)$ $\Delta H = -67.6$ k.cal

   (e) $HCl(aq) + KOH(aq) \longrightarrow KCl(aq) + H_2O(l)$ $\Delta H = -13.8$ k.cl

   (f) $C(s) + 2S(s) \longrightarrow CS_2(l)$ $\Delta H = +21.0$ k.cal

   (g) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + H_2O(l)$ $\Delta H = -210.8$ k.cal

2. State various types of heat of reactions.

3. State whether each statement is true or false.

   (a) Breaking chemical bond takes in energy from the surrounding.

   (b) For an exothermic reaction at constant volume and temperature, its $\Delta E = (E_P - E_R)$ will be positive.

   (c) Energy is evolved during the formation of chemical bond.

   (d) For an endothermic reaction, $\Delta H$ is negative.

## Summary

1. Chemical reactions involve changes in energy.
2. Chemical energetic deals with the energy changes and thermochemistry deals with quantities of heat released or absorbed during the chemical reactions.
3. The energy stored in a substance by virtue of its molecules is called its internal energy which is denoted by E.
4. Total internal energy of a substance is the sum of all the energies such as vibrational, rotational, translational, potential energy and kinetic energy from intermolecular forces. i.e. $E = E_v + E_r + E_t + E_p$
5. For all endothermic reactions taking place at constant volume and temperature, $\Delta E$ will have a positive value.
6. For all the exothermic reactions that take place at constant volume and temperature, $\Delta E$ will have a negative value.
7. According to the law of conservation of energy 'energy can neither be created nor destroyed but it can be converted from one form into another'.
8. The energy contained in a chemical bond that can be converted into heat is known as enthalpy.
9. The enthalpy change is the amount of heat evolved or absorbed in a reaction carried out at constant pressure and temperature which is given as,
   $\Delta H = H_P - H_R$
10. Negative value of $\Delta H$ indicates that endothermic reactions are accompanied by the release of heat energy in which the heat content of the reaction is lowered.
11. Positive value of $\Delta H$ indicates that exothermic reactions are accompanied by the absorption of heat energy in which the heat content of the reaction is raised.
12. The reactions which absorb energy from their surroundings, usually in the form of heat are called endothermic reaction.
13. The reactions which release energy to their surroundings, usually in the form of heat are called exothermic reaction.
14. The amount of heat released when 1 mole of a substance undergoes

complete combustion in the presence of excess of $O_2$ or air at a given temperature is called heat of combustion of that substance.

15. The amount of heat evolved when 1 gram equivalent of an acid is neutralized by 1 gram equivalent of a base when both of them are present in dilute aqueous solutions is called heat of neutralization.
16. The amount of heat absorbed or evolved when 1 mole of solute dissolves in so much amount of solvent that further addition of solvent to the solution produces no further change in heat content is called heat of solution.
17. The heat is absorbed during the dissolution of KI and on the other hand heat is evolved during the dissolution of LiCl in water.

## Exercise

**I. Fill in the blanks with correct word(s).**

1. The energy stored in a substance by virtue of its molecules is termed as _____
2. In a reaction where $H_R < H_P$, the heat energy is _____
3. The change in the heat content, during a chemical reaction at constant pressure is referred to as _____
4. The heat content of a chemical reaction during exothermic reaction is ______
5. In a reaction where $H_R > H_P$, the value of $\Delta H$ will be _____

**II. State whether each of the statement is True or False.**

1. For an exothermic reaction taking place at constant volume and temperature, $\Delta E$ will have a negative value.
2. The heat content of a reaction is raised during endothermic reaction.
3. The positive value of $\Delta H$ indicates that the reaction is exothermic.
4. Enthalpy is the energy contained in a chemical bond that can be converted into heat.
5. The amount of heat change, when 1 mole of solute dissolves in its solvent is called as the heat of formation.
6. The positive value of $\Delta H$ favours spontaneous reaction.

### III. Match the items of Column I with the corresponding items of Column II.

| Column I | Column II |
|---|---|
| 1. Stable compounds | a. $H_2O_2$ and $C_2H_2$ |
| 2. Formation of nitrogen oxide | b. $\Delta E = E_P - E_R$ |
| 3. Change in internal energy | c. endothermic reaction |
| 4. Dissolution of LiCl | d. $CO_2$ and $H_2O$ |
| 5. Unstable compounds | e. $\Delta H = H_P - H_R$ |
| | f. exothermic reaction |

### IV. Choose the most appropriate response from the given options in following questions.

Refer the Figure 6.6 and answer the questions (1 to 5) that follow.

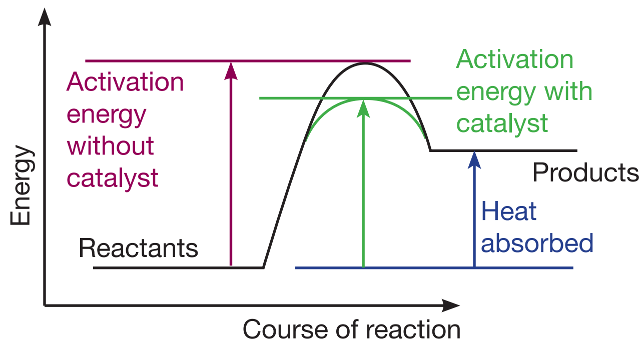



Figure 6.6 Thermochemical reaction

1. When catalyst is used, the activation energy of the reaction

   A increases. C escalates.

   B decreases. D is constant.

2. The type of thermochemical reaction is

   A endothermic reaction. C biochemical reaction.

   B exothermic reaction. D redox reaction.

3. The value of $\Delta H$ is

   A negative. C zero.

   B positive. D neutral.

4. The value of $\Delta E$ is

   A negative. C neutral.

   B positive. D zero.

5. When the chemical bonds are broken, the energy is

| | | | |
|---|---|---|---|
| A | evolved. | C | released. |
| B | stored. | D | absorbed. |

## V. Write answers for the following questions

1. Identify the following process as exothermic or endothermic reaction.

| | | | |
|---|---|---|---|
| A | Dissociation of molecules into ions | F | Sublimation of naphthalene ball |
| B | Formation of cations | G | Formation of halides acids |
| C | Formation of anions | H | Dissolution of KI in water |
| D | Neutralization of acid by a base | I | Decomposition of HgO |
| | | J | Evaporation of water |
| E | Melting of ice | K | Protein synthesis |

2. Explain with an equation why combustion of methane is an exothermic reaction, although it involves both breaking and making of the bond.
3. Describe at least five applications of energy changes that take place in chemical reaction.
4. Differentiate the following;
   (a) Change in internal energy and change in enthalpy of a substance.
   (b) Endothermic reaction and exothermic reaction.
   (c) Chemical energetic and chemical kinetics.
5. What sign is given to $\Delta E$ and $\Delta H$ values for endothermic and exothermic reaction? Use proper illustration to explain them.
6. Explain the following heat of reaction with an example each.

| | | | |
|---|---|---|---|
| A | Heat of combustion | C | Heat of neutralization |
| B | Heat of formation | D | Heat of solution. |

7. What is the standard heat of formation? What information is given by positive and negative value of standard heat of formation regarding the stability of the compound?
8. How is internal energy of a system related to an enthalpy?
9. Write down any two practical applications of endothermic and exothermic reactions in our day to day life.
10. What is the effect of endothermic and exothermic reaction on the surrounding?

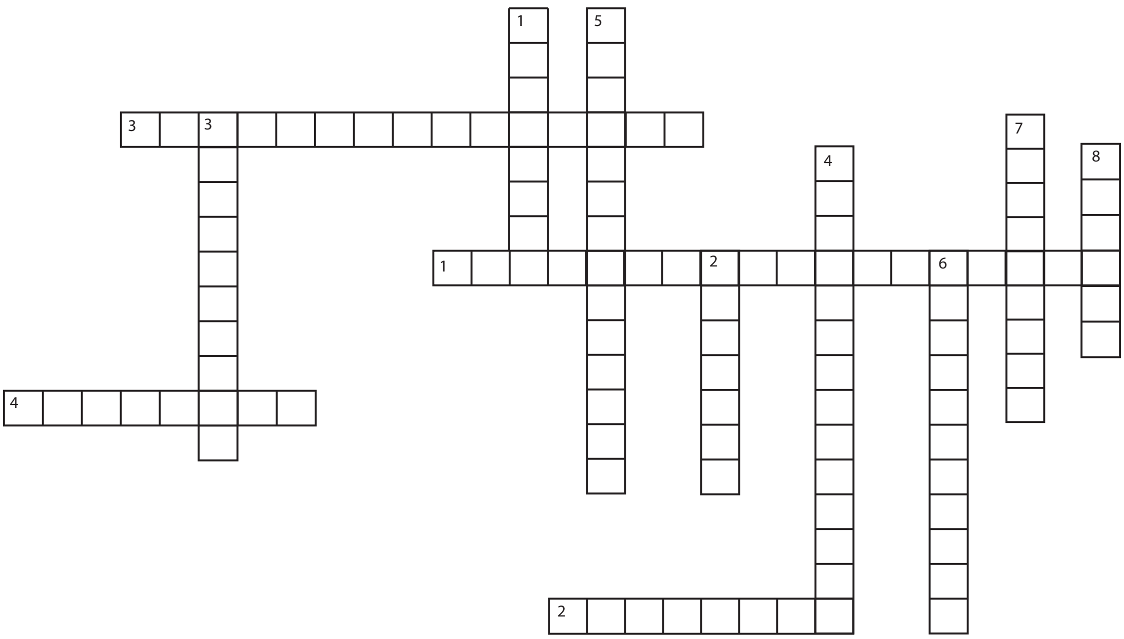
## VI. Solve the cross word puzzle

Across

1. The branch of Chemistry which deals with the energy changes in chemical reactions.
2. The heat contained within a molecule.
3. The branch of Chemistry which deals with the heat changes in chemical reactions.
4. The ΔH of an exothermic reaction is____.

Down

1. The ΔH of an endothermic reaction is______.
2. In exothermic reactions heat content of the reaction is_______.
3. The reaction in which the heat energy is released.
4. The energy stored in substances.
5. The reaction between acid and base to form salt and water.
6. The reaction in which the heat energy is absorbed.
7. The standard heat of formation has strong influence on the_____of substances.
8. In endothermic reactions heat content of the reaction is____



# 7 Reversible Reaction

## 7.1 Introduction

In many chemical reactions, reactants are transformed entirely into the reaction products. For example, a burning fuel results in new products as one of the reactants is completely used up. The reactions cannot be reversed as the resulting product does not react to form original substance. A reaction in which the reactants are completely converted into the products and the change cannot be reversed is called an irreversible reaction. However, all chemical reactions do not proceed to completion. Those chemical reactions in which the reactants are not completely used up have tendency to reverse the reaction. Thus, reversible reactions are those chemical reactions where the reactants form products, which in turn react together to give the reactants back.

In a reversible reaction, both forward and backward reaction occurs simultaneously to reach a stage where the rate of forward and backward reaction becomes equal. Thus, at this stage, the reaction is said to have attained a state of chemical equilibrium. A chemical equilibrium is an apparent state of rest at which the forward and the backward reactions are proceeding at the same rate. On changing the concentrations, temperature or pressure of the reaction, reversible reaction undergoes a shift to re-establish its equilibrium. The influence of various factors on a system was generalized by Le Chatelier and is known as Le Chatelier's Principle. The chemical industry makes use of Le Chatelier's Principle to increase the yield of product in a reaction.

Figure 7.1 Le Chatelier

## 7.2 Reversible reactions and equilibrium

**Learning objectives**

On completion of this topic, students should be able to:

» *explain reversible and irreversible reaction with examples.*
» *define chemical equilibrium.*
» *explain the factors influencing the direction of an equilibrium reaction.*
» *explain Le Chatelier's Principle in manufacturing processes.*
» *give some examples of equilibrium reactions in manufacturing processes.*
» *apply Le Chatelier's Principle to make and test predictions about how different factors affect a chemical system at equilibrium.*
» *explain Haber's process to manufacture ammonia.*

In an irreversible reaction, chemical reaction proceeds to completion by converting entire amount of the reactants into products. In such chemical reactions, reactants no longer retain their chemical identity. Few examples of such reactions are:

- Neutralisation reaction between strong acid and strong base.
  $NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$

- Double decomposition reactions or precipitation reactions.
  $BaCl_2(aq) + H_2SO_4(aq) \longrightarrow BaSO_4(s)\downarrow + 2HCl(aq)$

- Redox reactions.
  $AgNO_3(aq) + NaCl(aq) \longrightarrow NaNO_3(aq) + AgCl(s)$

A chemical reaction that proceeds in both forward and backward direction is a reversible reaction. In such chemical reactions, reaction reverses as entire amount of the reactants are not converted into products. The reaction which reverses automatically without any change is said to be in equilibrium. In most situations, chemical reactions exist in equilibrium when products and reactants are formed at the same rate and no further net chemical change occurs. Such a system in which the forward and the backward reactions are taking place at the same rate is said to be in a state of dynamic equilibrium as shown in the Figure 7.2.

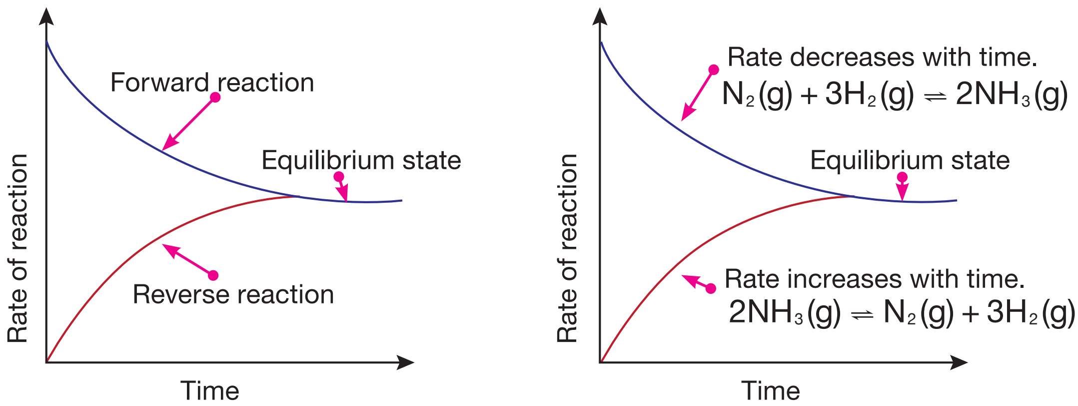



Figure 7.2 Attainment of equilibrium in reversible reaction.

At equilibrium, there is a constant ratio between the concentrations of the products and the reactants in any change. One example of a reversible reaction that reaches equilibrium is the formation and the decomposition of ammonia as explained in the Figure 7.2.

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

In above reaction the double arrow indicates that the reaction can go in either direction. $NH_3$ decomposes into $N_2$ and $H_2$ and the products formed, $N_2$ and $H_2$ combines to form $NH_3$.

Some examples of reversible reactions are:

$$CH_3COOH(l) + NaOH(l) \rightleftharpoons CH_3COONa(l) + H_2O(l)$$
$$FeCl_3(l) + 3H_2O(l) \rightleftharpoons Fe(OH)_3(s) + 3HCl(l)$$
$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$
$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$$
$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$$
$$3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$$
$$CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$$

**Activity 7.1 Action of heat on hydrated copper sulphate.**

**Materials required**

Test tube, test tube holder, delivery tube, one hole cork, beaker, spirit lamp or Bunsen burner, match box, copper sulphate, cold water and clamp stand.

**Procedure**

1. Add one spatula of powdered hydrated copper (II) sulphate in the test-tube

and set up the apparatus as shown in the Figure 7.3.

2. Heat the copper (II) sulphate until its colour changes completely.
3. Allow the content to cool and add the condensed water on to it.
4. Observe the changes.

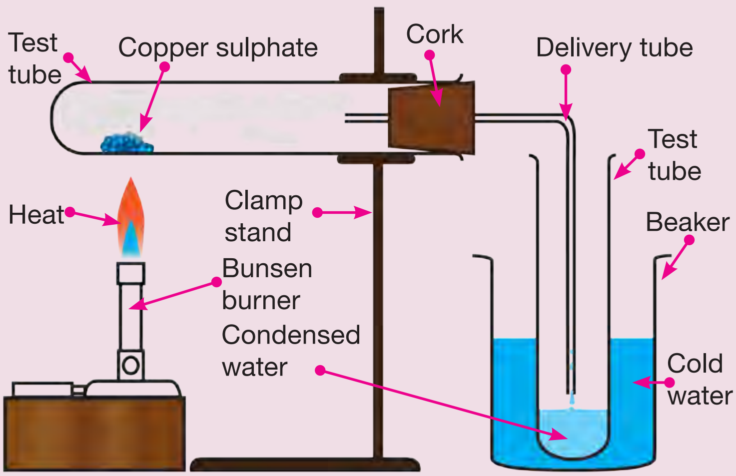



Figure 7.3 Action of heat on hydrated copper sulphate

**Precaution**

Ensure that the mouth of the delivery tube does not touch on the condensed liquid during heating. This is to prevent back suction.

**Questions**

1. Write observations when:
   (a) hydrated copper sulphate is heated.
   (b) water is added to heated copper sulphate.
2. Write the chemical equation to show the action of heat on hydrated copper sulphate.
3. What type of reaction is this?

**Activity 7. 2 Action of acid and base on copper sulphate solution.**

**Materials required**

Test tube, copper sulphate solution, concentrated hydrochloric acid, and ammonia solution.

**Procedure**

1. Add about 5 mL of copper sulphate solution into the test tube.
2. Slowly add the concentrated hydrochloric acid into the test tube containing copper sulphate and shake the solution until yellow green colour is obtained.
3. Now carefully add the ammonia solution into the test tube without shaking and observe the colour change.
4. Continue to add the ammonia solution with gentle shaking until the colour changes to dark blue.
5. Now slowly add concentrated hydrochloric acid to the dark blue solution with gentle shaking and observe the colour change.

**Precaution**

Both concentrated HCl and $NH_4OH$ are corrosive in nature. The experiment should be carried out under the supervision of a teacher.

Use PVC gloves and safty goggles.

**Questions**

1. What happened to the colour of dark blue solution on adding HCl?
2. Write the balanced chemical equation between copper sulphate and hydrochloric acid.
3. What type of reaction is this?

## 7.3 Le Chatelier's Principle

The qualitative influence of various factors on a system in equilibrium was generalized by Le Chatelier and is known as Le Chatelier's Principle. *It states that when a system in equilibrium is subjected to a change in temperature, pressure or concentration of a component, the equilibrium shifts in the direction of the reaction opposing the change*. Some examples of equilibrium reactions in manufacturing processes using Le Chatelier's Principle are:

### i. In Haber's process

The process combines nitrogen and hydrogen to form ammonia. The reaction is reversible and the production of ammonia is exothermic. On adding energy, the system responds to the stress by shifting the equilibrium to the left, and favours the formation of nitrogen and hydrogen. Thus, temperature must be carefully controlled in order to form more ammonia.

$$N_2(g) + H_2(g) \rightleftharpoons 2NH_3(g) + Energy$$

### ii. In Contact process

The process is employed in manufacturing of sulphuric acid from the reaction between sulphur dioxide and oxygen to form sulphur trioxide. The reaction is exothermic and reversible. The reaction proceeds with decrease in volume. Thus, low temperature favours the formation of sulphur trioxide.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + Energy$$

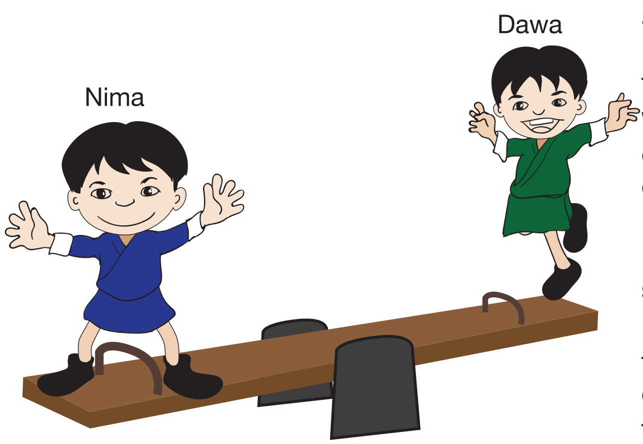
### iii. In Birkland and Eyde process

When nitrogen and oxygen reacts to form nitric oxide by this process, endothermic reversible reaction proceeds with no change in volume. The change in pressure will not affect its equilibrium as there is no change in volume during the reaction. The higher temperature will shift the equilibrium to the right in the direction of endothermic reaction.

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) - Energy$$

Le Chatelier's Principle is equally applicable to physical equilibria such as melting of ice, vaporization of water and solubility of substances.

**Analogy for equilibrium reaction.**




Suppose two brothers, Nima and Dawa played a bet of balancing themselves on the see-saw.

When Nima and Dawa have maintained equilibrium, the see-saw does not slant on either side.

If Dawa shift towards the fulcrum, then Nima too has to shift towards the fulcrum so as to maintain equilibrium.

Here, Dawa's shift can be compared to the stress or disturbance to a system in equilibrium and Nima's shift as the change to minimize the effect.

**Self Evaluation**

1. What is a chemical equilibrium?
2. State Le Chatelier's Principle.
3. What are the factors that affect the chemical equilibrium of a reaction?





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4. Give reasons for the following:
   (a) Low temperature favours the formation sulphur trioxide in Contact process.
   (b) High temperature favours the formation of nitric oxide in Birkland and Eyde process.

## 7.4 Factors affecting the systems at equilibrium

Chemical equilibria respond to three kinds of stresses; changes in the concentrations of reactants or products, changes in temperature and changes in pressure. When a stress is first applied to the system, chemical equilibrium is disturbed. As a result, rates of the forward and the backward reaction in the system are no longer equal. A system responds to the stress by forming either more products or more reactants. The following are the factors affecting the system at equilibrium.

### i. Concentration

On increasing the concentration of reactants, system undergoes a stress. The system tends to decrease the concentration of the reactant by converting some of the reactant into product. Therefore, the rate of the forward reaction is greater than the rate of the reverse reaction. As the rate of forward reaction increases, the equilibrium shifts to the right. The reactant concentration will continue to drop until the reaction attains the state of equilibrium. For example, in a reaction

$$A + B \rightleftharpoons C + D$$

any increase in the concentration of A or B at equilibrium will shift the reaction in forward direction and any increase in the concentration of C or D will shift the equilibrium in a reverse direction.

**Activity 7.3 Action of ammonium hydroxide on copper sulphate solution.**

**Material required**

Test tube, dropper, ammonium hydroxide, hydrated copper sulphate and water.

**Procedure**

1. Take about 5 mL of copper sulphate solution in a test tube.
2. Add 1 or 2 drops of ammonium hydroxide to it and observe the colour change.
3. To the above solution add excess of ammonium hydroxide and observe the colour change.

**Questions**

1. What change is observed when one or two drops of ammonium hydroxide

is added to copper sulphate solution. Explain.

2. Give balanced chemical equation for the reaction between copper sulphate and ammonium hydroxide.

On adding one or two drops of aqueous solution of ammonium hydroxide to a solution of copper sulphate, a sky blue precipitate of copper hydroxide is formed. The precipitate is soluble in excess of ammonium hydroxide to form deep blue solution forming tetramine copper (II) hydroxide.

On adding one or two drops of ammonium hydroxide solution, chemical equilibrium favours the formation of product. As the concentrations of the products increase, they begin to react together and favour the backward reaction. Eventually, the rate of forward and backward reactions reaches the equilibrium. When excess of ammonia solution is added at this stage, the system responds to the increase in the concentration of the reactants by forming more products. This can be observed by noticing the change in colour of a solution from sky blue to deep blue.

$$CuSO_4 + 2NH_4OH \longrightarrow \underset{\text{(blue)}}{Cu(OH)_2} \downarrow + (NH_4)_2SO_4$$

$$Cu(OH)_2 + 4NH_3 \rightleftharpoons \underset{\text{(deep blue)}}{[Cu(NH_3)_4](OH)_2} \downarrow$$

The equilibrium in a bottle of a carbonated liquid is explained in the equation.

$$H_3O^+(aq) + HCO_3^-(aq) \rightleftharpoons 2H_2O(l) + CO_2(aq)$$

On removing the cap of a carbonated liquid bottle, the dissolved carbon dioxide leaves the solution and enters the air. The forward reaction rate of this system will increase to produce more $CO_2$. This increase in the rate of forward reaction decreases the concentration of $H_3O^+$ ($H_3O^+$ ions make soda taste sharp) in the carbonated liquid. As a result, the drink tastes flat. If we increase the concentration of $CO_2$ in the bottle, the reverse reaction rate would increase to form $[H_3O^+]$ and $[HCO_3^-]$.

### ii. Temperature

The effect of temperature on a system at equilibrium depends on whether the reaction is endothermic or exothermic. In endothermic reaction, the system absorbs energy from the surrounding to proceeds the reaction from left to right. On contrary, the exothermic reaction proceeds from right to left by releasing the energy from the chemical system. For example,

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g) + \text{Energy} \qquad \Delta H = +197\,kJ$$

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{Energy} \qquad \Delta H = -92\,kJ$$

the reaction to produce sulphur dioxide and oxygen is endothermic, because the reactant absorbs heat to proceeds its reaction from left to right. On other hand, the reaction to produce ammonia is exothermic as the reaction releases the heat. More ammonia is produced if heat is removed by decreasing the temperature.

*The effect of temperature on the position of equilibrium can be summarized as:*

a) Endothermic change ($\Delta H > 0$). An increase in temperature shifts the equilibrium to the right, thereby forming more products. A decrease in temperature shifts the equilibrium to the left, thereby forming more reactants.

b) Exothermic change ($\Delta H < 0$). An increase in temperature shifts the equilibrium to the left, thereby forming more reactants. A decrease in temperature shifts the equilibrium to the right, thereby forming more products.

### iii. Volume and pressure

Increase of pressure decreases the volume of the gases. Le Chatelier's principle predicts a shift in equilibrium to relieve this change. Therefore, the shift must reduce the pressure of the gases.

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$$

There are two gas molecules on the left side of the equation and three gas molecules on the right side. If the equilibrium shifts to the left, the pressure of the mixture will decrease due to less gas molecules. Increase in pressure shifts the equilibrium in the direction in which there is decrease in volume so that the product of pressure and volume remains constant. The equilibrium will shift to the right at constant temperature. If there is the same number of gas molecules on both sides of the reaction equation, it has no effect on the position of equilibrium as long as there is no change in temperature.

In the conversion of nitrogen and hydrogen to ammonia by Haber's process:

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

The four moles of reacting gases occupy a larger volume than two moles of ammonia. If pressure on the reacting system is increased, the system will tend to reduce the pressure by reducing the number of particles. Increasing the pressure on the reaction increases the concentration of the gas molecules, so the reaction will go faster and more ammonia is formed.

### iv. Catalyst

A catalyst usually speeds up the rate of a reaction, either by allowing a different reaction mechanism or by providing additional mechanisms. The overall effect is to lower the activation energy which increases the rate of reaction. The activation energy is lowered in same amount for forward and reverse reactions, and there is the same increase in reaction rates for both the reactions. As a result, a catalyst does not affect the position of equilibrium. It only affects the time to achieve equilibrium. So, a catalyst does not affect the concentration of the reactants but the reaction takes less time to reach equilibrium.

**Table 7.1 The effects of changing conditions on a system at equilibrium.**

| Type of reaction | Change to system | Direction of change |
|---|---|---|
| All reactions | increasing any reactant concentration, or decreasing any product concentration | toward products. |
| | decreasing any reactant concentration, or increasing any product concentration | toward reactants. |
| | using a catalyst | no change. |
| Exothermic | increasing temperature | toward reactants. |
| | decreasing temperature | toward products. |
| Endothermic | increasing temperature | toward products. |
| | decreasing temperature | toward reactants. |
| Equal number of reactant and product gas molecules | changing the volume of the container, or adding a non-reacting gas | no change. |
| More gaseous product molecules than reactant gaseous molecules | decreasing the volume of the container at constant temperature | toward reactants. |
| | increasing the volume of the container at constant temperature, or adding a non-reacting gas at constant pressure | toward products. |
| Fewer gaseous product molecules than reactant gaseous molecules | decreasing the volume of the container at constant temperature | toward products. |
| | increasing the volume of the container at constant temperature | toward reactants. |

## Self Evaluation

1. Predict the effect of each of the following on the indicated equilibrium system in terms of which reaction (forward, reverse or neither) will be favoured.

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + 184\,kJ$$

(a) addition of $Cl_2$
(b) removal of HCl
(c) increased pressure
(d) decreased temperature
(e) removal of $H_2$
(f) decreased pressure
(g) increased temperature
(h) decreased system volume

## 7.5 Application of Le Chatelier's Principle.

This principle is applied

1. in the study of the **physical equilibrium** such as effects of temperature and pressure on solubility of solids on dissolution, melting of ice and vaporisation of water.
2. in the study of the **chemical equilibrium** during the formation of ammonia by Haber's process, nitric oxide by Birkland and Eyde process, sulphur trioxide by contact process, oxidation of carbon monoxide by steam etc.

In Haber's process, nitrogen from air and hydrogen from natural gas combines to produce ammonia with the evolution of heat.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{Heat}$$

In this process, forward reaction is favoured by high pressure, low temperature and excess of nitrogen and hydrogen as discussed in the Figure 7.4.

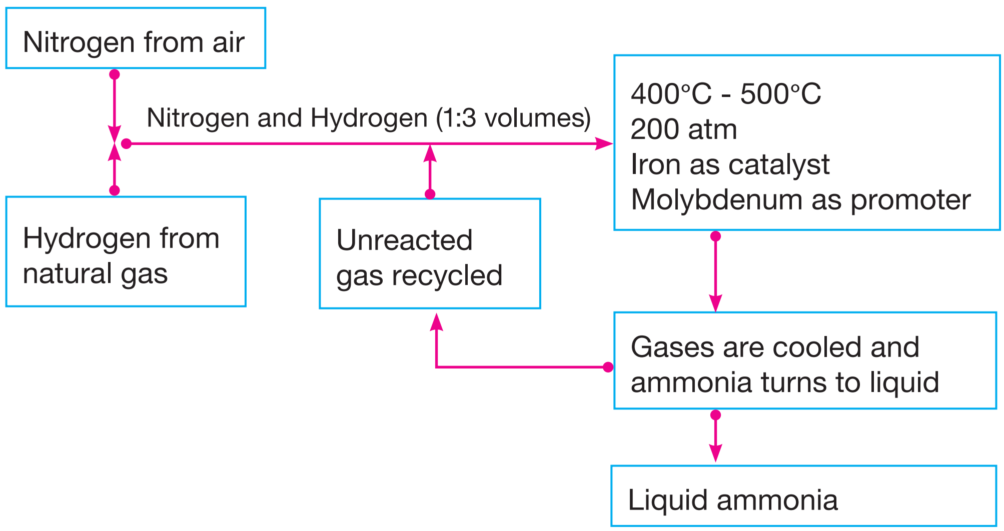



Figure 7.4 Haber's process

**Self Evaluation**

1. For the reaction

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

which direction will the equilibrium shift? Explain your answer in each case.

(a) If more nitrogen is added to the system.
(b) If the temperature is decreased.
(c) If the pressure is increased.
(d) If the volume of reactant is increased.
(e) If hydrogen is removed from the system
(f) If temperature is reduced and pressure is increased.
(g) If catalyst is not used.

## Summary

1. During irreversible reactions, products do not significantly reform reactants.
2. During reversible reactions, products reform the original reactants.
3. Reversible reactions can reach equilibrium.
4. The forward and reverse reaction rates are equal at chemical equilibrium.
5. Chemical reactions in which a reaction automatically reverses and there is no net overall change are said to be in equilibrium.
6. At chemical equilibrium, reactant and product concentrations remain unchanged.
7. Chemical equilibria are dynamic equilibria.
8. Le Chatelier's Principle states that chemical equilibria adjust to relieve applied stresses.
9. Stresses in a chemical reaction are changes in concentration, temperature, and pressure.
10. Temperature changes affect the values of equilibrium constants.
11. Chemical equilibria respond to three kinds of stress; changes in the concentrations of reactants or products, changes in temperature, and changes in pressure.
12. When a stress is first applied to a system, chemical equilibrium is disturbed.
13. Pressure changes have almost no affect on equilibrium reactions in solution.
14. Pressure changes can affect equilibrium reactions in the gaseous phase.
15. In endothermic change ($\Delta H > 0$). An increase in temperature shifts the equilibrium to the right, forming more products. A decrease in temperature

shifts the equilibrium to the left, forming more reactants.

16. In exothermic change ($\Delta H < 0$). An increase in temperature shifts the equilibrium to the left, forming more reactants. A decrease in temperature shifts the equilibrium to the right, forming more products.
17. The Haber process involves a synthesis reaction.

## Exercise

**I. Fill in the blanks with correct word(s).**

1. When a stress is first applied to a system _______ is disturbed.
2. ________ principle states that chemical equilibria adjust to relieve applied stresses.
3. Increase of pressure decreases the volume of the ______ .
4. In equilibrium, reactants are never fully used up because they are constantly being formed from _____.
5. The energy required to get a reaction started is called the ______.

**II. State whether the following statements are True or False. Correct the false statements.**

1. If products are removed from a reaction at equilibrium, more reactants will go on to form products.
2. The reaction shifts towards left on decreasing temperature in an exothermic reaction.
3. At chemical equilibrium, reactant and product concentrations remain unchanged.
4. The volume changes when inert gas is added to equal number of reactant and product of gas molecules.
5. When one product is a gas and other products and reactants are solid or liquid, it is easy to remove the gas from the reaction.

**III. Choose the most appropriate response from the given options.**

1. In which of these reactions is the formation of the products favoured by an increase in pressure?

A $2O_3(g) \longrightarrow 3O_2(g)$

B $C(s) + O_2(g) \longrightarrow CO_2(g)$

C $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

D $CO_2(g) + 2H_2O(l) \longrightarrow H_3O^+(aq) + HCO_3^-(aq)$

2. What is the effect of an increase in temperature on an exothermic reaction at equilibrium?
   - A It has no effect on the equilibrium.
   - B It shifts the equilibrium in favour of the forward reaction.
   - C It shifts the equilibrium in favour of the reverse reaction.
   - D It shifts the equilibrium in favour of both the forward and reverse reactions.
3. Which of the following factors does not affect the equilibrium of a solution?
   - A Temperature.
   - B Pressure.
   - C Concentration.
   - D Catalyst.
4. Which of the following conditions are applied in Haber's process to manufacture ammonia?
   - I By increasing the concentration of reactant.
   - II By decreasing the concentration of reactant.
   - III By increasing the temperature.
   - IV By increasing the pressure.
   - A I and II.
   - B II and IV.
   - C I and IV.
   - D II and III.
5. Chemical reactions in which a reaction automatically reverses and there is no net overall change is said to be
   - A equilibrium.
   - B endothermic.
   - C exothermic.
   - D completion.
6. Which type or types of change, if any, can reach equilibrium?
   - A A chemical change only
   - B A physical change only
   - C Both a chemical and a physical change
   - D Neither a chemical nor a physical change
7. Which statement correctly describes an endothermic chemical reaction?
   - A The products have higher potential energy than the reactants, and the $\Delta H$ is negative.
   - B The products have higher potential energy than the reactants, and the $\Delta H$ is positive.
   - C The products have lower potential energy than the reactants, and the $\Delta H$ is negative.
   - D The products have lower potential energy than the reactants, and the $\Delta H$ is positive.

8. Given the equation representing a reaction at equilibrium:

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

What happens when the concentration of $H_2(g)$ is increased?

A The equilibrium shifts to the left, and the concentration of $N_2(g)$ decreases.

B The equilibrium shifts to the left, and the concentration of $N_2(g)$ increases.

C The equilibrium shifts to the right, and the concentration of $N_2(g)$ decreases.

D The equilibrium shifts to the right, and the concentration of $N_2(g)$ increases.

## IV. Write answers for the following questions

1. Chickens do not have sweat glands. Thus, when the temperature rises they tend to breathe faster which results into lowering the concentration of carbonate ions in their blood. Since, the eggshells are mostly calcium carbonate, faster breathing chickens lay eggs with thinner shells. To solve the issues of high surrounding temperature, farmer supply carbonated water for their chickens rather than installing air conditioner. How does this relate to Le Chatelier's Principle?

2. Hydrogen and iodine gas combines in a reversible exothermic reaction to form hydrogen iodide gas.
   (a) Write a balanced chemical equation for this reaction.
   (b) If more iodine gas is added after the reaction reaches equilibrium, will the reaction be shifted to the left or the right?
   (c) If heat is added after the reaction reaches equilibrium, will the reaction be shifted to the left or the right?

3. Explain whether an exothermic reaction that is at equilibrium will shift to the left or to the right to readjust after each of the following procedures is followed.
   (a) Products are removed.
   (b) More reactants are added.
   (c) More heat is added.
   (d) Heat is removed.

4. The following equilibrium takes place in a close rigid container for the following reaction.

$$CO(g) + 3H_2(g) \longrightarrow CH_4(g) + H_2O(g)$$

The volume and temperature are kept constant, but the pressure on the system is increased. Explain how this affects the concentration of the reactants and products, and the direction in which the equilibrium shifts.

5. The following reaction is at equilibrium.

$$2SO_3(g) \longrightarrow 2SO_2(g) + O_2(g) + \text{Heat}$$

Which condition will shift the equilibrium to the right, a decrease in volume or a decrease in temperature? Explain.

6. Why are industrial processes operated at high temperature?
7. The reaction speeds up in forward direction on increasing the pressure in Haber process. Explain.
8. Several steps are involved in the industrial production of sulphuric acid. One step involves the oxidation of sulphur dioxide gas to form sulphur trioxide gas. A catalyst is used to increase the rate of production of sulphur trioxide gas. In a rigid cylinder with a movable piston, this reaction reaches equilibrium as represented by the equation below.

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g) + 392\,\text{kJ}$$

(a) Increase in the pressure of gases increases the rate of forward reaction. Explain in terms of collision theory.

(b) What is occurred when more oxygen gas is added to the reaction at equilibrium? State in term of the concentration of sulphur trioxide.

# 8 Rate of Reaction

## 8.1 Introduction

A chemical reaction is a process in which the reactants are converted into product(s) through the rearrangement of atoms. How fast or slow a chemical reaction occurs can be determined by its rate of reaction. The rate of chemical reaction is either the speed at which the reactants are used up or the products are formed. The rate of chemical reaction changes continuously with time. The rate of many chemical reactions can be altered by changing the conditions such as temperature, pressure, concentration, presence of catalyst, enzymes etc.

In living organisms almost all chemical reactions takes place with the help of enzymes. Enzymes are bio-catalyst found in nature and are used in biotechnology. The activity of an enzyme is affected by change in pH and temperature. Chemical industries apply the optimum conditions for various chemical reactions to maximise the yield. Some of the important and useful chemical reactions are, burning of fuel, cooking of food, respiration, digestion of food, ripening of fruits, photosynthesis etc.

## 8.2 Collision theory

**Learning objectives**

On completion of this topic, students should be able to:

» *explain collision theory.*

» *explain the conditions for effective collision.*

» *define activation energy in a chemical reaction.*

How do reactions occur? What are the factors responsible for the rate of reaction? This can be explained by the collision theory which is based on the following features:

1. A chemical reaction takes place only when the particles of reactants collide with each other. Higher the frequency of the collision, faster will be the rate of reaction.
2. Only the colliding particles with sufficient energy react to form product, while others remain unchanged.
3. Only the effective collisions bring about the chemical reaction.
4. For effective collision to take place, reactant molecules should
   (a) posses energy greater or equal to threshold energy and
   (b) must have proper orientation.

### 8.2.1 Threshold energy and activation energy

The reactants must collide with sufficient energy to break the bonds and to form the new bonds in the products. When the colliding particles do not possess sufficient energy, the bonds are not broken and new bonds are not formed. For a collision to be effective, the colliding molecules must possess a certain minimum amount of energy called threshold energy. If the reacting molecules have energy less than their threshold energy, there will be no effective collision between them and no product will be formed. Thus in order to form the product, combining molecules or atoms must possess energy equal to threshold energy either of its own or by supplying heat or light. The extra amount of energy which the reactant molecule has to absorb is called the activation energy and it must be equal to the threshold energy. Consider the burning of fuel as an example. For the fuel to burn, it needs a certain amount of heat to ignite. The heat provides the activation energy for this reaction. So activation energy is the minimum amount of energy which the reacting species must possess to start a chemical reaction.

The energy level diagram shown in the Figure 8.1 is an exothermic reaction. The energy curve goes up from the reactants energy level to begin

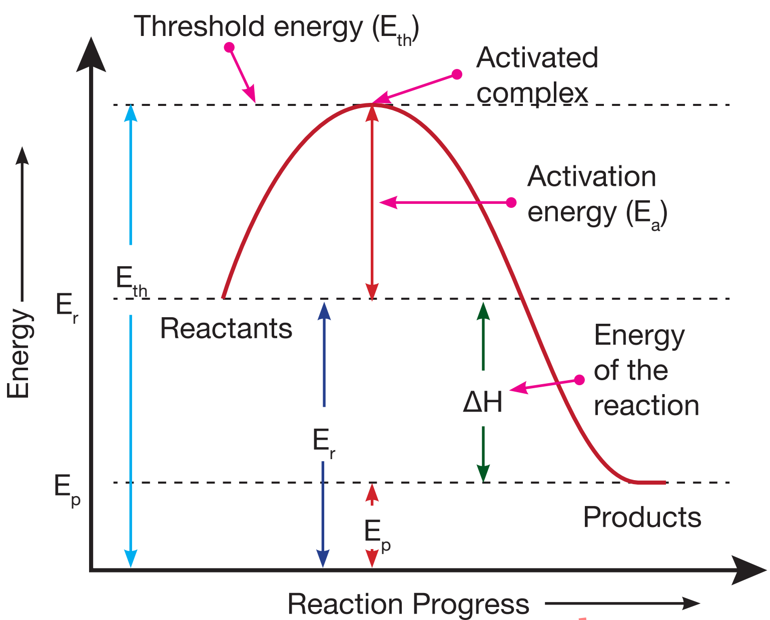


Figure 8.1 Energy changes in chemical reactions

with and then drops to the products energy level. This is because many reactions need activation energy to start the reaction. As the reaction proceeds, the heat content of the reaction is decreased. In an endothermic reaction, the energy is supplied to the reactants from the surrounding to obtain products and the heat content of the reaction is raised.

The collision energy depends on the kinetic energy of the colliding particles. When the molecules with less energy collide, they move apart unchanged. When the molecules having the necessary activation energy collide, their atoms are rearranged to form intermediate, a temporary product called activation complex. It is short lived and may reform the original bonds to form reactant molecules or it may form new bonds to give product.

## 8.2.2 Orientation of reactants

A large number of collisions become ineffective, if the colliding molecules are not properly oriented during collision. The proper orientation for the collision of the reacting molecules favours the formation of product. This is also known as having the correct collision geometry.

The Figure 8.2 illustrates the proper orientation for effective collisions between nitrogen atoms of one molecule with the nitrogen atom of another molecule. If nitrogen atoms of both $NO_2$ are properly oriented, it results in the formation of dinitrogen tetra oxide. However, no product is formed if a nitrogen atom of one $NO_2$ collides with oxygen atom of another $NO_2$.

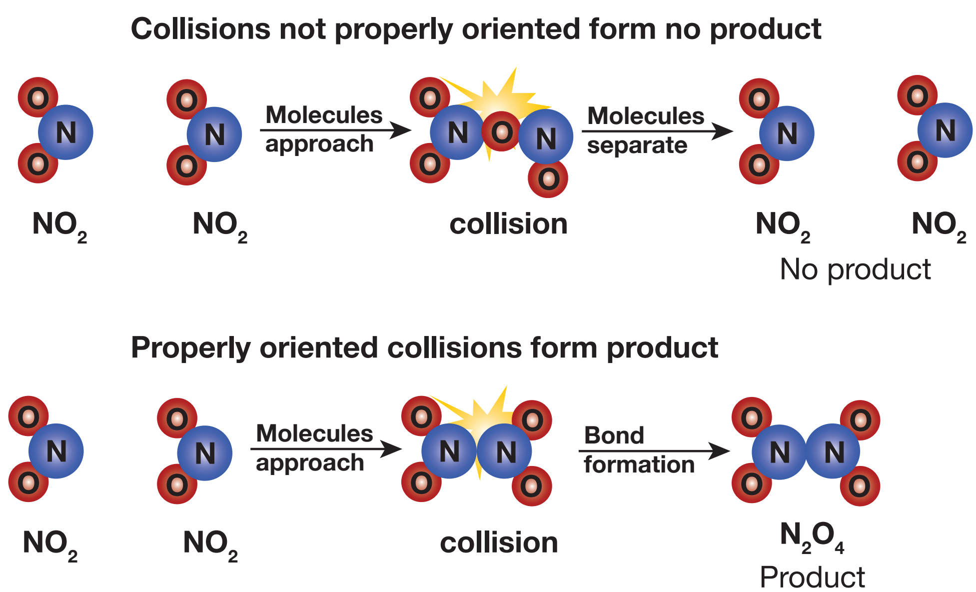



Figure 8.2 Effective and ineffective collisions of $NO_2$ molecules

**Table 8.1 Conditions that alters the rate of reaction.**

| Condition | Explanation |
|---|---|
| Temperature | As the temperature increases, the reactant molecules possess energy greater than the threshold energy to bring about effective collision per unit of time. |
| Concentration | Higher the concentration of reacting molecules, greater is the number of collision. |
| Area | The greater the surface area of a solid reactant, greater is the area available for the reaction to occur. |
| Catalyst | The presence of catalyst lowers the activation energy of the reactants by providing alternative pathway for reaction. |

## Self Evaluation

1. Explain collision theory of reaction.
2. Identify the condition(s) that increases the rate of reaction for the following
   (a) Digestion of food.
   (b) Burning of wood.
   (c) Making cheese from curd.
   (d) Photosynthesis.
   (e) Reaction between HCl and magnesium.
   (f) Ripening of fruits.
3. Define the term activation energy?
4. Why all collisions do not result into chemical reactions?
5. Which of the following two illustration; (A) and (B) lead to form required products? Give reason.

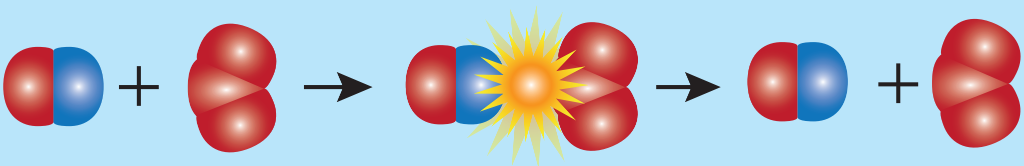



Illustration A

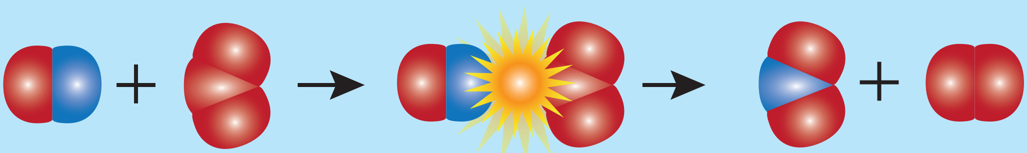



Illustration B

## 8.3 Reaction rates

**Learning objectives**

On completion of this topic, students should be able to:

» *explain the rate of reaction.*

» *write the expression for the rate of reaction.*

» *illustrate the factors affecting the rate of reaction.*

All chemical reactions proceed at different rates. The rates are always measured in a unit per time interval. The rate of a chemical reaction measures how fast reactants are converted into products. The rate decreases gradually as the reaction proceeds and become zero when the reaction is complete. Reactions can be slow or fast depending on the conditions such as temperature, pressure, concentration, presence of catalyst, etc. The speed of reaction can be determined by measuring its reaction rate.

### 8.3.1 Expressing reaction rates

The speed at which a chemical reaction proceeds is called the reaction rate. This can be expressed as a change in the amount of reactant consumed or product formed per unit change of time. So, the rate of a chemical reaction is expressed quantitatively in terms of the change in concentration of one of the reactants or products per unit change of time.

$$\text{Rate of reaction} = \frac{\text{change in concentration of reactant}}{\text{change in time}}$$

$$= \frac{\text{change in concentration of product}}{\text{change in time}}$$

Reaction rates are always positive by convention. A rate that is expressed as the change in concentration of a product is the rate at which the concentration of the product increases. The rate that is expressed in terms of the change in concentration of a reactant is the rate at which the concentration of the reactant decreases.

### 8.3.2 Factors affecting the rate of reaction

There are many factors which influence the rate of a chemical reaction. The reaction rate may depend on the nature of reactant, the structure of molecules, the collision between the molecules and the energy possessed by the reacting molecules. For example, the reactions of ions are usually fast as they have attraction for each other, while the reaction of covalent molecules is usually slow. It is important to note that a reaction does not have a single specific rate.

Reaction rates depend on the following four main factors as discussed below:

### i. Temperature

Most reactions are favoured at higher temperatures. For example, in the kitchen we increase the temperature to speed up the chemical processes of cooking food. Baking a cake speeds up the reactions that change the liquid batter into a spongy product. Lowering the temperature slows down most reactions. Photographic film and batteries stay longer if they are kept cool because the lower temperature slows the reactions to form products. Most manufacturing industries use either heating or cooling to control their processes for optimal performance. Adding heat to the reactants breaks the bonds and increases the speed of molecules or atoms. The faster they move, the more likely they will collide and react to form product. Removing heat decreases the rate of reactions. Thus, it is for this reason that the food does not get spoilt at faster rate on freezing.

At low temperature, the reactant particles move slowly and less collision occurs resulting into less products. Increase of temperature increases the kinetic energies of particles and they move faster and increase the frequency of collision. So, the number of effective collisions that form the products increases. Therefore, a rise in temperature increases reaction rate by increasing the collision frequency as well as the energy of the colliding particles. The rate of many reactions at room temperature is approximately doubled for every 10°C increase in temperature. Pressure has almost no effect on reactions taking place in the liquid or solid states. However, it does change the rate of reactions taking place in the gas phase.

**Activity 8.1 Investigating the reaction rate at different temperature.**

**Materials required**

Large test tube, test tube holder, measuring cylinder, water bath, acetic acid and calcium carbonate

**Procedure**

1. Take two large test tubes and label them as A and B.
2. Add 5 mL of acetic acid in both the test tubes.
3. Warm the acetic acid of the test tube A by placing it in water bath.
4. Remove the test tube A from the water bath and add one gram of calcium carbonate in each of the test tube A and B.
5. Observe what happens.

**Questions**

1. Based on your observation, in which acetic acid solution did the calcium carbonate react faster? Why?
2. What may happen on the reaction rate if sodium bicarbonate is used instead of calcium carbonate?
3. What can you conclude from the experiment?

### ii. Concentration

**Activity 8. 2 Investigating the reaction rate in different concentration.**

**Materials required**

HCl, two large test tubes, two small magnesium ribbons of equal size, measuring cylinder and a stop watch.

**Procedure**

1. Take two labelled large test tube, add 2 mL of Conc. HCl in one test tube and 2 mL of dil. HCl in another test tube.
2. Start a stopwatch at the moment you drop a small piece of magnesium ribbon into the first test tube.
3. Stop the stopwatch when the magnesium has finished reacting.
4. Repeat steps 2–3 with a second magnesium piece of same length or weight using the second test tube.

**Questions**

1. In which hydrochloric acid solution did the magnesium piece dissolves faster? Explain your observation.
2. Provide evidences to prove that the chemical reaction have occurred?
3. Name the gas produced during the reaction and write the balanced chemical equation for the reaction.
4. What conclusion can you draw from the experiment?

Changing the concentration of reactants can alter the rate of reaction. More the concentration of the reactants, the greater is the number of reactant particles in a unit volume which collide with each other. Therefore, if the solutions of reactants are concentrated, the reaction proceeds at faster rate giving more products. For example, if a fire is burning slowly, fanning the flames increases the concentration of oxygen and the fire burns faster.

In case of the gaseous reactant, concentration may be increased by increasing the pressure. Increasing of pressure does not increase the number of particles, but it brings the particles closer thereby increasing the collision frequency. For example, the Haber's process uses high pressure to increase the rate of reaction between hydrogen and nitrogen to form ammonia. Lowering the concentration decreases the rate of reaction.

### iii. Surface area

Chemical reaction involving a liquid and a solid reactants take place at the surface of the solid. The surface area of a solid can be increased by grinding them into small powder. When the surface area of reactants is increased, the frequency of collision between the reactant particles is also increased. This explains why campfires are started with paper and small twigs rather than logs, and salt dissolves faster in water when the mixture is stirred. Can you guess why our intestine is long?

### iv. Catalysts

A chemical reaction occurs only when the reactant molecules collide and attain the activation energy. A catalyst lowers activation energy by providing alternative pathway for reaction thereby allowing the larger portion of the reactant particles to collide. Adding catalyst to a reaction mixture increases the reaction rate, without itself undergoing any permanent changes. The process is called catalysis and it is widely used in the chemical industries.

**Analogy for catalysis**

Suppose there is a mountain between two valleys such that the only route for people to get from one valley to the other is over the mountain. In such case, only the most active people will manage to get from one valley to the other.

If a tunnel is cut through the mountain, more people will now manage to get from one valley to the other by this easier route. In this case, the tunnel has lower activation energy than going over the mountain.

Here, the tunnel has provided an alternative route without lowering the mountain.

Figure 8.3 Illustration for catalysed reaction

**Activity 8. 3 Investigating the role of catalyst**

**Material required**

Hydrogen peroxide, test tube, measuring cylinder and manganese dioxide.

**Procedure**

1. Take two test tubes and label them as A and B.
2. Add 1 mL of hydrogen peroxide to each of the test tube.
3. Add few crystals of manganese dioxide to test tube B.
4. Observe the reactions in both test tubes and answer the following questions.

**Questions**

1. In which test tube the decomposition of hydrogen peroxide is violent?
2. In which test tube the decomposition of hydrogen peroxide is negligible. Why?

Hydrogen peroxide solution is commonly used as a mild antiseptic and as a bleaching agent. It decomposes very slowly when stored in a bottle, forming oxygen as shown in the equation.

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

Adding few crystals of manganese dioxide $MnO_2(s)$ to hydrogen peroxide causes a violent decomposition. The manganese dioxide is one of the catalysts for the decomposition of hydrogen peroxide.

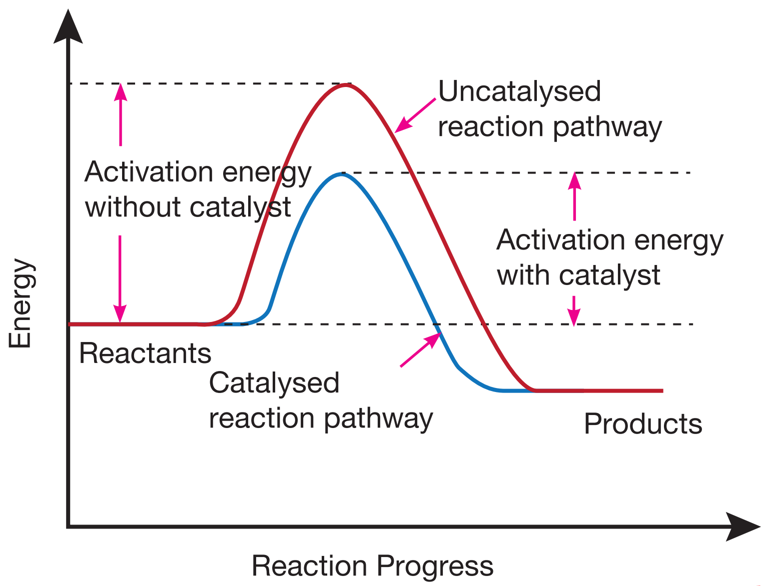



Figure 8.4 Uncatalysed and catalysed reaction

**Self Evaluation**

1. State the conditions that influence the rate of reaction.
2. What does the rate of a reaction indicate?
3. How is the reaction rate expressed?
4. Answer the following using the collision theory:
   (a) Why are catalyst used in a chemical reaction?
   (b) Explain the role of concentration of reactant in the reaction rate.
   (c) Increase in temperature increases the reaction rate. Explain the statement.
   (d) How does pressure influence the rate of a reaction in gas? Explain.
   (e) Could a catalysed reaction pathway have activation energy higher than the uncatalysed reaction? Explain.

## 8.4 Biological catalyst

**Learning objectives**

On completion of this topic, students should be able to:

» *explain enzyme catalysed reaction.*

» *explain the affect of temperature and pH on enzyme activity.*

» *state the importance of enzymes and their uses in biotechnology.*

Enzymes are proteins produced by the living cells. They are called biological catalysts as they enhance the rate of vital biochemical reactions such as breathing, digestion, pumping of heart, cell division, contraction of muscles, electrolyte balance etc. Enzymes function in a mild environment similar to a body temperature of a living organism to support a life. They are inactive at 0°C and destroyed by heat at 100°C. Enzymes are highly specific and can act on a single or a small group of closely related substances. These substances on which the enzymes act are called substrates. During catalytic action, the enzymes do not undergo any permanent modification and are regenerated at the end of the reaction. The general enzyme catalysed reaction takes place as per the equation:

$$\text{Enzyme} + \text{Substrate} \longrightarrow \text{Enzyme - Substrate complex}$$

$$\text{Enzyme - Substrate complex} \longrightarrow \text{Enzyme} + \text{Product}$$

Enzyme-substrate complex is an unstable complex. For example, amylase acts on starch and produce maltose units. In a given reaction amylase is the enzyme, starch is the substrate and maltose is the product.

$$\underset{(E)}{\text{amylase}} + \underset{(S)}{\text{starch}} \longrightarrow \underset{(ES)}{\text{amylase - starch}}$$

$$\underset{(ES)}{\text{amylase - starch}} \longrightarrow \underset{(E)}{\text{amylase}} + \underset{(P)}{\text{maltose}}$$

Where; E = Enzyme; S = Substrate; ES = Enzyme-substrate complex; P = product.

Emil Fischer, in 1894 postulated a lock and key analogy to explain the specific action of an enzyme with a single substrate. In his analogy, the lock is the enzyme and the key is the substrate. Only the correct size key (substrate) fits into the key hole (active site) of the lock (enzyme) as shown in the Figure 8.5 with respect to mechanism of an enzyme action.

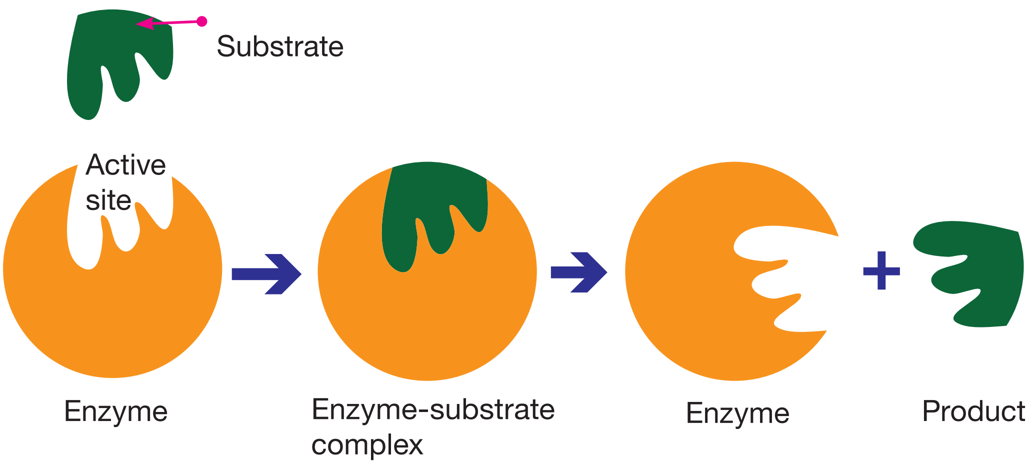



Figure 8.5 Mechanism of an enzyme action.

## 8.4.1 Factors influencing enzyme activity

Like all catalyst, enzymes take part in the reaction and alter the rate of reaction, but not the position of the equilibrium. The activity of an enzyme is affected by several factors like temperature, pH, enzyme concentration, substrate concentration, metal ions (activators), inhibitors, etc. However, the pH and temperature have significant influence on enzyme activity.

### i. pH

Enzymes are affected by changes in pH. All the enzymes have a particular pH at which their activity is maximum. The pH at which the enzyme shows maximum activity is known as optimum pH. Extremely high or low pH generally results in complete loss of activity for most enzymes because the changes in pH can make and break intra and intermolecular bonds, changing the shape and, its effectiveness. The optimum pH value will vary from one enzyme to another. Only in this optimum pH, ionisation of active amino acids in enzymes and substrate are favoured for enzyme – substrate complex formation. So at optimum pH, the rate of reaction is at an optimum. Any change in pH above or below the optimum pH will cause decrease in the rate of reaction.

### ii. Temperature

Rise in temperature causes increase in the rate of enzyme catalysed reactions up to a certain temperature of about 45°C. Above this temperature the activity declines due to denaturation of enzymes. Denaturation is the alteration of protein shape and structure due to an external stress such as temperature, high pH, etc. As the enzyme is denatured at high temperature and inactivated at low temperature, the reaction which it catalyses slow down and ultimately stops. So the temperature at which the enzyme shows maximum activity is known as optimum temperature. The optimum temperature of most of the enzymes is found to be 37°C. Every enzyme has optimum conditions at which its reaction rate is the fastest. A 10°C rise in the temperature will increase the activity of most enzymes by 50 to 100% because of increased kinetic energy. Why do you think it is advisable to drink warm water?

## 8.4.2 Importance of enzymes in biotechnology

a) Almost all reactions that take place in the human body are catalyzed by enzymes.

b) During digestion, enzymes speed up the breakdown of foods into smaller particles from mouth till the intestine which is absorbed by the cells.

c) Enzymes are even involved in the production of other enzymes in cells.

d) Enzymes are now used in medicine to treat disorders that result from their deficiency. For example, lactase pills are taken along with the dairy products to prevent from bloating and diarrhoea.

e) Enzymes catalyse many biological reactions and enhance the rate of product formation in metabolic pathways.

f) Some enzymes in blood are used as diagnostic indicators of various diseases. For example, the level of transaminase increases in blood during jaundice, a liver disorder. The level of transaminase in blood is used to determine liver function.

g) Some enzymes are used for therapeutic purposes like Penicillinase to treat patients allergic to penicillin, Asparaginase to treat leukaemia, Diastase to treat indigestion, Collagenase in treating skin ulcers etc.

### 8.4.3 Enzymes and their functions.

Some of the enzymes and their important functions are discussed in Table 8.2.

**Table 8.2 Enzymes and their functions**

| Enzyme | Function |
|---|---|
| Salivary amylase | Breaks down starch into smaller polysaccharides. |
| Pepsin | Breaks down proteins into small polypeptides |
| Pancreatic amylase | Breaks down starch and polysaccharides into disaccharides. |
| Lipase | Breaks down fats into glycerol, fatty acids. |
| Amino peptidase | Breaks down polypeptides into amino acids. |

**Self Evaluation**

1. What are biological catalysts? Give examples.
2. Why are enzymes highly substrate specific?
3. What happens to enzymes if they are heated above the optimum temperature?
4. Explain the role of pepsin in human body.
5. How do enzymes differ from other catalysts?

## Summary

1. A chemical reaction takes place only when the particles of reactants approach and collide with each other.
2. Only the colliding particles which possess sufficient energy react to form products. Other colliding particles simply rebound from each other unchanged.
3. The rate of a chemical reaction measures how quickly reactants are changed into products.
4. Reaction rates depend on factors such as temperature, pressure, concentration and the presence of catalyst.
5. Reaction rates generally increase with reactant concentration or, in the case of gases, pressure.
6. Increase of temperature increases the kinetic energies of particles and they

move faster and increase the frequency of collision.

7. The minimum amount of energy that the reacting molecule must possess for the collision to result in a chemical reaction is threshold energy.
8. Enzymes are biological catalysts that increase the rates of reactions important to an organism.
9. Enzymes catalyse many biological reactions and enhance the rate of product formation in metabolic pathways.
10. Some enzymes in blood are used as diagnostic indicators of various diseases.
11. Catalysts provide a pathway of lower activation energy.
12. pH and temperature has significant influence on enzyme activity.

## Exercise

**I. Fill in the blanks with correct word(s).**

1. _____ provide a pathway of lower activation energy.
2. Enzymes are _____ that increase the rates of reactions.
3. When the surface area is increased the ____of collision between the reactant particles is increased.
4. The colliding molecules must possess a certain minimum amount of energy called_____.
5. Pressure has almost no effect on reactions taking place in the liquid or _____ states.

**II. State whether the following statements are true or false. Rewrite the false statements correctly.**

1. The change in the amount of reactants or products over time is called the chemical kinetics.
2. The decrease in surface area of the reactants increases the rate of reaction.
3. The rate of a chemical reaction is increased by lowering the concentration the reactant.
4. Enzymes are affected by changes in pressure.
5. Heating the reactants breaks the bonds and increases the speed of molecules or atoms.

**III. Choose the most appropriate response from the given options.**

1. Which one of the following sets of reactants is most likely to have the slowest rate of reaction?

   A Two covalent compounds in solution.

   B Two ionic substances in solution.

   C Two solid ionic substances.

   D A covalent compound in solution and an ionic substance in solution.

2. The Figure 8.6 shows the volume of oxygen gas collected over the time for the decomposition of hydrogen peroxide. At which point on the graph has the reaction stopped?

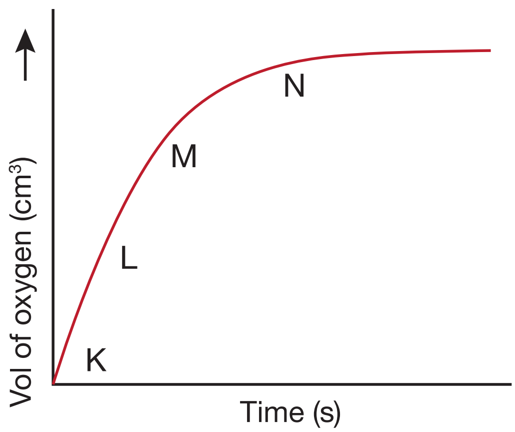



Figure 8.6

   A K. C M

   B L. D N.

3. An enzyme called diastase is use to treat

   A allergy. C leukaemia.

   B jaundice. D indigestion.

4. In solid and solution, the rate of a reaction is not affected by

   A catalyst. C pressure.

   B temperature. D concentration

5. Given the balanced equation representing a reaction:

$$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$$

   This reaction occurs more quickly when powdered iron is used instead of a single piece of iron of the same mass because the powdered iron

   A acts as a better catalyst than the single piece of iron.

   B absorbs less energy than the single piece of iron.

   C has a greater surface area than the single piece of iron.

   D is more metallic than the single piece of iron.

6. In a biochemical reaction, an enzyme acts as a catalyst, causing the
   A activation energy of the reaction to decrease.
   B potential energy of the reactants to decrease.
   C kinetic energy of the reactants to increase.
   D heat of reaction to increase.
7. A catalyst is added to a system at equilibrium. If the temperature remains constant, the activation energy of the forward reaction
   A decreases.
   B increases.
   C remains the same.
   D slightly increases.
8. Which one of the following is most likely to have the fastest rate of reaction?
   A Boiling of egg.
   B Setting of cement.
   C Burning of match stick.
   D Paint drying.
9. Which items correctly complete the following statement? A catalyst can act in a chemical reaction to
   I increases the equilibrium constant.
   II lower the activation energy.
   III decrease energy of the reaction.
   IV provide new path for the reaction.
   A Only I and II.
   B Only II and III.
   C Only III and IV.
   D Only II and IV.
10. Which of the following reactions will start with the highest rate?
   A 2g of limestone powder in 100mL of 0.1M HCl at 55°C.
   B 2g of limestone chips in 100mL of 0.1M HCl at 25°C.
   C 2g of limestone powder in 100mL of 0.1M HCl at 25°C.
   D 2g of limestone chips in 100mL of 0.1M HCl at 55°C.

## IV. Write answers for the following questions

1. A student wishes to investigate the change in reaction rate with the change in concentration of HCl (aq). Given the reaction:

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

   (a) Identify one variable that might affect the rate and should be held constant during this investigation.
   (b) Describe the effect of increasing the concentration of HCl(aq) on the reaction rate and justify your response in terms of collision theory.
2. How does pressure affect the rates of reactions in gas?
3. $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) + 890.4\,kJ$

Explain, in terms of collision theory, why a lower concentration of oxygen gas decreases the rate of this reaction.

4. The following are the data obtained by a student while performing the experiment on rate of a reaction. Study the data provided in the experiments and answers the questions that follow.

Experiment 1

| Solution A mL | Solution B mL | Water mL | Time in second |
|---|---|---|---|
| 10 | 10 | 0 | 16 |
| 9 | 10 | 1 | 20 |
| 8 | 10 | 2 | 22 |
| 7 | 10 | 3 | 24 |
| 6 | 10 | 4 | 26 |

Experiment 2

| Trial | Temperature °C | Time in second |
|---|---|---|
| 1 | 5 | 25 |
| 2 | 15 | 19 |
| 3 | 25 | 15 |
| 4 | 35 | 11 |
| 5 | 45 | 9 |

(a) Plot the data in a graph and draw line through the plotted point using time as 'y' axis and concentration of solution as 'x' axis.

(b) Based on the experimental data make a hypothesis about the effect of concentration of the reactants on reaction rate.

(c) Plot the data in a graph and draw line through the plotted point using time a 'y' axis and temperature as 'x' axis.

(d) Based on the experimental data make a hypothesis about the effect of temperature on the reaction rate.

(e) How does the collision theory relate to the rate of chemical reaction?

(f) What other factors affect the rate of reaction?

5. When lumps of calcium carbonate react with hydrochloric acid, carbon dioxide gas is released according to the equation

$$CaCO_3(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

What effect will the following have on the rate of the reaction?

(a) Increasing the temperature.
(b) Adding water to the acid.
(c) Using powdered calcium carbonate instead of lumps.

6. What are catalysts and how do they function?
7. Give an example of an enzyme-catalyzed reaction.
8. When hydrogen peroxide solution, used as an antiseptic is applied to a wound, it often bubbles. Explain.
9. Figure 8.7 shows activation energies for the decomposition of HI and HBr. Use the diagram to answer the questions that follow.

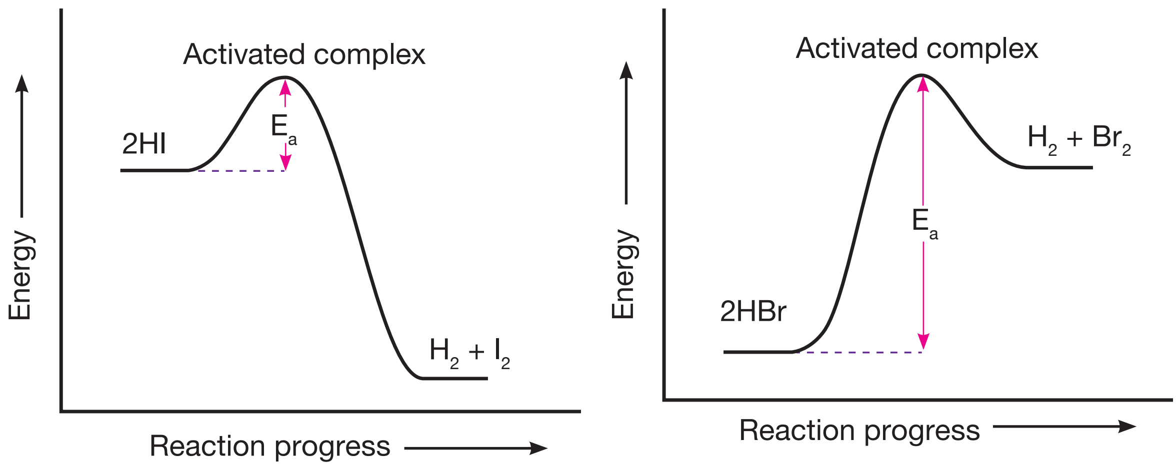



Figure 8.7 Activation energy diagram

(a) Which of these decomposition reactions is endothermic?
(b) Which of these reactions requires an input of energy to initiate the decomposition?
(c) Why does hydrogen bromide decompose more quickly than hydrogen iodide?
(d) How would each curve above change if a catalyst were added?

10. Explain why, even though a collision may have energy in excess of the activation energy, reaction may not occur.
11. Unlike gas-phase reactions, a reaction in solution is hardly affected by pressure. Explain.
12. Explain on each of the following statements by relating the factors that influences the rates of chemical reactions.

(a) Colours in the fabric of curtains in windows exposed to direct sunlight often fade.

(b) Meat is preserved longer when stored in a freezer rather than a refrigerator.

(c) Taking one or two aspirin will not harm most people, but taking at once can be fatal.

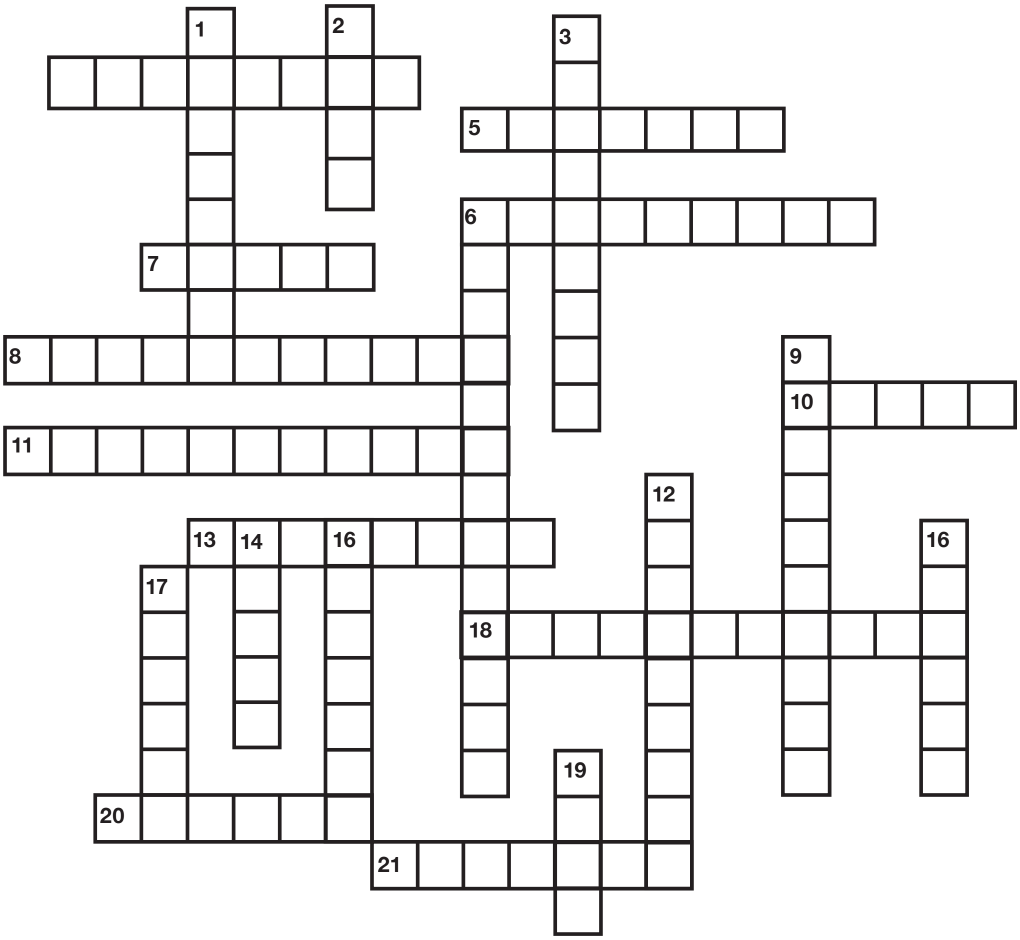
## V. Solve the crossword puzzle.

**Across**

4. Some enzymes in the blood are used as diagnostic indicators of various ____.
5. Particles have this kind of energy.
6. This helps reactions go faster.
7. This process uses high pressure to increase the rate of reaction of hydrogen and nitrogen to form ammonia.
8. In this reaction the energy is supplied to the reactant from the surrounding to form product
10. Enzymes are the proteins produced by the living _____.
11. The enzyme diastase is used to treat _____.
13. This has almost no effect on reaction taking place in the liquid or solid states.
18. Increasing this usually makes the reaction go faster.
20. Biological catalyst.
21. Particles must do this before they react.

**Down**

1. Catalyst provides the alternative pathway by lowering the activation energy of the ____.
2. When the reaction is complete the rate of reaction becomes ____.
3. At high temperature the enzymes are ______.
6. Increasing this usually causes a greater chance of collision.
9. The particles must have this energy to react.
12. The colliding molecules must possess minimum amount of this energy.
14. All chemical reactions proceed at different _____.
15. Chemical reactions involving a solid reactant take place at the _____.
16. Increasing the temperature gives the particles more ____.
17. This breaks down proteins into small polypeptides.
19. Reactions are like this at low concentration.


1 2 3 5 6 7 8 9 10 11 12 13 14 16 16 17 18 19 20 21

# 9 Alcohol

## 9.1 Introduction

An alcohol constitutes a member of organic compound in which a hydroxyl group (–OH) is bonded to a carbon atom of hydrocarbon chain. Therefore, alcohols are hydroxy derivates of alkanes, obtained by replacing the hydrogen atoms of alkanes by a hydroxyl group. The –OH is a functional group of alcohol. Alcohol, especially ethanol is prepared from starchy grains by fermentation both at small and large scale.

Alcohol and its properties have offered wide range of applications in industries and in daily life as an antiseptic, biofuel, solvent, paints and many more. The commercial production of alcohol, for example ethanol, in distilleries can boost up the nation's economy and mitigate the unemployment issues. However, the pattern in drinking habits can also bring negative impacts to health, society and economy.

## 9.2 Alcohol – structure, classes and nomenclature.

**Learning objectives**

On completion of this topic, students should be able to:

» *write the names of alcohol in its homologous series using common and IUPAC naming systems.*

» *derive alcohols from its corresponding alkane.*

» *differentiate the classes of alcohol based on number of -OH bonded to it.*

» *apply general formula of alcohol ($C_nH_{2n+1}$–OH) to write their formula.*

### 9.2.1 Homologous series and functional group.

The term homologous series is often used for a group of compounds that have the same functional group. Thus, a series of organic compounds with same functional group in which the members differ by one $–CH_2$ unit is referred to as homologous series. The individual members of such series are called homologues and the phenomenon is called homology. Each homologous series or families are represented by their general formula as shown in the Table 9.1, where 'n' is the number of carbon atom.

Table 9.1 General formula of some homologous series

| Homologous series | General formula |
|---|---|
| Alkane | $C_nH_{2n+2}$ |
| Alkene | $C_nH_{2n}$ |
| Alkyne | $C_nH_{2n-2}$ |
| Alcohol | $C_nH_{2n+1}OH$ |

Table 9.2 Homologous series of alcohol

| Name | Molecular formula ($C_nH_{2n-1}OH$) |
|---|---|
| Methanol | $CH_3OH$ |
| Ethanol | $CH_3CH_2OH$ |
| Propanol | $CH_3CH_2CH_2OH$ |
| Butanol | $CH_3CH_2CH_2CH_2OH$ |
| Pentanol | $CH_3CH_2CH_2CH_2CH_2OH$ |
| Hexanol | $CH_3CH_2CH_2CH_2CH_2CH_2OH$ |
| Heptanol | $CH_3CH_2CH_2CH_2CH_2CH_2CH_2OH$ |
| Octanol | $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2OH$ |

A functional group can be an atom or a group of atoms attached to some carbon atom in a hydrocarbon chain, which largely determines the chemical properties of an organic compound. Any organic compounds with same functional group belong to the same homologous series. For example, the saturated hydrocarbon (i.e., alkane) attached with hydroxyl (–OH) group is alcohol family. Thus, the term functional group is defined as an atom or a group of atoms in a molecule that gives the molecules its characteristic properties. According to IUPAC, each functional group has a characteristic name that often carry over in naming the individual compounds.

### Table 9.3 Functional group of some common family

| Family | Functional group | Suffix | General representation | Example |
|---|---|---|---|---|
| Alkane | - C - C - | -ane | | Ethane |
| Alkene | - C = C - | -ene | | Ethene |
| Alkyne | -C ≡ C - | -yne | | Ethyne |
| Alkanol (alcohol) | - OH | -ol | R - OH | Ethanol |
| Alkanoic acid (carboxylic acid) | — C(=O) — O — H<br>or -COOH | -oic acid | R - COOH | Ethanoic acid |
| Alkanoates (ester) | R — C(=O) — O — R' | -oates | R - $CO_2$ - R' | Ethyl ethanoate |
| Aldehydes | R — C(=O) — H<br>or -CHO | -al | R- CHO | Ethanal |

## 9.2.2 Alcohols – hydroxy derivates of alkane.

Alcohols are represented by the general formula ($C_nH_{2n+1}$-OH or ROH), where $C_nH_{2n+1}$ or R is an alkyl radical and –OH is a functional group. As represented in the general formula, alkyl radicals are group of atoms which are obtained by removing one hydrogen atom from its corresponding alkane. To name alkyl radicals, the letter ***–ane*** from alk**ane** is replaced by ***–yl***. Table 9.4 shows a list of alkyl group and its corresponding alkane.

### Table 9.4 Alkyl and its corresponding alkane

| Alkane ($C_nH_{2n+2}$) | Condensed formula | Alkyl ($C_nH_{2n+1}$) | Condensed formula |
|---|---|---|---|
| Methane | $CH_3$-H | Methyl | $CH_3$- |
| Ethane | $CH_3CH_2$-H | Ethyl | $CH_3CH_2$- |
| Propane | $CH_3CH_2CH_2$-H | Propyl | $CH_3CH_2CH_2$- |
| Butane | $CH_3CH_2CH_2CH_2$-H | Butyl | $CH_3CH_2CH_2CH_2$- |
| Pentane | $CH_3CH_2CH_2CH_2CH_2$-H | Pentyl | $CH_3CH_2CH_2CH_2CH_2$- |

The equation below shows how alcohol is derived from its corresponding alkane molecules. If one hydrogen atom from alkane is removed and then replaced by hydroxyl group (-OH), alcohol is obtained.

$$\underset{\text{(alkane)}}{C_nH_{2n+2}} \xrightarrow{+H} \underset{\text{(alkyl)}}{C_nH_{2n+1}} \xrightarrow{+OH} \underset{\text{(alcohol)}}{C_nH_{2n+1}OH}$$

For example, one hydrogen atom from methane is replaced by a hydroxyl group (-OH) to obtain methanol.

$$\underset{\text{(methane)}}{CH_4} \xrightarrow{-H} \underset{\text{(methyl)}}{CH_3-} \xrightarrow{+OH} \underset{\text{(methanol)}}{CH_3OH}$$

## 9.2.3 Structural representation.

Usually any organic compounds can be represented by more than one types of formula (i.e., molecular formula, graphic or structural formula and condensed formula). The Table 9.5 shows the representation of ethanol with different formula.

**Table 9.5 Representation of ethanol**

| Molecular formula | Graphic or structural formula | Condensed formula |
|---|---|---|
| $C_2H_5OH$ | $\begin{array}{ccccccc} & & H & & H & & \\ & & \vert & & \vert & & \\ H & - & C & - & C & - & O - H \\ & & \vert & & \vert & & \\ & & H & & H & & \end{array}$ | $CH_3CH_2OH$<br>Or<br>$H_3C - CH_2 - OH$ |


(a) Methanol

(b) Ethanol

Figure 9.1 3D structure of alcohol

**Activity 9.1 Worksheet**

**Instruction**

Copy and complete Table 9.6 in your note book. Alkane family is done for you. Using the same table answer the questions that follow:

**Table 9.6**

| Family / Property | Alkane | Alkene | Alkyne | Alcohol |
|---|---|---|---|---|
| Bond between carbon chains. | Single | | | |
| Functional group | C-C | | | |
| Saturated or unsaturated | Saturated | | | |
| General formula | $C_nH_{2n+2}$ | | | |
| First four members of the family | $CH_4$ | | | |
| | $C_2H_6$ | | | |
| | $C_3H_8$ | | | |
| | $C_4H_{10}$ | | | |

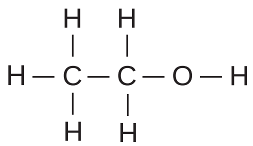
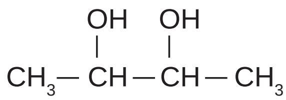
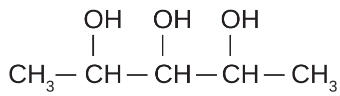
**Question**

1. Write down the common and the IUPAC name for first two members under each family.
2. On what basis do you classify each family as saturated or unsaturated? State the reason.
3. Predict the molecular formula for fifth and sixth members of each family.
4. Represent
   (a) butanol with its graphical formula.
   (b) methanol with its condensed formula.
   (c) propanol with its molecular formula.

## 9.2.4 Classification

Alcohols are classified as mono, di and tri hydric alcohol depending on the number of hydroxyl group attached in a molecule. Table 9.7 shows the differences between three classes of alcohol.

**Table 9.7 Classes of alcohol**

| Monohydric Alcohol | Dihydric Alcohol | Trihydric alcohol |
|---|---|---|
| These alcohols contain one –OH group attached to alkane. | These alcohols contain two –OH groups attached to alkane. | These alcohols contain three –OH groups attached to alkane |
| It is obtained by the replacement of one hydrogen atom of an alkane by a –OH group. | It is obtained by the replacement of two hydrogen atoms of alkane by two –OH group. | It is obtained by the replacement of three hydrogen atoms of alkane by three –OH group. |
|  Ethanol |  Butan-2,3-diol or 2,3-butanediol |  Pentan-2,3,4-triol or 2,3,4-pentanetriol |

## 9.2.5 Nomenclature

Generally, there are two systems of naming the alcohols as discussed below:

### i. Common naming system

In this system, the lower monohydric alcohols (R-OH) are named as alkyl alcohol. The alkyl group present in it is named first, followed by writing 'alcohol' as a separate word. Table 9.8 shows some common name for first four members of monohydric alcohols.

**Table 9.8 Common names of lower alcohols.**

| Alcohol | Name of alkyl group in alcohol | Common name |
|---|---|---|
| $CH_3OH$ | Methyl | Methyl alcohol or wood alcohol |
| $CH_3CH_2OH$ | Ethyl | Ethyl alcohol or alcohol |
| $CH_3CH_2CH_2OH$ | Propyl | Propyl alcohol or rubbing alcohol |
| $CH_3CH_2CH_2CH_2OH$ | Butyl | Butyl alcohol |

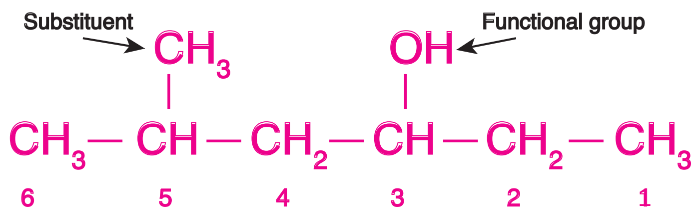
### ii. IUPAC system.

In this system, alcohols are generally named as alkanol by replacing the ending '-e' of alkane and adding the suffix '-ol'.

**Basic steps in IUPAC system of nomenclature.**

**STEP 1: Numbering the carbon atoms in a main chain or parent chain.**

Identify the longest continuous carbon chain that has –OH group attached to it followed by numbering of the carbon atoms. The numbering is done such that carbon chain containing –OH group gets the lowest possible number.




**STEP 2: Identifying the root word, suffix and prefixes.**

a) The number that indicates the position of the -OH is prefixed to the name of the parent alkane, and the -e ending of the alkane is replaced by the suffix -ol.

b) Name the substituent and the number that indicates the position of it is prefixed to it.

c) If more than one -OH group appears in the same molecule, suffixes such as -diol and -triol are used. In these cases, the -e ending of the alkane is retained.

**STEP 3: Naming the alcohol.**

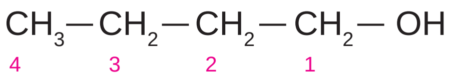
Use the sequence (**prefix + root word + suffix**) to name the alcohol.

In case of above structure, the name of alcohol is obtained as follow:

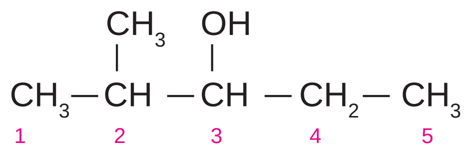
| | Prefix (indicates substituents) | Root word (indicates no. of C-atom in chain) | Suffix (indicates functional group) |
|---|---|---|---|
| | methyl | hexan (after dropping ending -e of 'hexane' | -ol |
| Position | $5^{th}$ carbon chain | | $3^{rd}$ carbon chain |
| Name | 5 – methyl - 3 - hexanol or 5 – methylhexan - 3 - ol | | |

**Solved examples**

Name the alcohol by studying the structures (a) and (b).



(a)



(b)

Solution (a)

| | Prefix | Root word | Suffix |
|---|---|---|---|
| | | butan | -ol |
| Position | | | $1^{st}$ carbon chain |
| Name | Butanol or 1-butanol or butan-1-ol | | |

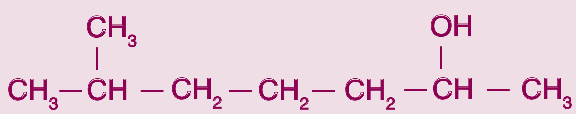
Solution (b)

| | Prefix | Root word | Suffix |
|---|---|---|---|
| | Methyl | pentan | -ol |
| Position | $2^{nd}$ carbon chain | | $3^{rd}$ carbon chain |
| Name | 2 –methyl -3-pentanol or 2-methylpentan-3-ol | | |

## Activity 9.2 Naming alcohol as per IUPAC system

**Instruction**

The structure given below is an alcohol. Study the structure and answer the questions that follow:



4. Name the root word and prefix for a given structure?
5. In which carbon atom is methyl group bonded? Write the carbon number.
6. Write the IUPAC name for the given structure.
7. To which class of alcohol does it belong? Explain.

## Self Evaluation

1. Write IUPAC name for the following structures.

(a) $CH_3 - CH_2 - CH_2 - OH$

(b)
$$\begin{array}{ccccc} CH_2 & - & CH_2 & - & CH_2 \\ | & & & & | \\ OH & & & & OH \end{array}$$

(c) $CH_3CH = CH_2$

(d)
$$\begin{array}{ccccccccc} & & & & OH & & & & \\ & & & & | & & & & \\ CH_3 & - & CH & - & CH & - & CH_2 & - & CH_3 \\ & & | & & & & & & \\ & & HO & & & & & & \end{array}$$

(e)
$$\begin{array}{ccccc} CH_3 & - & C & = & CH_2 \\ & & | & & \\ & & CH_3 & & \end{array}$$

2. Write structural formula for the compounds whose IUPAC names are as follow:
   (a) 2-methylbutan-2-ol.
   (b) Pentan-2,3-diol or 2,3-pentanediol.
3. How is functional group different from homologous series?
4. Define the term alkyl. Name the corresponding alkyl for butane and propane.
5. Write down the importance of homologous series.
6. Name and write the structure of one monohydric alcohol other than ethanol.

## 9.3 Properties of alcohol

**Learning objectives**

On completion of this topic, students should be able to:

» *explain the variation in physical and chemical properties of alcohols in its homologous series.*
» *apply the concept of hydrogen bonding to explain some of the properties of alcohol.*
» *state the conditions required to convert ethanol to ethene and ester.*
» *relate the use of breathalyzer with oxidizing property of alcohol.*
» *differentiate between denatured alcohol and spurious liquor.*

## 9.3.1 Physical properties

### i. Physical state

The lower members of aliphatic alcohol are colourless liquid at room temperature with a characteristic alcoholic odour and burning taste. However, the alcoholic odour and burning taste for higher members of alcohol gradually decreases with increase in molecular weight. The higher alcohols with more than twelve carbon atoms are colourless, odourless and waxy solids.

### ii. Density

All liquid alcohols have densities of approximately 0.8g/mL. Therefore, alcohols are less denser than the water whose density is 1g/mL at 4°C.

### iii. Hydrogen bonding

The term hydrogen bond is defined as a weak chemical bond formed between hydrogen atom having slight positive charge and an electronegative atom having slight negative charge. There are two types of hydrogen bonding:

a) intramolecular hydrogen bonding

b) intermolecular hydrogen bonding

| Intramolecular hydrogen bonding | Intermolecular hydrogen bonding |
|---|---|
| The bond is formed between hydrogen and some electronegative atom within the same molecules. | The bond is formed between hydrogen and electronegative atom of different molecules of the same substance or different substances. |
| These types of bonding takes place in aromatic compounds containing two functional groups.<br>Example; nitrophenol, nitro-benzaldehyde and salicylic acid. | These types of bonding takes place in following compounds to form associated molecules.<br>Example; Hydrogen fluoride, Water, Ammonia, lower alcohols, carboxylic acid. |

Hydrogen bonds are responsible for bonding in water molecules in liquid and solid state. Hydrogen bonds are stronger than van der Waal's force but weaker than covalent and ionic bonds. The hydrogen bonding is illustrated in the Figure 9.2 with respect to the association of several molecules of water.

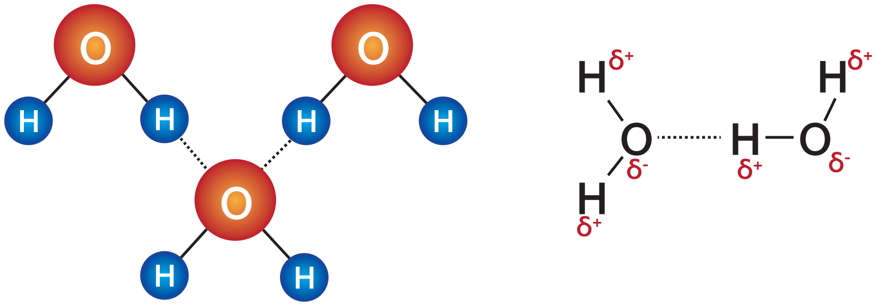
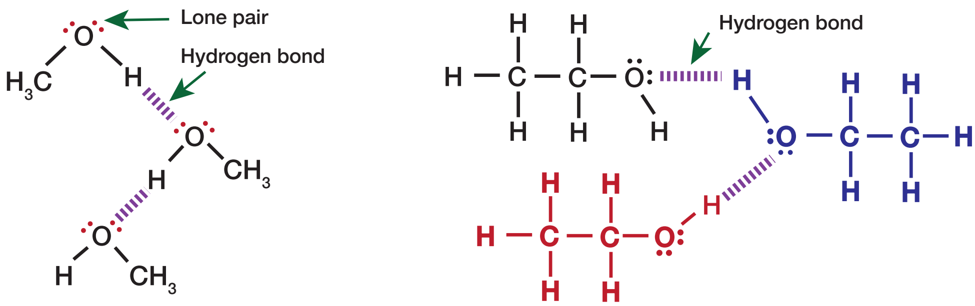



Figure 9.2 Hydrogen bonding in water molecules

In lower alcohols hydrogen bond is formed between hydrogen and electronegative oxygen atom of other molecule as shown in the Figure 9.3. The lower alcohols undergoing hydrogen bonding exhibits following characteristics:

- complete miscibility or solubility in water
- increased viscosity and surface tension
- higher melting point and boiling point




(a) Methanol

(b) Ethanol

Figure 9.3 Hydrogen bonding in alcohol

## iv. Boiling point and volatility

The boiling points of alcohols are higher than those of corresponding alkanes, haloalkanes and ethers due to the presence of strong intermolecular hydrogen bonding as shown in the Table 9.9. The hydrogen bonding results to associate alcohol molecules into larger aggregate, thereby requiring extra heat energy to separate the molecules from one another.

**Table 9.9 BP of alcohol with respect to its corresponding molecules.**

| Family | Alkane | Haloalkane | Ether | Alcohol |
|---|---|---|---|---|
| Compound | $C_2H_6$ (Ethane) | $C_2H_5Cl$ (Chloroethane) | $C_2H_5$- O –$C_2H_5$ (Ethoxyethane) | $C_2H_5OH$ (Ethanol) |
| Boiling point | -89°C | 12.3°C | 34.6°C | 78°C |

Within the alcohol series, there is increase in boiling point and melting point as the number of carbon atoms in the alcohol increase due to increase in its molecular mass. Hence, the lower members of alcohol in its series will have lower boiling point than those of higher members.

The term volatility of a substance is its tendency to convert into gaseous state easily. The substances with lower boiling point have higher volatility. Therefore, alcohols are less volatile than their corresponding alkane molecules due to the presence of hydrogen bonding. However, some members of alcohol are more volatile than the water and hence, alcohol is collected before water during distillation. For example, during the distillation of a mixture of ethanol and water, ethanol is collected before water.

### Activity 9.3 Molecular mass versus boiling point.

**Instruction**

1. Use the data from Table 9.10 to plot a line graph to study the trends in boiling point of alcohols and their corresponding alkanes.
2. Plot the boiling point on y –axis and the molecular mass on x –axis.
3. Use same grid paper to plot for both alcohol and alkane.

**Table 9.10 Boiling point and molecular mass of alcohol and alkane.**

| Alkane | Molecular mass | Boiling point |
|---|---|---|
| Methane | 16 | -161 |
| Ethane | 30 | -89 |
| Propane | 44 | -42 |
| Butane | 58 | -0.5 |
| Pentane | 72 | 36 |

| Alcohol | Molecular mass | Boiling point |
|---|---|---|
| Methanol | 32 | 65 |
| Ethanol | 46 | 78 |

| Alcohol | Molecular mass | Boiling point |
|---|---|---|
| Propanol | 60 | 97 |
| Butanol | 74 | 117 |
| Pentanol | 88 | 137 |

**Questions**

1. Describe the trends in boiling points of
   (a) alcohol
   (b) alkane family from lower to higher members in their series.
2. Why is the boiling point of ethanol higher than ethane although both have the same number of carbon atoms?
3. Write all the possible conclusions drawn from the graph.

## v. Solubility in water

The alcohols of lower members up to three carbon atoms in its homologous series are completely miscible in water in all proportion. This is because of the formation of intermolecular hydrogen bonding, when alcohols and water are mixed.

However, the solubility rapidly decreases for those alcohols which have more than three carbon atoms due to increase in molecular mass with their longer carbon chain. With the increase in chain length, the non polar alkyl group of alcohol dominates and hinders the formation of hydrogen bond which causes the decrease in miscibility of alcohols.

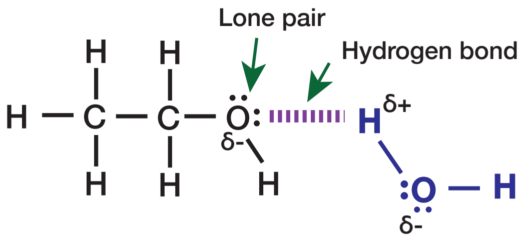



Figure 9.4 Intermolecular hydrogen bonding between ethanol and water

**Activity 9.4 Investigating the solubility of alcohol in water**

**Materials required**

Methanol, ethanol, propanol, test tubes, measuring cylinders and test tube rack.

### Procedure

1. Place about 10 mL of methanol into test tube, and add 2 mL of water. Repeat using 10 mL of methanol and the volume of water given in Table 9.11.

**Table 9.11**

| | | | | | |
|---|---|---|---|---|---|
| Volumes of methanol in mL | 10 | 10 | 10 | 10 | 10 |
| Volumes of water in mL | 2 | 4 | 6 | 8 | 10 |

2. Shake and examine the contents of each test tube to see whether one or two distinct layers are present.
3. Record your observation as shown in the observation Table 9.12.
4. Repeat procedure 1 to 3 using ethanol and propanol.

**Table 9.12 Observation table**

| Alcohol in water | No of layers formed | Miscible or immiscible |
|---|---|---|
| Methanol | | |
| Ethanol | | |
| Propanol | | |

### Safety precautions

Alcohols are toxic to organs.

Ensure that all ignition sources are removed from the area near the alcohol.

### Questions

1. Methanol and ethanol are miscible in water in all proportions. Do you agree with this statement as per the result obtained from the experiment?
2. Compare the solubility of propanol with the alcohols of lower members.
3. Predict the solubility of butanol and subsequent alcohol. Give reasons for answer.
4. What structural feature of alcohols is responsible for solubility of alcohol in water?

## vi. Viscosity

Viscosity is the property of a fluid that resists the force tending to cause the fluid to flow. Within alcohol series, the viscosity increases with increase in its molecular size. Thus, the viscosity is high for higher members of alcohol in homologous series due to strong intermolecular force.

## 9.3.2 Chemical properties of alcohol

### i. Combustion

Lower alcohols like methanol and ethanol burns with a blue flame producing carbon dioxide and water vapour in the presence of atmospheric oxygen. Such types of reactions are called combustion. During combustion, it releases a fair amount of energy. For higher alcohols it becomes increasingly more difficult to burn due to increase in their molecular weight.

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1368\,KJ/mol$$

When one mole of ethanol burns completely in air, it releases 1368 kilojoules of heat. The negative sign in a value indicates that the reaction is exothermic.

### ii. Oxidation

Alcohols on oxidation gives ethanal, commonly called acetaldehyde (a second members of aldehyde series) which on further oxidation produces ethanoic acid, commonly called acetic acid (a second members of carboxylic acid series). Potassium dichromate is used as an oxidizing agent as shown in the equation.

$$\underset{\text{(ethanol)}}{C_2H_5OH} \xrightarrow{K_2Cr_2O_7} \underset{\text{(ethanal)}}{CH_3CHO} + H_2O$$

$$\underset{\text{(ethanal)}}{CH_3CHO} \xrightarrow{K_2Cr_2O_7} \underset{\text{(ethanoic acid)}}{CH_3COOH}$$

If a bottle of wine is opened and exposed to the air, it will, after a while, starts to smell and taste like sour apples due oxidation of ethanol to ethanal. Likewise, if bottle is left open little longer, it will start to smell and taste like vinegar, due to further oxidation of ethanal to ethanoic acid by atmospheric oxygen in the presence microorganism such as ***Mycoderma aceti***.

A device called breathalyzer, to inspect the drunken driving is based on oxidizing property of ethanol. The breathalyzer contains an orange potassium dichromate (IV) which when reacted with ethanol forms ethanal. The orange potassium

Figure 9.5 Breathalyzer

dichromate (IV) itself gets reduced to green chromium compound. The presence of alcohol in breath is confirmed by observing the change in colour of potassium dichromate (IV) from orange to green on blowing into breathalyzer as shown in the Figure 9.5.

### iii. Esterification

The reaction in which alcohol reacts with carboxylic acid in the presence of concentrated sulphuric acid to form esters is called esterification. In this reaction, conc. $H_2SO_4$ acts as a protonating as well as a dehydrating agent.

Esters are the derivatives of organic acids which have pleasant, fruitlike odours and are responsible for flavors and fragrances of many fruits and flowers.

$$\underset{\text{(ethanoic acid)}}{CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-OH} + HO-CH_2CH_3 \xrightleftharpoons{\text{conc. } H_2SO_4} \underset{\text{(ethyl ethanoate)}}{CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-O-CH_2-CH_3} + H_2O$$

Esterification is an example of reversible reaction. However, the equilibrium can be shifted in the forward direction by removing water as soon as it is formed.

### Activity 9.5 Ester formation

#### Materials required

Test tube, sample of ethanol and butanol, acetic acid, dropper, conc. $H_2SO_4$, water bath and cold water.

#### Procedure

1. Mix thoroughly about $5cm^3$ each of ethanol and acetic acid in a test tube.
2. Carefully add $1cm^3$ of conc. $H_2SO_4$ to the mixture of ethanol and acetic acid, and warm the mixture at about 60°C in hot water bath for about 10 minutes.
3. Pour the mixture in a beaker containing ice-cold water. Record all your observation.
4. Repeat the above procedure using butan-1-ol and record all the observation including the smell.

#### Questions

1. Write equation for the reaction between ethanol and acetic acid.
2. Which substance is responsible for the production of fruity smell in the alcohol?
3. Why is the reaction mixture poured into cold water
4. Why is conc. $H_2SO_4$ used in the experiment?

### iv. Dehydration

The term dehydration is defined as the removal of water from a substance in the presence of the dehydrating agents.

**Case (i)** When an alcohol is heated up to 170°C in the presence of conc. $H_2SO_4$, the corresponding alkenes and water as by product are formed. For example, propanol reacts with conc. $H_2SO_4$ at 170°C to undergo dehydration to give propene and water as shown in the equation.

$$\underset{\text{(propanol)}}{CH_3CH_2CH_2OH} \xrightarrow{\text{excess of conc. } H_2SO_4 \text{ at } 170^\circ C} \underset{\text{(propene)}}{CH_3-CH=CH_2} + H_2O$$

**Case (ii)** If excess of propanol is heated to 140°C in the presence of conc. $H_2SO_4$, dipropyl ether is formed as shown in the equation. This reaction is often referred as condensation reaction.

$$\underset{\text{(propanol)}}{2CH_3CH_2CH_2OH} \xrightarrow{\text{conc. } H_2SO_4 \text{ at } 140^\circ C} \underset{\text{(dipropyl ether)}}{CH_3-CH_2-CH_2-O-CH_2-CH_2-CH_3} + H_2O$$

### v. Reaction with alkali metals

Alcohols are slightly acidic in nature as it has tendency to liberate hydrogen with the formation of alkoxides when it is reacted with active metals of group 1. For example, ethanol reacts steadily with sodium to liberate hydrogen and leaves a colourless solution of sodium ethoxide as shown in the equation.

$$\underset{\text{(ethanol)}}{2CH_3CH_2OH} + 2Na \longrightarrow \underset{\text{(sodium ethoxide)}}{CH_3CH_2ONa} + H_2\uparrow$$

This is the characteristic reaction of alcohol and is often used in organic analysis to identify the alcohol.

## 9.4 Denatured alcohol or Methylated spirit

Not all alcohols are safe for consumption except ethanol, after undergoing different stages of purification. Ethanol used for making alcoholic beverages is highly taxed, however, the ethanol supplied to industries and hardware store is cheaper. The ethanol supplied to industries and hardware stores are made unfit for consumption by mixing with one or more poisonous substances like methanol, pyridine and copper sulphate. This is called denatured alcohol. The poisonous chemical substances added to ethanol are called denaturant. Denatured alcohol containing 5 – 10 % methanol is referred to as methylated spirit. Methylated spirit is not even recommended to be used for making cosmetics such as perfume or toiletries.


Figure 9.6 Denatured alcohol and its statutory warning

### 9.4.1 Spurious liquor or illicit alcohol

Spurious liquor is prepared by diluting the denatured alcohol with water and then adding necessary colour and flavours. The alcohol obtained by this method is fatal for human consumption due to presence of impurities like methyl alcohol and acetaldehyde. Large number of deaths occurs due to consumption of spurious liquor. Thus, such alcohol is meant to be used as solvent for paint and varnishes.

### 9.4.2 Identification

It is difficult to identify the denatured alcohol from the consumption-grade ethanol, unless the containers are labeled. Denatured alcohols are colourless, but some are coloured by adding aniline to make it blue or purple for identification.

**Self Evaluation**

1. Explain the trend in boiling point of alcohols in their homologous series.
2. The solubility of alcohol decreases with the increasing number of carbon atoms. Explain.
3. Why open flames near to ethanol can lead to serious fires?
4. What are the constituents of methylated spirit? Where it is used?
5. On which property of alcohol does breathalyzer work? State its uses.
6. Name the products obtained on dehydration of ethanol. Write the chemical equation.
7. How can we identify the denatured alcohol from the ethanol?

## 9.5 Preparation and uses of ethanol

Although fermentation is the oldest chemical process for the preparation of ethanol from carbohydrates, it is one of the most widely used methods for industrial preparation even today. Fermentation is a slow decomposition of large molecules of certain organic compounds into simpler one under the catalytic influence of enzymes.

**Learning objectives**

On completion of this topic, students should be able to:

» *explain the principles of manufacture of alcohol in the distilleries.*

» *describe the stages involved in preparation of ethanol by different processes.*

» *explain the conditions required for fermentation.*

» *investigate the fermentation of glucose solution.*

» *state the uses of alcohols, especially ethanol*

### 9.5.1 Ethanol from starch by fermentation

The starch required for this process is obtained from maize, rice, barely, potato, etc. The starches present in those raw materials are converted into maltose in following steps:

#### i. Malting

Initially the moist barley is allowed to germinate in dark at 7°C to 17°C for about two days. This germinated barley is then dried, and is called Malt which is then heated up to about 62°C to stop further germination. It is then crushed and extracted with water. This Malt extract contains an enzyme called diastase.

#### ii. Mashing

In this process the starchy substances are steamed at 140°C to150°C under pressure to break the cell walls. The steaming of starch results to liberate a solution containing starch which is called a Mash. A mash is a fermentable starchy mixture from which alcohol or spirits can be distilled.

#### iii. Hydrolysis

In this process the mashed and the malted extract are treated together at 47°C to 57°C to obtain maltose.

$$\underset{\text{(starch)}}{2(C_{12}H_{10}O_5)_n} + nH_2O \xrightarrow{\text{diastase in malt extract}} \underset{\text{(maltose)}}{nC_{12}H_{22}O_{11}}$$

### iv. Alcoholic fermentation

Once the maltose is obtained from a starch, it is then fermented in presence of yeast. Yeast contains two enzymes (i.e., maltase and zymase) which perform different function as shown in the equation.

$$\underset{\text{(maltose)}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{maltase}} \underset{\text{(glucose)}}{2C_6H_{12}O_6}$$

$$\underset{\text{(glucose)}}{C_6H_{12}O_6} \xrightarrow{\text{zymase}} \underset{\text{(ethanol)}}{2C_2H_5OH} + 2CO_2$$

#### Activity 9.6 Investigating the fermentation of a glucose solution

Fermentation is a slow process and hence, the activity is designed to complete in three days. However, the number of days required for complete fermentation may vary from place to place depending upon the surrounding temperature.

**Day 1**

**Materials required**

Conical flask (100 mL), measuring cylinder (50 mL), digital balance, cotton wool, thermometer, warm water of 30°C–40°C, filter paper, 5 g of glucose and 1 g of yeast.

**Procedure**

1. Prepare a glucose solution by mixing 5 g of glucose and 50 mL of warm water in a conical flask. Ensure that the glucose is dissolved completely in water by swirling the flask.
2. Weigh 1 g of yeast and add it to the solution. Then loosely plug the top of a conical flask with cotton wool.
3. After completing the procedure 1 and 2, weigh the conical flask and record its total content as the initial mass.
4. Normally, the fermentation takes place within three days depending upon the reaction conditions. Therefore, secure the flask in a warm place, with a temperature of approximately 37°C for three days or until fermentation completes (as indicated when no more bubbles are released).

**Day 2**

**Materials required**

25 mL of lime water, beaker, conical flask, rubber stopper and rubber tube.

**Procedure**

1. Before the conduct of test for $CO_2$, measure and record the mass of a flask along with its content.
2. Then remove the cotton wool and replace with rubber stopper such that the rubber tube passing through it is immersed into a beaker containing lime water as shown in the Figure 9.7.

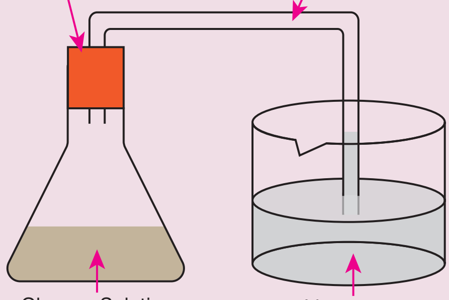



Figure 9.7 Experimental set up for testing $CO_2$

3. Carefully observe and record any changes in the appearance of lime water.
4. After finishing the test for $CO_2$, close the conical flask with a cotton wool. Secure in a place where solution is maintained at ideal temperature for fermentation.

**Procedure (Day 3)**

1. Once again re-weigh and record the final mass of a conical flask.

**Table 9.13 Observation table**

| Sl.No. | Mass of flask | Change in mass |
|---|---|---|
| Day 1 | | |
| Day 2 | | |
| Day 3 | | |

**Questions**

1. Why do we measure the mass of content in flask at various times throughout the investigation?

2. Why is rubber tube from flask connected to a beaker containing lime water? Write chemical equation to support your answer.
3. Why is yeast added to the glucose solution?
4. Why is temperature important factor during the process?
5. What is precipitate? What is the name of a white precipitate formed upon passing bubbles to the lime water?
6. Why is it necessary to close the content in conical flask with cotton wool?

## 9.5.2 Ethanol from ethene by hydration.

Hydration is employed for commercial production of ethanol from ethene. The ethene required is obtained by cracking of petroleum. The cases of hydration are:

### i. Indirect addition of water

In this case the ethene is heated in the presence of conc. $H_2SO_4$ at 75°C, and then treated with water to produce ethanol.

$$\underset{\text{(ethene)}}{H_2C = CH_2} + H - OH \xrightarrow{\text{conc.}H_2SO_4, 75^\circ C} \underset{\text{(ethanol)}}{CH_3 - CH_2 - OH}$$

### ii. Direct addition of water

In this case the ethene and the steam are compressed at 60 - 70 atmospheric pressure and then passed over phosphoric acid.

$$\underset{\text{(ethene)}}{H_2C = CH_2} + H - OH \xrightarrow{60\text{-}70\,\text{atm}, 300^\circ C} \underset{\text{(ethanol)}}{CH_3 - CH_2 - OH}$$

Ethanol obtained by this method may not be a sustainable practice, since the raw material used is non-renewable resource.

## 9.5.3 Ethanol from molasses - commercial production

In distilleries any fruits, grains or vegetables that can be fermented are the raw materials for alcohol production. In case of molasses and some fruits where sugar is already present, fermentation can be started directly. However, if grains are used as raw materials where sugar is not primarily present, fermentation can take place only after converting the starch into sugar.

Molasses are dark coloured syrupy mass and contains about 60% of fermentable sugars like sucrose and invert sugar (a mixture of glucose and fructose). It is obtained from sugar industry after separation of cane or beet sugar crystals from concentrated sugar cane juice. The steps employed during manufacturing process are:

## i. Dilution

The molasses is diluted with water to bring down its concentration up to 10% and then acidified with a small amount of sulphuric acid to avoid bacterial growth. Some nutritive solutions of ammonium salts like $(NH_4)_2SO_4$ or $(NH_4)_3PO_4$ are added in case of insufficient yeast.

## ii. Fermentation

After dilution, yeast is added and the temperature of a mixture is maintained at 30°C for a few days to undergo fermentation. The function of an enzyme (invertase and zymase) in yeast is shown in the equations. The carbon dioxide formed is allowed to escape in air by avoiding the entry of air into content as its presence would oxidize ethanol to acetic acid.

$$\underset{\text{(sucrose)}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{invertase}} \underset{\text{(glucose)}}{C_6H_{12}O_6} + \underset{\text{(fructose)}}{C_6H_{12}O_6}$$

## iii. Distillation

Fermented liquid contains about 8 to 10% ethanol and are called wash. It is distilled in a distillation plant to remove water and other impurities present in wash. The distillate contains about 90% ethanol and the residue left is used as cattle feed.

## iv. Rectification

The alcohol obtained after distillation still contains impurities other than water. These impurities are further removed by fractional distillation. The process of repeated distillation to obtain pure alcohol is called rectification.

Figure 9.8 Exterior view of distillation plant.

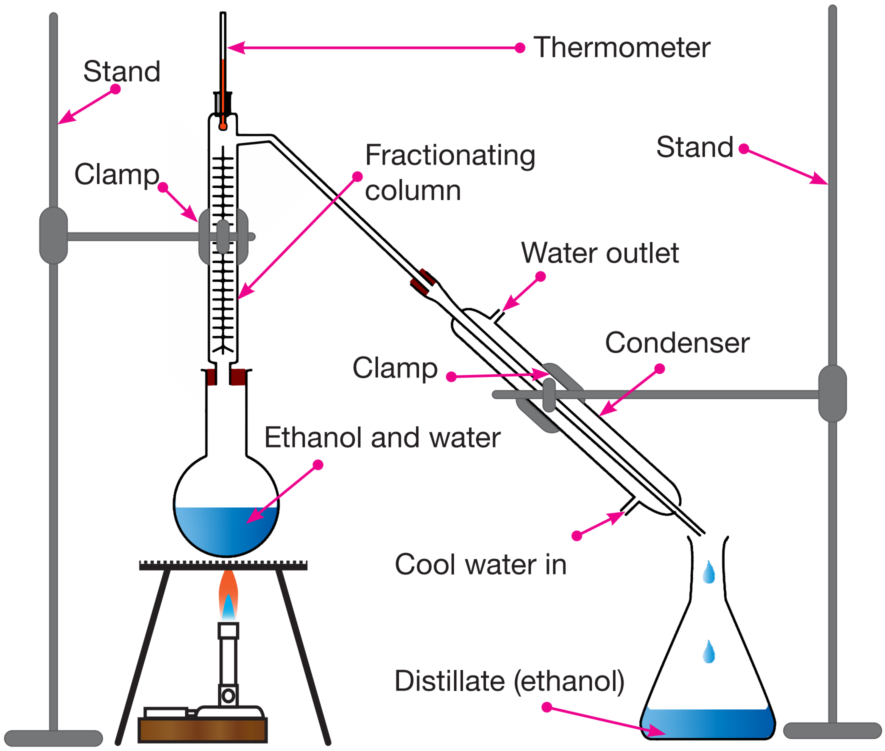



Figure 9.9 Fractional distillation

### 9.5.4 Uses of ethanol

#### i. as biofuel.

When 30% ethanol is blended with petrol, the resulting mixture is called gasohol and it is used as a fuel to power automobiles. The use of gasohol reduces our dependence on non-renewable petroleum products and pollution to environment.

#### ii. in medicine.

1. for preparation of pharmaceutical products. Liquid based medicines such as cough syrups, vitamin tonics etc., contain ethanol.
2. it is used in intravenous (IV) injection.
3. it is used for its antiseptic properties. It is often found in antibacterial wipes and hand sanitizers as it is effective in fighting against the spread of bacteria, fungi and viruses.

### iii. as beverages.

1. it is main ingrident in preparation of alcoholic beverages.
2. it has great cultural and cermonial significance.

### iv. as a solvent.

for fats, gums, resins and many other organic compounds.

Many covalent solutes which do not dissolve in water are soluble in ethanol. Quick drying lacquers for example, often contain ethanol as solvent. Ethanol is volatile, so it evaporates quickly, leaving the dry lacquer. Ethanol is used as a solvent in perfumes.

### v. as a fluid for scientific apparatus.

e.g., dyed ethanol is used as fluid in thermometer, spirit levels, etc.

### vi. as a reagent.

in manufacturing of chemicals like acetyl dehyde (dye), acetic acid (manufacture of veniger), chloroform (antiseptics), diethyl ether (anesthesia) etc.

**Self Evaluation**

1. Explain:
   (a) Ethanol should be rectified after distillation.
   (b) Ethanol is a good solvent for most of the non polar covalent substances.
2. Name the raw materials used to produce alcohol in distilleries.
3. State whether the fermentation process is a chemical or physical change with a valid reason to support your answer.
4. Which industrial methods of preparing alcohol are more economical? Explain.
5. Name the byproduct obtained during fermentation of glucose.
6. Ethanol is widely used in antibacterial wipes and hand sanitizer. Explain.

# 9.6 Ethanol and its impacts

**Learning objectives**

On completion of this topic, students should be able to:

» *compare the advantages and disadvantages of ethanol production to economy and environment.*

» *interpret ethanol cycle.*

» *describe the societal and health issues of drinking alcohol.*

## 9.6.1 Impact on environment

Both ethanol and petrol undergo combustion producing $CO_2$ and heat.

The following equations show a complete combustion of ethanol and octane (petrol):

$$CH_3CH_2OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O \qquad \Delta H = -1367\ KJ/mol$$

(ethanol)

$$C_8H_{18} + 12.5\,O_2 \longrightarrow 8CO_2 + 9H_2O \qquad \Delta H = -1367\ KJ/mol$$

(octane)

In both the cases, carbon dioxide gas is released. However, the good thing in using ethanol fuel is that the $CO_2$ emitted during combustion is literally considered as neutral to the environmental. This is because the amount of carbon dioxide released is same as the amount absorbed by the crops grown as feedstock for ethanol production. The principle so called 'carbon neutral' during burning of biofuel sounds good, but there are still some underpinning environmental issues such as deforestation and disturbances to the ecosystem in the process of mass cultivation of feedstock for ethanol.

When petroleum products are used as a fuel, the incomplete combustion produces pollutants such as carbon monoxide, soot and polyaromatic hydrocarbon which would pose a threat to the environment and the health of people. The gasohol, a fuel which is mixture of petrol and ethanol are considered more eco-friendly. This is because the oxygen in ethanol ensures the cleaner and complete combustion of petrol (octane).

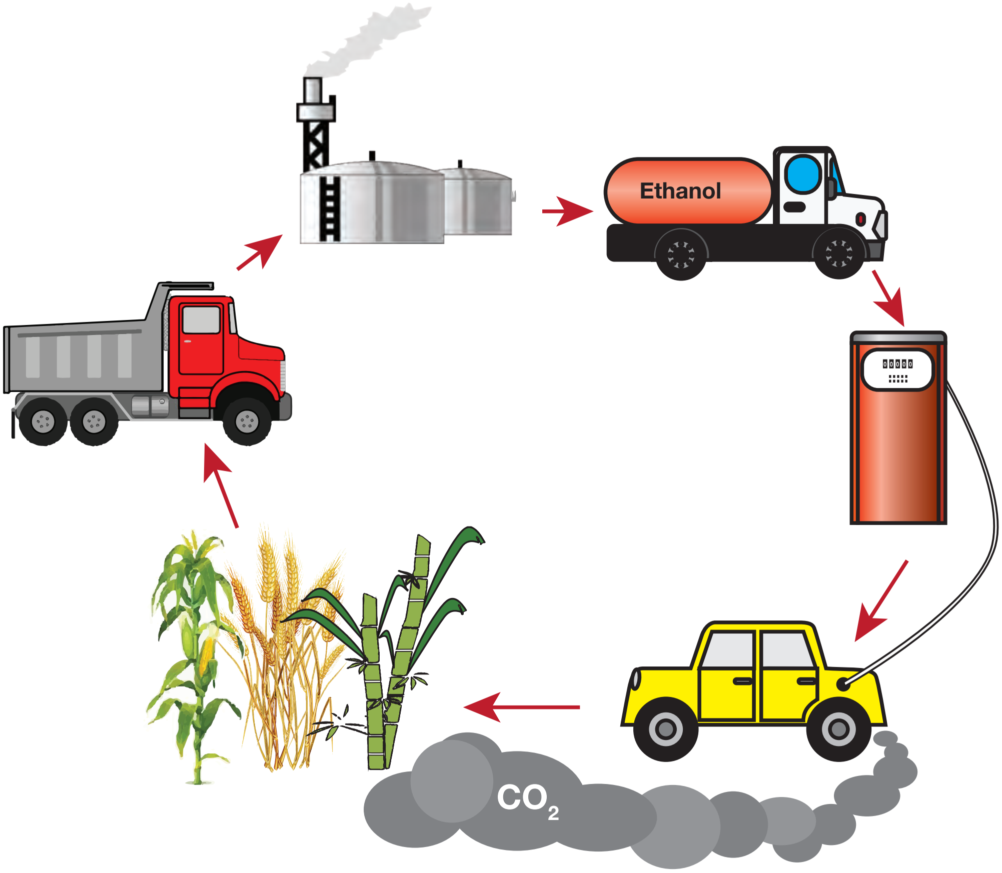



Figure 9.10 Ethanol cycle

## 9.6.2 Impact to economy, society and health.

Ethanol is well known alcohol in its homologous series, since it is used in the preparation of alcoholic beverages and fuel. An alcohol based fuels are called biofuel and are mainly obtained from the starchy agricultural produce such as corn, potato, wheat etc. Ethanol production is not only environmental friendly but also creates a potential for economic growth. However, the drinking pattern among the consumers is known to bring serious threat to health and socio-economic issues.

## Activity 9.7 Case study

**This excerpt is adapted from article titled "The Myth behind Alcohol Happiness" authored by Dr. Chencho Dorji, consultant psychiatrist and technical advisor to the National Mental Health Programme - Bhutan.**

### Effect on economy

Economists have calculated direct and indirect costs of drinking in Western countries, but it is difficult to determine a clear economic impact of alcohol use in the Bhutanese context because of the lack of reliable information and a database. Moreover, factors such as economic loss due to reduced efficiency and productivity, job loss, and other social and relationship problems are really impossible to estimate. However, given the widespread use of alcohol in the country, the economic cost is clearly enormous. Two studies by the Ministries of Agriculture, Trade and Industry in rural villages point out that as much as 50 percent of the grain harvests of households are used to brew alcohol each year. These findings have prompted local governments and the National Assembly to issue resolutions to ban the sale and consumption of homemade alcohol in public places.

It is presumed that homemade alcohol production represents more than industrial production since 80 percent of the population who live in rural villages consume mainly this type of alcohol. Not only it is cheaper (no tax or excise is payable) and readily available, it is more popular among drinkers. Some argue that while regular drinkers spend from half to three-quarters of their earnings on alcohol, heavy drinkers spend all their income or even borrow money to do so.

In a drive to increase domestic revenue, however, the Royal Government liberalized the sale and cost of bar licenses in 1999. Now there is one bar for every 250 Bhutanese and 10 bottles of alcohol per year for every man, woman and child in Bhutan. This is an alarming news to Bhutanese, whose national goal is to achieve Gross National Happiness. Assuming these figures are correct, Bhutan has perhaps one of the highest per-capita alcohol consumption rates in the developing world. While revenue from alcohol sales has reached an unprecedented high-close to US dollars 2.5 million annually – and accounts for one of the top 10 revenue-generating industries, increased alcohol revenues do not cover the enormous cost of alcohol-related problems in the country including loss of productivity, premature deaths, increased treatment costs and other social problems.

### Social impact of alcohol

At the same time, the adverse social impact of alcohol although easy to see, is likewise difficult to measure. Unemployment, poverty, relationship problems, divorce and parental separation, neglect and abuse of children, drunken brawls and domestic violence, crime, accidents and deaths are commonly associated with heavy drinking. Alcohol is also held responsible for high-risk behaviors such as unsafe sex, sexual promiscuity, and use of other psychoactive substances. Men experience more alcohol-related problems than women, but women are often direct victims of the consequences of men's drinking. For example, women who live with heavy drinkers are more at risk of serious violence, when compared to women who live without heavy drinkers. Drinking by women of childbearing age may also increase the risk of unwanted pregnancies and other social complications. Children are invariably affected directly or indirectly.

**Alcohol-related health problems**

Data from hospitals and health centres, as well as from community surveys, indicate that alcohol is a leading cause of mortality and morbidity in middle-aged Bhutanese men and women. According to these health statistics, alcohol is one of the five leading causes of deaths (all age groups) in Bhutan and responsible for as many as 30 percent of deaths in the adult hospital wards. It is the number one killer of adult men in Bhutan today.

Prevalence studies in the country show that as much as 50 percent of the population drinks alcohol (mainly homemade), and nearly 20 percent drink regularly, with an average consumption of five bottles per week. Up to 40 percent of school children even admitted to drinking alcohol at least once. Police sources further indicate that drunk driving is the top cause of motor vehicle accidents in the country. Many people mistakenly believe that homemade alcohol is less harmful to health than the industrial variety. Actually, scientists have found that homemade and cheaper variety alcohol is more damaging to liver because of its higher aldehyde content. Heavy drinkers all over the world drink mainly cheap alcohol because of their poor economic situations. Alcohol can damage nearly every organ and system in the body; its psychoactive action can alter the functioning and structure of the brain. Its use contributes to more than 60 diseases, including cirrhosis of the liver, heart disease, and cancer.

Research has shown that low or moderate consumption of alcohol is beneficial to people who are 40 years and older because of its protective effects for coronary or ischaemic heart disease. However, the patterns of drinking, often with heavy episodic consumption among many consumers are likely to increase rather than decrease the occurrence of coronary heart disease. Drinking to intoxication is a significant cause of alcohol-related injuries and accidents.

The World Health Organization (WHO) has pointed out that alcohol is one of the most important risks to health in the world today, responsible for almost two million deaths (3.2 percent of total deaths) and accounting for four percent of the global disease burden in 2002. Alcohol is the leading cause of disability among men in developed countries and the fourth leading cause of disability in developing countries. Therefore, not only many precious lives are lost in their prime to alcohol, but also the direct and indirect costs of treatment of alcohol





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related health problems are staggering.

**Questions**

1. Why is the alcoholic beverages produced in distilleries safer for consumption than locally brewed alcohol?
2. The import of foreign alcoholic beverages is highly taxed with respect to those alcoholic beverages produced within Bhutan. Explain.
3. Discuss some advantages and disadvantages, assuming that the production of alcohol in distilleries of Bhutan is stopped for a year.
4. What was basis to the statement 'alcohol is number one killer of adult men in Bhutan today'?
5. How many types of alcohol related diseases are known? Name at least five diseases.
6. What can you do to minimize alcohol related issues as a responsible citizen of GNH country?

**Self Evaluation**

1. Although ethanol fuel is preferred over petrol to power the automobiles, petrol and diesel are being used extensively to run automobiles. Justify.
2. State some environmental impacts of biofuel production?
3. Can biofuels help mitigate climate change? Write your opinion.
4. How could an environmentally sustainable biofuel production be ensured?
5. What changes to agricultural land would biofuel production require?

## Summary

1. An atom or a group of atoms in organic molecules that gives the molecules its characteristic properties is called functional group.
2. Hydroxy (–OH) is a functional group of alcohol family.
3. A series of organic compounds with same functional group in which two adjacent members are differed by one $-CH_2$ unit is referred to as homologous series.
4. Alcohols are classified as mono, di, tri and poly hydric depending upon the number of hydroxyl group present in it.

5. The physical properties of alcohol such as its high melting and boiling point, complete solubility in water, high viscosity and surface tension are due its ability to undergo intermolecular hydrogen bonding.

6. The boiling points of alcohols are higher than those of its corresponding alkane molecules.

7. Breathalyzer is a device used for measuring the amount of alcohol breath. It is based on oxidizing property of alcohol.

8. Esterification is a chemical reaction resulting in the formation of at least one ester product when alcohol reacts with carboxylic acid in the presence of conc. $H_2SO_4$.

9. Alcohols are combustible in nature. Ethanol is widely used to power the automobiles because of its combustible property.

10. Denatured alcohol is ethanol that has additives to make it poisonous or nauseating, to discourage recreational consumption.

11. The term fermentation is referred to a slow decomposition of large molecules of certain organic compounds into simpler one under the catalytic influence of enzymes.

12. Rectification is process of purification of ethanol by repeated distillation.

13. Ethanol production and its uses as biofuel are considered as carbon neutral to the environment.

14. Alcohol are widely use as beverages, biofuel, industrial solvent, antiseptic in hospitals, etc.

## Exercise

**I. Correct the following statement by changing the letter in bold only.**

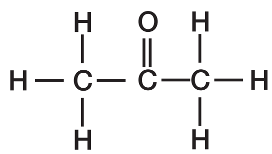
1. Alkanol is **common** system of naming alcohols in its series.
2. An enzyme in the yeast that converts sugar to glucose is **diastase**.
3. Alcohols are example of **non polar** compound.
4. Ethene is converted to ethanol by **decomposition** reaction.
5. An alcohol that has its boiling point just below propanol is **methanol**.

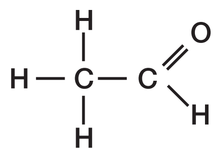
## II. Match the items of Column I with the corresponding items of Column II.

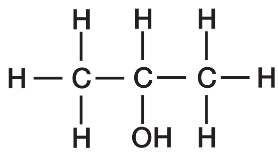
| Column I | Column II |
|---|---|
| 1. Aldehydes | a. –OH |
| 2. Carboxylic acid | b. Ethyl ethanoate |
| 3. Alkane | c. –CHO |
| 4. Alcohol | d. –ane |
| 5. Alkanoates | e. Ethanoic acid |
| | f. –ene |

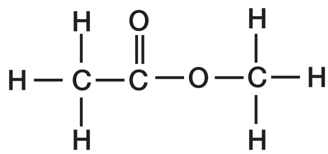
## III. Choose the most appropriate response from the given options.

1. Which graphic formula represents an alcohol?

A. 

B. 

C. 

D. 

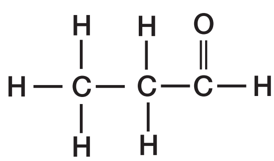
2. What is the missing product during incomplete combustion of ethane?

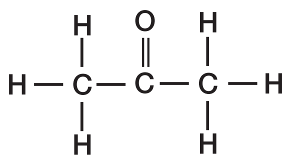
| | | | |
|---|---|---|---|
| A | $CH_3OH$ | C | HCOOH |
| B | $H_2O$ | D | $H_2O_2$ |

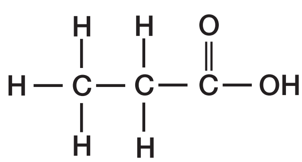
3. Which pair is obtained from the fermentation of glucose?

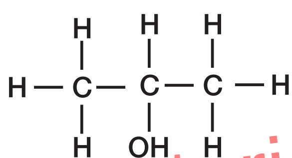
| | | | |
|---|---|---|---|
| A | Cellulose and water | C | Carbon dioxide and water |
| B | Ethanol and oxygen | D | Ethanol and carbon dioxide |

4. Study the organic structures below to answer the question.

i. 

ii. 

iii. 

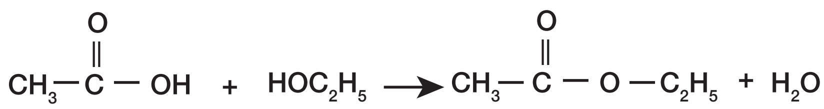
iv. 

Which pair of structure contains alcohol and acid?

A (i) and (ii) C (i) and (iii)

B (ii) and (iv) D (iii) and (iv)

5. The chemical reaction below is an example of



A fermentation C saponification

B hydrogenation D esterification

6. Which substance when dissolved in water forms a solution that conducts an electric current?

A $C_2H_5OH$ C $C_6H_{12}O_6$

B $C_{12}H_{22}O_{11}$ D $CH_3COOH$

7. Which general formula represents the compound, $CH_3CH_2CCH$?

A $C_nH_n$ C $C_nH_{2n}$

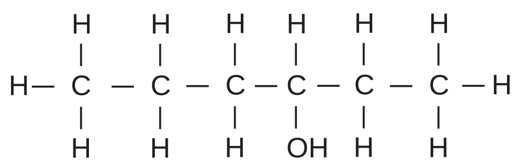
B $C_nH_{2n-2}$ D $C_nH_{2n+2}$

8. Which of the following is a renewable resource?

A Ethanol C Petroleum

B Uranium D Aluminium

9. What is the IUPAC name for a compound represented by the structure below?



A Hexan-3-ol C Heptan-3-ol

B Hexan-4-ol D Heptan-5-ol

## IV. Write answers for following questions.

1. To which homologous series does $CH_3CH_2CH_2CH_3$ belong?
2. Ethanol ($C_2H_5OH$) is a volatile and flammable liquid with a distinct odour at room temperature. Ethanol is soluble in water. The boiling point of ethanol is 78.2°C at 1 atmosphere. Ethanol can be used as a fuel to produce heat energy, as shown by the balanced equation.

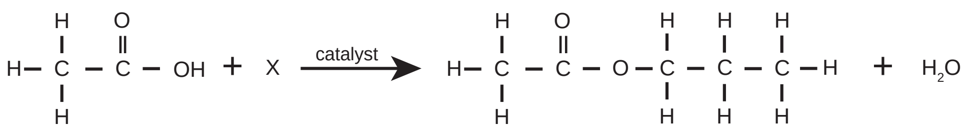
$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O + 1367\,kJ$$

(a) State one use of ethanol based on its volatile nature.

(b) Define the term flammable substances.

(c) Identify one physical property of ethanol, stated in the passage that can be explained in terms of chemical bonds and intermolecular forces.

3. The incomplete equation below represents an esterification reaction.



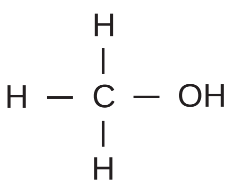

(a) Identify and write functional groups present in the reactant and product molecules.

(b) Write down the IUPAC name for the reactant represented by its structure in the equation.

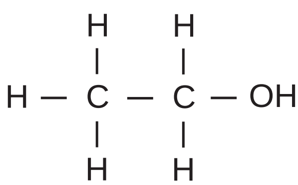
(c) Write condensed formula for the reactant molecule represented by its structure.

(d) Write name and structural formula of the reactant represented by X.

4. The structures shown below are first two members in a homologous series of alcohol. On the basis of these structures, answer the questions that follow:



methanol



ethanol

(a) Represent the next two succeeding members in series with its structural formula.

(b) Name the corresponding alkanes for two members represented in (a).

(c) Which of the two alcohols represented in (a) is more miscible in water? Explain.

5. Petroleum and sugar cane are both raw materials used for the production of ethanol.

(a) Construct separate flow diagrams for the production of ethanol from each raw material.

(b) Compare the environmental sustainability of producing ethanol from these two raw materials.

6. List down all the possible advantage and disadvantages of ethanol as a fuel.

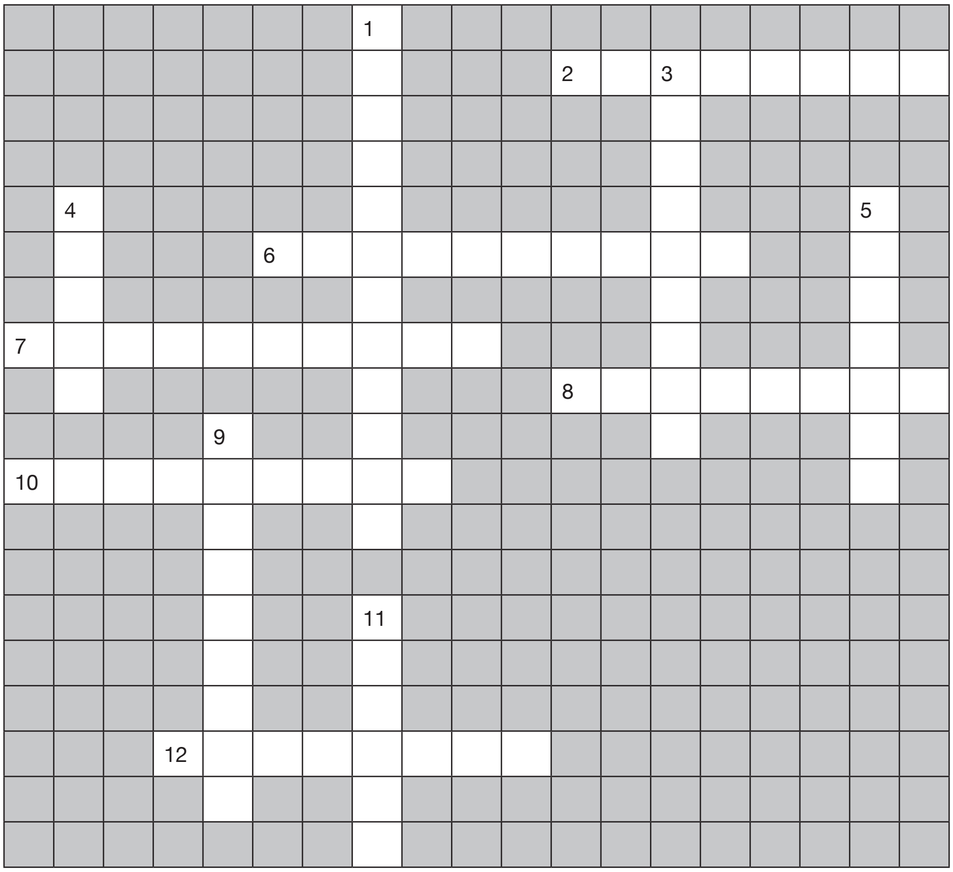
## V. Complete the word puzzle.

**Down**

1. The chemical process in which micro organism such as yeast act on carbohydrates to produce ethanol and carbon dioxide.
3. The use of breathalyzer is based on _____ property of alcohol.
4. Ethanol burns in air to produce carbon dioxide and___.
5. The functional group of alcohol family.
9. The hydrocarbon species which has single bond between carbon atoms is called _____ hydrocarbon.
11. The enzyme complex in yeasts that catalyzes the breakdown of sugar into alcohol and carbon dioxide.

**Across**

2. The IUPAC name of monohydric alcohol which has three carbon atoms.
6. Alcohol containing two replaceable atoms of hydrogen.
7. The poisonous substance added to ethanol in denaturing alcohol is called ____.
8. The next member of alcohol after butanol in homologous series.
10. The chemical process or reaction that favors the conversion of ethanol to ethene.
12. An alcohol with the lowest boiling point in its homologous series.


1
2
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# Specimen Question Paper

**Chemistry** | **Time: 2 Hours**

**Class X** | **Total Marks: 100**

**READ THE FOLLOWING DIRECTIONS CAREFULLY:**

1. Do not write during first fifteen minutes. This time is to be spent on reading the questions. After having read the questions, 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 each question or parts of questions, are given in a bracket [ ].
4. Read the direction to each question carefully and write all 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 × 25 = 25]**

(i) The function of fluorspar in the electrolytic reduction of alumina dissolved in fused cryolite ($Na_3AlF_6$) is, it

A lowers the rate of dissociation of ions.

B increases the temperature of the electrolytic cell.

C acts as a catalyst to speed up the reaction.

D lowers the temperature of electrolytic cell.

(ii) In the equation PV = nRT what value does the 'R' always have if the pressure is measured in atmosphere and the temperature measured in Kelvin?

A $2.66 \times 10^{23}$ C 26.6

B 0.0821 D 3.14

(iii) A gas sample contains 16.0 g of $CH_4$, 16.0 g of $O_2$, 16.0 g of $SO_2$ and 33.0 g of $CO_2$. What is the total number of moles of gas in the sample?

A 2.25 moles. C 2.75 moles.

B 2.50 moles. D 3.00 moles.

(iv) Which of the following orders is not in accordance with the property stated against it?

A $F_2 > Cl_2 > Br_2 > I_2$ : electronegativity.

B $F_2 > Cl_2 > Br_2 > I_2$ : oxidizing power.

C $F_2 > Cl_2 > Br_2 > I_2$ : electron affinity.

D $F_2 < Cl_2 < Br_2 < I_2$ : boiling point.

(v) The combined gas law is shown in the equation: $P_1 \times \frac{V_1}{T_1} = P_2 \times \frac{V_2}{T_2}$. What must remain constant for this to be true?

A Pressure. C Temperature.

B Number of Moles. D Volume.

(vi) A compound has an empirical formula of $C_2H_4O$. An independent analysis gave a value of 132 g for its molar mass. What is the molecular formula of the compound?

A $C_4H_4O_5$ C $C_2H_4$

B $C_{10}H_{12}$ D $C_6H_{12}O_3$

(vii) Which transition element has its melting point below 1000°C?

A Mercury. C Iron.

B Scandium. D Silver.

(viii) The enthalpy of formation of compounds K, L, M and N are –84.00, +24.00, +6.00 and –16.00 k.cal mole$^{-1}$ respectively. The correct order of increasing stability of the compound is

A K < L < M < N. C L < M < N < K.

B N < M < K < L. D M < N < L < K.

(ix) Which reduction process is employed, when metals are required to be obtained with high degree purity?

A Carbon reduction. C Aluminium reduction.

B Hydrogen reduction. D Electrolytic reduction.

(x) The equation given below represents

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

A exothermic reaction as it releases heat.
B endothermic reaction as it releases heat.
C exothermic reaction as it absorbs heat.
D endothermic reaction as it absorbs heat.

(xi) In which of reactions is the formation of the products favored by increase in pressure?

A $2O_3(g) \longrightarrow 3O_2(g)$

B $2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$

C $C(s) + O_2(g) \longrightarrow CO_2(g)$

D $CO_2(aq) + 2H_2O(l) \longrightarrow H_3O(aq) + HCO_3(aq)$

(xii) An unknown saturated hydrocarbon X has three carbon atoms in it. If one hydrogen atom from this saturated hydrocarbon is replaced by –OH group, product obtained is

A methanol. C propan-1-ol.

B ethanol. D butan-1-ol.

(xiii) Which are the materials mixed with the ore, before it is subjected for smelting during the extraction of iron?

A Coke and silica. C Silica and lime stone.

B Coke and limestone. D Coke, silica and limestone.

(xiv) During the production of ethanol by fermentation, the starchy substances are steamed at 140°C to150°C under pressure to liberate a solution containing starch. This stage is called

A dilution. C distillation.

B malting. D mashing.

(xv) When the atoms of halogen combine with the atoms of alkali metal, the nature of bond in a compound is

A covalent bond. C coordinate bond.

B electrovalent bond. D hydrogen bond.

(xvi) Which property is NOT common for silver and gold?

A Valence electronic configuration is $3d^{10}\ 4s^1$.

B Cannot displace hydrogen from acids.

C Exhibit variable valency of +1 and +2.

D Weakly electropositive.

(xvii) During the formation of a chemical bond the energy is

A evolved. C increased.

B stored. D absorbed.

(xviii) The relative atomic mass of Ne is 20. Its relative molecular mass is

A 10 C 30

B 20 D 40

(xix) Given the reaction at 50°C

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

The rate of this reaction can be increased by using 2.0 g of powdered zinc instead of a 2.0 g strip of zinc because the powdered zinc has

A lower kinetic energy.
C more surface area.
B lower concentration.
D more zinc atoms.

(xx) The heat of formation of $CO_2$ is – 394.0 kJ. What will be the heat of combustion of carbon?

A +94.1 k.cal
C -394.0 kJ
B 00.00 kJ
D 394.0 kJ

(xxi) Which event must always occur for a chemical reaction to take place?

A Formation of a precipitate.
B Formation of a gas.
C Effective collisions between reacting particles.
D Addition of a catalyst to the reaction system.

(xxii) Which is the correct order for the trend in boiling point of alcohol in its homologous series with respect to the number of carbon atom present in the alcohol?

A methanol < ethanol < butanol < propanol.
B methanol > ethanol > propanol > butanol.
C methanol > ethanol > butanol > propanol.
D methanol < ethanol < propanol < butanol.

(xxiii) An alloy which does not contain copper is

A invar.
C german silver.
B gun metal.
D aluminium bronze.

(xxiv) How many litres of steam will be formed from 2L of $H_2$ and 1 L of $O_2$, if all volumes are measured at the same temperature and pressure?

A 1L.
C 3L.
B 2L.
D 4L.

(xxv) Ethanol reacts with sodium to form two products. These are

A sodium ethanoate and hydrogen.
B sodium ethoxide and hydrogen.
C sodium ethanoate and oxygen.
D sodium ethoxide and oxygen .

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 |
|---|---|
| 1. Boyles law | a. Ammonia |
| 2. Haber process | b. Equilibrium |
| 3. Avogadro's law | c. Sulphur trioxide |
| 4. Aufbau's Principle | d. 6.023 x $10^{23}$ particles |
| 5. Le Chatelier's Principle | e. Filling of orbitals |
| | f. Pressure-volume relationship |
| | g. Volume-temperature relationship |

c. **Fill in the blanks by writing suitable word(s). [5]**

(i) An enzyme that converts glucose to ethanol during fermentation is ______.

(ii) A reaction in which the resulting product does not react to form original substance is called _____ reaction.

(iii) As it descend down the electrochemical series, the tendency of the cations to get discharged at cathode _____.

(iv) The catalyst used during hydrogenation of vegetable fats and oil is _____.

(v) In a reaction where $H_R > H_P$, the heat energy is _____.

d. **State whether the following statements are True or False. Rewrite the false statements correctly. [5]**

(i) The heat content of a reaction is lowered during endothermic reaction.

(ii) The oxidation state of chlorine is one and that of fluorine can be as high as seven.

(iii) When a system is at equilibrium, the rate of forward reaction is greater than the rate of reverse reaction.

(iv) There are more atoms in 12 grams of carbon than in 12 grams of copper.

(v) When the temperature of gas filled balloon is doubled, the volume of balloon is doubled.

e. **Answer the following questions: [5]**

(i) Define Charles law.

(ii) Write two applications of electrolysis.

(iii) Suggest a reagent for conversion of ethyl alcohol to acetaldehyde.

(iv) Why is the electron affinity of fluorine lesser than the chlorine?

(v) Why does heat of combustion of an element or the compound have a positive value?

**f. Write down one difference between the following points.** **[5]**

(i) Paramagnetic and ferromagnetic substance.

(ii) Endothermic and exothermic reaction.

(iii) Empirical and molecular formula.

(iv) Oxidation and reduction reaction.

(v) Monohydric and dihydric alcohol.

## Section B (50 Marks)

*Attempt any **five** Questions.*

### Question 2

a. Answer the following questions with respect to the electrolysis of $CuSO_4$ solution using carbon electrode.

(i) Write down the cathodic reaction. [½]

(ii) Write down the anodic reaction. [½]

(iii) Draw the well labeled diagram to show the electrolysis of $CuSO_4$.[2]

b. Write the IUPAC name for the following alcohols and indicate them as mono, di or trihydric alcohol.

(i) $CH_3 - CH_2 - CH_2 - OH$ [1]

(ii) [1]

$$\begin{array}{ccc} CH_2 & — & CH_2 \\ | & & | \\ OH & & OH \end{array}$$

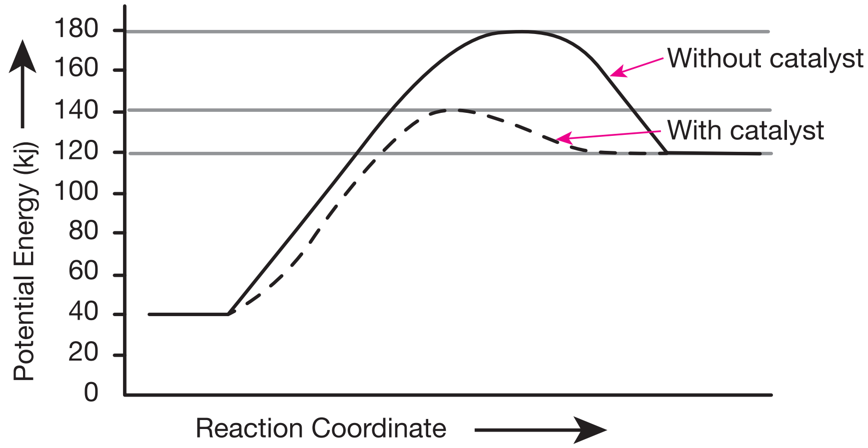
c. Which property of halogen makes chlorine a good bleaching agent? [1]

d. Write the electronic configuration of ( $_{29}Cu$) in s, p, d, f notation. [1]

e. A gas occupies a volume of 120.0 mL at a pressure of 0.75 atm and a temperature of 22°C. What will the volume be at a pressure of 1.25 atm and a temperature of 70°C? [3]

## Question 3

a. Given the potential energy diagram for a chemical reaction:




(i) What is the activation energy for the forward reaction without the catalyst? [1]

(ii) Why curves on the potential energy diagram with and without using the catalyst are different. [1]

(iii) What is a catalyst? [1]

b. Define the term coordination number. The coordination number of $[Fe(H_2O)_6]^{2+}$ ion is 6. What does it mean? [2]

c. Predict the effect of each of the following on the indicated equilibrium system in terms of which reaction (forward, reverse, or neither) will be favored.

$$H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + 184\,kJ$$

(i) Removal of HCl. [½]

(ii) Decreased temperature. [½]

(iii) Removal of $H_2$. [½]

(iv) Increased pressure. [½]

d. Calculate the percentage of phosphorus in the fertilizers, superphosphate $Ca\,(H_2PO_4)_2$. [3]

## Question 4

a. An organic compound X is widely used as a preservative in pickle and has a molecular formula, $C_2H_4O_2$. This compound reacts with ethanol to form a sweet smelling compound Y.

(i) Identify the compound X. [1]

(ii) Write the chemical equation for its reaction with ethanol to form compound Y. [1]

(iii) What is the name given to reaction obtained in (ii)? [1]

b. Why the d-block elements are called transition metals. Name two elements of 3d-seriess. [2]

c. What is the standard heat of formation? What is the significance of the positive and negative value of $\Delta H°f$ of a given compound regarding the stability of a substance? [2]

d. When vaporized, 64 g of $CH_3OH$ would occupy 44.8 L at STP. What is the vapour density of $CH_3OH$? [3]

## Questions 5

a. Answer the following questions with reference to Hall-Heroult process.

(i) What is the function of fluorspar fused with cryolite? [1]

(ii) Name the cathode and anode. [1]

(iii) Why is anode replaced from time to time during the course of process? [1]

b. Define the term effective nuclear charge. What is the effective nuclear charge experienced by the electrons of chlorine atom in its valance shell? [2]

c. A gas is confined in a cylinder with a moveable piston at one end. When the volume of a cylinder is 684 mL, the pressure of the gas is 1.32 atm. What will be the pressure if the volume of the cylinder is reduced to 513 mL? Assume that the temperature is constant. [3]

d. What is breathalyzer? How is the presence of alcohol in breath confirmed? [2]

## Question 6

a. What are amphoteric substances? Give two examples. [2]

b. Identify whether following reactions are endothermic reaction or exothermic reaction. Give reason(s). [2]

$$K(s) + H_2O(l) \longrightarrow KOH(aq) \qquad \Delta H = -48.0\,k.cal$$

$$N_2(g) + O_2(g) \longrightarrow 2NO(g) \qquad \Delta H = +21.65\,k.cal$$

c. What is denatured alcohol? [1]

d. Concentrated $HNO_3$ oxidizes P to $H_3PO_4$ (phosphoric acid) according to the following equation [3]

$$P + 5HNO_3 \longrightarrow H_3PO_4 + H_2O + 5NO_2$$

(i) How much $H_3PO_4$ is formed from 6.2 g of P?

(ii) What is the mass of $HNO_3$ consumed at the same time?

(iii) What would be the volume of the steam produced at the same time if measured at STP? (H = 1, N = 14, O = 16, P = 31)

e. Why is ethanol production considered as carbon neutral to the environment? Explain. [2]

## Question 7

a. Blast furnace is used for extraction of iron from its ore, heamatite. Write the equation(s) for each of the following reaction. [3]

(i) the formation of carbon monoxide in the furnace.

(ii) two reactions in which heamatite is reduced to iron.

(iii) the function of limestone in the furnace.

b. The empirical formula of a compound is $C_2H_5$. It has a vapour density of 29. Determine the relative molecular mass of the compound and its molecular formula. [3]

c. Define metal activity series. Why do gold and silver do not corrode easily? [2]

d. In the Haber process, which conditions favour a high yield of ammonia at equilibrium? [1]

e. State any two applications of energy changes that take place in chemical reaction. [1]

# Glossary

**Activation complex**: A temporary or intermediate product formed when atoms are rearranged in a chemical reaction.

**Activation energy, $E_a$**: The minimum amount of energy that the reacting molecule must possess for the collision to result in a chemical reaction.

**Aldehyde**: An organic compound containing -CHO group, formed by the oxidation of alcohols.

**Alloy**: It is a homogeneous mixture of metals or a metal and a non-metal, which cannot be separated by physical means.

**Aluminothermy**: The ignition of thermite (mixture of metal oxide and Al powder) with a magnesium wire embedded in a mixture of Mg powder and $BaO_2$ at a very high temperature.

**Amphoteric Substance**: Those substance (especially a metal oxide and hydroxide) that can react both as an acid as well as a base.

**Auto-reduction**: The reduction of some substance like $Cu_2O$ in absence of any reducing agents.

**Avogadro's law**: Law stating that at the same temperature and pressure, equal volumes of different gases contain an equal number of particles.

**Avogadro's number**: It is a dimensionless quantity, and has the same numerical value of the Avogadro constant i.e., $6.023 \times 10^{23}$, representing the number of atoms, molecules, or ions in one mole of a substance.

**Biocatalysts**: Substances formed in living cells that accelerate or slow down the chemical processes in the body.

**Biochemical reactions**: It is the transformation of one molecule to a different molecule inside a cell which is mediated by enzymes, which are biological catalysts.

**Biotechnology**: It is the use of biological processes, organisms, or systems to manufacture products intended to improve the quality of human life.

**Breathalyzer**: It is a device used for measuring the amount of alcohol in a breath.

**Carboxylic acid**: An organic acid containing a carboxyl group. The simplest examples are methanoic (or formic) acid and ethanoic (or acetic) acid.

**Chemical energetic**: The study which deals with the energy changes during chemical reactions.

**Chlorofluorocarbon (CFCs)**: It is an organic compound that contains carbon, chlorine, and fluorine, produced as a volatile derivative of methane and ethane.

**Collision theory**: It explains the rate of reactions on the basis of effective collision between molecules.

**Dalton's law**: A law stating that the pressure exerted by a mixture of gases in a fixed volume is equal to the sum of the pressures that would be exerted by each gas alone in the same volume.

**Dynamic equilibrium**: The chemical equilibrium between a forward reaction and the reverse reaction where the rate of the reactions is equal.

**Electrolysis**: The process of decomposition of a chemical compound in its molten state or in aqueous solution by the passage of electricity.

**Effective nuclear charge**: It is the net charge an electron experiences in an atom with multiple electrons.

**Empirical formula**: A formula for a compound which shows the simplest ratio present.

**Electron cloud**: It is a region of negative charge surrounding an atomic nucleus that is associated with an atomic orbital.

**Enthalpy (H)**: It is the total heat content of a system.

**Enzymes**: Biological catalysts that increase the rates of reactions important to an organism. They are produced by the living cells.

**Enzyme–substrate complex**: The intermediate product formed when a substrate molecule interacts with the active site of an enzyme. It undergoes a chemical reaction and is converted into a new product.

**Equilibrium reaction**: The chemical reactions in which a reaction automatically reverses and there is no net overall change.

**Esters:** An organic compound obtained by replacing the hydrogen of an acid by an alkyl or other organic group. They often have a characteristic pleasant, fruity odor.

**Esterification**: The general name for a chemical reaction in which two reactants (typically an alcohol and an acid) form an ester as the reaction product.

**Flux**: A substance added to get rid of the gangue.

**Fuel gases/Blast furnace gas**: The mixture of waste gases which comes out of the blast furnace and mainly consists of about 25% of CO, 10% of $CO_2$ and 6% of $N_2$.

**Functional group**: It is an atom or a group of atoms in a molecule that gives the molecules its characteristic properties.

**Gangue**: The impurities like sand, rocky materials, earthy particles etc.associated with the ores.

**Gas constant**: The gas constant, also known as the universal molar gas constant, is a physical constant that appears in an equation defining the behavior of a gas under theoretically ideal conditions. The gas constant is, by convention, symbolized with R.

**Gas laws**: The physical laws that describe the properties of gases, including Boyle's and Charles's laws.

**Enthalpy of combustion**: The amount of heat released when 1 mole of a substance undergoes complete combustion at a given temperature.

**Enthalpy of formation**: The amount of heat absorbed or evolved when 1 mole of substance is formed from its element at STP.

**Heat of reaction**: The amount of heat released or absorbed in a chemical reaction when the numbers of moles of reactants have completely reacted to form the products.

**Heat of neutralization**: The amount of heat evolved when 1 mole of an acid is neutralized by 1 mole of a base in dilute solution.

**Heat of solution**: The amount of heat absorbed or evolved when 1 mole of solute dissolves in so much amount of solvent that further addition of solvent to the solution produces no further change in heat content.

**Homologous series**: It is a series of organic compounds with same functional group in which the successive members differ by one $-CH_2$ unit.

**Ideal gas equation**: It is the equation of state of a hypothetical ideal gas.

**Ideal gas law**: A physical law describing the relationship of the measurable properties of an ideal gas. It is derived from a combination of the gas laws of Boyle, Charles and Avogadro.

**Internal energy (E)**: The energy associated with the random and disordered motion of molecules.

**Ions**: The charged particles which are formed either by the loss or gain of electrons.

**Isotopes**: Those atoms with the same number of protons, but differing numbers of neutrons.

**Kelvin scale**: It is a temperature scale designed so that zero degree K is defined as absolute zero and the size of one unit is same as the size of one degree Celsius.

**Leaching**: It is a widely used metal extraction technique which converts metals into soluble salts in aqueous media from the desired ores.

**Levigation**: The washing of powdered ore with an upward stream of running water in a hydraulic classifier whereby, the lighter gangue particles are washed away while heavier ore particles settle down and are removed from the base.

**Le Chatelier's Principle**: When a system in equilibrium is subjected to a change in temperature, pressure or concentration of a component, the equilibrium shifts in the direction of the reaction opposing the effect of change and a new equilibrium condition is established.

**Malt**: Barley or other grain that has been steeped, germinated, and dried, used for brewing or distilling and vinegar-making.

**Metallurgy**: A science of metal and the various processes involved in the extraction of metals from their respective ores.

**Metal activity series**: The arrangement of metal in descending order of their reactivity.

**Minerals**: It is a naturally occurring chemical substances found in the earth's crust.

**Molar volume**: The volume occupied by a mole of a gas at STP. At STP 1 mole of gas occupies 22.4 liters.

**Molasses**: It is a thick, dark brown juice obtained from raw sugar during the refining process.

**Mole**: The amount of a chemical substance that contains as many elementary entities, e.g., atoms, molecules, ions, electrons, or photons, as there are atoms in 12 grams of carbon-12, the isotope of carbon with relative atomic mass 12.

**Nomenclature**: It is a set of rules to generate systematic names for chemical compounds.

**Nuclear charge**: The total charge of all the protons in the nucleus, which has same value as the atomic number.

**Ores**: The minerals from which desired metals can be extracted.

**Orbit**: It is a fixed path of an electron around the nucleus of an atom. Bohr's model of the hydrogen atom is based on the concept of orbit.

**Orbital**: It is a region of space around the nucleus where the probability of finding an electron is maximum.

**Partial pressure**: The pressure that one component of a mixture of gases would exert if it were alone in a container.

**Pulverization**: The process in which the smaller pieces of ores are reduced to fine powder with the help of stamp mill or ball mill.

**Real gases**: A gas that does not behave as an ideal gas due to interactions between gas molecules. It is also known as non-ideal gas.

**Redox reaction**: It is a chemical reaction in which simultaneous oxidation and reduction take place.

**Reaction rate**: The change in the amount of reactants or products over time.

**Slag**: The product obtained by the combination of gangue with the flux.

**Smelting**: It is a process of heating metal oxide in the presence of carbon.

**Substrate**: The material or substance on which an enzyme acts.

**Thermochemical equation**: It is a chemical equation that includes the quantity of heat released or absorbed during the chemical reaction.

**Thermo chemistry**: The branch of chemistry that deals with the quantities of heat evolved or absorbed during chemical reactions.

**Total pressures**: The sum of the partial pressure of each individual gas.

**Threshold energy**: The minimum amount of energy required by a molecule for the reaction to takes place.

**van der Waals force**: A weak force of attraction between electrically neutral molecules that collide with or pass very close to each other. Or, van der Waals forces are the intermolecular forces that cause molecules to cohere in liquid and solid states of matter, and are responsible for surface tension and capillary action.

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