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CHEMISTRY STUDENT TEXTBOOK GRADE 10

CHEMISTRY

STUDENT TEXTBOOK

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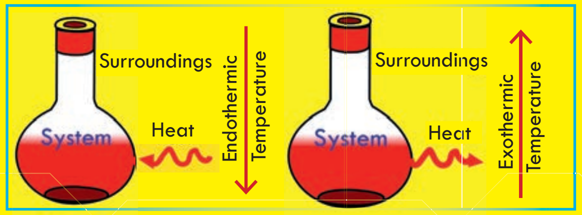

FEDERAL DEMOCRATIC REPUBLIC OF ETHIOPIA
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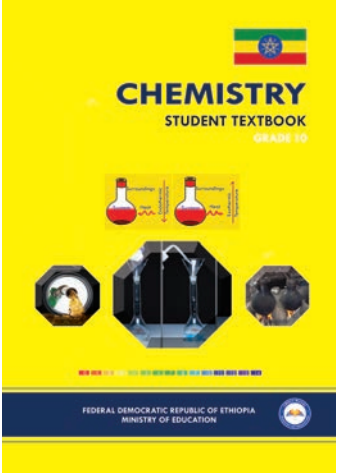

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CHEMISTRY

STUDENT TEXTBOOK
GRADE 10

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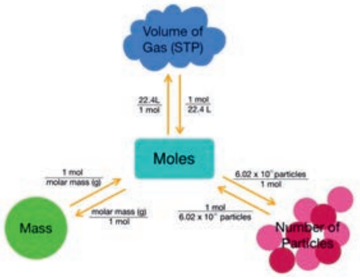
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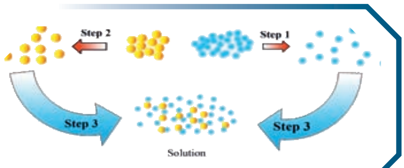
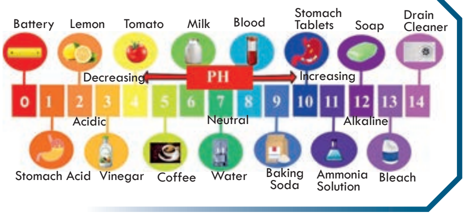
22.4L / 1 mol
1 mol / 22.4 L
Moles
1 mol / molar mass (g)
molar mass (g) / 1 mol
6.02 x 10 particles / 1 mol
1 mol / 6.02 x 10 particles
Mass
Number of Particles


Step 2
Step 1
Step 3
Step 3
Solution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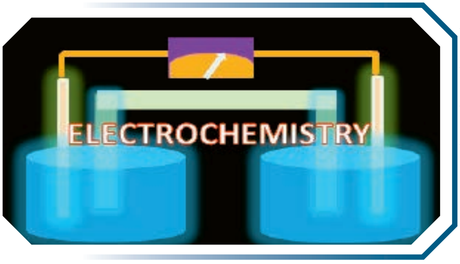
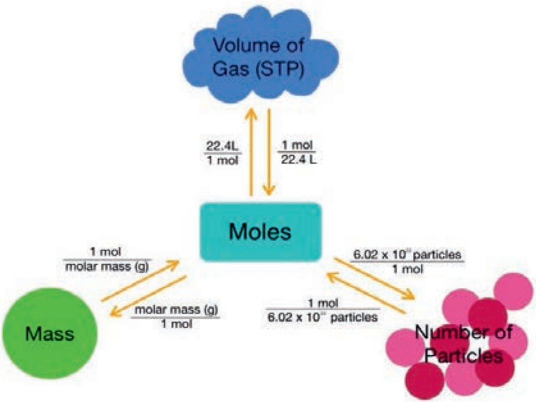

ELECTROCHEMISTRY


Lead
Gold
Diamond
Carbon
Silver
Copper
Sulfur
Phosphorus

UNIT 1



CHEMICAL REACTIONS AND STOICHIOMETRY

Unit Outcomes

At the end of this unit, you will be able to

- define the basics of chemical reaction;
- describe the four major types of reactions;
- develop skills in writing and balancing chemical equations;
- define oxidation - reduction reactions
- analyse redox reactions by specifying the oxidizing agent, the reducing agent, the substance reduced or oxidized;
- explain molecular mass, formula mass, molar mass, empirical and molecular formulas;
- determine molecular mass, formula mass, molar mass, empirical and molecular formulas
- describe the mole concept
- solve problems related to moles of substances
- develop skills in solving problems based on chemical equations (mole-mole, mass - mass, volume - volume and mass - volume problems);
- develop skills in determining the limiting reactant, theoretical yield, actual yield and percentage yield;
- demonstrate scientific inquiry skills: observing, inferring, predicting, classifying comparing and contrasting, communicating, measuring, asking questions designing experiments, interpreting data, drawing conclusions, applying concepts, relating cause and effect and problem - solving.

Start-up Activity

Discuss the following questions in groups of three or four students, then share your findings with the class.

1. Changes in our daily lives include fruit ripening, which involves the transformation of raw fruit (containing starting substance) into ripe fruit (containing new substance), and the fermentation of *injera* (new material) from *teff* flour (starting material).
    a. What are the differences between the raw and ripened fruit?
    b. What are the differences between the fermented injera and the teff flour?
    c. Are these changes reversible? Make a list of at least three more changes that occur in your day-to-day activities, and note the differences between the starting materials and new materials in each.
2. Are the following changes reversible?
    a. Dough is changed to bread through baking.
    b. Firewood changed to ash through burning.
3. Explain whether the following refer to the quality or quantity of substances.
    a. A farmer determines the amount of fertilizer for his farm.
    b. A nurse calculates the dose of a medicine for a patient.
    c. A chef mixes the right amounts of ingredients for cooking food.

1.1 Introduction

At the end of this section, students will be able to

☞ define chemical reaction;

☞ give some examples of chemical reactions.

☞ explain physical and chemical changes using examples

Change is the law of nature. We observe various types of changes around us. The growth of seed in to a plant, the burning of wood in a fire place, the rusting of iron articles in moist air, the rotting of food and the evaporation of liquids are some of the changes that we see taking place around us. Scientists categorize changes as physical and chemical changes. A physical change is a change that does not involve the formation of a new substance with a new chemical composition. Examples of physical changes are evaporation of liquids, powdering of sugar and melting of ice. The change which results in the formation of one or more new substances with new chemical composition is known as a chemical change. Examples of chemical changes are turning of milk into curd, photosynthesis by green plants and rotting of egg by bacteria. Chemical changes occur due to chemical reactions between substances. In general, a change is called a chemical reaction if it exhibits all or part of the following characteristics.

i. Formation of new substances
ii. Production of heat or light or both
iii. Change in color
iv. Change in temperature and etc.

Chemical reaction is the process in which reacting substances, called reactants, are converted into new substances, called products. The characteristics of the products are completely different from those of the reactants.

For example, if you burn sulphur with oxygen, the sulphur and oxygen are completely converted to sulphur dioxide. Sulphur dioxide is a colourless gas with a pungent odour. These characteristics of sulphur dioxide are completely different from the characteristics of the original substances, sulphur and oxygen. A chemical reaction is represented by a short hand notation called chemical equation as:

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

A chemical equation uses chemical symbols to show what happens during a chemical reaction. A balanced chemical equation can be used to describe the relationships between the amounts of reactants and products. The quantitative study of reactants and products in a chemical reaction is called reaction's stoichiometry. In this unit you will study, the types of chemical reactions, chemical equations, redox reactions, chemical formulas, masses of compounds, mole concepts, and reaction's stoichiometry.

Exercise 1.1

a. What is a chemical reaction?
b. Explain the meanings of chemical and physical changes using examples.
c. Give two examples of evidences for chemical reactions not mentioned in the text.

1.2 Chemical Equations

At the end of this section, students will be able to

☞ explain the conventions used to write chemical equation;
☞ write the chemical equations of varieties of chemical reactions;

☞ balance chemical equations using;
  - The Inspection method,
  - The Least Common Multiple (LCM) method,
  - The Algebraic method.

Activity 1.1

Form a group and discuss each of the following and present your conclusion to the rest of the class.

1. How can you show chemical symbols, physical states and quantities of substances involved in a chemical reaction?
2. What is the difference between a chemical equation and a chemical reaction?
3. Which law is applied when a chemical equation is balanced?

A chemical equation is a shorthand representation of an actual chemical reaction in terms of chemical symbols and formulas. In a chemical equation the starting substances are called reactants; and the new substances produced are known as products.

In a chemical equation, reactants are written on the left side and products on the right side of the equation. An arrow (→) is placed between the two sides to indicate transformation of reactants into products.

$$\text{Reactants} \rightarrow \text{products}$$

1.2.1 Writing Chemical Equation

In writing chemical equation, instead of using words, chemical symbols and formulas are used to represent the reaction.

Steps to write a chemical equation

1. Write a word equation: A word equation is stated in words. For example, the word equation for the reaction between hydrogen and nitrogen to produce ammonia is written as:

$$\text{Hydrogen} + \text{Nitrogen} \rightarrow \text{Ammonia (Word equation)}$$

Note that we read the '+' sign as 'reacts with' and the arrow can be read as 'to produce', 'to form', 'to give' or 'to yield'.

2. Write the symbols and formulas for the reactants and products in the word equation.

$$H_2 + N_2 \rightarrow NH_3 \text{ (Unbalanced equation)}$$

3. Balance the equation

$3H_2 + N_2 \rightarrow 2NH_3$ (Balanced equation)

Chemically reactive substances might be solids, liquids, gases, or they can dissolve in a solvent. Particularly ionic compounds commonly react in aqueous solution, or dissolve in water. Sometimes this information is added to an equation by placing the appropriate symbols after the formulas:

(s) - Solid (l) - Liquid (g) - Gas (aq) - Aqueous solution

Example 1.1: Write the chemical equation for the reaction that occurs between calcium carbonate and sulfuric acid.

1. Word equation:

Calcium carbonate + Sulphuric acid → Calcium sulphate + Water + Carbon dioxide

2. Chemical equation: $CaCO_3(s) + H_2SO_4(l) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$
3. The equation is balanced

Example 1.2: Write the chemical equation for the reaction of sodium chloride and silver nitrate.

1. Word equation:

Silver nitrate + Sodium chloride → Silver chloride + Sodium nitrate

2. Chemical equation: $AgNO_3(aq) + NaCl\ (aq) \rightarrow AgCl(s) + NaNO_3\ (aq)$
3. The equation is balanced.

Generally, any chemical equation must fulfill the following conditions:

A. An equation must represent a true and possible chemical reaction.
B. The symbols and formulas must be written correctly. The elements – hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine and iodine exist as diatomic molecules. For example, these elements should be written as molecules in the equation. The equation must be balanced.

A chemical equation has both qualitative and quantitative meanings. Qualitatively, a chemical equation indicates the types of the reactants and products in the reaction.

Quantitatively, a chemical equation expresses the relative number (amount) of moles, molecules or masses of the reactants and products. For example, the balanced chemical equation : $2H_2 + O_2 \rightarrow 2H_2O$ has both qualitative and quantitative meaning. Qualitatively, it shows that the reaction between hydrogen and oxygen forms water. Quantitatively, it indicates that two units of hydrogen combine with one unit of oxygen to form two units of water. The quantities may be expressed in terms of grams, number of moles or number of molecules (see for more details in subsection 1.6.1).

1.2.2 Balancing Chemical Equations

At the end of this section, students will be able to

☞ balance chemical equations using:
- The Inspection method
- The Least Common Multiple (LCM) method
- The Algebraic method

According to the Law of Conservation of Mass, atoms are neither created nor destroyed during a chemical reaction (refer to grade 9 chemistry). As a result, the number of atoms of each element should remain the same before and after the reaction. To balance a chemical equation means to equalize the number of atoms on both sides of the equation by putting appropriate coefficients in front of the formulas.

A. Balancing Chemical Equations by Inspection Method

The most common way to balance chemical equations is to use inspection method. It is exactly what the name states. You balance the equation by inspecting it. It is also known as a trial and error method or a hit and trial method.

Follow four easy steps, given below, to balance a chemical equation by inspection method:

1. Write the word equation.
2. Write the correct symbols or formulas for the reactants and products (unbalanced equation).
3. Start with the most complicated compound count and tabulate the number of each type of atoms on the two sides of the unbalanced equation.
4. Make the number of each type of atoms on the left side equal to the number of corresponding atoms on the right side of the equation.

Example 1.3: Balance the chemical reaction that takes place between iron and water to form iron (IV) oxide and hydrogen gas by the inspection method.

Step 1: Iron + Water → Iron (IV) oxide + Hydrogen (word equation)

Step 2: $Fe + H_2O \rightarrow Fe_3O_4 + H_2$ (Unbalanced equation)

Let us count and tabulate the number of various types of atoms on the reactant and product sides of the expression

$$Fe + H_2O \rightarrow Fe_3O_4 + H_2$$

Atom	Reactant side	Product side	Equation
Number of Fe atoms	1	3	
Number of O atoms	1	4	
Number of H atoms	2	2	
Total no. of atoms	4	9	Unbalanced

Balancing Fe atoms: There is one Fe atom on the left side while there are three Fe atoms on the right side. Therefore, a suitable coefficient of Fe on the left side is 3. Thus,

$$\underline{3}\mathrm{Fe} + \mathrm{H_2O} \rightarrow \mathrm{Fe_3O_4} + \mathrm{H_2} \qquad (2)$$

Balancing O atoms. There is one O atom (in H_2O) on the left side and there are four O atoms on the right side. Therefore, a proper coefficient of H_2O is 4 as H_2O.

$$\underline{3}\mathrm{Fe} + \underline{4}\mathrm{H_2O} \rightarrow \mathrm{Fe_3O_4} + \mathrm{H_2} \qquad (3)$$

Balancing H atoms. There are eight H atoms (in $4H_2O$) on the left side, but only two H atoms (in H_2) on the right side. Therefore, an appropriate coefficient of H is 4 as $4H_2$. Thus,

$$\underline{3}\mathrm{Fe} + \underline{4}\mathrm{H_2O} \rightarrow \mathrm{Fe_3O_4} + \underline{4}\mathrm{H_2} \qquad (4)$$

Equation 4 is a balanced chemical equation. Note that number of atoms of each element is conserved (no. of atoms reactant side = no. of atoms product side). How?

These are shown below:

Atom	Reactant side	Product side	Equation
Number of Fe atoms	3	3	
Number of O atoms	4	4	
Number of H atoms	8	8	
Total number of atoms	15	15	Balanced

NB: Inspection method works best for simple equations.

B. Balancing chemical equations by the LCM method

By taking the LCM of the total valency of the reactants and products and dividing it by the total valency of the reactants and products, the coefficients for the balanced chemical equation are derived. This is known as The LCM method. The following examples show all of the steps required to balance a chemical equation using the LCM Method.

Examples:

i. When aluminum reacts with oxygen, aluminum oxide is formed. Write the balanced chemical equation for the reaction

Step 1: Represent the reaction by a word equation:

Aluminum + Oxygen → Aluminum Oxide

Step 2: Write the correct symbols or formulas for the reactants and products

$$\mathrm{Al} + \mathrm{O_2} \rightarrow \mathrm{Al_2O_3}$$

Step 3: Place the total valency of each atom above it.

$$\overset{3}{Al} + \overset{4}{O_2} \rightarrow \overset{6\ \ 6}{Al_2O_3}$$

From the equation we see that the valency of aluminum is 3. The total valency of oxygen is $2 \times 2 = 4$. The total valency of aluminum in Al_2O_3 is $3 \times 2 = 6$. The total valency of oxygen in Al_2O_3 is $2 \times 3 = 6$, LCM is 12.

Step 4: Find the LCM of each total valency and place it above the arrow, here LCM is 12.

$$\overset{3}{Al} + \overset{4}{O_2} \xrightarrow[LCM]{12} \overset{6\ \ 6}{Al_2O_3}$$

Step 5: Divide the LCM by each total valency number to obtain the coefficients for each of the reactants and products. Place the obtained coefficients in front of the respective chemical formulas

$$4Al + 3O_2 \rightarrow 2Al_2O_3 \text{ (Balanced)}$$

Checking: There are 4 aluminum and 6 oxygen atoms on both sides of the equation. Hence, the chemical equation is correctly balanced

ii. The reaction of sodium sulphate with aluminum nitrate would form aluminum sulphate and sodium nitrate. Write the balanced chemical equation.

Step 1: Sodium sulphate + Aluminum nitrate → Aluminum sulphate + Sodium nitrate (word equation)

Step 2: $Na_2SO_4 + Al(NO_3)_3 \rightarrow Al_2(SO_4)_3 + NaNO_3$

Step 3: $\overset{2\ \ 2}{Na_2SO_4} + \overset{3\ \ 3}{Al(NO_3)_3} \rightarrow \overset{6\ \ 6}{Al_2(SO_4)_3} + \overset{1\ \ 1}{NaNO_3}$

Step 4: $\overset{2\ \ 2}{Na_2SO_4} + \overset{3\ \ 3}{Al(NO_3)_3} \xrightarrow[LCM]{6} \overset{6\ \ 6}{Al_2(SO_4)_3} + \overset{1\ \ 1}{NaNO_3}$

Step 5: $3Na_2SO_4 + 2Al(NO_3)_3 \rightarrow Al_2(SO_4)_3 + 6NaNO_3$ (Balanced)

There are 6 nitrogen, 6 sodium, 3 sulphur, 30 oxygen and 2 aluminum atoms on both sides of the equation. Thus, the equation is correctly balanced.

iii. The reaction of magnesium nitrate with ammonium phosphate would form ammonium nitrate and magnesium phosphate. Write the balanced equation.

Step 1: Magnesium nitrate + Ammonium phosphate → Magnesium phosphate + Ammonium nitrate

Step 2: $Mg(NO_3)_2 + (NH_4)_3PO_4 \rightarrow NH_4NO_3 + Mg_3(PO_4)_2$

Step 3: $\overset{2\ \ 2}{Mg(NO_3)_2} + \overset{3\ \ 3}{(NH_4)_3PO_4} \rightarrow \overset{1\ \ 1}{NH_4NO_3} + \overset{3\ \ 3}{Mg_3(PO_4)_2}$

Step 4: $\overset{2\ \ 2}{Mg(NO_3)_2} + \overset{3\ \ 3}{(NH_4)_3PO_4} \xrightarrow[LCM]{6} \overset{1\ \ 1}{NH_4NO_3} + \overset{3\ \ 3}{Mg_3(PO_4)_2}$

Step 5: $3\ Mg(NO_3)_2 + 2\ (NH_4)_3PO_4 \rightarrow 6NH_4NO_3 + Mg_3(PO_4)_2$ (Balanced)

There are 12 nitrogen, 24 hydrogen, 3 magnesium, 2 phosphorous and 26 oxygen atoms on both sides of the equation. Thus, the equation is correctly balanced.

B. Balancing Chemical Equations Using Algebraic Method

This method of balancing chemical equations involves assigning algebraic variables as stoichiometric coefficients to each species in the unbalanced chemical equation. These variables are used in mathematical equations and are solved to obtain the values of each stoichiometric coefficient. Consider the formation of ammonia from hydrogen and nitrogen as an example.

Step 1: Write the unbalanced equation with the correct symbols of the reactants & products:

$$N_2 + H_2 \rightarrow NH_3 \text{ (unbalanced equation)}$$

Step 2: Assign algebraic variables to each species as coefficients (a, b, c) in the unbalanced equation:

$$aN_2 + bH_2 \rightarrow cNH_3$$

Then, set equations for each element so that it is equal in both the right and left hand side.

N: $2a = c$ (subscript of N is 2 in the left side and 1 in the right side)
H: $2b = 3c$ (subscript of H is 2 in the left side and 3 in the right side)

Step 3: Choose the smallest variable and assign arbitrary number in order to determine the remaining variables.

In the above case, a is the smallest coefficient. Assuming $a = 1$, the values of b and c can be obtained as follows. $c = 2 \times 1 = 2$, $2b = 3c = 3 \times 2 = 6$; $b = 6/2 = 3$
Since a, b, and c have no common multiples, they can be substituted into the equation as follows.

$$N_2 + 3H_2 \rightarrow 2NH_3 \text{ (balanced equation)}$$

Example 1.4: Balance the reaction: $Al + O_2 \rightarrow Al_2O_3$ using the algebraic method. Assign variables for each element

$$aAl + bO_2 \rightarrow cAl_2O_3$$

The equation for Aluminum: $a = 2c$ (subscript of Al is 2 in the left side and 2 in the right side). The equation for oxygen: $2b = 3c$ (subscript of O is 2 in the left side and 3 in the right side)

Assuming a = 1, we get:

$$c = a/2;\ c = 1/2$$

$$2b = 3 \times (1/2) = 3/2\ ;\ b = 3/4$$

Since fractional values of b and c are obtained, the lowest common multiple denominator between the variables a, b, and c must be found and multiplied with each variable.

Since the lowest common multiple denominator is 4, each of the variables must be multiplied by 4.

Therefore, a = 4 x 1 = 4; b = (3/4) x 4 = 3; c = (1/2) x 4 = 2

Substituting the values of a, b, and c in the unbalanced equation, the following balanced chemical equation is obtained.

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

Example 1.5: Balance the equation: $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$

Step 1: $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$

Step 2: $aPCl_5 + bH_2O \rightarrow cH_3PO_4 + dHCl$

P: a = c (subscript of P is 1 in the left side and 1 in the right side)

Cl : 5a = d (subscript of Cl is 5 in the left side and 1 in the right side)

H: 2b = 3c + d (subscript of H is 2 in the left side and 3 in H_3PO_4 in 1 in HCl)

O : b = 4c (subscript of O is 1 in the left side and 4 in the right side)

Assume a =1 , therefore c=1, 5a = d ⇒ d =5, b =4.

Substituting the values of a, b, c and d in the unbalanced equation:

$$PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl \text{ (balanced)}$$

Exercise 1.2

1. Balance the following chemical equations, using the inspection method:
   a. $Na + H_2O \rightarrow NaOH + H_2$
   b. $KClO_3 \rightarrow KCl + O_2$
   c. $H_2O_2 \rightarrow H_2O + O_2$
   d. $Al + H_3PO_4 \rightarrow AlPO_4 + H_2$
   e. $HNO_3 + H_2S \rightarrow NO + S + H_2O$
2. Write the balanced chemical equation to represent the following reactions.
   a. Sodium bromide reacts with chlorine to form sodium chloride and bromine

b. Hydrochloric acid reacts with sodium carbonate to form sodium chloride, water and carbondioxide
c. Potassium chlorate when heated produces potassium chloride and oxygen
d. Calcium carbonate reacts with hydrochloric acid to form water, carbon dioxide and calcium chloride
e. Silver oxide decomposes to silver and oxygen gas.

3. Balance the following equations using the LCM method.
   a. $Al + HCl \rightarrow AlCl_3 + H_2$
   b. $Fe_2(SO_4)_3 + KOH \rightarrow K_2SO_4 + Fe(OH)_3$
   c. $CaCl_2 + Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + NaCl$
   d. $FeCl_3 + NH_4OH \rightarrow Fe(OH)_3 + NH_4Cl$
4. Balance the following equations using the algebraic method.
   a. $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$
   b. $Mg + H_2O \rightarrow Mg(OH)_2 + H_2$
   c. $Zn(NO_3)_2 \rightarrow ZnO + NO_2 + O_2$
   d. $H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$
   e. $NH_3 + O_2 \rightarrow NO + H_2O$
   f. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
   g. $FeCl_3 + MgO \rightarrow Fe_2O_3 + MgCl_2$
   h. $BaCl_2 + K_3PO_4 \rightarrow Ba_3(PO_4)_2 + 6KCl$
   i. $P_4O_{10} + H_2O \rightarrow H_3PO_4$

1.3 Types of Chemical Reactions

At the end of this section, students will be able to

- list the four types of chemical reactions;
- define combination reaction and give examples;
- conduct experiments on combination reactions in groups;
- define decomposition reaction and give examples;
- conduct some experiments on decomposition reactions in groups;
- define single - displacement reaction and give examples;
- conduct some experiments on simple displacement reactions in groups;
- define double - displacement reaction and give examples;
- conduct some experiments on double - displacement reactions in group.

Activity 1.2

In your group, discuss about the following processes and share your conclusion with the rest of the class.

1. What are the final products of digestive protein and starch? Is protein and starch digestion a building up or breaking down process ?
2. What is alcohol fermentation and its process? What are the two important chemical compounds produced during the fermentation of “Teji” or “Tella”? What is the main ingredient (starting material) for the production of “Teji” and “Tella”?
3. What are the two major chemical reactants and the two chemical products of photosynthesis? Is photosynthesis a breaking down or a building up process? Compare this process with cellular respiration.
4. What are the combustion products of kerosene? Is the combustion of kerosene a building up or a breaking down process?

Different elements and compounds react in different ways to produce different kinds of new substances. Many chemical reactions can be classified as one of the four basic types of reactions. By analyzing the reactants and products of a particular reaction, you can classify them into one of these categories. Some reactions may be categorized into more than one category. A complete understanding of these reaction types will help predict the products of unknown reactions. The four basic types of chemical reactions are combinations, decompositions, single displacement, and double displacement reactions, which are described in the next section.
Such reactions are generally classified as shown in *Figure 1.1*.

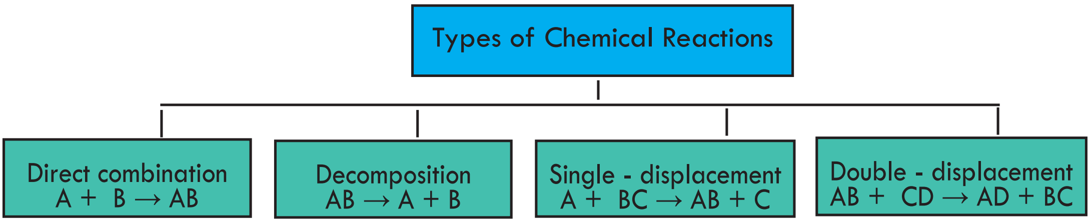


Figure 1.1 Classification of chemical reactions.

A. Direct Combination Reactions

Combination reactions are those reactions in which two types of pure substances react directly and form a single substance. In a combination reaction, two elements, two compounds, or an element and a compound react to form a single compound. In general, a combination reaction may be represented by the equation.

$$A + B \longrightarrow AB$$

where the reactants A and B are elements or compounds, the product AB is a compound. Such type of reaction is also known as synthesis or composition reaction.

Examples

$2Na + Cl_2 \longrightarrow 2NaCl$
Element + Element Compound
$CaO + CO_2 \longrightarrow CaCO_3$
Compound + Compound Compound

Note that in the above reactions, there is a single product.

Experiment 1.1

Investigation of Combination Reaction

Objective: To investigate the reaction between sulphur and iron.
Apparatus: Test tube, stand, burner, watch glass
Chemicals: Sulphur powder, iron filings

Procedure:

1. Mix about 3 g of iron filings and 2 g of powdered sulphur in a watch glass.
2. Transfer the mixture in a glass test tube.
3. Mount the test tube in a sloping position on a stand as shown in ***Figure 1.2***.
4. Heat the test tube until the mixture in the glass glows red hot.
5. Remove the test tube from the flame and observe the result.

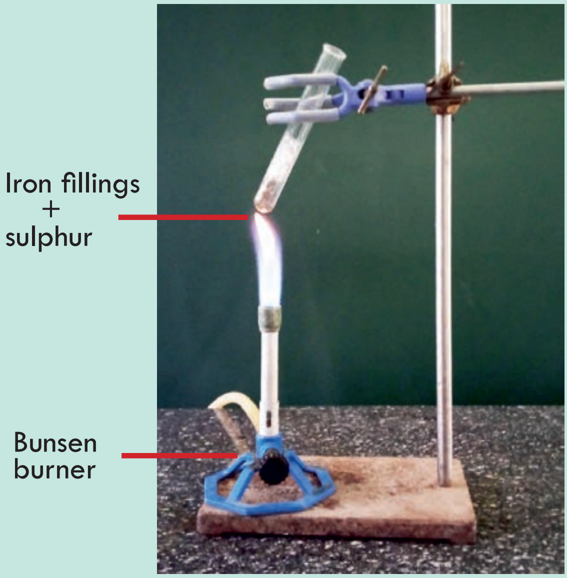


Figure 1.2 The reaction between iron and sulphur.

Questions

1. What were the colors of iron filings and sulphur before the reaction?
2. What was the color of the resulting compound after the reaction?
3. Write a balanced chemical equation for the reaction.

B. Decomposition Reactions

A decomposition reaction (or analysis) is a reaction that involves the breaking down of a single compound into two or more elements or simpler compounds. A decomposition reaction can be carried out using heat (thermal decomposition), light (photo decomposition), electricity (electrical decomposition) or a catalyst. But most decomposition reactions are carried out when heat is supplied and this heat energy is indicated by a 'delta' (Δ) symbol above the arrow. The general form of equation for a decomposition reaction is:

$$AB \rightarrow A + B$$

where the reactant AB must be a compound and the products A and B could be elements or compounds.

The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. For example, mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas, similarly a colorless potassium perchlorate decomposes to potassium chloride and oxygen gas.

$$2HgO(s) \xrightarrow{\Delta} 2Hg(l) + O_2(g)$$

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

Experiment 1.2

Investigation of Decomposition Reaction

Objective: To investigate the decomposition of copper (II) carbonate($CuCO_3$).
Apparatus: Test tube, stand, burner, cork, delivery tube.
Chemicals: $CuCO_3$ and lime water.

Procedure:

Put 1g of $CuCO_3$ powder in a glass test-tube. Mount the test tube in a sloping position on a stand as shown in ***Figure 1.3***. Fit a cork and a delivery tube to the test tube. Put another test tube containing lime water at the end of the delivery tube. Heat the $CuCO_3$ with a burner.

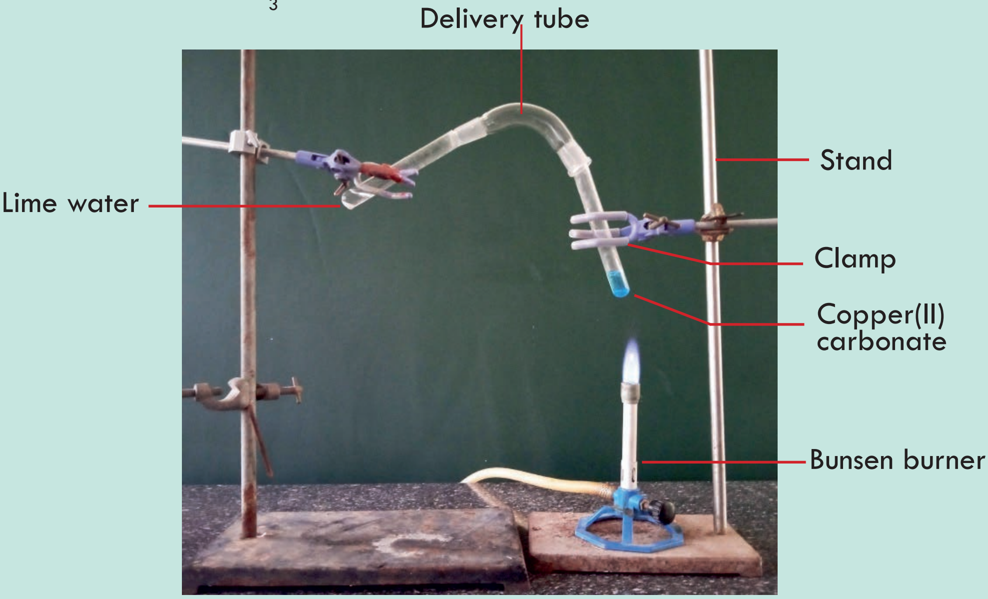


Figure 1.3 Decomposition of copper (II) carbonate.

Questions

1. What was the color of copper (II) carbonate before heating?
2. What was the color during heating and after cooling?
3. What change did you observe in the lime water?
4. Write a balanced chemical equation for the reaction.

C. Single Displacement Reactions

A reaction in which one element displaces another element from its compound is known as single displacement or replacement reaction. Such a reaction is represented by the following two general forms.

$$A + BC \rightarrow B + AC$$

If A is a metal, it will displace B to form AC, provided A is a more active metal than B.

$$A + BC \rightarrow BA + C$$

If A is a non-metal, it will displace C to form BA, provided A is a more active non-metal than C. In general, a more reactive element displaces a less reactive element from a compound.

Examples of single-displacement reactions

1. **Active metals displace hydrogen from acids**

Reactive metals such as potassium, calcium, sodium, and zinc displace hydrogen gas from dilute acids. For example, zinc is an active metal, and it displaces hydrogen from hydrochloric acid; but copper metal cannot do so.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$
$$Cu + HCl \rightarrow \text{No reaction}$$

2. **Reactive metals,** such as potassium, calcium, and sodium react vigorously with water to displace hydrogen:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$
$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

3. **A more active metal displaces a less active metal**

Zinc displaces copper from copper (II) sulphate solutions

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

Iron displaces copper from copper (II) sulphate solutions

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

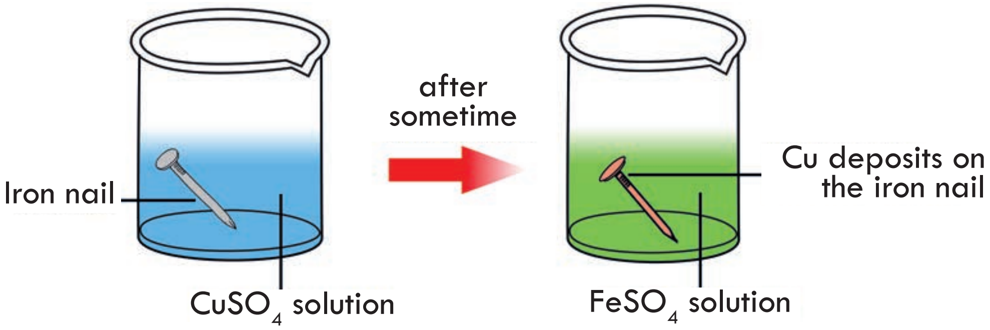


Figure 1.4 Replacement of copper by iron metal from copper sulphate solution.

Experiment 1.3

Investigation of Single Displacement Reaction

Objective: To investigate the displacement reaction between iron and copper (II) Sulphate ($CuSO_4$).

Apparatus: Iron rod and beaker

Chemicals: $CuSO_4$.

Procedure:

1. Clean a piece of iron rod or iron knife with emery paper to remove any rust.
2. Take saturated $CuSO_4$ solution in a beaker.
3. Dip the iron rod into the $CuSO_4$ solution as shown in ***Figure 1.5*** and wait for a few minutes. What did you observe on the iron rod?
4. Allow the reactants to stand for one day and observe any change on the iron rod.

Questions

1. What did you observe on the iron rod after one day?
2. Write a balanced chemical equation for the reaction.
3. Write the conclusion for the experiment.

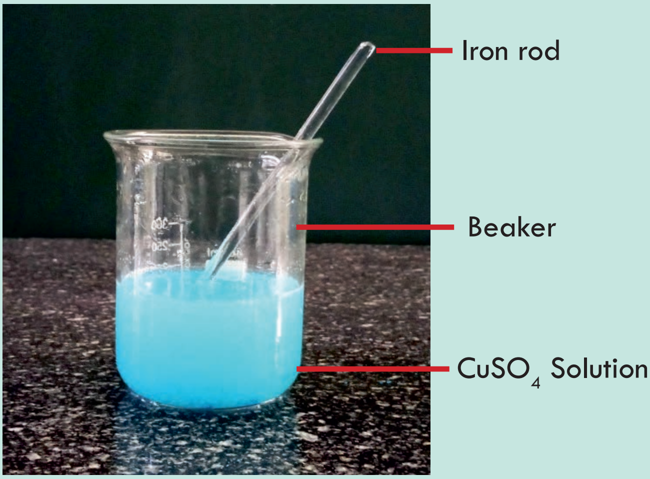


Figure 1.5 Reaction between iron and copper (II) sulphate.

Exercise 1.3

Complete the following single displacement reactions:

1. $Zn + H_2SO_4 \rightarrow$
2. $Zn + Cu(NO_3)_2 \rightarrow$
3. $Cu + AgNO_3 \rightarrow$
4. $Cl_2 + KI \rightarrow$
5. $F_2 + CaCl_2 \rightarrow$

D. Double Displacement Reactions

A double displacement reaction is a reaction in which two compounds react together to form two new compounds by exchange of the positive and negative ions of each reactant. Such a reaction is also known as double replacement reaction or metathesis. This type of reaction can be written in the following general form of equation:

$$AB + CD \rightarrow AD + CB$$

Experiment 1.4

Investigation of Double Displacement Reaction

Objective: To observe the displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$.
Apparatus: Beaker, stirrer, filter paper, filter funnel.
Chemicals: Na_2SO_4, $Ba(NO_3)_2$ and water.

Procedure:

1. Take solution of $Ba(NO_3)_2$ into a beaker and add drop-wise Na_2SO_4 solution as shown in *Figure 1.6*. Then stir it continuously.
2. Filter the precipitate using a filter paper and funnel. Collect the filtrate or the solution in a clean beaker.

Questions

1. Write the names of the compounds that are formed as a precipitate and as solution at the end of the reaction.
2. What was the colour of the precipitate.
3. Write the balanced chemical equation for the reaction.

Figure 1.6 Double displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$

Example 1.6:

The two soluble compounds $AgNO_3$ and NaCl react to produce an in soluble precipitate of AgCl and a soluble $NaNO_3$ solution.

$$\underset{\text{Soluble}}{AgNO_3} + \underset{\text{Soluble}}{NaCl} \rightarrow \underset{\text{Insoluble}}{AgCl} + \underset{\text{Soluble}}{NaNO_3}$$

When aqueous solutions of $BaCl_2$ and Na_2SO_4 react, a precipitate of $BaSO_4$ is formed.

$$\underset{\text{Soluble}}{BaCl_2} + \underset{\text{Soluble}}{Na_2SO_4} \rightarrow \underset{\text{Insoluble}}{BaSO_4 \downarrow} + \underset{\text{Soluble}}{2NaCl}$$

When dilute hydrochloric acid is added to sodium carbonate double displacement reaction takes place. In this reaction carbondioxide gas is produced.

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2\uparrow$$

Exercise 1.4

Give the correct answers for the following questions.

1. What type of reaction does usually take place in each of the following reactions?
   a. An active metal reacting with water
   b. A metal reacting with a non-metal
   c. An acid reacting with a metal hydroxide
   d. Heating of a metal hydrogen carbonate
2. Classify the following reactions as combination, decomposition, single or double displacement reactions.
   a. $FeO + C \rightarrow Fe + CO$
   b. $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$
   c. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$
   d. $2Cu(NO_3)_2 \rightarrow 2CuO + 4NO_2 + O_2$
   e. $2Na_3PO_4 + 3Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6NaOH$
   f. $CuSO_4.5H_2O \rightarrow CuSO_4 + 5H_2O$

1.4 Oxidation and Reduction Reactions

Activity 1.3

Form a group and discuss the following phenomenon and present your conclusion to the class.

When you dry your meal dishes with a towel, the towel can be termed as a drying agent and the dish as wetting agent. Assume that the meal dishes were wet with water, which one would gain water, the towel or the wetting agent and which one would lose the water, the wetting agent or the towel, after cleaning? Relate this idea with oxidizing agent and reducing agent, oxidation and reduction of substances.

At the end of this section, students will be able to

- define redox reactions;
- define the terms oxidation and reduction in terms of electron transfer;
- define oxidation number (oxidation state);
- state oxidation number rules.;
- determine the oxidation number of an element in a given formula.

In our day to day activity, we are familiar with the chemical processes like rusting of iron, burning of substances, respiration, digestion of food and so on. All such types of processes or reactions are known as oxidation and reduction or redox reactions. Additionally, most metallic and nonmetallic elements are obtained from their ores by the process of oxidation or reduction. What are oxidation and reduction reactions?

Oxidation-Reduction

An oxidation is defined as the loss of one or more electrons by an atom, and a reduction is the gain of one or more electrons. Oxidation and reduction occur simultaneously. Thus, an oxidation reduction reaction, or redox reaction, is one in which electrons are transferred from one atom to another.

For example, when metallic zinc is added to a solution containing copper(II) sulphate ($CuSO_4$), zinc reduces Cu^{2+} by donating two electrons to it. Therefore, Zn is oxidized and copper is reduced and oxidation – reduction or redox reaction takes place as shown in the following chemical equation:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$
$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

Oxidation and reduction reactions can also be defined in terms of oxidation number.

Oxidation is an increase in the oxidation number of an element and reduction is a decrease in the oxidation number. For example,

$$Cu^0 + 2Ag^+ \rightarrow Cu^{2+} + 2Ag^0$$

The oxidation number of copper is increased from 0 to +2 and thus copper is oxidized. The oxidation number of silver is decreased from +1 to 0, and therefore silver is reduced.

1.4.1 Oxidation Number or Oxidation State

At the end of this section, students will be able to

- define oxidation number (oxidation state);
- state oxidation number rules.;
- determine the oxidation number of an element in a given formula.

Oxidation number or oxidation state is the number of electrons that an atom appears to have gained or lost when it is combined with other atoms. Oxidation number could be integers including zero and fractional numbers.

Rules for Assigning Oxidation Numbers

Rule 1: The oxidation state of an uncombined element is zero. This applies for poly atomic molecules, S_8, P_4 and large structures of carbon or silicon each have an oxidation state of zero.

Example 1.7: The oxidation number of Be = 0, Cu = 0, Br in $Br_2 = 0$, O in

$O_3 = 0$, S in $S_8 = 0$. P in $P_4 = 0$, O in $O_2 = 0$

Rule 2: The oxidation number of a monatomic ion is equal to the charge on the ion.

Example 1.8: $Na^+ = +1$, $Mg^{2+} = +2$, $S^{2-} = -2$.

Rule 3: The oxidation number of oxygen in a compound is usually −2 except in the following cases:

Exceptions

The oxidation number of oxygen in:

A.	peroxides is −1.	Example: Na_2O_2
B.	superoxides is −1/2.	Example: KO_2
C.	oxygen diflouride is +2	Example: OF_2

Rule 4: The oxidation number of hydrogen in its entire compounds is +1 except in metal hydrides, (like NaH, CaH_2 and AlH_3), where its oxidation number is −1.

Rule 5: The sum of the oxidation number of all the atoms in a neutral compound is zero.

Example 1.9:

$\overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4}$, $(+2) + (+6) + (-8) = 0$

Rule 6: In a polyatomic ion, the sum of the oxidation numbers of the constituent atoms equals the charge on the ion.

Example 1.10:

$(\overset{+6}{S}\overset{-2}{O_4})^{-2}$, $(+6) + (-8) = -2$

Rule 7: Elements of group IA have +1 and group IIA have +2 oxidation states in all of their compounds

Rule 8: In a compound, the more electronegative element is assigned a negative oxidation number, and the less electronegative element is assigned a positive oxidation number.

Example 1.11: Assign oxidation number to all elements in

i. HNO_3
ii. $Cr_2O_7^{2-}$
iii. MnO_4^-
iv. $Ca(H_2PO_4)_2$.

i. HNO_3: According to rule 4 oxidation number of H = +1. Thus the other group (the nitrate ion) must have a net oxidation number of −1. Oxygen has an oxidation number of −2, and if we use x to represent the oxidation number of nitrogen, then the nitrate ion can be written as

$$[N^{(x)}O_3(^{2-})]^- \quad \text{or} \quad x + -6 = -1 \Rightarrow x = +5$$

ii. $Cr_2O_7^{2-}$: From rule 6 we see that the sum of the oxidation numbers in the dichromate ion $Cr_2O_7^{2-}$ must be −2. We know that the oxidation number of O is −2, so what remains is to determine the oxidation number of Cr, which we call y. The dichromate ion can be written as

$$[Cr_2^{(y)}\, O_7(^{2-})]^{2-}, \quad 2y + 7 \times -2 = -2 \Rightarrow y = +6$$

iii. MnO_4^-: Let the oxidation number of Mn be x. $Mn^xO_4^{2-}$. The sum of the oxidation numbers of Mn and O in MnO^-_4 is −1 (Rule 6)

$$x + (-2 \times 4) = -1, \quad x - 8 = -1 \Rightarrow x = +7$$

Therefore, the oxidation number of Mn in MnO_4^- is +7.

iv. $Ca(H_2PO_4)_2$: The oxidation number of Ca is +2. Let, the oxidation number of P be x.

$$\overset{+2}{Ca}(\overset{+1}{H}_2\overset{x}{P}\overset{-2}{O}_4)_2$$

$$+2 + (4 \times (+1)) + (2 \times x) + (8 \times (-2)) = 0$$
$$2 + 4 + 2x - 16 = 0$$
$$2x - 10 = 0 \text{ or } x = +5$$

Hence, the oxidation number of P in $Ca(H_2PO_4)_2$ is +5.

Exercise 1.5

1. Find the oxidation numbers to the underlined species for the following compounds or ions.
   - a. $K_4[\underline{Fe}(CN)_6]$
   - b. $K_2\underline{Cr}_2O_7$
   - c. $H_2\underline{Pt}Cl_6$
   - d. $Na_2\underline{S}_4O_6$
   - e. $\underline{S}_2O_8^{-2}$
   - f. $H_2\underline{P}_2O_7^{2-}$
   - g. $H_2\underline{S}O_4$
   - h. $H\underline{Au}Cl_4$
   - i. $Fe_2(\underline{S}O_4)_3$
2. Determine whether the following processes are oxidation or reduction reactions
   - a. $Cu^{2+} + 2e^- \rightarrow Cu$
   - b. $K \rightarrow K^+ + e^-$
   - c. $O + 2e^- \rightarrow O^{2-}$
   - d. $S^{2-} \rightarrow S + 2e^-$
   - e. $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   - f. $N + 3e^- \rightarrow N^{3-}$

1.4.2 Oxidizing and Reducing Agents

At the end of this section, students will be able to

- define the terms reducing and oxidizing agent;
- identify reducing and oxidizing agents from a given redox reaction;
- compare and contrast oxidizing and reducing agents.

Whenever one substance loses an electron (is oxidized), another substance must gain that electron (be reduced). The substance that gives up an electron and causes reduction is called a reducing agent. The substance that gains an electron and causes the oxidation is called an oxidizing agent. The following comparison shows the characteristics of reducing and oxidizing agents.

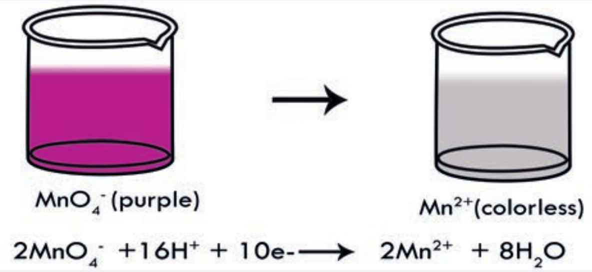
Reducing agent	**Oxidizing agent**
1. Loses one or more electrons	1. Gains one or more electrons
2. Causes reduction	2. Causes oxidation
3. Undergoes oxidation	3. Undergoes reduction
4. Becomes more positive	4. Becomes more negative

Tests for an oxidizing agent are accomplished by mixing it with a substance that is easily oxidized to give a visible color change when the reaction takes place.

Example 1.12:

I. Permanganate ion (MnO_4^-) in acidic solution changes color from purple to colorless

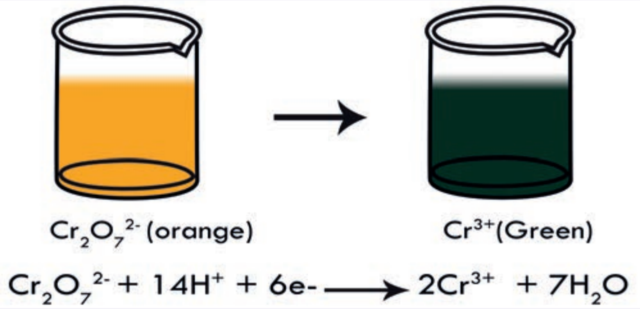
$$MnO_4^- \rightarrow Mn^{2+}$$



Figures 1.7 The reduction of Mn^{7+} to Mn^{2+} in acidic media.

II. Dichromate in acidic solution changes color from orange to green

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$



Figures 1.8 The reduction of Cr^{6+} to Cr^{3+} in acidic media.

Other common oxidizing agents are chlorine, potassium chromate, sodium chlorate and manganese (IV) oxide. Similarly, certain reducing agents undergo a visible colour change with a substance which is easily reduced.

I. A moist starch solution changes potassium iodide paper to blue-black to show that iodine is formed, $2I^- \rightarrow I_2$. That is potassium iodide is a reducing agent.
II. Hydrogen sulphide bubbled through a solution of an oxidizing agent forms a yellow precipitate, $S^{2-} \rightarrow S$. Thus, H_2S is a reducing agent.

Other common reducing agents are carbon, carbon monoxide, sodium thiosulphate, sodium sulphite and iron (II) salts. The oxidizing or reducing ability of substances depend on many factors. Some of these are:

Electronegativity: Elements with high electronegativity such as F_2, O_2, N_2 and Cl_2 are good oxidizing agents. Elements with low electronegativity for example, metallic elements like Na, K, Mg and Al are good reducing agents.

Oxidation states: In a compound or ion, if one of its elements is in its higher oxidation state, then it is an oxidizing agent. Similarly, if an element of a compound or ion is in its lower oxidation state, then it is a reducing agent.

Example 1.13: Oxidizing agents $KMn^{+7}O_4$, $NaCl^{+7}O_4$, $K_2Cr_2{}^{+6}O_7$
Reducing agents $Fe^{+2}S$, $C^{+2}O$, $Na_2S^{+4}O_3$

1.4.3 Analyzing Redox Reactions

At the end of this section, students will be able to explain why every reduction reaction is accompained by oxidation reaction.

Oxidation and reduction or redox reactions occur simultaneously in a given reaction.

Activity 1.4

Form a group and discuss the following and present your conclusion to the class.

1. How can you identify an oxidizing agent and reducing agent in a reaction?
2. Why a reducing agent undergoes oxidation and an oxidizing agent undergoes reduction?
3. Why must every redox reaction involve both an oxidizing agent and a reducing agent?

Example 1.14: Identify the reducing and oxidizing agents in the following balanced redox reactions:

$$2Fe + 3O_2 \rightarrow Fe_2O_3$$

Assign oxidation states of the reactants and products, then identify the species oxidized and reduced.

$$\overset{0}{2Fe} + 3\overset{0}{O_2} \rightarrow 2\overset{+3}{Fe_2}\overset{-2}{O_3}$$

Fe is oxidized from 0 to +3
O is reduced from 0 to –2

Therefore, Fe is reducing agent and O_2 is oxidizing agent.

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$
$$Mg^0 + 2H^{+1}Cl^{-1} \rightarrow Mg^{+2}Cl_2^{-} + H_2^{0}$$

Mg is oxidized from 0 to +2, therefore Mg is reducing agent.
H is reduced from +1 to 0, therefore HCl is the oxidizing agent.

Exercise 1.6

1. Where do the most easily reduced and oxidized elements found in the periodic table of the elements?
2. For the following substances, tell whether the oxidation number increases or decreases in a redox reaction:
   a. An oxidizing agent
   b. A reducing agent
   c. A substance undergoing oxidation
   d. A substance undergoing reduction
3. In the following reactions, label the oxidizing agent and the reducing agent.
   a. $ZnO + C \rightarrow Zn + CO$
   b. $8Fe + S_8 \rightarrow 8FeS$
   c. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
   d. $PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$

1.4.4 Balancing Redox Reactions: Oxidation-Number-Change Method

At the end of this section, students will be able to balance redox reactions using oxidation change method.

One way to balance redox reactions is by keeping track electron transfer, by using the oxidation numbers of each of the atoms. It is based on the difference in oxidation number of oxidizing agent and the reducing agent.

For the **oxidation-number-change method**, start with the unbalanced equation. The example below is for the reaction of iron (III) oxide with carbon monoxide.

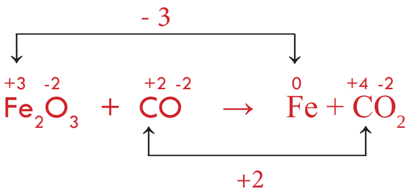
$$Fe_2O_3 + CO \rightarrow Fe + CO_2$$

Step 1: Assign oxidation numbers to each of the atoms in the equation and write the numbers above the atom.

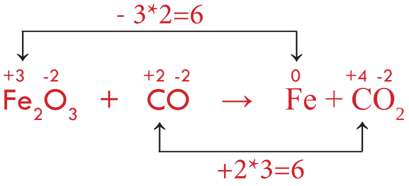
$$\overset{+3}{Fe_2}\overset{-2}{O_3} + \overset{+2}{C}\overset{-2}{O} \rightarrow \overset{0}{Fe} + \overset{+4}{C}\overset{-2}{O_2}$$

Step 2: Identify the atoms that are oxidized and those that are reduced. In the above equation, the C atom is being oxidized from +2 to +4 . The Fe atom is being reduced from +3 to 0.

Step 3: Use a line to connect the atoms that are undergoing a change in oxidation number. On the line, write the oxidation-number change.



Step 4: Use coefficients to make the total increase in oxidation number equal to the total decrease in oxidation number. In this case, the least common multiple of 2 and 3 is 6. So the oxidation-number increase should be multiplied by 3, while the oxidation-number decrease should be multiplied by 2. The coefficient is also applied to the formulas in the equation. So 3 is placed in front of the CO and in front of the CO_2. 2 is placed in front of the Fe on the right side of the equation. The Fe_2O_3 does not require a coefficient because the subscript of 2 after the Fe indicates that there are already two iron atoms.



Step 5: Check whether each element is balanced. Occasionally, a coefficient may need to be placed in front of a molecular formula that was not involved in the redox process. In the current example, the equation is now balanced.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Example 1.15: Balance the chemical equation:

$$HNO_3 + H_3AsO_3 \rightarrow NO + H_3AsO_4 + H_2O$$

1. Identify the oxidation number of each atom.

   Reactant side: H= +1; N= +5; O = -2; As= +3
   Product side: N= +2; O = -2; H= +1; As= +5

2. Determine the change in oxidation number for each atom that changes.

   N: +5 → +2; Change = -3
   As: +3 → +5; Change = +2

3. Make the total increase in oxidation number equal to the total decrease in oxidation number. We need 2 atoms of N for every 3 atoms of As. This gives us total changes of -6 and +6.
4. Place these numbers (2 & 3) as coefficients in front of the formulas containing those atoms

$$2HNO_3 + 3H_3AsO_3 \rightarrow 2NO + 3H_3AsO_4 + H_2O \text{ (balanced)}$$

Exercise 1.7

Balance the following using oxidation number change method

a. $Al + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2$
b. $KClO_3 \rightarrow KCl + O_2$
c. $MnO_2 + Al \rightarrow Mn + Al_2O_3$
d. $Cu + H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$

Non-redox Reactions

At the end of this section, students will be able to identify redox and non redox reactions.

What are non redox reactions? non redox reactions are chemical reactions where the oxidation states of chemical elements remain unchanged in reactants and products. Therefore, reactions in which neither oxidation nor reduction takes place or no species either gains or loses electrons are non-redox reactions. Neutralization and double displacement reactions are examples of non-redox reactions.

Example 1.16:

$$Na_2SO_4 + CaCl_2 \rightarrow CaSO_4 + 2NaCl$$
$$CaCO_3 \rightarrow CaO + CO_2$$
$$KOH + HNO_3 \rightarrow KNO_3 + H_2O$$

1.5 Molecular and Formula Masses, the Mole Concept and Chemical Formulas

Activity 1.5

Discuss the following in your group and present your conclusion to the class.

1. Is it possible to count 'teff' seeds for practical purposes? If not, why? How do you quantify 'teff' for practical purposes such as in markets?
2. Molecules, atoms or ions are extremely small entities. How can we express their quantities for practical purposes such as in the laboratory reactions and in industrial applications?

At the end of this section, students will be able to

- explain and determine molecular mass, formula mass, molar mass, empirical and molecular formulas;
- describe the mole concept and solve problems related to moles of substances.

1.5.1 Molecular mass (MM) and Formula Mass (FM)

At the end of this topic, students will be able to explain and determine molecular mass, formula mass.

The molecular mass or formula mass of a compound is used to calculate the percentage of its constituents. For example, the molecular mass of H_2O is used to determine the percentage of hydrogen, and the formula mass of NaCl is used to determine the percentage of sodium. What is molecular mass or formula mass?

Molecular Mass (MM) is the sum of masses of all the atoms present in a molecule.

Example 1.17:

MM of H_2O = 2 x 1 amu (2 hydrogen atoms) + 16 amu (1 oxygen atom) = 18 amu,

MM of CO_2 =12 amu (1carbon atom) +2 x 16 amu (2 oxygen atoms) = 44 amu

Ionic compounds are not comprised of discrete molecules, but rather, ions. The smallest unit of an ionic compound that retains the identity of an ionic compound is called a formula unit. The mass of a formula unit is called its formula mass.

Fomula Mass (FM) is the sum of the atomic masses of all atoms present in the formula unit of the compound, whether it is molecular or ionic. But FM is used mostly for ionic compounds.

Example 1.18:

FM of NaCl = 23 amu (1Na) + 35.5 amu (Cl) = 58.5 amu.

FM of $Ca(OH_2)_2$ = 40 amu (1Ca) + 2 x [16 amu(1O) + 2 amu (2H)] = 76 amu

1.5.2 The Mole Concept

At the end of this topic, students will be able to

- define mole and explain the concept using examples;
- define molar mass;
- calculate the molar masses of substances.

It is the concept adopted as a convenient way to deal with the tremendous amount of molecules or ions. A mole is defined as the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12. 12 g of carbon-12 contains $6.022 x 10^{23}$ atoms. This number is known as Avogadro's number (NA), in honor of Amedeo Avogadro (1776-1856, who first proposed the concept). The term mole, like a dozen or a gross, refers to a particular number of things. A dozen of eggs equals 12 eggs, a gross of pencils equals 144 pencils, and a mole of any substance equals $6.02 x 10^{23}$ units of the substance.

Example 1.19: 1 mole of H_2O equals 6.02×10^{23} molecules of H_2O
1 mole of NaCl equals 6.02×10^{23} formula units of NaCl

Molar mass is the molecular mass or formula mass expressed in grams. It is the mass of 1 mole of a substance.
Molar mass of substance A = mass of one mole of A = $6.02 x 10^{23}$ units of substance A

Example 1.20: molar mass of H_2O = 18 g = mass of 1 mole of H_2O = 6.02×10^{23} molecules of H_2O

Molar mass of NaCl = 58.5 g = mass of one mole of NaCl =6.02×10^{23} formula units of NaCl

1.5.3 Chemical Formulas - Empirical and Molecular Formulas

At the end of this topic, students will be able to

- define empirical and molecular formulas;
- determine empirical and molecular formulas;
- calculate percentage composition of an element in a substance.

Chemical formulas are used to express the composition of compounds in terms of chemical symbols. By composition we mean not only the elements present but also the ratios in which the atoms are combined. Here, you deal with two types of formulas, empirical formulas and molecular formulas.

Empirical formula (or simplest formula) for a compound is the formula of a substance written with the smallest ratio (whole number ratio) subscripts. The molecular formula is the actual formula that tells you the exact number of atoms of different elements present in a molecule.

Example 1.21:	Empirical formula	Molecular formula
Benzene	CH	C_6H_6
Glucose	CH_2O	$C_6H_{12}O_6$

Note that the empirical formula, merely tells you the ratio of numbers of atoms in the compound.

Percent Composition by Mass (%)

Percent composition by mass tells you what percent of each element is present in a compound. Thus, it helps in chemical analysis of the given compound. The percentage composition of a given compound is defined as the ratio of the amount of individual elements present in the compound multiplied by 100. Mathematically,

$$\text{Percent composition of an element} = \frac{\text{n} \times \text{molar mass of the element}}{\text{molar mass of compound}} \times 100\%$$

where n is the number of moles of the element in 1 mole of the compound. For example, in 1 mole of hydrogen peroxide (H_2O_2) there are 2 moles of H atoms and 2 moles of O atoms. The molar masses of H_2O_2, H, and O are 34.02 g, 1.008 g, and 16.00 g, respectively. Therefore, the percent composition of H_2O_2 is calculated as follows:

$$\%\ \text{H} = \frac{2 \times 1.00\ \text{g H}}{34.02\ \text{g}\ H_2O_2} = 5.926\%$$

$$\%\ \text{O} = \frac{2 \times 16}{34.02\ H_2O_2} = 94.06\%$$

Determination of Empirical and Molecular Formulas

Steps to determine empirical formula are:

Step 1: Derive the number of moles of each element from its mass. Masses of elements may be given in terms of percent composition of elements or grams.

Step 2: Divide each element's molar amount by the smallest molar amount to yield subscripts for a tentative empirical formula.

Step 3: Multiply all coefficients by an integer, if necessary, to ensure that the smallest whole number ratio of subscripts is obtained

Example 1.22: What is the empirical formula of a compound that contains 43.6% P and 56.4% O ?

Step 1: Derivation of moles of each element,

$$\text{Number of moles each element} = \frac{\text{Given mass of the element}}{\text{Molar mass of the element}}$$

$$\text{Number of moles of P} = \frac{43.6\ \text{g}}{31.0\ \text{g/mol}} = 1.41\ \text{mol}$$

$$\text{Number of moles of O} = \frac{56.4\ \text{g}}{16.0\ \text{g/mol}} = 3.53\ \text{mol}$$

Step 2: Smallest molar amount is 1.41

P: 1.41/1.41=1.00,   O: 3.53/1.41 = 2.50

Step 3. Multiply both by 2 to convert into whole number, P: 2 x 1 =2, O : 2 x 2.5 = 5
Empirical formula is P_2O_5

Example 1.23: A compound contains 79.9% carbon and 20.1 % of hydrogen by mass. What is the empirical formula of the compound?

Step 1: C= (79.9 g)/(12 g/mol) = 6.7 mol, H = 20.1g/(1g/mol) = 20.1 mol
Step 2: C= 6.7/6.7 = 1,   H = 20.1/6.7 = 3
Step 3: Empirical formula = CH_3

Molecular formula = Empirical formula x n

$$n = \frac{\text{molar mass of molecular formula}}{\text{molar mass of empirical formula}}$$

Example 1.24: What is the molecular formula of the oxide of phosphorus that has the empirical formula P_2O_5 if it's molecular mass is 284?

Molar mass of P_2O_5 = 142,   n = 284/142 = 2
Therefore, molecular formula = $(P_2O_5)_2 = P_4O_{10}$

Example 1.25: The compound ethylene glycol is often used as antifreeze. It contains 38.7% carbon, 9.75% hydrogen, and the rest oxygen. The molecular weight of ethylene glycol is 62.07 g. What is the molecular formula of ethylene glycol?

1. Calculate the empirical formula. Assume 100 g of the compound, which will contain 38.70 g carbon, 9.75 g hydrogen and the rest oxygen.

   Mass of O = 100 g - (38.7 g of C + 9.75 g H) = 51.55 g

2. Calculate the moles of each element present:

$$\text{No of moles of C} = \frac{38.70 \text{ g C}}{12.01 \text{ g C}} = 3.22 \text{ mol C}$$

$$\text{No of moles of H} = \frac{9.75 \text{ g H}}{1.008 \text{ g H}} = 9.67 \text{ mol H}$$

$$\text{No of moles of O} = \frac{51.55 \text{ g O}}{16.01 \text{ g O}} = 3.22 \text{ mol O}$$

Next, calculate the ratio of molecular weight to empirical formula weight. The molecular weight is given. The empirical formula is CH_3O, so the empirical formula weight is 12.01 + 3(1.008) + 16.00 = 31.03.

$$n = \frac{\text{Molecular mass}}{\text{empirical formula mass}} = \frac{62.07}{31.03} = 2$$

Therefore, the molecular formula is twice the empirical formula: $C_2H_6O_2$.

Exercise 1.8

1. Define and discuss each of the following terms: mole, Avogadro's number, molar mass, molecular mass, formula mass, percentage composition
2. Distinguish between
   a. Formula mass and molecular mass of a compound
   b. Empirical and molecular formula
3. How many atoms and ions are there in one formula unit of
   a. K_2CO_3?
   b. $Ba_3(PO_4)_2$?
4. How many moles of each element are present in
   a. 1.00 mole of $Cu_3(PO_4)_2$
   b. 2.5 mole of Na_2CO_3
   c. 2 moles of $Al_2(SO_4)_3$
5. Calculate the empirical formula for:
   a. 39.3% Na, 60.7% Cl
   b. 56.5% K, 8.7% C, 34.8% O
6. A. A compound of carbon and hydrogen contains 92.3% C and has a molar mass of 78.1g/mol. What is its molecular formula?
   B. A compound of carbon, hydrogen and chlorine has a molar mass of 99 g/mol. Analysis of a sample shows that it contains 24.3% carbon and 4.1% hydrogen. What is its molecular formula?
7. Calculate the molar mass of a compound if 0.372 mole of it has a mass of 152 g.
8. Calculate the molecular mass or formula mass (in amu) of each of the following substances:
   a. CH_4
   b. NO_2
   c. SO_3
   d. C_6H_6
   e. Na_2SO_4
   f. $Ca_3(PO_4)_2$
9. Calculate the percentage composition by mass of the following compounds:
   a. C_2H_5OH
   b. Na_3PO4
   c. H_2SO_4
   d. $CaCO_3$
   e. $K_2Cr_2O_7$

1.6 Stoichiometry

Activity 1.6

Refer to grade 9 chemistry and revise the three fundamental laws of chemical reactions. Discuss in class using the following and give your own examples for each law.

a. Carbon forms two types of oxides namely carbon monoxide and carbon dioxide.
b. The ratio of the mass of hydrogen to the mass of oxygen in water is always 1:8.
c. The total masses of calcium, carbon and oxygen are the same before and after the decomposition of calcium carbonate.

At the end of this section, students will be able to

- deduce mole ratios from balanced chemical equations;
- solve mass-mass problems based on the given chemical equation;
- define molar volume;
- state Avogadro's principle;
- solve mass-mass problems based on the given chemical equation;
- solve mass-volume problems based on the given chemical equation;
- solve volume-volume problems based on the given chemical equation;
- describe limiting and excess reactants.

Stoichiometry is the study of the quantitative composition of substances and the relationship that exists between the reactants and the products involved in chemical reactions. In other words, stoichiometry is the study of the amount or ratio of moles, mass, or volumes (for gases) of reactants and products. Stoichiometric calculations are based on the following two major principles.

A. The composition of any substance in the chemical equation should be expressed by a definite formula.
B. The law of conservation of mass must be obeyed (the mass of reactants equals the mass of products)

Activity 1.7

Form a group and discuss the following phenomenon. When wood burns, the ash weighs much less than the original wood. Where did the "lost mass" go? How can you estimate the mass of the wood that is no longer present? Present your conclusion to the class.

Note that stoichiometric calculations are done for balanced chemical equations.

1.6.1 Molar Ratios in Balanced Chemical Equation

At the end of this section, students will be able to deduce mole ratios from balanced chemical equations.

Stoichiometry problems of this type are called either mass-mass or mole-mole problems. A balanced chemical equation provides information about the nature, masses, number of moles, number of molecules/atoms of reactants and products.

For example, consider the reaction of hydrogen with nitrogen to produce ammonia. The equation shows that 3 molecules of hydrogen and 1 molecule of N_2 react to produce 2 molecules of ammonia. The equation can also represent the reaction of 3 moles ($3 \times 6.02 \times 10^{23}$ molecules) of H_2 with 1 mole (6.02×10^{23} molecules) of N_2 to yield 2 moles ($2 \times 6.02 \times 10^{23}$ molecules) of NH_3. The same equation can also tell us the combination of 6 g of H_2 with 28 g of N_2 to yield 34 g of NH_3. This can be further interpreted as follows:

$$N_2 \quad + \quad 3H_2 \quad \rightarrow \quad 2NH_3$$

1 molecule N_2 + 3 molecules H_2 → 2 molecules NH_3 (Molecular interpretation)
1mol N_2 + 3 mol H_2 → 2 mol NH_3 (Molar interpretation)
28.0 g N_2 + 3 x 2.02 g H_2 → 2 x 17.0 g NH_3 (mass interpretation)

Calculations based on chemical equations (stoichiometric problems) are classified into mass-mass, volume-volume and mass-volume problems.

1.6.2 Mass – Mass Relationships

At the end of this section, students will be able to solve mass-mass problems based on the given chemical equation or using the conversion factor.

There are two methods for solving mass-mass problems. These are mass-ratio and mole-ratio methods.

A. The Mass - ratio Method

In this method, the mass of one substance is determined from the given mass of the other substance using the following steps.

Step 1: Write the balanced chemical equation.

Step 2: Place the given mass above the corresponding formula, and x above the formula of the substance whose mass is to be determined

Step 3: Write the total molar mass of the substances below the formula of each substance. (Total molar mass is the molar mass of the substance multiplied by its coefficient).

Step 4: Set up the proportion.

Step 5: Solve for the unknown mass, x.

Example 1.26: How many grams of aluminum metal must be heated to produce 20.4 g of aluminium oxide?

Solution:

Step 1: $4Al + 3O_2 \rightarrow 2Al_2O_3$

Step 2: $\overset{x}{4Al} + 3O_2 \rightarrow \overset{20.4\ g}{2Al_2O_3}$

Step 3: $\underset{108}{\overset{x}{4Al}} + 3O_2 \rightarrow \underset{204}{\overset{20.4\ g}{2Al_2O_3}}$

Step 4: $x/(108\ g) = (20.4\ g)/(204\ g)$

Step 5: $\Rightarrow x = (20.4\ g)/(204\ g) \times 108\ g = 10.8\ g$

Therefore, 10.8 g of aluminum metal is needed to produce 20.4 g of aluminum oxide.

Note that stoichiometric problems can also be solved using the mole ratio or molar mass as conversion factors. The numbers in a conversion factor come from the coefficients of the balanced chemical equation.

In general, for mole – mole problems,

$$\text{mole A} \times \text{mole ratio} = \text{mole B}$$

Mass-Mass problems,

$$\text{gram A} \times \frac{\text{1mole A}}{\text{molar mass A}} \times \frac{\text{mole B}}{\text{mole A}} \times \frac{\text{molar mass B}}{\text{mole B}} = \text{gram B}$$

In the above example, coefficients of Al and Al_2O_3 are 4 and 2 respectively Therefore, conversion factor = mole ratio = 4 mol Al/2 mol O_2. Given molar masses, 1 mole Al_2O_3 = 102 g, 1 mole Al = 27 g

The conversion steps are summarized as follows:

grams of Al_2O_3 → moles of Al_2O_3 → moles of Al → grams of Al
↓ ↓ ↓
Conversion factor: molar mass of Al_2O_3 mole ratio molar mass of Al

20.4 g x 1 mol Al_2O_3/102 g Al_2O_3 x 4 mol Al /2 mol Al_2O_3 x 27 g Al/1mol Al = 10.8 g

Example 1.27: What mass of magnesium sulfate will be produced if 4.8 g of magnesium reacts with excess sulfuric acid?

Step 1: $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

Step 2: $\overset{4.8g}{Mg} + H_2SO_4 \rightarrow \overset{x}{MgSO_4} + H_2$

Step 3: $\overset{4.8g}{\underset{24g}{Mg}} + H_2SO_4 \rightarrow \overset{x}{\underset{120g}{MgSO_4}} + H_2$

Step 4: $\Rightarrow x = (4.8\ g)/(24\ g)\ \times 120\ g = 24\ g$

Therefore, 13.6 g $MgSO_4$ will be produced. Check the result using conversion factor. **Hint:** mole ratio = 1mole Mg/1mole $MgSO_4$

B. The Mole - ratio Method

In this method, calculations are made in terms of moles; therefore the given mass is converted into mole. The obtained mole can be converted back to mass if required. Follow the steps given below to solve problems of mass-mass relationships using mole ratio method:

Step 1: Write the balanced chemical equation.

Step 2: Convert the given mass to moles and write the obtained moles and the required quantity, x, above the formulas of the respective substances.

Step 3: Place the coefficients as the number of moles under the formula of each substance involved.

Step 4: Set up the proportion.

Step 5: Solve for the unknown value, x; and convert the moles obtained into mass.

Example 1.28: How many grams of calcium oxide are needed to react completely with 22.0 g of carbon dioxide?

Solution:

Step 1: $CaO + CO_2 \rightarrow CaCO_3$

Step 2: moles of CO_2 = given mass/molar mass = 22 g/44 g/mol = 0.5 mol

Step 3: $\overset{x}{\underset{1mol}{CaO}} + \overset{0.5mol}{\underset{1mol}{CO_2}} \rightarrow CaCO_3$

Step 4: $x/1mol = 0.5mol/1mol \qquad \Rightarrow x = 0.5\ mol$

Step 5: Convert the moles into grams of CaO

Mass of CaO produced = No. of moles CaO x molar mass CaO

$$= 0.5mol \times 56\ g/mol = 28\ g$$

Solve the above problem using conversion factor,

gram $CO_2 \rightarrow$ Mole $CO_2 \rightarrow$ mole CaO $\rightarrow$ gram CaO

$$22\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44\text{ g } CO_2} \times \frac{1\text{ mol CaO}}{1\text{ mol } CO_2} \times \frac{56\text{ g CaO}}{1\text{ mol CaO}} = 28\text{ g of CaO}$$

Example 1.29: How many grams of carbon monoxide must react with excess iron oxide to produce 28 grams of iron?

Step 1: $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Step 2: No. of moles Fe $= \frac{\text{given mass}}{\text{molar mass}}$

$= 28$ g/ 56 g/mol = 0. 5 mol

Step 3: $Fe_2O_3 + \underset{3\text{ mol}}{\overset{x}{3CO}} \rightarrow \underset{2\text{ mol}}{\overset{0.5\text{ mol}}{2Fe}} + 3CO_2$

Step 4: x/(3 mol) = 0.5 mol/(2 mol) ⇒ x = 0.5 x 3 mol/2 = 0.75 mol of CO

Step 5: Convert the moles into grams of CO

Mass of CO required = no. of moles CO x molar mass CO

= 0.75 mol x 28 g/mol =21 g (check the result using the conversion factor)

Note that although mass-mass and mole-mole stoichiometric problems are more common, there are also mass-mole and mole-mass problems.

In mole-mass problems,the amount of one substance is given in moles and the mass of another substance is determined, usually in grams and in mass-mole problems, the mass of one substance is given, usually in grams and the amount of another substance is determined in moles. Refer to relevant sources and solve ***Exercise 1.9*** (a,b). You may use the general sequence of conversions

for a mole-mass calculation follow the following sequence:

mol 1st substance → mol 2nd substance → mass 2nd substance

Example: What mass of Fe_2O_3 must be reacted to generate 3 moles of Al_2O_3?

$$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$$

Given: Moles of Al_2O_3 = 3, Molar mass of Al_2O_3 = 102 g Required: Mass of Fe_2O_3,

molar mass of Fe_2O_3 = 160 g

Solution: Follow the sequence,Gram Al_2O_3 → mole Al_2O_3 → mole Fe_2O_3 → Gram Fe_2O_3

$$3\text{ moles of } Al_2O_3 \times \frac{1\text{ mole } Fe_2O_3}{1\text{ mol } Al_2O_3} \times \frac{160\text{ g}}{1\text{ mol } Fe_2O_3} = 480\text{ g of } Fe_2O_3$$

for mass- mole calculation follow the following sequence:

mass 1st substance → mol 2nd substance → mol 2nd substance

Example: How many moles of C_2H_5OH are needed to generate 106.7 g of H_2O?

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

Given: mass of H_2O = 106.7 g, molar mass of H_2O = 18 g Required: No. of mole of C_2H_5OH

Solution: follow the steps: Gram H_2O → moleH_2O → mole C_2H_5OH

$$106.7\ g \times \frac{1 mole\ H_2O}{18g} \times \frac{mole\ C_2H_5OH}{3mol\ H_2O} = 1.97 \text{ moles of } C_2H_5OH$$

Exercise 1.9

1. Which of the following statements is correct for the equation shown here?

   $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

   a. 6 g of H_2O are produced for every 4 g of NH_3 reacted.

   b. 1 mole of NO is produced per mole of NH_3 reacted.

   c. 2 moles of NO are produced for every 3 moles of O_2 reacted

2. How many moles of H_2O are required to produce 4.5 moles of HNO_3 according to the following reaction?

   $$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

3. How many moles of CaO are needed to react with excess water to produce 370 g of calcium hydroxide?
4. In the decomposition of $KClO_3$, how many moles of KCl are formed in the reaction that produces 0.05 moles of O_2?
5. How many grams of $CaCO_3$ are needed to react with 15.2 g of HCl according to the following equation?

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

6. What mass of nitrogen dioxide is produced by the decomposition of 182 g of magnesium nitrate?
7. Solve the following using either mass-mole or mole-mass relations

   a. What mass of hydrogen sulphide gas is burned in a reaction that produces 4 mol SO_2 and water vapor: $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$

   b. How many moles of CO_2 will be produced when 50 g of $CaCO_3$ are reacted according to the chemical equation:

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

1.6.3 Volume – Volume Relationships

At the end of this section, students will be able to solve volume -volume problems based on the given chemical equation.

At Standard Temperature (0^0C) and Pressure (1atm) or STP, one mole of any gas occupies 22.4 liters. This, 22.4 liters is known as molar volume at STP.

The volume of a gas and its number of molecules are related and explained by Avogadro's Law. Avogadro's Law states that equal volumes of different gases, under the same conditions of temperature and pressure, contain equal number of molecules. According to this law, the volume of a gas is proportional to the number of molecules (moles) of the gas at STP as shown below:

$V \sim n$; where V is the volume and n is the number of moles of gas

In Volume – volume problems, the volume of one substance is given and the volume of the other substance is calculated.

Example 1.30: What volume of oxygen will react with methane to produce 44.8 liters of carbon dioxide at STP?

Step 1: $CH_4 + 2O_2 \rightarrow 2H_2O + CO_2$

Step 2: Place the given volume and the required volume, x above the corresponding formulas.

$$CH_4 + \underset{44.8\ L}{\overset{x}{2O_2}} \rightarrow 2H_2O + \underset{22.4\ L}{\overset{44.8\ L}{CO_2}}$$

Step 3: Write the total molar volume (22.4 L multiplied by any coefficient) below the formulas.

Step 4: x/44.8 L=(44.8 L)/(22.4 L) ⟹ x = 44.8 /22.4 x 44.8 L= 89.6 L

Step 5: Solve for the unknown volume, x

x = 89.6 L of oxygen is needed.

Exercise 1.10

1. What volume of nitrogen reacts with 33.6 litres of oxygen to produce nitrogen dioxide?
2. How many litres of sulphur trioxide are formed when 4800 cm^3 of sulphur dioxide is burned in air?
3. How many litres of ammonia are required to react with 145 litres of oxygen according to the following reaction?

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

4. Calculate the volume of oxygen produced in the decomposition of 5 moles of $KClO_3$ at STP?
5. How many moles of water vapour are formed when 10 litres of butane gas, C_4H_{10} is burned in oxygen at STP?

1.6.4 Mass – Volume Relationships

At the end of this section, students will be able to solve mass -volume problems based on the given chemical equation.

In mass-volume problems, either the mass of one substance is given and the volume of the other is required or the volume of one substance is given and the mass of the other one is required. The steps to solve such type of problems are the same as the previous steps except putting the masses on one side and the volumes on the other side of the equality sign.

Example 1.31: How many grams of magnesium should react with sulphuric acid to produce 5.6 L of hydrogen at STP?

Solution:

Step 1: $Mg + H_2SO_4 \rightarrow H_2$

Step 2: $\overset{x}{Mg} + H_2SO_4 \rightarrow \overset{5.6\ L}{MgSO_4} + H_2$

Step 3: $\underset{24\ g}{\overset{x}{Mg}} + H_2SO_4 \rightarrow \underset{22.4\ L}{\overset{5.6\ L}{MgSO_4}} + H_2$

Step 4: $x/24\ g = 5.6\ L/22.4\ L$

$\Rightarrow x = 6\ g$

Step 5: x = 6 g of magnesium is needed

Exercise 1.11

1. How many litres of oxygen are required to react with 156 g of benzene according to the chemical equation:

$$2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$$

2. What mass of iron metal would be completely oxidized by 89.6 L of oxygen to produce Fe_2O_3 at STP?
3. Calculate the mass of calcium carbide that is needed to produce 100 cm^3 of acetylene according to the equation:

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

4. How many milliliters of sulphur dioxide are formed when 12.5 g of iron sulphide ore (pyrite) reacts with oxygen (at STP) according to the equation

$$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$$

1.6.5 Limiting and Excess Reactants

At the end of this section, students will be able to describe limiting and excess reactants. calculate the limiting and excess reactants of a reaction.

Activity 1.8

Consider a dance contest at a club. If there are 14 men and only 9 women, then only 9 female/male pairs can compete.

1. How many men will be left without partners?
2. Who (men or women) are in excess?
3. Who (men or women) limits the number of female/male pair that can dance in the contest?
4. How can you relate this analogue with the chemical reactions involving two or more reactants?

Discuss in your group and present in class your conclusion.

In a chemical reaction involving two reactants, the reaction will stop when all of one reactant has been completely consumed no matter how much of the second reactant remains. The reactant completely consumed first in a reaction is called the limiting reactant, because it limits or determines the amount of product that can be formed. When the limiting reactant is completely consumed, no more product can be formed and thus the other reactant remains excess. Excess reactants are the reactants present in quantities greater than necessary to react with the quantity of the limiting reactant.

Suppose you want to make some cheese sandwiches. Each is made from two slices of bread and a slice of cheese.

2 slices bread + 1 slice cheese → 1 cheese sandwich

If you have 7 slices of bread and 2 slices cheese, you can make 2 cheese sandwiches. 3 slices of bread is excess. The number of sandwiches is limited by the number of slices of cheese. Hence, slices of cheese are the limiting ingredient and slices of bread are excess ingredient.

Note that: Yield of the product is the one calculated using the limiting reactant and it is the smallest.

Example 1.32: Magnesium metal reacts with hydrochloric acid by the following reaction:

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

If 0.30 mol Mg is added to hydrochloric acid containing 0.52 mol HCl, how many moles of H_2 are produced?

Solution:
Here there are 0.3 mol of Mg and 0.52 mol of HCl. Therefore, first determine the limiting reactant. Calculate the yield of H_2 (product) using both reactants, the reactant that gives lower yield is the limiting reactant and the amount obtained by the limiting reactant is the yield of the product .

A. Using the quantity of Mg
0.3mol/1mol = x/1mol

$$\underset{1\ \text{mol}}{\overset{0.3\ \text{mol}}{Mg}} + 2HCl \rightarrow MgCl_2 + \underset{1\ \text{mol}}{\overset{x}{H_2}}$$

$0.3\ \text{mol}/1\ \text{mol} = x/1\ \text{mol}$
$\Rightarrow x = 0.3$ mol of H_2

B. Using the quantity of HCl
0.52 mol/2 mol = x/1mol

$$Mg + \underset{2\ \text{mol}}{\overset{0.52\ \text{mol}}{2HCl}} \rightarrow MgCl_2 + \underset{1\text{mol}}{\overset{x}{H_2}}$$

$0.52\ \text{mol}/2\ \text{mol} = x/1\ \text{mol}$
$\Rightarrow x = 0.26$ mol of H_2

H_2 produced using HCl is less than H_2 produced using Mg.

Hence, HCl is the limiting reactant and Mg is the excess reactant. The number of moles of H_2 produced is equal to the quantity of H_2 obtained using HCl (limiting reactant) which is 0.26 mol .

Example 1.33: In a process for producing acetic acid, oxygen gas is bubbled into acetaldehyde, CH_3CHO, containing manganese (II) acetate (catalyst) under pressure at 60 ^{0}C.

$$2CH_3CHO + O_2 \rightarrow 2HC_2H_3O_2$$

In a laboratory test of this reaction, 20.0 g CH_3CHO and 10.0 g O_2 were put into a reaction vessel. A. How many grams of acetic acid can be produced? B. How many grams of the excess reactant remain after the reaction is complete?

Solution

1. Convert grams in to moles

$$\text{No. of moles of } CH_3CHO = \frac{\text{given mass}}{\text{molar mass}}$$

$$= \frac{20\text{ g}}{44\text{ g/mol}} = 0.454\text{ mol}$$

$$\text{No. of moles of } O_2 = \frac{\text{given mass}}{\text{molar mass}}$$

$$= \frac{10\text{ g}}{32\text{ g/mol}} = 0.312\text{ mol}$$

2. Calculate the amount of $HC_2H_3O_2$ produced by 0.454 mol. CH_3CHO and 0.312 mol O_2 in order to decide the limiting reactant.

A. Using the quantity of CH_3CHO	**B. Using the quantity of O_2**
$\underset{2\text{ mol}}{\overset{0.454\text{ mol}}{2CH_3CHO}} + O_2 \rightarrow \underset{2\text{ mole}}{\overset{x}{2HC_2H_3O_2}}$	$2CH_3CHO + \underset{1\text{ mol}}{\overset{0.312\text{ mol}}{O_2}} \rightarrow \underset{2\text{ mol}}{\overset{x}{2HC_2H_3O_2}}$
$0.454\text{ mol}/2\text{ mol} = x/2\text{ mol}$	$0.312\text{ mol}/1\text{ mol} = x/2\text{ mol}$
$\Rightarrow x = 0.454$ mol of $HC_2H_3O_2$	$\Rightarrow x = 0.624$ mol of $HC_2H_3O_2$

Therefore, the limiting reactant is acetaldehyde, CH_3CHO. It is the reactant that gives lower yield in the above reaction. Oxygen is the excess reactant in the above reaction since it gives higher yield.

Number of moles of $HC_2H_3O_2$ produced = 0.454 mol

A. Amount of $HC_2H_3O_2$ produced in grams = 0.454 mol x 60 g/mol = 27.2 g

B. Calculate the amount of oxygen consumed using the product obtained by the limiting reactant

$$2CH_3CHO + \underset{1\text{ mol}}{\overset{x}{O_2}} \rightarrow \underset{2\text{ mol}}{\overset{0.454\text{ mol}}{2HC_2H_3O_2}}$$

Amount of O_2 consumed, x = 0.454 mol x 1 mol/2 mol = 0.227 mol
We started with 10 g of O_2 or 0.312 mol, so the excess O_2 = 0.312 - 0.227 = 0.085mol or 2.7 g.

Exercise 1.12

1. If 6.5 g of zinc reacts with 5.0 g of HCl, according to the following reaction.

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

   a. Which substance is the limiting reactant?
   b. How many grams of the reactant remain unreacted?
   c. How many grams of hydrogen would be produced?
2. What mass of Na_2SO_4 is produced if 49 g of H_2SO_4 reacts with 80 g of NaOH?
3. If 20 g of $CaCO_3$ and 25 g of HCl are mixed, what mass of CO_2 is produced?

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

4. If 3 moles of calcium react with 3 moles of oxygen, then
   a. Which substance is the limiting reactant?
   b. How many moles of calcium oxide are formed?
5. For the reaction: $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$, how many grams of hydrogen are produced if 0.8 mole of aluminum reacts with 1.0 mole of sulphuric acid?

1.6.6 Theoretical, Actual and Percentage Yields

At the end of this section, students will be able to

- describe theoretical, Actual and percentage yields;
- calculate theoretical, Actual and percentage yields.

Activity 1.9

By Considering the preparation of 'Shiro' powder from chickpeas at home, discuss in group the following questions and present your conclusion to the class.

a. Have you noticed the amount of chickpeas before and after the preparation of Shiro powder? Is it possible for the amount of chickpeas to be equal after and before making the Shiro powder? If not, what is the reason?
b. How could the quantity of the Shiro powder preparation be maximized?
c. How can you compare this analogy with the yield of products in a chemical reaction?

Many chemical reactions do not proceed to give 100 % yield for a number of reasons. Some reactions generate unwanted products due to side reactions. Others are, by nature incomplete. Some products are difficult to collect without loss which reduces the yield. As a result, the actual yield (experimentally determined yield) of a product is usually less than the theoretical yield (calculated yield). The theoretical yield is the calculated amount of product that would be obtained if the reaction proceeds completely. The measured amount of product obtained in any chemical reaction is known as the actual yield.

The percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100.

$$\text{Percentage yield} = \frac{\overset{\swarrow\ \text{Measured amount}}{\text{Actual yield}}}{\underset{\leftarrow\ \text{Calculated}}{\text{Theoretical yield}}} \times 100\%$$

Example 1.34: What is the percent yield if 24.8 g of $CaCO_3$ is heated to give 13.1 g of CaO?

Solution

The actual yield of CaO (given in the problem!) is 13.1 g

Determine the theoretical yield using mass-mass relationship

$$\underset{100\text{ g}}{\overset{24.8\text{ g}}{CaCO_3}} \rightarrow \underset{56\text{ g}}{\overset{x}{CaO}} + CO_2$$

$x/(56\text{ g}) = (24.8\text{ g})/(100\text{ g}) \Longrightarrow\ x = 13.9\text{ g}$ is the theoretical yield of CaO

$$\text{Percentage yield of CaO} = \frac{\text{Actual yield of CaO}}{\text{Theoretical yield of CaO}} \times 100$$

Percentage yield of CaO = $(13.1\text{ g})/(13.9\text{ g}) \times 100 = 94.2\%$

Exercise 1.13

1. When 21.6 g of aluminum reacts with oxygen, 20.8 g of aluminum oxide is produced. What is the percentage yield of aluminum oxide?
2. When 14.5 g of SO_2 reacts with 21 g of O_2, what will be the theoretical yield and percentage yield of the reaction if the actual yield is 12 g?
3. In the reaction: $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2$, when 52.7 g of octane (C_8H_{18}) burns in oxygen, the percentage yield of carbon dioxide is 82.5%. What is the actual yield in grams?

Checklists for Reference

- Balancing chemical equations
- Chemical reaction
- Chemical equation
- Combination reaction
- Decomposition reaction
- Double displacement reaction
- Excess reactant
- Empirical formula
- Formula mass
- Limiting reactants
- Mass-mass problems
- Mass-volume problems
- Molecular formula
- Mole concept
- Oxidation
- Oxidizing agents
- Percentage yield
- Products
- Percentage composition
- Redox reaction
- Reducing agents
- Reduction
- Reactants
- Single displacement reaction
- Stoichiometry
- Theoretical and actual yield
- Volume-volume problems

Unit Summary

- Chemical equations are representations of chemical reactions using chemical symbols and formulas.
- In any reaction, the total mass of reactants equals the total mass of products according to the law of conservation of mass. Thus, we should always balance chemical equations.
- Chemical reactions can be classified as combination, decomposition, single displacement or double displacement reactions.
- When a substance loses electron (s) or increases its oxidation number in a reaction, the process is oxidation and when a substance gains electron (s) or decreases its oxidation number, the process is reduction.
- Oxidizing agents are substances reduced, and reducing agents are the substances oxidized in reduction – oxidation (redox) reaction.
- Reactions in which neither oxidation nor reduction takes place are non-redox reactions. Neutralization and double displacement reactions are examples of

non-redox reactions.

- ☞ Molecular mass or formula mass expressed in grams is called molar mass. Molar mass is the mass of one mole of a substance.
- ☞ A mole is defined as the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12. 12 g of carbon-12 contains $6.022x10^{23}$ atoms. This number is known as Avogadro's number (NA).
- ☞ Empirical formula (or simplest formula) for a compound is the formula of a substance written with the smallest ratio (whole number ratio) subscripts. The molecular formula is the actual formula that tells you the exact number of atoms of different elements present in a molecule.
- ☞ In chemical reactions, the quantitative relationship between reactants and products is called stoichiometry. There are mainly mass-mass, mass volume and volume-volume stoichiometric relations
- ☞ The reactant completely consumed first in a reaction is called the limiting reactant. The reactants present in quantities greater than necessary to react with the quantity of the limiting reactants are excess reactants
- ☞ In chemical reactions, the actual (experimentally determined) yield of a product is usually less than the theoretical (calculated) yield. The percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100.

Review Exercises

Part I: True-false type questions.

1. In any chemical reaction, each type of atoms is conserved.
2. In a balanced chemical equation, both sides of the equation have the same number of moles.
3. The oxidizing agent is oxidized by the reducing agent.
4. Most metallic elements are strong reducing agents, whereas most non-metallic elements are strong oxidizing agents.
5. The higher the percentage yields of a chemical reaction, the more efficient the reaction.

Part II: Write the missing words for each of the following.

6. The oxidation number of manganese in $KMnO_4$ is ______________ .
7. The maximum yield of product is determined by __________________ reactant.
8. Reactions which do not involve electron transfer are known as ______________ .
9. The sum of masses of all the atoms present in a molecule is ___________________ .
10. The sum of the atomic masses of all atoms present in the formula unit of an ionic compound is ______________ .
11. ____________ is the formula of a substance written with the smallest integer (whole number) subscripts.
12. ____________ is the actual formula that tells you the exact number of atoms of different elements present in a molecule.

Part III: Multiple choice questions.

13. Which of the following contains the largest number of H_2O molecules ?
    - a. 1 mole of H_2O
    - b. 18 g of H_2O
    - c. 6.02×10^{23} molecules of H_2O
    - d. 1 kg of H_2O
14. The formula mass of $Ca_3(PO_4)_2$ is
    - a. 310
    - b. 87
    - c. 200
    - d. 120
    - e. none
15. One mole of two different substances A and B have equal _______
    - a. number of particles
    - b. molar mass
    - c. formula mass
    - d. molecular mass
    - e. percentage composition of the elements

Part IV: Problem - solving questions.

Solve the following problems appropriately.

16. Calculate the molecular masses (in amu) of the following compounds:
    - a. sulphur dioxide (SO_2), a gas that is responsible for acid rain, and
    - b. caffeine ($C_8H_{10}N_4O_2$), a stimulant present in tea, coffee, and cola beverages.
17. How many moles of CO_2 are present in 176 g of CO_2 ?
18. How many grams of oxygen can be prepared by the decomposition of 12 grams of mercury (II) oxide?
19. 25 g of NH_3 is mixed with 4 moles of O_2 in the given reaction:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

    - a. Which is the limiting reactant?
    - b. What mass of NO is formed?
    - c. What mass of H_2O is formed?

20. Consider the following reaction: $2N_2O_5 \rightarrow 4NO_2(g) + O_2(g)$

When 40 g of N_2O_5 decomposes, 4.5 g of O_2 is formed. What is the percent yield?

21. Two compounds have the same composition: 85.62% C and 14.38% H.
    - a. Obtain the empirical formula corresponding to this composition.
    - b. One of the compounds has a molecular mass of 28.03 amu; the other, of 56.06 amu. Obtain the molecular formulas of both compounds.
22. A borane has the empirical formula BH_3 and a molecular mass of 28 amu. What is its molecular formula?
23. The percentage composition of acetic acid is found to be 39.9% C, 6.7% H, and 53.4% O and its molecular mass was determined by experiment to be 60 amu. Calculate the empirical and molecular formula of acetic acid.

24. Balance the following equations using any method :

a. $H_3BO_3 \rightarrow H_4B_6O_{11} + H_2O$

b. $P_2I_4 + P_4 + H_2O \rightarrow PH_4I + H_3PO_4$

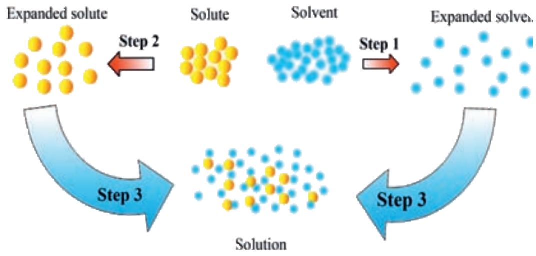
c. $HClO_4 + P_4O_{10} \rightarrow H_3PO_4 + Cl_2O_7$

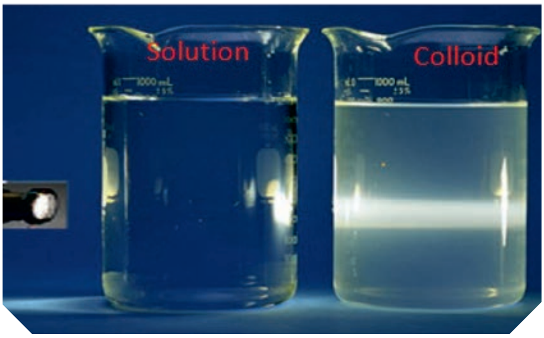
d. $H_2S + HNO_3 \rightarrow S + NO + H_2O$

e. $H_2SO_4 + HBr \rightarrow SO_2 + Br_2 + H_2O$

f. $KMnO_4 + HCl \rightarrow MnCl_2 + Cl_2 + KCl + H_2O$

g. $K + KNO_3 \rightarrow N_2 + K_2O$





UNIT 2

SOLUTIONS

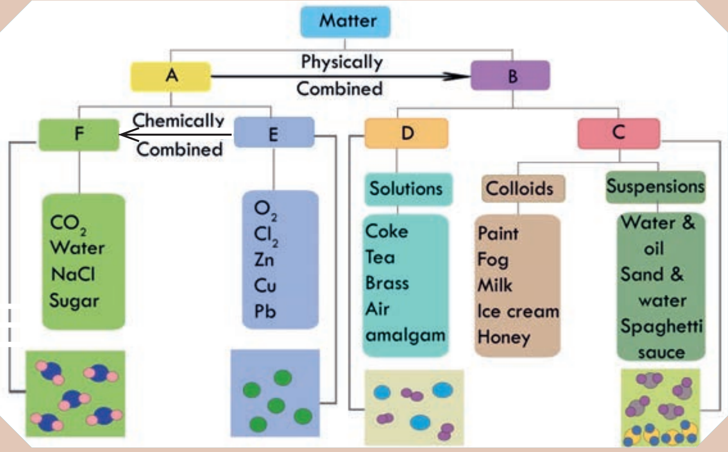
Unit Outcomes

At the end of this unit, you will be able to

- explain the types of solutions;
- describe the solution formation process, the rate of dissolution, the heat of solution and solubility;
- describe the dependence of solubility on temperature & pressure of solution;
- solve problems involving concentration of solutions;
- express the concentration of solutions in various units;
- describe, using the concept of equilibrium, the behavior of ionic solutes in solutions that are unsaturated, saturated and supersaturated;
- prepare solutions of required concentration by dissolving a solute or diluting a concentrated solution;
- demonstrate scientific inquiry skills along this unit: observing, classifying, comparing & contrasting, communicating, measuring, asking questions, drawing conclusion, applying concept and problem solving.

Start-up Activity

1. Consider a patient whose health condition is deteriorated and require a quick recovery.
   a. If you were a physician treating the patient, would you recommend an IV fluid or a capsule?
   b. Discuss the differences among isotonic, hypotonic, and hypertonic solutions. Which solution would you administer to a patient who is suffering from dehydration due to excessive diarrhea? Discuss in a group and present your conclusions to the class.
2. Recall from the previous grades that matter is classified into **pure substances** (elements and compounds) and **mixtures**. To refresh your memory of substances and mixtures from the general sciences, form a group of four students each, discuss and complete the missing information labeled as A, B, C, D, E, and F in the concept map shown below.



Note: Mixtures have the following general properties

- In a mixture each component keeps its original properties.
- The separation of components can be easily done.
- The proportion of the components is variable.
- Mixtures can be converted into component substances by physical process and vice versa. Whereas a substance is a single kind of matter that has definite physical and chemical properties regardless of its source or the way it is made and cannot be separated into other kinds of matter by any physical process. Thus, substances (compounds and elements) are inter-converted through chemical process only.

2.1 Heterogeneous and Homogeneous Mixtures

At the end of this section, students will be able to

- define the terms mixture, homogeneous and heterogeneous mixtures;
- distinguish between homogeneous and heterogeneous mixtures.

Activity 2.1

Form a group, perform the following activity and present your conclusions to the class.

1. Classify each of the mixtures- sand + water, coffee, tea; seawater, air, brass; steel, natural gas, pizza, vinegar, vegetable salad, and fruit punch as
   a. Homogeneous if it has a uniform composition or
   b. Heterogeneous if it is an insoluble mixture of substances and has non-uniform composition.
   c. Where do you categorize blood, milk, butter and cloud: to homo or hetero?

Mixtures are combinations of two or more substances in which the individual substances keep their own properties. Components of mixtures (unlike those of compounds) can have variable proportions, can be separated physically, and retain their properties. Mixtures are part of your daily life. The air you breathe the food you eat, most of the medications you use, the skin and hair products such as body lotion and shampoo, etc. are all mixtures. There are two types of mixtures: homogeneous and heterogeneous mixtures.

A homogeneous mixture is a mixture in which the composition is uniform throughout the mixture. i.e., it has no visible boundaries the components are evenly distributed throughout the entire mixture. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same/constant regardless of the source or the way it is made. NaCl, for instance, is always formed in a ratio of 23 is to 35.5 by mass. But a homogeneous mixture of NaCl and water may be formed by combining them in various proportions; 5 g in 100 g or 10 g in 100 g water; would all give a homogeneous mixture of NaCl and water called a solution. When a spoonful of sugar dissolves in water, the composition of the mixture, after sufficient stirring, is the same throughout the solution. This solution is a homogeneous mixture.

A heterogeneous mixture is a mixture that consists of physically distinct parts, each with different properties. This type of mixtures has at least two visible phases. Mixture of oil and water, wood ash and water, sand and water, etc. are good examples.

2.1.1 Suspensions, Solutions, Colloids

At the end of this section, students will be able to

- describe a suspension;
- define the terms solute, solvent, solution;
- explain the different types of solutions;
- give examples for each types of solutions;
- present a report on how jewelry gold is made to class after a visit to nearby goldsmith;
- describe a colloid.

A. Suspensions

Mixtures can further be classified into suspensions, colloids, and solutions. A suspension is a heterogeneous mixture in which the solid particles are spread throughout the liquid without dissolving in it. A mixture of chalk and water, sand and water, mixture of slaked lime (calcium hydroxide) and water, muddy water are some examples.

B. Solutions

Solutions are all around us. The air we breathe and much of what we drink for example, soda, coffee, and tea, and the fluids in our body are all solutions. A solution is a homogeneous mixture of a solute dissolved in a solvent. Because of dissolution of solute in solvent, the particles in solution are of the size of individual atoms, molecules or ions. The solvent part of the solution is usually a liquid but can be a gas or a solid (see types of solutions in subsequent section). Generally, the major component of the solution is called solvent, and the minor component is called solute. If both components in a solution are 50%, the term solute can be assigned to either component. When a gaseous or solid material dissolves in a liquid, the gas or solid material is called the solute. When two liquids dissolve in each other, the major component is called the solvent and the minor component is called the solute.

Note that all solutions are homogeneous mixtures but not all homogeneous mixtures are solutions.

Properties of Solutions

A solution has a single uniform phase and has no visible boundaries because the components are individual atoms, ions, or molecules. These components do not settle upon standing or separate by filtration.

- The particles of a solution are very small that they can pass through the pores of even fine filter paper. So, solute cannot be separated from the solvent by filtration.
- The solutions are very stable; the components do not separate, neither coagulate upon standing.

Note: Although the components of a solution may not be separated by settling (sedimentation) or filtration the solute can be separated from the solvent by other physical processes. For example, a solution of sodium chloride and water can be separated by distillation or evaporation.

Types of Solutions

Although we usually think of solutions as liquid, they can exist in all three physical states: solid, liquid, and gas.

i. **Gaseous solutions:** Gaseous solutions are usually described as gas-gas solutions where both solute and solvent are gases. The atmosphere is a gaseous solution that consists of nitrogen, oxygen, argon, carbon dioxide, water (in the form of vapor), methane, and some other minor components. The major component (Nitrogen) is regarded as a solvent and the other components are regarded as solute. Other example is natural gas (See unit 6 for details).

ii. **Liquid solutions:** Solids, liquids and gases dissolve in a liquid solvent to form liquid solutions. Some examples include carbonated beverage (gas and liquid), alcoholic beverage (liquid in liquid), and sea water (solid in liquid). Note that in liquid solutions the solute can be a solid, a liquid, or a gas but the solvent is always a liquid.

iii. **Solid solutions:** Many alloys, ceramics, and polymer blends are solid solutions. Solid solutions have no restriction on the state of the solute but the solvent has to be solid. Alloys are solid-solid solutions. Dental filling solution is a good example of liquid (mercury)-solid (silver) solution. Other examples of solid solutions include wax, and H_2 in Pd (gas in solid; used for hydrogen storage).

Project 2.1

Do you know how jewellery gold is made? Form a self-manged group, pay a visit to the nearby goldsmith, and report your experiences to the class.

Exercise 2.1

1. Complete the following table using examples that are not previously mentioned in this textbook.

Solute	Solvent	State of the solution	Example
Gas	Gas	Gas	
Gas	Liquid	Liquid	
Liquid	Liquid	Liquid	
Solid	Liquid	Liquid	
Solid	solid	solid	
Gas	solid	solid	
liquid	solid	solid	Mercury in zinc amalgam

2. Classify each of the following as homogeneous or heterogeneous mixtures and discuss their composition.
 a) glass b) sand c) Cement d) water gas e) producer gas.

C. Colloids

A colloid is a heterogeneous mixture in which insoluble particles of one or more substances are suspended uniformly throughout another substance.

Properties of Colloids

- A colloid (or colloidal solution) appears to be homogeneous but actually, it is heterogeneous upon closer inspection.
- The size of particles in a colloid is bigger than those in a true solution but smaller than those in a suspension. It is between 1 nm and 100 nm in diameter.
- The particles of a colloid can pass through a filter paper. So, a colloid cannot be separated by filtration.
- There are two phases in colloidal solution. They are known as the dispersed phase and the dispersion medium. The component present in smaller proportions is the dispersed phase while the one present in greater proportion is the dispersion medium.
- Colloids are quite stable. The particles do not separate on keeping.
- A colloid (or a colloidal solution) scatters a beam of light passing through it (because its particles are fairly large). As a result, these rays as well as colloidal particles become visible. The phenomenon in which the particle in a colliod scatter the beams of light that are directed at them and make the path of the light beam visible is called Tyndall effect (***Figure 2.1***). This effect is not observed in true solutions. You may expect suspensions to display Tyndall effect. But the Tyndall effect in suspensions disappear as the particles settle down.

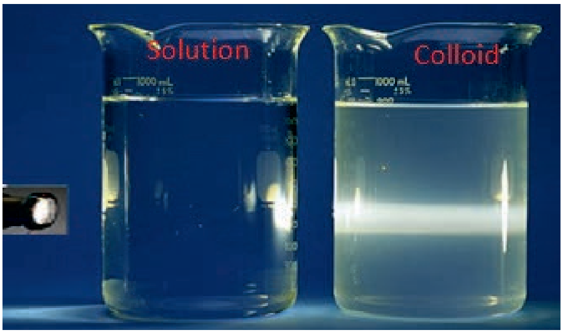


Figure 2.1 The light beam is not visible as it passes through true solution (left) but it is clearly visible when it passes through a colloid (right).

Types of Colloids

Depending upon the state of matter present in the dispersed (internal) and dispersion (external or continuous) medium there are 8 different types of colloidal solutions (*Table 2.1*).

Table 2.1 Types of colloids.

Dispersed Phase	Dispersing Medium	Type	Example
Liquid	Gas	Aerosol	Fog, clouds, mist
Solid	Gas	Aerosol	Smoke, automobile exhaust
Gas	Liquid	Foam	Shaving cream
Liquid	Liquid	Emulsion	Milk, face cream
Solid	Liquid	Sol	Milk of magnesia, mud
Gas	Solid	Foam	Foam, rubber, sponge, pumice
Liquid	Solid	Gel	Jelly, cheese, butter
Solid	Solid	Solid sol	Colored Gemstone, milky glass

Activity 2.2

Form a group and discuss the following phenomenon and present your conclusions to the class.

1. The sky appears blue at day time but red at sunrise and sunset. Why?
2. How is rainbow formed?
3. Is soap a substance or a mixture?

Coagulation

Coagulation is the process by which the dispersed phase of a colloid is made to aggregate and thereby separate from the continuous phase. The curdling of milk when it sours is an example of coagulation.

Project 2.2

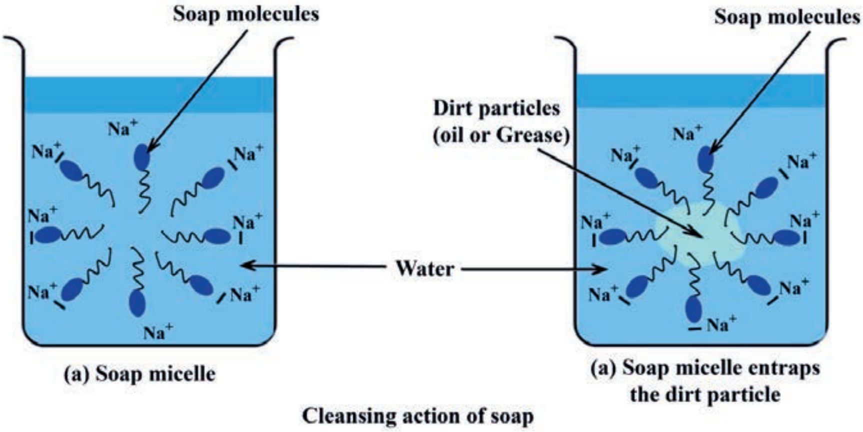
Read on "cloud seeding" technology and understand the mechanism. Then, prepare for a conversation to argue for or against on the intervention of this technology on the natural water cycle to produce artificial rain. You may seek information from Internet. Your teacher will organize the conversation session.

Association of colloids

When molecules or ions that have both a hydrophobic (nonpolar, water hating) and a hydrophilic (polar, water loving) end are dispersed in water, they associate, or aggregate, to form colloidal-sized particles called micelles. The hydrophobic end point inward toward one another and the hydrophilic ends are on the outside of the micelle facing the water molecules. A colloid in which the dispersed phase consists of micelles is called an association colloid. Ordinary soap in water provides an example of an association colloid. Soap consists of compounds such as sodium stearate, $C_{17}H_{35}COONa$. The stearate ion has a long hydrocarbon end that is hydrophobic (because it is nonpolar) and a carboxyl group (COO-) at the other end that is hydrophilic (because it is ionic).

Stearate

In water solution, the stearate ions associate into micelles in which the hydrocarbon end point inward toward one another and away from the water, and ionic carboxyl groups are on the outside of the micelle facing the water (*Figure 2.2*). The cleansing action of soap occurs because oil and grease can be absorbed into the hydrophobic center of the soap micelles and washed away.



(a) Soap micelle

(a) Soap micelle entraps the dirt particle

Cleansing action of soap

Figure 2.2 The cleansing action of soap. The soap molecule is represented by a polar head and zigzag hydrocarbon tail. An oily spot can be removed by soap because the nonpolar tail dissolves in the oil, and the entire system becomes soluble in water because the exterior portion is now ionic..

Project 2.3

Make your own hair gel using locally available materials. You may be assisted with internet. Write all the procedures you followed including the name of the materials, the source, the amount and the processing procedures and steps. You may consider a combination of gelatin or flaxeed with some kind of essential oil. For the essential oil, ethanol ("*areki*") extracts from lavender, rosemary, lomengrass, thyme, or tea tree can be used.

Exercise 2.2

1. Explain the cleansing action of soap and detergents.
2. Discuss the difference between
   - a. suspension and colloid
   - b. solution and colloid
   - c. pure substance and mixtures
   - d. compound and homogeneous mixtures
3. When a corona virus infected patient coughs or sneezes, fine droplets of respiratory fluid containing the viral particles get dispersed in the atmosphere as a colloidal particle called______
   - a. emulsion
   - b. aerosol
   - c. fog
   - d. smoke (choose correct answer)
4. Classify each of the following as suspension, solution, or colloid using a vein diagram.
   - a. Paint
   - b. Yogurt
   - c. toothpaste
   - d. honey
   - e. soda water
   - f. sea water
   - g. blood
   - h. ruby
   - i. urine
   - j. soap suds
   - k. juice
   - l. vinegar
5. Colloidal sulphur particles are negatively charged with thiosulphate ions, $S_2O_3^{2-}$, and other ions on the surface of the sulphur. Indicate which of the following would be most effective in coagulating colloidal sulphur, NaCl, $MgCl_2$, or $AlCl_3$.

2.2 The Solution Process

At the end of this section students will be able to explain how the "like dissolves like" rule depends on inter-particle forces of interactions.

A solution is formed when the solute particles are completely dissolved in the solvent.

We have noted previously that gases mix freely. But the formation of liquid and solid solutions requires overcoming the solute-solute and solvent-solvent inter-particle forces of attraction before the mixing step. Because the solute as well as the solvent particles are held to each other by intermolecular forces in liquid and solid state. The solution process thus involves three basic steps;

i. breaking up the solute-solute inter-particle forces
ii. breaking up the solvent-solvent inter-particles forces
iii. formation of solute-solvent interaction.

Thus, the extent of dissolution of a solute in a solvent is governed by intermolecular forces. The energy change associated with the dissolution process and entropy (the increase in disorder) that result when molecules of the solute and solvent mix to form a solution are the forces driving the solution process.

Note that: the formation of a solution from a solute and a solvent is a physical process, not a chemical change.

2.2.1 Liquid Solutions and Inter-particle Forces of Attractions

When one substance (the solute) dissolves in another (the solvent), particles of the solute disperse throughout the solvent. The solute particles occupy positions that are normally taken by solvent molecules.

The ease with which a solute particle replaces a solvent molecule depends on the relative strengths of inter-particle forces in the three types of interactions (*Figure 2.3*):

- ☞ solute-solute interaction
- ☞ solvent-solvent interaction
- ☞ solute- solvent interaction

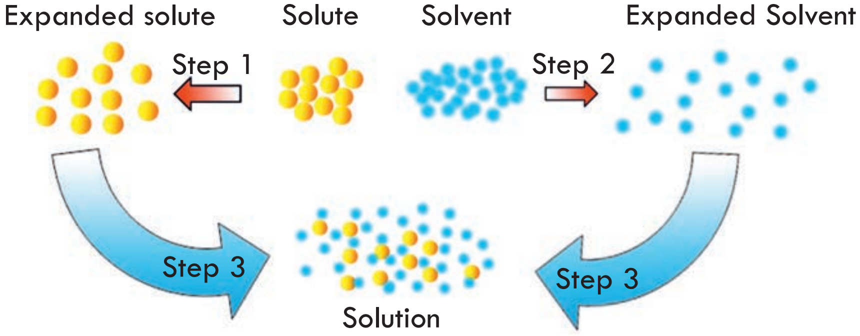


Figure 2.3 A molecular view of the solution process, portrayed as taking place in three steps.

Solutes dissolve in a solvent when the inter-particle forces of interaction between solute molecules and solvent molecules are replaced by solute-solvent interactions. The useful rule-of-thumb "like dissolves like" says that substances with similar types of intermolecular forces dissolve in each other. That means, when the inter-particle forces within the solute are similar to those within the solvent, the forces replace each other and a solution forms. Thus, by knowing the forces, we can often predict whether a solute will dissolve in a solvent.

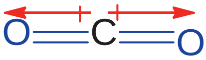
Thus, ionic compounds like NaCl (that are extremely polar) and polar compounds like ethanol (C_2H_5OH) dissolve in polar solvents (e.g. H_2O) and nonpolar compounds and elements (e.g. hexane, CCl_4, I_2, oil) dissolve in nonpolar solvents like hexane (C_6H_{14}), carbon tetrachloride (CCl_4), and benzene.

How can we Predict the Likeness of a Substance?

One useful classification of materials is polarity. As you read about covalent and ionic compounds in Grade 9, you learned that ionic compounds have the highest polarity forming full cations and anions within each formula unit as electrons are transferred from one atom to another. You also learned that a covalent bond is polar when the bonded atoms have different electronegativities and, thus, share the bonding electrons unequally. In diatomic molecules, such as HF, there is only one bond that is polar and the molecule is also polar. In larger molecules, molecular polarity is determined by the net dipole moment (imbalance of charge over the molecule); which in turn depends on both shape and bond polarity. Dipole moment (μ) is a measure of molecular polarity, given in the unit debye (D) derived from SI units of charge (coulomb, C) and length (m): $1 \text{ D} = 3.34 \times 10^{-30}$ C.m.

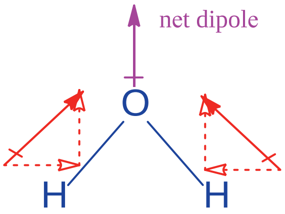
Substances with zero or low electronegativity difference in between the bonding atoms and having zero net dipole moment such as H_2, O_2, N_2, CH_4, CCl_4 are nonpolar whereas H_2O, NH_3, CH_3OH, NO, CO, HCl, H_2S, PH_3, etc. having high electronegativity difference in between the bonding atoms and having net dipole moment are polar compounds. See illustrations below where the dipole moments cancels out in carbon dioxide (CO_2) but add-up to give a net dipole in water molecule (H_2O). Note that the presence of polar bonds does not always result in a polar molecule; we must also consider shape and the atoms surrounding the central atom.

(a) CO_2



Adding these vectors together cancels them out. Therefore, CO_2 is non-polar.

(b) H_2O



The horizontal components of the two dipoles cancels out each other. But the vertical components add-up to give a net dipole moment in H_2O molecules.

Inter-particle/Intermolecular Forces

You studied in grade 9 that atoms in a molecule (e.g. Cl_2, H_2O, NH_3, etc) are held together by covalent bond/s. The covalent bonds are called intramolecular forces-force within a molecule. You are also familiar with the three states of matter-solid, liquid and gas. The gaseous state is characterized by absence of interparticle forces and hence the molecules of a gas are free to move independent of the other molecules.

However, in liquids and solids, molecules are held together by intermolecular forces between molecules. Intermolecular forces are also called noncovalent interactions. They are attractive or repulsive forces between neighboring molecules.

These forces include:

- dipole-dipole interaction
- dipole-induced dipole interaction
- dispersion forces or London forces
- ion-dipole interaction
- ion-induced dipole
- hydrogen bonding

The relative strengths of these forces in order of weakest to strongest is:

Dispersion forces < dipole-induced dipole < dipole-dipole < H - bonding < ion dipole bond

Note: Dipole-dipole, dipole-induced dipole, and dispersion forces make up what chemists commonly refer to as *van der Waals forces*, after the Dutch physicist Johannes van der Waals. Ions and dipoles are attracted to one another by electrostatic forces called ion-dipole forces, that are not van der Waals forces. Hydrogen bonding is a particularly strong type of dipole-dipole interaction. These forces play a central role in the formation of solutions. So let's have a brief look at each of them:

Dipole-Dipole Forces

Dipole–dipole forces occur between polar molecules as a result of electrostatic interactions among dipoles (***Figure 2.4(a)***). e.g. Solubility of HBr in H_2O.

Dipole–induced Dipole Force

Arise when a polar molecule distorts the electron cloud of a nonpolar molecule (***Figure 2.4(b)***). The solubility in water of atmospheric O_2, N_2, and noble gases, while limited, is due in part to these forces. Paint thinners and grease solvents also use them.

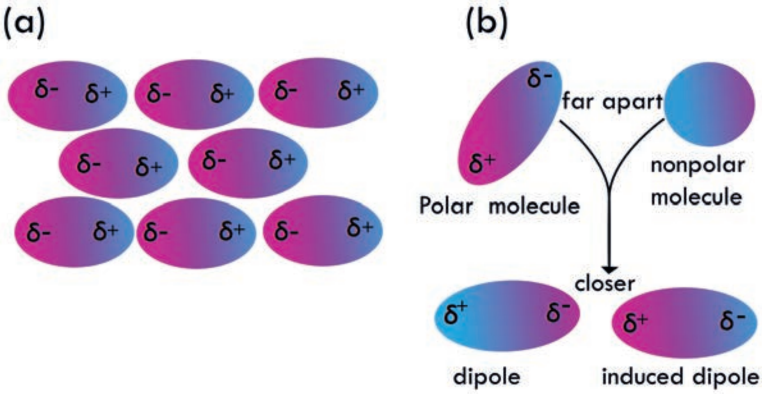


Figure 2.4 Dipole–dipole (a) and dipole-induced dipole (b) forces of attraction between polar and nonpolar molecules.

Dispersion Forces (London forces)

Dispersion forces occur between all neighboring molecules and arise because the electron distribution within molecules is constantly changing. This type of force is also called instantaneous dipole-induced dipole force. It is common to all types of molecules-polar as well as nonpolar. But it is more important to account for properties of nonpolar molecules.

Ion-Dipole Forces

Ion-dipole forces are the principal force involved when an ionic compound dissolves in water (*Figure 2.5*). e.g. NaCl dissolves in H_2O via ion-dipole interaction.

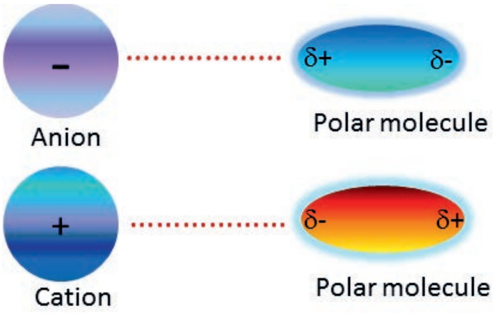


Figure 2.5 Two types of ion-dipole interaction

Ion–induced dipole forces, one type of charge-induced dipole force, rely on polarizability. They arise when an ion's charge distorts the electron cloud of a nearby nonpolar molecule (**Figure 2.6**). This type of force initiates the binding of the Fe^{2+} ion in hemoglobin to an O_2 molecule entering a red blood cell.

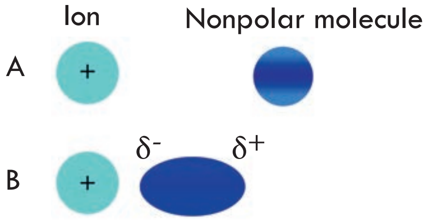


Figure 2.6 Ion-induced dipole interactions. (A) when the ion and the nonpolar molecules are far apart, (B) when the ion and the nonpolar molecule get closer-the nonpolar molecule gets polarized.

The Hydrogen Bond

Definition: The hydrogen bond is an attractive interaction between a hydrogen bonded to an electronegative O, N, or F atom in one molecule and an unshared electron pair on O, N, or F atoms of another molecule, *Figure 2.7*.

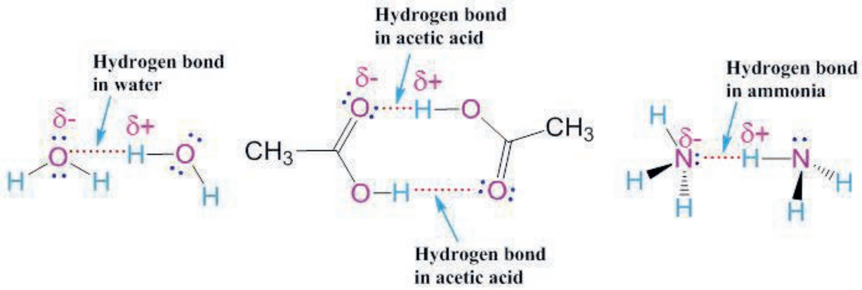
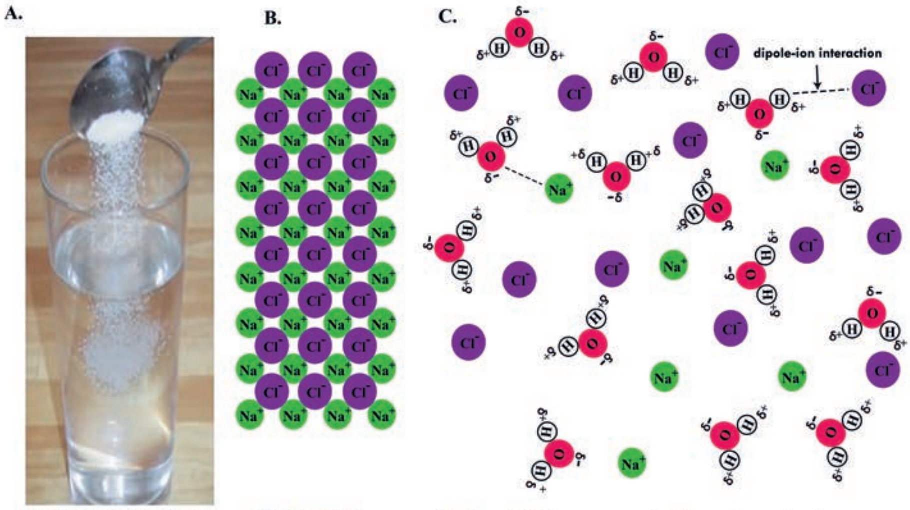


Figure 2.7 Hydrogen bonding in water, acetic acid, and ammonia. Solid lines represent covalent bonds, and dotted lines represent hydrogen bonds.

2.2.2 Solutions of Ionic Solids in Water

Many ionic compounds are soluble (at least 1 g dissolves in 100 mL water) in water; however, not all ionic compounds are soluble. Chemists refer to substances as soluble, slightly soluble, or insoluble in a qualitative sense. For instance, KCl is readily soluble in water whereas AgCl is insoluble. $Ca(OH)_2$ is slightly soluble. Ionic compounds that are soluble in water exist in their ionic state as hydrated cations and anions within the solution (See *Figure 2.8*).



NaCl and water mixture NaCl crystal Na^+ and Cl^- ions sorrounded by water molecules

Figure 2.8 Hydration of Na^+ and $Cl^–$ ions as NaCl dissolves in water.

2.2.3 The Solubility Rules for Ionic Solids

At the end of this section students will be able to predict relative solubilities.

Solubility is defined as the maximum amount of a substance that can be dissolved in a given volume (usually 100 mL) of solvent at a given temperature. The solubility rules for common ionic compounds are listed below.

i. All common salts of the Group 1A elements (such as Na^+, K^+, Li^+) and ammonium, NH_4^+, are soluble

ii. All common salts containing acetate (CH_3COO^-), or nitrate (NO_3^-), and most perchlorates (ClO_4^-) are soluble

iii. All common chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are soluble, except those of Ag^+, Pb^{2+}, Cu^{2+}, and Hg_2^{2+}. All common fluorides (F^-) are soluble, except those of Pb^{2+} and Group 2A.

iv. Most hydroxide salts are only slightly soluble. Hydroxides of GI elements are soluble. Hydroxides of GII elements (Ca, Sr, and Ba) are slightly soluble. Hydroxide salts of transition metals and Al^{3+} are insoluble. Thus, $Fe(OH)_3$, $Al(OH)_3$, are not soluble.

v. All compounds containing sulfate (SO_4^{2-}) are soluble, except those of barium (Ba), strontium (Sr), lead (Pb), calcium (Ca), silver (Ag), and mercury (Hg)

vi. Except for those compounds following rule 1, compounds containing carbonate (CO_3^{2-}), sulfides (S^{2-}), oxides (O^{2-}), and phosphates (PO_4^{3-}) are insoluble.

Activity 2.3

Form a group, discuss the following activity and present your conclusions to the class. Fluoride (F^-) has a significant alleviating effect against dental caries if the concentration is approximately 1 mg/l. However, continuing consumption of higher concentrations (>1.5 mg/l) can cause dental fluorosis and in extreme cases even skeletal fluorosis.

One of the methods used for the removal of excess fluoride from drinking water is precipitation by employing lime [quicklime (CaO) or slacked lime ($Ca(OH)_2$)] followed by sedimentation and/or filtration. Discuss how lime removes fluorides from water.

Dental fluorosis

Skeletal fluorosis

The dissolution of molecular solids also follows the same rule. Molecular solids are made up of atoms or molecules held together by London or dispersion forces, dipole-dipole forces, or hydrogen bonds. Examples of molecular solids include ice, solid carbon dioxide, sucrose, sulfur, solid hydrocarbons (e.g. octadecane, $C_{18}H_{38}$), etc. Polar molecular solids such as ice, glucose, and sucrose dissolve in polar solvents like water. Nonpolar molecular solids such as sulfur, solid hydrocarbons, etc dissolve in nonpolar solvents like benzene, hexane, etc.

2.2.4 Solution of Liquids in Liquids

In addition to the dissolution of ionic compounds in water, solutions can also be made by mixing two compatible liquids. The liquid in the lower concentration is termed as solute, and the one in higher concentration is the solvent. For example, grain alcohol (CH_3CH_2OH) is a polar covalent molecule that can mix with water. When two similar liquids (in terms of polarity) are placed together and are able to mix into a solution, they are said to be miscible. Liquids that do not share similar characteristics and cannot mix together, on the other hand, are termed immiscible. For example, oil and water are immiscible because oil is nonpolar whereas water is polar.

Exercise 2.3

1. Which of the following exists predominantly in the water (H_2O) molecule?
   a. van der Waal's force
   b. ion-dipole force
   c. hydrogen bond
   d. none
2. Which of the following has the highest induced dipole interactions in its molecules?
   a. He
   b. Ar
   c. Ne
   d. Kr
3. Oxygen (O) and sulphur (S) are in the same group in the periodic table. They form compounds with hydrogen, H_2O and H_2S. However, H_2O is a liquid whereas H_2S is a gas at a room temperature. Explain.
4. Classify each of the following as soluble or insoluble in water:
   a. $BaCO_3$
   b. $AlCl_3$
   c. MgO
   d. $Al(OH)_3$
   e. KI
   f. $Al_2(SO_4)_3$
   g. $Ca_3(PO_4)_2$
   h. $PbCl_2$
   i. $AgNO_3$
   j. $(NH_4)NO_3$
   k. NH_4Cl
5. Classify each of the following pairs as miscible or immiscible:
   a. Water and grain alcohol
   b. “Tej” and water
   c. Water and oil
   d. Water and honey
   e. Water and petrol
   f. Gasoline and "naphtha"
6. What type (s) of intermolecular (interparticle) forces exist when the following pairs interact:
   a. HBr and H_2S b. Cl_2 and CBr_4
   c. I_2 and NO_3^-, d. NH_3 and C_6H_6(Benzene)

2.2.5 The Rate of Dissolution

At the end of this section students will be able to define rate of dissolution

Dissolution is defined as the process through which a solute dissolves in a solvent (e.g. water) to produce a solution. How fast does a solute dissolve in a given solvent? What are the major factors that affect the rate of dissolution? The rate of dissolution is the speed with which a solute dissolves in a solvent to form a solution.. This largely depends upon two factors-the inter-particle forces discussed so far and, to a lesser extent on conditions such as the surface area of the solid solute, and the temperature and the pressure (for gaseous solutes) of the system. When the solvent-solute interactions are stronger than those between solute-solute and solvent-solvent particles, the dissolution process becomes faster. A good example is dissolution of ethanol (alcohol) in water. The increasing surface area of the solute will increase the rate of dissolution because it increases the number of solute particles in contact with the solvent. Stirring or agitation of the mixture can also increase rate of dissolution.

2.2.6 Energy Changes in Solution Process

At the end of this section students will be able to

- define heat of solution, solvation energy and hydration energy;
- apply the concept of heat of solution to the solution of ammonium nitrate crystal;
- apply the concept of heat of solution to the solution of sodium hydroxide crystal;
- explain how heat of solution is influenced by the inter particle interaction forces.

For instance, if you mix ammonium chloride in water in a flask, the flask gets cold. But when sodium hydroxide (NaOH) is mixed in water, the flask gets hot. Why? The reason is attributed to the energy changes that accompany the dissolution process. As discussed earlier, the solution process involves three steps: breaking solute-solute interaction, breaking solvent-solvent interaction, and forming solute-solvent interactions or the mixing step. When the energy released during the third step in solution process, i.e. the mixing step is larger than the energy required to break the solute-solute and solvent-solvent forces, energy is released-the process is exothermic. This energy heats up the surrounding and that is why the flask gets hot. On the contrary, if the energy released in the mixing step is less than the energy required to break the solute-solute and solvent-solvent forces, the dissolution process absorbs energy from the surrounding and the surrounding (the flask) gets cold-the dissolution process is endothermic. The energy involved in the mixing step is called solvation energy. Solvation refers to the process of attraction and association of molecules of a solute and a solvent. Solvation energy is the energy released when solute molecules are solvated by solvent molecules. Solvation in water is called hydration. Thus, hydration energy is the energy released when solute molecules or ions are solvated by water molecules.

The heat of solution or enthalpy of solution is defined as the amount of heat released or absorbed during the dissolution process and can be calculated using the equation

$$q = m \times Cg \times \Delta T \qquad \text{Eq 2.1}$$

Where q is the amount of energy released or absorbed in Joules (J), m is mass of the solution, ΔT is the temperature change (T_i-T_f), Cg is the specific heat of solution (J/g.°C) (can be assumed to be the same as that of water, 4.184 J/g.°C for dilute solutions). The unit of heat of solution is Joule or kJ. Heat (Enthalpy) of solution can either be positive (endothermic) or negative (exothermic) depending on the observed temperature change.

The molar heat of solution (ΔH_{soln}) of a substance is the heat absorbed or released when one mole of the substance is dissolved in water at constant pressure:

$$\Delta H_{solution} = \frac{q}{\text{the number of moles of solute}} \qquad \text{Eq 2.2}$$

The enthalpy changes are expressed in kJ/mol for a reaction taking place at standard conditions (a temperature of 298.15 K and a pressure 1 atm).

Example 2.1: When 5.19 g of $NaCO_3$ was dissolved in 75.0 g of water, the temperature of the water rose by 3.8°C. Calculate the heat of solution and the molar heat of solution. Take c of the solution as 3.820 J/g.°C and molar mass of $NaCO_3$ as 105.99 g/mol.

Given	**Required**	**Solution**
Mass solution = mass solute + mass solvent = 5.19 g + 75.0 g = 80.19 g An increase of temperature means Tf > Ti. Thus, ΔT = Ti – Tf = -3.80°C Number of moles of solute = given mass divided by molar mass = 5.19 g/105.99 g/mol = 0.0489 mole	q and H_{sol}	q = mcΔT = 80.19 g x 3.820 J/g.°C x (-3.80°C) = -1164.0 Joules ΔH_{sol} = q/n =-1164.0 J/0.0489 moles = **-23,803.70 J/mole** ⇒Exothermic

To understand the enthalpy changes involved during the dissolution, let's consider the hypothetical three-step process happening between two substances: solute and solvent. The three processes are solute-solute, solvent-solvent, and solute-solvent interactions each accompanied by an enthalpy change:

Step 1: Solute particles separate from each other. This step involves overcoming intermolecular attractions, so it is endothermic.

Solute (aggregated) + Heat → Solute (separated) $\Delta H_{solute} > 0$

Step 2: Solvent particles separate from each other. This step also involves overcoming attractions, so it is endothermic, too.

Solvent (aggregated) + Heat → Solvent (separated) $\Delta H_{solvent} > 0$

Step 3: Solute and solvent particles mix and form a solution. The different particles attract each other and come together, so this step is exothermic:

Solute (separated) + Solvent (separated) → Solution + Heat $\Delta H_{mix} < 0$

We combine the three individual enthalpy changes to find the enthalpy of solution

(ΔH_{soln}), the total enthalpy change that occurs when solute and solvent form a solution:

$$\Delta H_{soln} = \Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mix}$$

Endothermic and Exothermic Dissolution Processes

Depending on the relative magnitude of $\Delta H_{solute} + \Delta H_{solvent}$ & ΔH_{mix}, the ΔH_{soln} can be zero (ideal solution), negative (exothermic) or positive (endothermic).

Ideal Solutions ($\Delta H_{soln} = \Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mix} = 0$)

When the strengths of the intermolecular forces of attraction in solute-solute and solvent-solvent are similar to solute-solvent the solution is formed with no accompanying energy change. Such a solution is called an ideal solution. A mixture of ideal gases (or gases such as helium and argon, which closely approach ideal behavior) is an example of an ideal solution, since the entities comprising these gases experience no significant intermolecular attractions. Ideal solutions may also form when structurally similar liquids are mixed. For example, mixtures of the alcohols - methanol (CH_3OH) and ethanol (C_2H_5OH) form ideal solutions, as do mixtures of the hydrocarbons pentane, C_5H_{12}, and hexane, C_6H_{14}.

Enthalpy of Hydration (Ionic Solids in Water)

Therefore, $H_{soln} = H_{solute} + H_{solvent} + H_{mix} = H_{lat} + H_{hydr}$
Where ΔH_{lat}, Lattice energy, is the energy required to break up the ions apart from crystal lattice. ΔH_{solv}, solvation energy, is the energy released or absorbed when solute particles are completely surrounded by solvent molecules. When water is used as the solvent, we use the term hydration energy, ΔH_{hyd}, rather than the more general term solvation energy. Hydration is usually exothermic, so heat is released when water molecules completely surround solute particles.

How does Lattice Energy affect the Solubility of Ionic Solids?

Breaking up the lattice is an endothermic process. Hydration of ions favors the dissolution of an ionic solid in water. Thus, Lattice energy works against the solution process (*Figure 2.9*), so an ionic solid with relatively large lattice energy is usually insoluble.

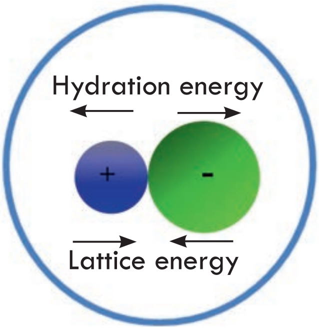


Figure 2.9 The effects of lattice energy and hydration energy on the solution process of ionic solids.

Lattice energies depend on the charge on the ions and also the distance between the centers of the neighboring positive and negative ions. As the magnitude of the charge on the ions increases the lattice energy also increases. For this reason, you can expect substances with single charged ions to be more comparatively soluble, and those with multiple charged ions to be less soluble in water. Lattice energy is inversely proportional to the inter-particle distance between the centers of the two ions. Thus, CsCl has less lattice energy than NaCl.

Examples of Endothermic and Exothermic Processes

When sodium hydroxide is dissolved in water, the solution becomes hot (the solution process is exothermic). On the other hand, when ammonium nitrate is dissolved in water, the solution becomes very cold (the solution process is endothermic). This cooling effect from the dissolving of ammonium nitrate in water is exploited in instant cold packs used in hospitals. An instant cold pack consists of a bag of NH_4NO_3 crystals inside a bag of water (*Figure 2.10*). When the inner bag is broken, NH_4NO_3 dissolves in the water. Heat is absorbed, so the bag feels cold. Hot packs, by contrast, containing either $CaCl_2$ or $MgSO_4$, produce heat when the salts dissolve in water (*Figure 2.10*).

Figure 2.10 The instant cold compress and hot compresses

Note: Enthalpy of solution is only one part of the driving force in the formation of solutions; the other part is the entropy of solution. You will learn about entropy at higher levels.

Enthalpy (heat) of solution can be determined in the laboratory by measuring the temperature change of the solvent when solute is added. A calorimeter is a device used to measure the amount of heat involved in a chemical or physical process.

Activity 2.4

Discuss the following questions and present finding to the class

1. Discuss how people traditionally alleviate inflammation (tenderness), and pains caused by muscle or joint damage in relation to the medical uses of cold packs and hot packs.
2. The process of dissolution is always accompanied by energy change. (True/False).

Exercise 2.4

1. Which one of the following pairs of ions has greater enthalpy of hydration?
   - i. Na^+ or Cs^+?
   - ii. Mg^{2+}or Cs^+?
   - iii. F^- or Cl^-? Explain.
2. A student added 4.00 g of NaOH(s) to 100 g of water in a polystyrene foam cup. The temperature of the water rose by 10.0°C. Assuming the polystyrene foam cup is well insulated and the specific heat capacity of water is 4.18 J/°C.g, determine the molar enthalpy of solution of sodium hydroxide in kJ mol^{-1}. Answer: ΔH_{soln} = -43.5 kJ mol^{-1}, verify this!
3. The molar heat of solution, ΔH_{soln}, of NaOH is -445.1 kJ/mol. In a certain experiment, 5.00 g of NaOH is completely dissolved in 1.000 L of water at 20.0°C in a foam cup calorimeter. Assuming no heat loss, calculate the final temperature of the water. Answer: T_f = 33.2°C, verify this!

Experiment 2.1

Investigation of Heat of Solution

Objective: To investigate the heat of solution of calcium chloride and ammonium nitrate.

Apparatus: Erlenmeyer flask, thermometer, cork.

Chemicals: Calcium chloride, ammonium nitrate and water.

Procedure:

1. Weigh out accurately about 4 g of calcium chloride in a weighing boat (or weighing paper) and record the mass on your results sheet.
2. Using graduated cylinder, accurately measure 40 mL of distilled water; record your measurement to one decimal place on your results sheet.
3. Pour the water into an Erlenmeyer flask and stopper it with a cork as shown in *Figure 2.11*.

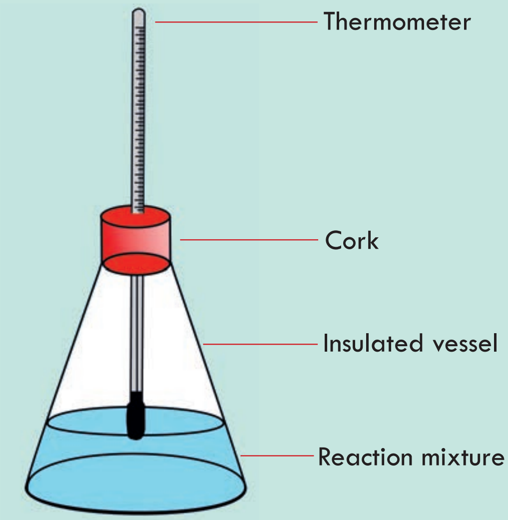


Figure 2.11 Apparatus to measure the heat of solution.

4. Insert a thermometer through a hole in the cork and record initial temperature, again recording your results to one decimal place in the table below.
5. Remove the cork along with the thermometer and quickly transfer the weight calcium chloride and replace the cork; not the thermometer.
6. Insert the stirrer and stir to dissolve the salt.
7. Insert the thermometer and record the final temperature of the solution (highest or lowest temperature attained) in the table

Chemical	Initial temperature of the solvent	Maximum temperature of the solution	Heat of solution
$CaCl_2$			
NH_4NO_3			

8. Repeat the procedure using ammonium nitrate.
9. Calculate the total amount of heat in Joules (J) absorbed or liberated.
10. Calculate the amount of heat in Joules (J) absorbed or liberated per mole of salt and determine whether the process is endothermic or exothermic.

Insulate the vessel, if you can, or assume the heat exchange with surrounding negligible. Assume that the specific heat of the solution is the same as the specific heat of water (4.184 J/g.°C) and molar mass of calcium chloride 110.98 g/mol.

Project 2.4

Your teacher will arrange a visit to nearby hospital or other health center. Your role is to contact the professionals on the usage of instant cold packs and instant hot packs. Ask them to get one each, discuss the usage, working principle, and the composition. Write the dissolution equation in each case. Report your answers to the class

2.3 Solubility as an Equilibrium Process

At the end of this section students will be able to

- define solubility;
- describe the distinctions among unsaturated, saturated and supersaturated solutions;
- prepare unsaturated and saturated solutions of sodium sulphate;
- prepare supersaturated solution of sodium thiosulphate;
- explain the equilibrium nature of saturated solution.

We have already defined solubility as the maximum amount of a substance that can be dissolved in a given volume (usually 100 mL) of solvent at a given temperature. When an ionic solid dissolves, ions leave the solid and enter the solvent. Suppose that you have a beaker of water to which you add some salt (*Figure 2.12*) and stir the mixture until the NaCl dissolves. You keep adding more and more salt and continue stirring. Eventually, you reach a point where no more of the salt will dissolve no matter how long or how vigorously you stir it. Why? At the molecular level, we know that action of the water causes the individual ions to break apart from the salt crystal and enter the solution, where they remain hydrated by the water molecules. What also happens is that some of the dissolved ions collide back again with the undissolved solute and crystallize. Recrystallization is the process of dissolved solute returning to the solid state. At some point the rate at which the solid salt is dissolving becomes equal to the rate at which the dissolved solute is recrystallizing. When that point is reached, the total amount of dissolved salt remains unchanged-equilibrium is established. Solution equilibrium is the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate.

$$\text{Solid (solute)} \underset{\text{Crystallization}}{\overset{\text{Dissolution}}{\rightleftharpoons}} \text{Dissolved ions}$$

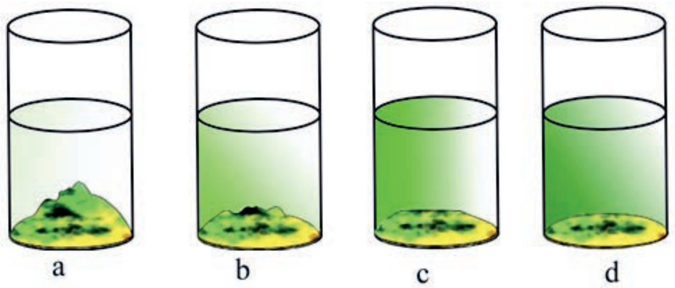


Figure 2.12 Illustrates the formation of saturated solution. (a) excess solid added to a solvent (b) The solute starts to dissolve (c) The solute continues to dissolve (d) No more solute is dissolving indicating that the process has reached a dynamic equilibrium where the rate of dissolution is equal to the rate of crystallization.

2.3.1 Unsaturated, Saturated, and Supersaturated Solutions

A saturated solution is a solution that is at equilibrium and contains the maximum amount of dissolved solute at a given temperature in the presence of undissolved solute. Therefore, if you filter off the solution and add more solute, it doesn't dissolve. A simple analogy is "when you have eaten your feel, you are saturated!"

Unsaturated solution is the solution that has not reached its maximum solubility. This means that more solute could still be added to the solution and dissolving would still occur.

Supersaturated solutions are solutions that have dissolved solute beyond the normal saturation point. Usually a condition such as increased temperature or pressure is required to create a supersaturated solution. For example, sodium acetate has a very high solubility at 273 K. When cooled, such a solution stays dissolved in what is called a meta-stable state. However, when a seeding crystal is added to the solution, the extra solute will rapidly solidify, leaving a saturated solution (*Figure 2.13*). During the crystallization process, heat is evolved, and the solution becomes warm. Common hand warmers use this chemical process to generate heat.


(a) (b) (c)

Figure 2.13 (a) a hot supersaturated solution (formed at elevated temperature) (b) the scenario when the supersaturated solution is cooled (c) cold supersaturated solution after seeding.

Experiment 2.2

Preparation of Unsaturated, Saturated and Supersaturate Solutions

Objectives: To prepare unsaturated, saturated and supersaturate solutions

Materials required: 100 g of $Na_2S_2O_3$, sodium thiosulfate, beaker, balance, filter paper, spatula, stirring rod

Procedure:

1. Pour 100 mL water in a beaker.
2. Add some crystal (about 5 g) of $Na_2S_2O_3$ into the water using spatula and stir until it dissolves.
   a) What do you call this type of solution?
3. Continue adding more and more $Na_2S_2O_3$ while stirring to dissolve.
   b) What do you observe after addition of large (about 70 g) amount of solute?
   c) Why does the excess solute remains undissolved?
4. Filter the undissolved solute. Collect the filtrate or the solution.
   d) What is the name of such a solution?
5. Add spatula-full more solute to the filtrate while heating the mixture and stir.
   e) Does the additional solute dissolve in absence of heating? What about while heating?
   f) What is the name of such a solution?
   Prepare your report and discuss your findings with rest of the class.

Activity 2.5

Discuss the following questions in group and present your conclusions to the class.

1. Discuss the differences among saturated, unsaturated, and supersaturated solutions in relation to your life experiences.
2. Consider three beakers A, B, and C containing 100 mL each of unsaturated, saturated, and supersaturated solutions of a salt. What would happen if a few crystals of the salt is added to each? Which solution would be able to dissolve, which not? In which case would you expect even more than added amount of solid to precipitate?

2.3.2 Factors Affecting Solubility of Substances

At the end of this section students will be able to

- describe the factors that affect solubility of substances;
- investigate the effect of temperature on solubility of sodium sulphate;
- conduct an experiment to determine solubility of table salt and sugar.

Solubility of a substance in a given solvent depends on

- inter-particle forces,
- temperature, and
- pressure (for gases only).

A. Effect of Temperature on Solubility of Solids

You know that more sugar dissolves in hot tea than in cold tea. Generally, the solubility of most of solid solutes increases with temperature. For instance, the solubility of sodium sulfate, Na_2SO_4, rises more than tenfold when the temperature rises from 0°C to 32.4°C, where it reaches a maximum of 49.7 g Na_2SO_4 per 100 g water. The solubility of NaCl is 36 g/100 mL of water at 25°C and 39 g/100 mL of water at 100°C.

Experiment 2.3

Determination of the Solubility of NaCl

Objective: To determine the solubility of NaCl

Apparatus: Beaker, evaporating dish, measuring cylinder, glass rod, filter paper, analytical balance, and Bunsen burner.

Chemicals: Sodium chloride and water.

Procedure:

1. Take an evaporating dish and weigh it.
2. Take 100 mL of water in a beaker and add about 20 g of sodium chloride to it. Stir the solution vigorously with a glass rod, add more sodium chloride while stirring and continue this process until undissolved sodium chloride is left in the beaker.
3. Take 50 mL of the supernatant saturated solution (assume that 50 mL of solution is equal to 50 mL of the solvent) and transfer it to an evaporating dish.
4. Heat the solution in the evaporating dish as shown in the Figure, till all the water has evaporated and dry sodium chloride is left in the evaporating dish.
5. Cool the evaporating dish containing dry sodium chloride to room temperature and weigh it again.

Observations and analysis:

1. Volume of the NaCl solution ….
2. Weight of the empty evaporating dish ….
3. Weight of the evaporating dish + NaCl collected after evaporating the solvent….
4. Calculate the solubility of NaCl in water and express the results in grams of NaCl/100 g of water (assume that 100 mL of water = 100 g of water).
5. Repeat the experiment with sugar.

B. Effect of Temperature on Solubility of Gases

The solubility of gases in liquids decreases with increasing temperature. The molecules in the gaseous state are so far apart because attractive intermolecular interactions in the gas phase are essentially zero for most substances. When a gas dissolves, it does so because its molecules interact with solvent molecules. Heat is released when these new attractive forces form. Thus, if external heat is added to the system, it breaks the attractive forces between the gas and the solvent molecules and decreases the solubility of the gas.

Activity 2.6

Perform the following activities in group and present your conclusions to the class.

(Effect of climate change and pollution on environment)

a. Discuss the effect of climate change (rise in temperature of the environment) on aquatic life.

b. As nations shift from agriculture-led to industry-led economy, environmental pollution becomes a serious concern. Water bodies, for instance, can be polluted by chemicals released to the environment by industries along with wastes. Discuss how pollution of water bodies by chemicals can affect aquatic life such as fish and its productivity. From this lesson, what precautions would you make when disposing wastes of personal hygiene materials such as sanitary pads (Modess) or hair gel containers and other plastic materials?

C. Effect of Pressure on Solubility of Gases: Henry's Law

At the end of this section students will be able to

- state Henry's law
- use Henry's law to calculate concentration of gaseous solute in a solution.

Activity 2.7

1. Do you know why (a) Carbonated beverages such as coke, Pepsi, fizzes (sparkles, Figure to the right) when the cap is removed (b) Multicellular organisms need hemoglobin in red blood cells to bind and carry O_2 to support the energy needs. Why this mechanism of replenishing O_2 in biological fluid is necessary? What would happen if hemoglobin is absent?

Pressure has virtually no effect on the solubility of solids or liquids in liquid solvents because they are incompressible. In contrast, the solubility of gases increases as the partial pressure of the gas above a solution increases (Henry's Law). As you can see in *Figure 2.14*, more and more gaseous particles enter the solution when the pressure is increased from P1 to P_2 and then P_3.

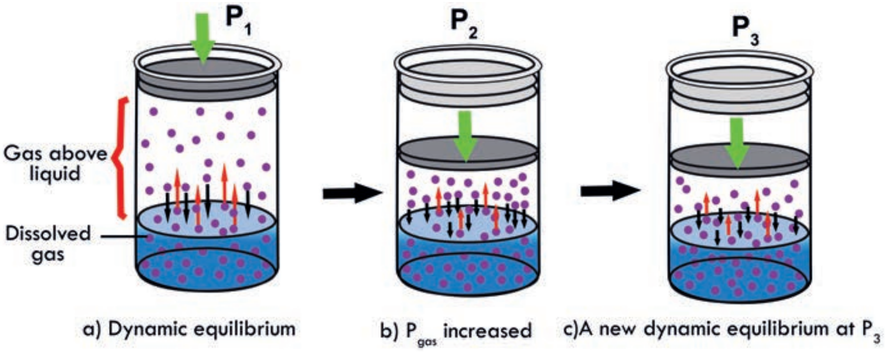


Figure 2.14 A model depicting why the solubility of a gas increases as the partial pressure increases at constant temperature.

Henry's Law: Relating Pressure to the Solubility of a Gas in a Liquid

The quantitative relationship between pressure and the solubility of a gas is described quantitatively by Henry's law, which is named for its discoverer, the English chemist, William Henry (1775 – 1836):

$$C \propto P \quad \text{or} \quad C = kP$$

where C is the concentration of dissolved gas at equilibrium, P is the partial pressure of the gas, and k is the Henry's law constant, which must be determined experimentally for each combination of gas, solvent, and temperature. Although the gas concentration may be expressed in any convenient units, we will use molarity here. The units of the Henry's law constant are therefore mol/(L·atm) = M/atm.

The partial pressure of a gas can be expressed as concentration by writing Henry's Law as

$$P = \frac{C}{k} \qquad \text{Eg 2.3}$$

Since partial pressure and concentration are directly proportional, if the partial pressure of a gas changes while the temperature remains constant, the new concentration of the gas within the liquid can be easily calculated using the following equation:

$$\frac{C_1}{P_1} = \frac{C_2}{P_2} \qquad \text{Eg 2.4}$$

Where C_1 and P_1 are the concentration and partial pressure, respectively, of the gas at the initial condition, and C_2 and P_2 are the concentration and partial pressure, respectively, of the gas at the final condition.

Note: Gases that form strong intermolecular forces such as H-bonding or react chemically with water, such as HCl and the other hydrogen halides, H_2S, and NH_3, do not obey Henry's law; all of these gases are much more soluble than predicted by Henry's law.

Example 2.2

The partial pressure of carbon dioxide gas inside a bottle of cola is 4 atm at 25°C. What is the solubility of CO_2? The Henry's law constant for CO_2 in water is 3.3 x 10^{-2} mol/L.atm at 25°C.

Given	Required	Solution
P_{CO_2} = 4 atm k_H = 3.3x10^{-2} mol/L.atm Temperature = 293 K = constant	Solubility of carbon dioxide	The relation between solubility of a gas and its partial pressure is given by Henry's law: $C = k_H \times P$ Substituting the values of k_H and P in the above equation gives: $\Rightarrow$ C = 3.3 x 10^{-2}(mol/(L.atm)) x 4 atm = 0.132 mol/L = **0.1mol/L;** when rounded to one significant figure.

Example 2.3

The solubility of nitrogen gas at 25°C and 1 atm is 6.8x10^{-4} mol/L. What is the concentration of nitrogen dissolved in water under atmospheric conditions? The partial pressure of nitrogen gas in the atmosphere is 0.78 atm.

Given	Required	Solution
P_1 = 1 atm C_1 = 6.8 x 10^{-4} mol/L P_2 = 0.78 atm **Note:** atmospheric condition means the pressure is 1 atm. Since air is 78% N_2, 78% of 1 atm is exerted by N_2. 78% x 1 atm = 0.78 atm Therefore, the partial pressure of N_2 under atmospheric condition is 0.78 atm	C_2	$\Rightarrow C_1/P_1 = C_2/P_2$ $\Rightarrow P_2 \times C_1/P_1 = C_2$ $C_2 = \dfrac{0.78 \text{ atm} \times \dfrac{6.8\times10^{-4} \text{ mol}}{\text{L}}}{(1 \text{ atm})}$ **C_2=5.3 x10^{-4} mol/L**

Note that the decrease in solubility is the result of lowering the pressure from 1 atm to 0.78 atm.

Exercise 2.5

1. Why are carbonated beverages such as *Coca Cola* are packed under pressure?
2. The concentration of CO_2 in a solution is 0.032 M at 3.0 atm. What is the concentration of CO_2 at 5.0 atm of pressure?
3. Which of the following gases has the greatest Henry's law constant in water at 25°C?
   a. CH_4
   b. Ne
   c. NH_3
   d. H_2

2.4 Ways of Expressing Concentration of Solutions

At the end of this section students will be able to

- define concentration of a solution;
- define mass percentage, ppm and ppb of a solute in a solution;
- calculate the mass percentage, ppm and ppb of a solute in a solution from a given information.

Solution Concentration

In chemistry, concentration is defined as the relative quantity of a solute against a total quantity of solution or solvent. There are different ways of expressing concentration of solutions, e.g. percent by mass or volume, mole fraction, molarity, molality, and normality.

2.4.1 Percent by Mass/Volume

The percent by mass (also called the percent by weight or the weight percent) is defined as

$$\text{Percent by mass of solute} = \frac{\text{mass of solute}}{\text{mass of solute + mass of solvent}} \times 100\%$$

$$= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \qquad \text{Eq 2.5}$$

The percent by mass is a unitless number because it is a ratio of two similar quantities.

Example 2.4: A solution is made by dissolving 13.5 g of glucose, $C_6H_{12}O_6$, in 0.100 kg of water. What is the mass percentage of solute in this solution?

Given	Required	Solution
Mass of solute (glucose) = 13.5 g Mass of solvent = 0.100 kg =100 g	%(w/w)	$\%(w/w)_{solute}$ $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$ But $\text{Mass}_{solution} = \text{mass}_{solute} + \text{mass}_{solvent}$ Substituting the values of masses of glucose (solute) and solution in the above equation gives: $\%(w/w)_{solute} = (13.5\ g/113.5\ g) \times 100\%$ $= \mathbf{11.9\%}$

Parts by Volume: The most common parts-by-volume term is volume percent, % (v/v), the volume of solute in 100 volumes of solution:

$$\text{Volume percent} = \frac{\text{Volume of solute}}{\text{Volume of solution}} \times 100\%$$

Example 2.5

A rubbing alcohol usually contains 70 mL of isopropanol in 100 mL of solution. What is the %(v/v) of this solution

%(v/v) isopropanol = (70/100) x 100% = 70%

Example 2.6

An iostonic solution contains 0.9 g of NaCl in 100 mL of aqueous solution. What is the % (w/v) of NaCl of this solution?

Given	Required	Solution
Mass (w) of NaCl = 0.9 g Volume solution = 100 mL	$\%(w/v)_{solute}$	$\%(w/v)_{solute}$ $= \frac{\text{Mass of solute}}{\text{Volume of solution}} \times 100\%$ $= (0.9\ g)/(100\ mL) \times 100\%$ $= \mathbf{0.9\%}$

Note that mass of 100 mL of solution is equal to 100 g if density of solution is assumed to be equal to density of water; the case for dilute solutions.

When making a percent solution, it is important to indicate what units (w/w, or, v/v, or, w/v) are being used, so that others can also make the solution properly. Also, recall that mass of solution is the sum of both the solvent and the solute when you are performing percent calculations.

For more dilute solutions, parts per million (10^6 ppm) and parts per billion (10^9 ppb) are used. These terms are widely employed to express the amounts of trace pollutants in the environment.

The mass-based definitions of ppm and ppb are given here:

$$\text{ppm} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6 \qquad \text{Eq 2.6}$$

$$\text{ppb} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^9 \qquad \text{Eq 2.7}$$

Note that, 1ppm = 1 g solute per 1,000,000 g solution = 1/(1,000.000)

Example 2.7

A 2.5 g sample of ground water was found to contain 5.4 μg of Pb^{2+}. What is the concentration of Pb^{2+}, in parts per million?

Given	Required	Solution
Mass of solute (Pb^{2+}) = 5.4 μg = 5.4×10^{-6} g Mass solution = 2.5 g	Concentration in ppm	$\text{ppm} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$ $= [(5.4 \times 10^{-6}\ \text{g})/(2.5\ \text{g})] \times 10^6$ $=$ **2.16 ppm**

Note that the sum of 2.5 g and 5.4 μg is approximately equal to 2.5 g. When reporting contaminants like lead in drinking water, ppm and ppb concentrations are often reported in mixed unit values of mass/volume. This can be very useful as it is easier for us to think about water in terms of its volume, rather than by its mass. In addition, the density of water is 1.0 g/mL or 1.0 mg/0.001 mL which makes the conversion between the two units easier. For example, if we find that there is lead contamination in water of 4 ppm, this would mean that there are:

$$4\ \text{ppm} = \frac{4\ \text{mg lead}}{1{,}000{,}000\ \text{mg solution}}$$

This is a very dilute solution, so the sum of masses of solute and solvent can be approximately equal to the mass of the solvent

Now, let's convert 1,000,000 mg of water into volume:

$$V = \frac{\text{mass}}{\text{density}}$$

$$= [(1{,}000{,}000\ \text{mg})/(1\ \text{mg}/0.001\ \text{mL})] = 1{,}000\ \text{mL} = 1\ \text{L}$$

Substituting in the above equation yields,

$$4 \text{ ppm} = \frac{4 \text{ mg lead}}{1 \text{L solution}}$$

Therefore, for dilute solutions

$$1 \text{ ppm} = 1 \text{ mg/L}$$

Exercise 2.6

1. A sample of 0.892 g of potassium chloride, KCl, is dissolved in 54.6 g of water. What is the percent, by mass, of KCl in the solution?
2. a) If 150 g of orange juice contains 120 mg of ascorbic acid (Vitamin C), what is the concentration of ascorbic acid, expressed in ppm?
   b) Express the concentration of ascorbic acid given in (a) in ppb.
3. What is the (w/v)% of a solution if 24.0 g of sucrose is dissolved in a total solution of 243 mL?
4. How many grams of NaCl are required to make 625 mL of a 13.5% solution?
5. Find the concentration of calcium ion (in ppm) in a 3.50 g pill that contains 40.5 mg of Ca^{2+}.
6. The label on a 300 mL beer bottle indicates 5.0 % alcohol by volume. How many milliliters of alcohol does the bottle of beer contain?

2.4.2 Mole Fraction

At the end of this section students will be able to

- define mole fraction;
- calculate mole fraction of a solute and a solvent in a solution.

A mole fraction of a component substance A (X_A) in a solution is defined as the moles of component substance divided by the total moles of solution (that is, moles of solute plus solvent).

$$\text{Mole fraction } (X_A) = \frac{\text{moles of substance A}}{\text{total moles of solution}} \qquad \text{Eq 2.8}$$

$$\text{Mole fraction } (X_B) = \frac{\text{moles of substance B}}{\text{total moles of solution}} \qquad \text{Eq 2.9}$$

Note: Usually a solution composed of solute and solvent. So $X_{solute} + X_{solvent} = 1$,

Therefore, if mole fraction of the solute is 0.25, then mole fraction of solvent = 1 - 0.25 = 0.75. In other words, if solute is 25% by mole, then solvent is 75%.

Put another way, the mole fraction gives the proportion of solute (or solvent) particles in solution. The mole percent is the mole fraction expressed as a percentage:

$$\text{Mole \%} = \frac{\text{amount (mol) of solute}}{\text{amount (mol) of solute} + \text{amount (mol) of solvent}} \times 100\% \qquad \text{Eq 2.10}$$

Example 2.8

What is the mole fraction of I_2 in a solution containing 30 g of I_2 in 400 g of CCl_4?

Given	**Required**	**Solution**
Mass solute (I_2) 30 g Mass solvent (CCl_4) = 400 g	mole fraction of I_2	Number of moles of $I_2 = \frac{\text{Mass } I_2}{\text{Molar mass of } I_2}$ = (30 g)/(254 g/mol) = 0.12 mol I_2 Number of moles of $CCl_4 = \frac{\text{Mass of } CCl_4}{\text{Molar mass of } CCl_4}$ = (400g)/(154g/mol) = 2.6 mol CCl_4 $X_{I_2} = \frac{n_{I_2}}{n_{I_2} + n_{CCl_4}}$ = (0.12 mol)/(0.12 mol + 2.6 mol) = 0.044 X_{CCl_4} = 1 - 0.044 = **0.96**

Exercise 2.7

1. A sample of rubbing alcohol contains 142 g of isopropyl alcohol (C_3H_7OH) and 58.0 g of water. What are the mole fractions of alcohol and water? M.wt C_3H_7OH = 60.0 g/mol, H_2O = 18 g/mol.
2. In a solution composed of A, B, and C, if we have 1 mole of A , 1 mole of B and 2 moles of C. Find the mole fraction of each of them.

2.4.3 Molarity

At the end of this section students will be able to

- define molarity;
- prepare molar solutions of different substances;
- calculate molarity of a solution from a given information.

The most common unit of concentration is molarity, which is also the most useful for calculations involving the stoichiometry of reactions in solution.

Definition: The molarity (M) of a solution is the number of moles of solute present in exactly 1 L of solution.

$$\text{Molarity} = \frac{\text{number of moles of solute}}{\text{volume of solution in litre}} \qquad \text{Eq 2.11}$$

$$1\ \text{M} = (1\ \text{mole of solute})/(1\ \text{L of solution})$$

$$\text{Number of moles of a solute} = \frac{\text{Given mass of solute or mass of solute dissolved}}{\text{Molar mass of the solute}} \qquad \text{Eq 2.11}$$

The units of molarity are therefore moles per liter of solution (mol/L), abbreviated as M.

Molarity has two drawbacks that affect its use in precise work:

- Effect of temperature. A liquid expands when heated, so a unit volume of hot solution contains less solute than that of cold solution; thus, the molarity is different.
- Effect of mixing. Because of solute-solvent interactions that are difficult to predict, volumes may not be additive: adding 500 mL of one solution to 500 mL of another may not give 1000 mL of final solution.

Example 2.9

1. What is the molar concentration of a solution containing 16.0 g of CH_3OH in 200 mL of solution?

Given	Required	Solution
Mass of solute (CH_3OH) = 16 g Volume of solution = 200 mL = 0.20 L Molar mass of CH_3OH = 32 g/mol	Molarity	Molarity $= \frac{\text{Number of moles of } CH_3OH}{\text{Volume of solution in liter}}$ $= \frac{\text{Mass of } CH_3OH}{\text{Molar mass of } CH_3OH \text{ x Volume of solution in L}}$ = (16 g)/(32 g/mol x 0.2 L) = 2.50 mol/L = **2.50 M**

2. Calculate the number of moles of sodium hydroxide (NaOH) needed to make 2.50 L of 0.100 M NaOH.

Given	Required	Solution
Molarity =0.1 M Volume solution = 2.50 L	Number of moles of solute (NaOH)	Molarity $= \frac{\text{Number of moles of } CH_3OH}{\text{Volume of solution in liter}} = \frac{n}{V}$ Rearranging the above equation gives: n = M x V = 0.1Mx2.5 L = 0.1 mole/L x2.5 L = **0.25 mole**

Exercise 2.8

1. 5.85 g of sodium chloride (NaCl) is dissolved in in water to make 250 mL of solution. Assuming the density of the solution as 1 g/mL, calculate
   a. the molarity of the solution.
   b. the mass percentage of the solute.
2. How would 250 ml of 0.15 M KNO_3 solution be prepared?
3. A solution of hydrochloric acid contains 36 percent HCl, by mass. Calculate the mole fraction of HCl in the solution.

2.4.4 Molality

At the end of this section students will be able to

- define molality;
- prepare molal solutions of different substances;
- calculate molality of a solution from a given information.

Molality (m) does not contain volume in its ratio; it is the number of moles of solute dissolved in a given mass of solvent (Kg):

$$\text{Molality (m)} = \frac{\text{amount (mol) of solute}}{\text{(mass (kg) of solvent}} \qquad \text{Eq 2.13}$$

Note that molality includes the quantity of solvent, not solution. Molality has two advantages over molarity for precise work:

- Effect of temperature. Molal solutions are based on masses of components, not volume. And since mass does not change with temperature, neither does molality.
- Effect of mixing. Unlike volumes, masses are additive: adding 500 g of one solution to 500 g of another does give 1000 g of final solution.

For these reasons, molality is the preferred term when temperature, and hence density, may change, as in a study of physical properties. Note that, in the case of water, 1 L has a mass of 1 kg, so molality and molarity are nearly the same for **dilute** aqueous solutions.

Example 2.10

Calculate the molality of a solution prepared by dissolving 32.0 g of $CaCl_2$ in 271g of water.

Given	Required	Solution
Mass of $CaCl_2$ = 32.0 g Mass of H_2O = 271 g = 0.271 kg	Molality(m) =?	$\text{Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$ $= \frac{\text{mass of } CaCl_2}{\text{Molar mass } CaCl_2 \text{ x Mass (kg)of of } H_2O}$ $= (32.0 \text{ g})/(111 \text{ g/mol x } 0.271 \text{ kg})$ $= \mathbf{1.06\ m\ CaCl_2}$

Exercise 2.9

1. How many grams of glucose ($C_6H_{12}O_6$) must be dissolved in 563 g of ethanol (C_2H_5OH) to prepare a 2.40×10^{-2} m solution? Ans: mass of glucose = 2.43 g (Verify this)

2.4.5 Normality

At the end of this section students will be able to

- define the terms equivalent weight, number of equivalents and normality;
- calculate normality of a solution from a given information.

The normality of a solution is the gram equivalent weight of a solute per liter of solution. One gram equivalent of a solute per liter of solution is expressed as 1 normal. It is indicated using the symbol N, eq/L, or meq/L (= 0.001 N) for units of concentration.

$$N = \frac{\text{number of gram equivalents of solute}}{\text{volume of solution in litre}} \qquad \text{Eq 2.14}$$

Where number of gram equivalents of solute $= \frac{\text{mass of solute}}{\text{equivalent weight}}$ Eq 2.15

but equivalent weight$= \frac{\text{molar mass of solute}}{z}$ Eq 2.16

Substituting eq 2.16 in eq 2.15 gives:

$$\text{number of gram equivalents of solute} = \frac{\text{mass of solute}}{\text{molar mass}} \times z \qquad \text{eq 2.17}$$

Substituting eq 2.17 in eq 2.14 gives:

$$N = \frac{\text{mass of solute}}{\text{molar mass} \times \text{volume of solution in litre}} \times Z \qquad \text{Eq 2.18}$$

Where ‘z’ is the number of transferable H^+ or OH^- ions in acid-base reaction, positive or negative charges carried by the cations or anions in precipitation reaction, or electrons in redox reactions.

Thus, z = 1 for HCl, 2 H_2SO_4 (because 1 H_2SO_4 releases 2 H^+ ions in aqueous solution), 3 for H_3PO_4, 1 for 1 NaOH, 2 for $Ca(OH)_2$, 2 for Na_2CO_3, and 6 for $Al_2(SO_4)_3$ (because $Al_2(SO_4)_3$ dissociates to give $2Al^{3+}$ = 6 positive charges).

Example 2.11

Calculate the number of gram equivalent or simply equivalents in each of the following

(a) 18.25 g HCl (b) 20 g NaOH (c) 98 g H_2SO_4

Given	Required	Solution
(a) Mass solute (HCl) = 18.25 g Molar mass (HCl) = 36.5 g/mol	eq	The reaction: $HCl \rightarrow H^+ + Cl^-$ HCl produces one H^+; thus z =1. number of gram equivalents of solute =(8.25/36.5) x 1= **0.5 eq**
(b) Mass solute (NaOH) = 20 g Molar mass (NaOH) = 40 g/mol	eq	The reaction: $NaOH \rightarrow Na^+ + OH^-$ NaOH produces one OH^-; z =1. number of gram equivalents of solute =(20/40) x 1 = **0.5 eq**
(c) Mass solute (H_2SO_4) = 98 g Molar mass = 98 g/mol	eq	The reaction: $H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$ H_2SO_4 releases two H^+ ions in this case; z =2. number of gram equivalents of solute =(98/98) x 2 = **2 eq**

Example 2.12

(i) Calculate the number of equivalents present in 0.50 mol H_3PO_4 if the acid is

a. completely neutralized to give PO_4^{3-}

b. converted to $H_2PO_4^{-}$

c. converted to HPO_4^{2-}

(ii) Calculate the Normality of each of the solutions described in Q(i) a to c if the solution volume was 100 mL each.

(iii) What is the mass of Na_2CO_3 required to prepare 100 mL of 0.1 N Na_2CO_3? (Given: Molar mass Na_2CO_3= 106 g/mol).

Given	**Required**	**Solution**
(a) Mole H_3PO_4 = 0.5	number of equivalents	The appropriate equation is $H_3PO_4 \rightarrow 3H^+ + PO_4^{3-}$ H_3PO_4 releases 3 H^+ ions in this case; z = 3 equivalents = (mass of solute)/(molar mass) x z equivalents = n x z= 0.5 x 3 =**1.5 eq** Where, n = (mass solute)/(molar mass solute) =moles of solute
(b) Mole H_3PO_4 = 0.5	eq	The appropriate equation is $H_3PO_4 \rightarrow H^+ + H_2PO_4^{1-}$ H_3PO_4 releases 1 H^+ ion in this case; z =1 equivalents = (mass of solute)/(molar mass) x z equivalents = n x z=0.5 x 1=**0.5 eq**
(c) Mole H_3PO_4 = 0.5	eq	The appropriate equation is $H_3PO_4 \rightarrow 2H^+ + H_2PO_4^{2}$ H_3PO_4 releases 2 H^+ ion in this case; z =2 equivalents=(mass of solute)/(molar mass) x z equivalents=n x z=0.5 x 2=**1 eq**

Normality is the only concentration unit that is reaction dependent. For example, in a multistep dissociation of diprotic (e.g. H_2SO_4) and polyprotic acids (e.g. H_3PO_4), the value of 'z' depends on the actual number of H^+ released; not just the one in the formula.

ii) Calculate the Normality of each of the solutions described in Q(i) a to c if the solution volume was 100 mL each.

Given	**Required**	**Solution**
(a) Mole H_3PO_4 = 0.5 V = 100 mL = 0.1 L	Normality	a) N=Eq/L= (1.5 Eq)/(0.1 L)=15 Eq/L=15 N b) N=Eq/L= (0.5 Eq)/(0.1 L)=5 Eq/L=5 N c) N=Eq/L= (1 Eq)/(0.1 L)=15 Eq/L=**10 N**

(iii) What is the mass of Na_2CO_3 required to prepare 100 mL of 0.1 N Na_2CO_3? (Given: Molar mass Na_2CO_3= 106 g/mol)

Given	**Required**	**Solution**
N = 0.1 eq/L V = 100 mL = 0.1 L Molar mass Na_2CO_3 = 106 g/mol	Mass?	Na_2CO_3 release $2Na^+$; z = 2 Thus $N = \frac{\text{mass of solute}}{\text{molar mass x volume of solution in litre}} \times Z$ ⇒ Mass = [0.1 eq/Lx106 g/mol x 0.1 L]/2 = **0.53 g**

The easiest way to find normality is from molarity. All you need to know is the value of ‘z’.

N = M.z; where N is normality and M is molarity

Example 2.13

Calculate the normality of 0.5 M H_2SO_4 solution (assume complete dissociation).

Given	**Required**	**Solution**
M = 0.5	N	N = Mz = 0.5 x 2 = **1 N**

Exercise 2.10

1. How many equivalents of solute is contained in1 L of 2 N solution?
2. Calculate the mass of $Al_2(SO_4)_3$ in 250 mL of solution if the concentration is 0.48 N.
3. Calculate the molarity and normality of a solution that contains 16.2 g of the salt $Fe_2(SO_4)_3$ in 200 mL of solution.
4. Calculate the normality of
    a. 0.1381 M NaOH
    b. 0.0521 M H_3PO_4 assuming complete dissociation.

2.4.6 Conversion of Concentration Units

At the end of this section students will be able to inter convert various concentration expressions.

The interconversion of concentration units is very simple. Because they are intrinsic units. That means, any amount of liquid taken from a 5% (v/v) ethanol solution, for instance, always contains ethanol to solution in ratio of 5 is to100 regardless of the volume taken. Therefore, 5% (v/v) read as 5 mL of ethanol per 100 mL of solution. 5 ppm glucose means 5 g of glucose per 10^6 g of solution. Similarly, 5 ppb of glucose means 5 g of glucose per 10^9 g of solution. If percent is given in (v/v), convert the mL of solute into moles by converting volume to mass by multiplying with density followed by calculating moles using the formula (mole = mass/molar mass). And convert mL of solution into liter.

Example 2.14

Converting Mole Fractions to Molality

The mole fractions of glucose, $C_6H_{12}O_6$, and water are 0.150 and 0.850, respectively. What is the molality of glucose in the solution?

Given	**Required**	**Solution**
$X_{solute(glucose)} = 0.15$ $X_{solvent(water)} = 0.85$	Molality = m	$\text{Molality} = m = \dfrac{\text{number of moles of solute}}{\text{mass of solvent in kg}}$ Let's consider 0.15 moles of glucose and 0.85 moles of water which is consistent with the given mole fractions. Thus, to calculate the mass of solvent in kg we convert the mole solvent into mass solvent as follows: $\text{Mass}_{H_2O} = \text{mole x molar mass}$ $= 0.85 \text{ mole x} 18 \text{ g/mole}$ $= 15.3 \text{ g} = \mathbf{0.0153 \ kg}$ Substituting the values in the above equation gives: $\text{Molality} = m = \dfrac{0.15 \text{ mole}}{0.0153 \text{ kg}} = \mathbf{9.8 \ m}$

Example 2.15

1. An aqueous solution is 0.273 m KCl. What is the molar concentration of potassium chloride, KCl? The density of the solution is 1.011 x 10^3 g/L.

Given	Required	Solution
Molality = 0.273 m Density solution = 1.011 x 10^3 g/L solute = KCl Molar mass $_{KCl}$ = 74.6 g/mole	Molarity = M	Molarity = M = $\frac{\text{number of moles of solute}}{\text{Volume of solution in litre}}$ but molality = 0.273 m means 0.273 mole of KCl per 1 kg (=1000 g) of solution. Now, we know number of moles of solute. To determine volume of solution in liter we proceed as follows: $V_{solution} = mass_{solution}/d_{solution}$ $= (mass_{solute} + mass_{solvent})/d_{solution}$ = (0.273 mole x 74.6 g/mole) + (1000g)/ (1.011 x 10^3 g/L) =1.01 L Substituting the values in the above equation gives: Molarity = M = $\frac{0.273 \text{ mole}}{1.01 \text{ L}}$ =**0.270 M**

Example 2.16: Converting (w/w) percent to Molarity

Calculate the molarity of 36.5%(w/w) HCl solution (density of solution = 1.20 g/mL; molar mass HCl = 36.5 g/moL)

Given	Required	Solution
%(w/w) = 36.5 Density solution = 1.20 g/mL Molar mass (HCl) = 36.5 g/mole	Convert % into Molarity	Molarity = M = $\frac{\text{number of moles of solute}}{\text{Volume of solution in litre}}$ but %(w/w) = 36.5% means 36.5 g of HCl in 100 g of solution. Thus, number of moles of HCl is calculated as follows: Number of moles of HCl = mass/molar mass = 36.5 g/36.5 g.mol^{-1}=1 mol volume of solution = m_{sol}/d_{soln} = 100 g/1.2 = 83.3 mL = 0.0833 L Substituting the values in the above equation gives: Molarity = M = $\frac{1 \text{ mole}}{0.0833 \text{ L}}$ =**12.0 M**

Exercise 2.11

Hydrogen peroxide is a powerful oxidizing agent: it is used in concentrated solution in rocket fuel. But in dilute solution in hair bleach. An aqueous solution of H_2O_2 is 30.0% by mass and has a density of 1.11 g/mL; molar mass of H_2O_2 is 34.02 g/mol. Calculate the

i. molality

ii. mole fraction of H_2O_2

iii. molarity

2.5 Preparation of Solutions

At the end of this section students will be able to prepare molar, molal, or normal solutions of different substances.

Note that in the example above, we still don't have enough information to actually make the solution in the laboratory. There is no piece of equipment that can measure out the moles of a substance. For this, we need to convert the number of moles of the sample into the number of grams represented by that number. We can then easily use a balance to weigh the amount of substance needed for the solution. In this section, we will consider preparation of stock solutions using a standard solid compounds and preparation of dilute solutions using stock solutions.

2.5.1 Preparation of Stock Solutions from Pure Compound

A) preparation of liquid solution using a solid solute

The solution preparation generally involves 6 steps; part of which are shown in (***Figure 2.15***):

Step 1: Calculate the required amount (mass). When you are asked to prepare a solution that consisted of a dissolved solid, you usually need to calculate the mass required to be dissolved in a suggested volume. But mass can be calculated by using number of moles (n) as;

$$\text{Mass required} = n \times \text{molar mass}$$

Step 2: Weigh the required mass using balance.

Step 3: Transfer the measured amount into a beaker of enough capacity. Add some solvent and mix (shake) to dissolve. The amount of solvent added at this stage should not exceed half the final volume of the solution to be prepared. Heating over a hot plate may be required if the solid is not readily soluble.

Step 4: Transfer the dissolved mixture into volumetric flask of desired capacity and shake to mix.

Step 5: Dilute to the mark by filling the remaining portion of the flask with the solvent. The solvent is usually water as we often prepare aqueous solutions.

Step 6: Cap it and label your flask containing the prepared solution with the key information (concentration, name of solute, date of preparation, and owner (your name). This step is very important even if you need the solution immediately.

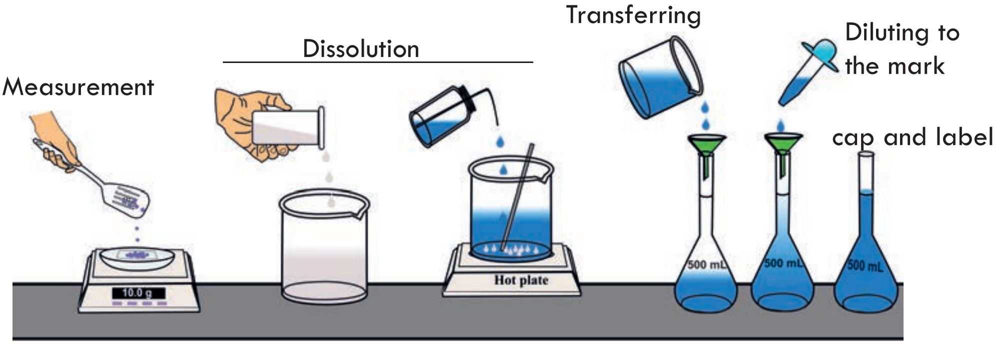


Figure 2.15 Illustration of steps involved in making a 500 mL liquid solution of given concentration by dissolving a solid solute in a liquid solvent.

B) Preparation of a liquid solution using a liquid solute

A solution may also be prepared by dissolving a liquid in another liquid. In this case you need to calculate the volume of the solute that is required. If number of moles is known, then volume (v) can be calculated using;

$V = m/d = (n \times \text{molar mass})/d$; where d density of the solute.

Similar steps are followed. But in this case we use measuring cylinder instead of balance to measure the volume of liquid and the measured volume of liquid is directly transferred to a volumetric flask of desired capacity using a funnel followed by step 5 and step 6.

Example 2.17

Prepare 100 mL of 1 M NaOH using pure NaOH solid.

Given	**Required**	**Solution**
Molarity = 1 M = 1 mol/L Volume of solution = 100 mL = 0.1 L molar mass of NaOH = 40 g/mol	Mass NaOH	$\text{Molarity} = M = \frac{\text{number of moles of solute}}{\text{Volume of solution in litre}} = \frac{n}{V}$ ☞ This is a liquid solution to be prepared by dissolving appropriate mass of the solute (NaOH) in distilled water to form 0.1 L of solution. ☞ Thus, we need to calculate mass of NaOH required to be dissolved in water to form 0.1 L of the solution. ☞ But mass = number of moles x molar mass ☞ Number of moles is obtained by rearranging the molarity equation as: ☞ Number of moles = MV ☞ mass = (1 mol/L)x0.1Lx40 g/mole = **4 g**

Therefore, the solution is prepared by dissolving 4 g of NaOH in 100 mL volumetric flask, and diluting to the mark by filling the remaining portion of the flask with distilled water.

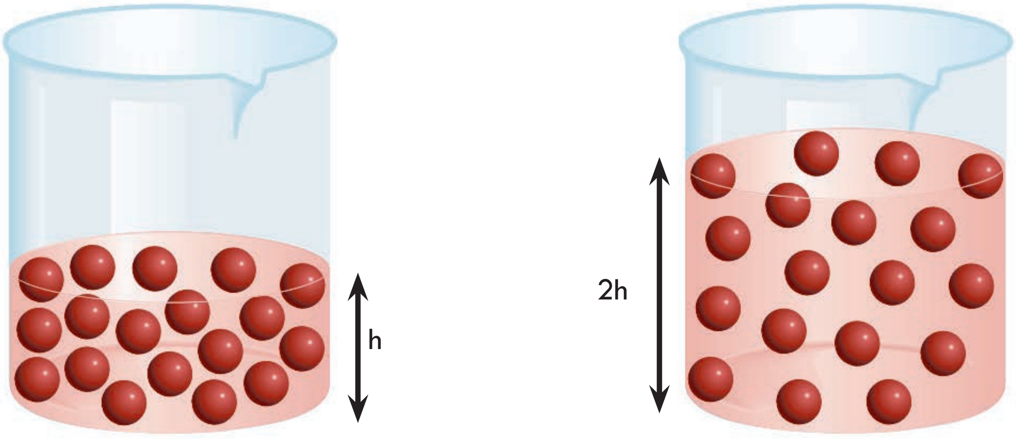
2.5.2 Diluting a Solution

At the end of this section students will be able to

- explain dilution process;
- calculate the volume or concentration changes during dilution of solution;
- prepare a dilute solution from concentrated solution.

Concentrated solutions are often stored in the laboratory stockroom for use as needed. Frequently we dilute these "stock" solutions before working with them.

Dilution is the procedure for preparing a less concentrated solution from a more concentrated one by simply adding more solvent. Often, a small portion (Vi) of stock solution is withdrawn from the more concentrated solution and diluted to final volume Vf by simply adding more solvent. Imagine we have an aqueous salt solution with 18 particles in a certain volume V1 of solution and that it has a certain concentration M1. That means we have 18 particles (18/(6.02x10^{23} moles)) dissolved in the volume V^1 of the solution. Let's dilute this solution by adding more water to double the volume of the solution to $V_2 = 2V_1$ as illustrated in *Figure 2.16*. What stays the same and what is changed when more water is added? Did the number of particles (number of moles) change? What about the concentration?



Solution I
Before dilution; volume = V_1,
Concentration = M_1

Solution II
After dilution; volume = V_2,
Concentration = M_2

Figure 2.16. The dilution of a more concentrated solution (solution I) to a less concentrated one (solution II) by doubling the volume of solution by adding more solvent does not change the total number of solute particles.

The molarity of solution I is
$M_1 = \text{Moles}_1/V_1$

And the molarity of solution II is

$M_2 = Moles_2/V_2$

Rearranging the equations to find moles we obtain

$Moles_1 = M_1\ V_1$

And

$Moles_2 = M_2\ V_2$

As you can see in ***Figure 2.16***, adding more solvent changed the volume of the solution. Doing so we also changed its concentration because the number of particles (number of moles) per unit volume (per unit litre or per unit mL) is changed. However, the number of particles (number of moles) of solute did not change. So,

$Moles_1 = Moles_2$

$M_1\ V_1 = M_2\ V_2$

This formula is known as dilution formula. It is used to prepare dilute solutions. For instance, to prepare a dilute solution having a volume V_2 and concentration M_2 using a more concentrated solution having concentration M_1, we withdraw the volume V_1 from the more concentrated (stock) solution having the concentration M_1 and add more water to raise the volume from V_1 to V_2. We often need to calculate V_1 using the formula:

$$V_1 = M_2 V_2/M_1$$

Caution! When preparing dilutions of concentrated acids such as sulfuric acid, the directions usually call for adding the acid slowly to water with frequent stirring. If water were added to acid, the water would quickly heat and splatter, causing harm to the person making the solution.

Example 2.18

Describe how you would prepare 2.50×10^2 mL of a 2.25 M H_2SO_4 solution, starting with a 7.41 M stock solution of H_2SO_4.

Given	Required	Solution
$M_1 = 7.41$ M $V_1 = ?$ $M_2 = 2.25$ M $V_2 = 250$ mL	V_1	$M_1\ V_1 = M_2 V_2$ $V_1 = \frac{M_2 V_2}{M_1}$ $V_1 = (2.25\ M \times 250\ mL)/(7.41\ M)$ $= \mathbf{75.9\ mL}$

Thus, we must withdraw 75.9 mL from the 7.41 M H_2SO_4 solution, transfer it to a 250 mL volumetric flask half-filled with distlled water, and dilute it with sufficient amount of water to the mark.

Exercise 2.12

1. How would you prepare 2.00 x 10^2 mL of a 0.866 M KOH solution, starting with a 5.07 M stock solution?
2. What is the difference between dilute and unsaturated? Or concentrated and saturated?
3. Prepare a 100 mL of 0.1 M sucrose from 1M sucrose solution by dilution.

Exercise 2.13

Outline the preparation of 100 mL of (a) hypotonic (0.1 M) and (b) hypertonic (0.25 M) aqueous NaCl solution and (c) 0.1 M H_2SO_4 using a stock solution of 98% (w/w) H_2SO_4. Density = 1.84 g/ml. (Hint: First calculate the molarity of the 98%(w/w) H_2SO_4 solution as shown in example 2.16. Then use the dilution formula). Show all the necessary steps and let your teacher check it!

2.6 Solution Stoichiometry

At the end of this section students will be able to

- use stoichiometrically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substance.

In Unit 3 you studied about stoichiometric calculations in terms of the mole method, which treats the coefficients in a balanced equation as the number of moles of reactants and products. Solving stoichiometry problems for reactions in solution requires the additional step of converting the volume of reactant or product in solution to amount (mole) using the relationship MV = number of moles of solute.

Generally, we follow these steps:

1. Balance the equation. Note that the coefficients in chemical equation indicate number of moles.
2. Find the amount (mol) of one substance from the volume and molarity.
3. Relate it to the stoichiometrically equivalent amount of another substance.
4. Convert to the desired units.

Mole-Mole Calculation

Example 2.19

What is the concentration of sodium hydroxide required to react completely with equal volume of 0.104 M sulfuric acid?

Given

$V_{NaOH} = V_{H_2SO_4}$

$M_{H_2SO_4} = 0.104$ M

Required

$M_{NaOH} = ?$

Solution

Equation: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

Stoichiometry	1 mole	2 mole			
Equation	H_2SO_4	+ 2NaOH	→	Na_2SO_4	+ $2H_2O$
Given	$nH_2SO_4 = 0.104$ mol/L x VH_2SO_4	nNaOH=MNaOH x VNaOH			

Stoichiometry:= $n_{NaOH}/n_{H_2SO_4} = 2/1$

$= M_{NaOH}V_{NaOH}/M_{H_2SO_4}V_{H_2SO_4}$

$= M_{NaOH}V_{NaOH}/0.104MV_{H_2SO_4}$

Since volume is the same, it cancels out and the equation becomes: $2/1 = M_{NaOH}/0.104$ M

$\Rightarrow M_{NaOH} = 0.104$ M x 2

$=$ **0.208 M**

Exercise 2.14

What is the concentration of sodium hydroxide that is required to react completely with equal volume of 0.104 M hydrochloric acid?

Mole - Mass

Example 2.20

What mass of solid magnesium hydroxide can be produced if 45 mL of a 0.63 M $Mg(NO_3)_2$ solution reacts completely with excess NaOH?

Given

$V_{Mg(NO_3)_2} = 45$ mL

$M_{Mg(NO_3)_2} = 0.63$ M

Note that M and V are given means number of moles (n) is known: n=MV

Required

$V_{AgNO_3} = ?$

Solution

Stoichiometry		1 mole		58.321 g	
Equation	$2NaOH(aq) + 1Mg(NO_3)_2(aq)$		→	$Mg(OH)_2(s) + 2NaNO_3(aq)$	
Given		0.63 M, 45 mL		Mass =?	

Stoichiometry:=$n_{Mg(NO_3)_2}/mass_{Mg(OH)_2}$

=1 mole/58.321 g

=0.63 mol/L x 0.045 L/$mass_{Mg(OH)_2}$

$\Rightarrow mass_{Mg(OH)_2}$ = 0.63 mol/L x 0.045L x 58.321 g/1 mole

=1.65g

Mass - Volume

Example 2.21

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

Given

$w_{Ag} = 3.66$ g

$M_{Al(NO_3)_3} = 0.0995$ M

Required

$V_{Al(NO_3)_3} = ?$

Solution

Stoichiometry	3x107.97 g	1 mole			
Equation	3Ag(s)	+ $Al(NO_3)_3$(aq)	→	3 $AgNO_3$(aq)	+ Al(s)
Experimental	3.66 g	$n_{Al(NO_3)_3} =$ $M_{Al(NO_3)_3} \times V_{Al(NO_3)_3}$			

Stoichiometry: $= w_{Ag}/n_{Al(NO_3)_3} = 3 \times 107.97 \text{ gAg}/1\text{mole}_{Al(NO_3)_3}$
$= 3.66\text{g}/M_{Al(NO_3)_3} V_{Al(NO_3)_3} = 3.66\text{g}/0.0995\text{M} \times V_{Al(NO_3)_3}$

$\Rightarrow$ $3 \times 107.97\text{g}/1\text{mole} = 3.66\text{g}/(0.0995\text{M} \times V_{Al(NO_3)_3})$; where MV = mole

$V_{Al(NO_3)_3} = 3.66 \text{ g} \times 1 \text{ mole}/3 \times 107.97 \text{ g} \times 0.0995 \text{ mol/L} =$ **0.114 L**

Exercise 2.15

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4$(s) according to the following chemical equation?

$$H_2C_2O_4(s) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(l)$$

Mole - Volume Calculation

Example 2.22

What volume of a 0.35 M $AgNO_3$ is required to react completely with 55mL of a 0.24 M NaCl solution?

Given

$V_{NaCl} = 55$ mL

$M_{NaCl} = 0.24$ M

Note that M and V are given means number of moles (n) is known: n=MV

Required

$V_{AgNO_3} = ?$

Solution

Stoichiometry	1 mole	1 mole			
Equation	NaCl(aq) + $AgNO_3$(aq)		→	$NaNO_3$(aq) + AgCl(s)	
Experimental	0.24 M, 55 mL	0.35 M, V=?			

Stoichiometry: $= n_{NaCl}/n_{AgNO_3} = 1/1$

$= (0.24 \text{ mol/L} \times 0.055\text{L})/(0.35 \text{ mol/L} \times V_{AgNO_3})$

$=1$

$V_{AgNO_3} = (0.24 \text{ mol/L} \times 0.055\text{L})/(0.35 \text{ mol/L}) = 0.0377$ L=**37.7 mL**

Mole - Number of Particles

Example 2.23

1. How many HCl particles are there in 0.1 mole of the HCl solution?

Strategy: 1 mole HCl contains Avogadro's number of HCl particles, then 0.1 mole contains how many?

Answer: 0.6022×10^{23} particles = 6.022×10^{22} HCl particles

Exercise 2.16

What volume of 0.5 M glucose ($C_6H_{12}O_6$) contains 3.01×10^{22} molecules of glucose?

2.7 Describing Reactions in Solution

At the end of this section students will be able to

- explain the relationship between reacting ions, spectator ions, precipitation and solubility;
- write net ionic equations.

Molecular Equation

A molecular equation is a chemical equation in which the reactants and the products are written as if they were molecular substances, even though they may actually exist in solution as ions. The state (solid, liquid, gas) of each substance is indicated in parentheses after formula. (aq) is used to indicate the that molecule is dissolved in water.

Consider the reaction between silver nitrate, $AgNO_3$, and sodium chloride, NaCl, in aqueous solution to give solid silver chloride, AgCl, and aqueous sodium nitrate. The equation for this reaction may be written as:

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

This is just a molecular equation. Note from the solubility rules (***Table 2.1***) that all common nitrates and chlorides of Group 1A are soluble. That is why (aq) is used after these formulae. The solubility rule also taught us that AgCl is insoluble in water. That is why (s) to indicate that it is precipitated out from the solution to the bottom of the reaction flask/test tube as a solid.

Ionic Equation

Although a molecular equation is useful in describing the actual reactants and products, it does not give any information about what is happening at the level of ions. Since this kind of information is very useful, you often need to write the molecular equation as an ionic equation. Consider the reaction of silver nitrate and sodium chloride. Both are soluble ionic substances and are strong electrolytes. When they dissolve in water, they go into solution as ions giving on the reactant side of the reaction:

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq)$$

On the product side of the equation, AgCl(s) is an ionic compound that doesn't dissolve in water, but the ions are held together in particular sites in the crystalline solid. We leave the formula as AgCl(s) to convey this information in the equation. On the other hand, $NaNO_3$ is a soluble ionic compound and is a strong electrolyte. Also it dissolves in aqueous solution to give freely moving ions. Therefore, we can write it as

$$Na^+(aq) + NO_3^-(aq)$$

The complete equation is

$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq),$$

The ions appearing on both sides of the equation (Na^+ and NO_3^-) are called spectator ions, as they do not take part in the reaction and they can be canceled on both sides to express the essential reaction that occurs.

$$Ag^+(aq) + \cancel{NO_3^-}(aq) + \cancel{Na^+}(aq) + Cl^-(aq) \rightarrow AgCl(s) + \cancel{Na^+}(aq) + \cancel{NO_3^-}(aq)$$

The resulting equation is:

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

This net ionic equation, is without spectator ions and the reaction that actually occurs at an ionic level is between silver ions and chloride ions which forms solid silver chloride.

Steps for writing a net ionic equation

1. Write the balanced molecular equation
2. Write the balanced complete ionic equation
3. Cancel out the spectator ions. What remains is the net ionic equation.

Exercise 2.17

1. Discuss the difference between molecular equations and ionic equations? What are spectator ions?
2. Consider the reaction between sodium carbonate and calcium chloride in aqueous solution:
   a. How are sodium carbonate and calcium chloride found in water solution, dissociated to give ions or as solid? (Refer to solubility rule). Write dissociation reaction for each of them.
   b. Write the ionic equation for the reaction between sodium carbonate and calcium chloride. Are the products formed soluble in water?
   c. Write the soluble product in ionic form. Are there the same ions in the reactant and product side? If yes, cancel them out!
   d. Write the net ionic equation.
3. For the following molecular equation, write ionic equation, net ionic equation, and identify the spectator ions:

$$2AgNO_3\ (aq) + Na_2CrO_4\ (aq) \rightarrow Ag_2CrO_4\ (s) + 2NaNO_3\ (aq)$$

Key Terms of the Unit

- Concentrated
- Colloids
- Dilute
- Electrolyte solution
- Henry's law
- Heterogeneous solution
- Homogeneous solution
- Mass percentage
- Miscible
- Molality
- Molarity
- Mole fraction
- Nonelectrolyte
- Normality
- Parts per billion
- Saturated
- Solvation
- Supersaturated
- Suspension
- Unsaturated

Unit Summary

- To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.
- **A heterogeneous mixture** is a mixture that consists of physically distinct parts, each with different properties.
- **A homogeneous mixture** is a mixture in which the composition is the same throughout. i.e., it has no visible boundaries because the components are mixed as individual atoms, ions and molecules.
- **A suspension** is a heterogeneous mixture that consists of a dispersion of fine solid particles in a liquid or gas, removable by filtration.
- **A colloid** is a heterogeneous mixture in which insoluble particles of one or more substances are suspended uniformly throughout another substance.
- Scattering of a beam of light by colloidal particles is called **Tyndall effect.**

- A colloid can be a gel, an emulsion, a sol, foam, or aerosol.
- **Coagulation** is the process by which the dispersed phase of a colloid is made to aggregate and thereby separate from the continuous phase.
- A colloid in which the dispersed phase consists of micelles is called an **association colloid**.
- **A solution** is a homogeneous mixture. The major component is the solvent, while the minor component is the solute. The terms miscible and immiscible, instead of soluble and insoluble, are used for liquid solutes and solvents. The statement like dissolves like is a useful guide to predicting whether a solute will dissolve in a given solvent.
- For liquid solutions formed from a solvent A and solute B, an ideal solution results when the A-B attractions are the same as the A-A and B-B attractions and for such a solution, $\Delta H_{soln} = 0$.
- When the A-B attractions are greater than the A-A and B-B attractions, $\Delta H_{soln} < 0$ and the solution process is exothermic.
- When the A-B attractions are weaker than the A-A and B-B attractions, $\Delta H_{soln} > 0$ and the solution process is endothermic.
- Dissolving occurs by solvation, the process in which particles of a solvent surround the individual particles of a solute, separating them to make a solution. For water solutions, the word hydration is used. If the solute is molecular, it dissolves into individual molecules. If the solute is ionic it dissolves into ions. The individual ions separate from each other, forming a solution that conducts electricity. Such solutions are called **electrolytes**. If the dissociation of ions is complete, the solution is a strong electrolyte. If the dissociation is only partial, the solution is a weak electrolyte. If a solution is made up of molecules only, it does not conduct electricity and is called **nonelectrolyte**.
- The amount of solute in a solution is represented by the concentration of the solution. The maximum amount of solute that can be dissolved in a given amount of solvent is called the **solubility** of the solute. Such solutions are **saturated**. Solutions that have less than the maximum amount that it can dissolve are **unsaturated**. Most solutions are **unsaturated**, and there are various ways of stating their concentrations.
- A dilute solution is one in which there is a relatively small amount of solute dissolved in the solution.
- A concentrated solution contains a relatively large amount of solute.
- These two terms do not tell us whether or not the solution is saturated or unsaturated. They are used for comparing concentrations of two solutions in a more general qualitative sense.
- **Solubility** depends on two factors, predominantly **intermolecular forces** of attraction and to some extent conditions (surface area, temperature, pressure).
- **Increasing the temperature** usually **increases the solubility of solids** in water and liquids in water while it **decreases the solubility of gases** in water. Increasing

pressure increases solubility of gases in liquid. Henry's law says that the solubility of a gas is directly proportional to its partial pressure above the solution.

Review Exercise

Part I: Multiple Choice Questions

1. Soda water is an example of:
   a. liquid-liquid solution
   b. liquid-gas solution
   c. solid-liquid solution
   d. gas-gas solution
2. Which of the following type of matter can exhibit Tyndall effect?
   a. aerosol b. soda water c. brass
3. A solution that is formed when NaCl dissolves in water has the properties of-----
   a. Sodium Chloride
   b. Water
   c. both NaCl and water
   d. neither NaCl nor water
4. A simple way of determining whether a mixture is colliodial or not is through the use of ______
   a. particle size b. Tyndall effect c. non-uniform appearance d. a & b
5. A suspension is ……
   a. a homogenous mixture of a solute and a solvent
   b. a heterogeneous mixture of one phase dispersed in another phase
   c. a heterogeneous mixture that exhibits Tyndall effect
   d. a mixture that contain smaller particle size than a solution
6. Which properties of colloids is applied in the cloud seeding mechanism to produce artificial rain?
   a. Association
   b. Dispersion
   c. Coagulation
   d. Scattering
7. The cleansing action of soaps and detergents are related to which property of a colloid?
   a. Association
   b. Dispersion
   c. Coagulation
   d. Scattering
8. The continuous and dispersed phase in fog is _____ and _____ , respectively.
   a. gas, solid
   b. liquid, gas
   c. gas, liquid
   d. liquid, liquid
9. How many grams of NaOH is required to prepare 100 mL of 1 M NaOH solution?
   a. 4 b. 40 c. 0.4 d. 400
10. A solution that contain 4.9 g H_2SO_4 in 1 liter of solution has the same _____ and _____. Assume the solution has the same density as water.
    a. molarity and molality
    b. molarity and normality
    c. normality and molality
    d. none of the above
11. When the molarity of a sulfuric acid solution is doubled, its normality is _____
    a. halved b. doubled c. quadrupled d. remain the same
12. Mole fraction and percentage differ by a factor of _____
    a. 1 b. 10 c. 100 d. 0.1

13. Which of the following substances are not readily miscible within each other?
    a. C_6H_6 and CCl_4
    c. C_2H_5OH and H_2O
    b. H_2O and CCl_4
    d. CH_3OH and H_2O
14. Rate of dissolution is largely dependent upon:
    a. the inter-particle forces
    b. the surface area of solid solute
    c. the temperature of the system
    d. the pressure of the system
15. Which of the following statements is not correct?
    a. pressure has little effect on the solubility of liquids and solids
    b. the solubility of most solids increases with increasing temperature
    c. the solubility of gases in water increases with increasing temperature.
    d. none of the above
16. The quantitative relationship between gas solubility and pressure is given by:
    a. Raoult's law
    c. Henry's law
    b. Hess's law
    d. Dalton's law
17. A solution that has a capacity to dissolve more solute is called ____________ solution
    a. dilute
    c. supersaturated
    b. unsaturated
    d. concentrated
18. A concentrated solution is ____________
    a. same as supersaturated solution
    b. contain less solute than solvent
    c. contains relatively large amount of solute
    d. a saturated solution
19. Which of the following solutions is the most dilute?
    a. 0.1 M NaCl
    c. 0.001 M NaCl
    b. 0.01 M NaCl
    d. 10.0 mM NaCl
20. Which of the following is a saturated NaCl solution at 20°C? Solubility of NaCl is 36 g/100 mL.
    a. 0.615 M
    c. 36.0 M
    b. 6.15 M
    d. 0.0615 M
21. The ratio of the number of moles of solute divided by the total number of moles gives:
    a. the mole fraction of the solute
    c. the molality of the solution
    b. the molarity of the solution
    d. the normality of the solution
22. Phosphoric acid reacts with calcium hydroxide to form calcium hydrogen phosphate ($CaHPO_4$) and water. The equivalent weight of phosphoric acid in this reaction will be:
    a. 98 g/eq
    c. 32.6 g/eq
    b. 49 g/eq
    d. 14 g/eq

23. Which of the following is the balanced net ionic equation for the reaction between H_2SO_4 (a) and NaOH (aq)?
    a. $HSO_4^-(aq) \rightarrow H^+(aq) + SO_4^{2-}(aq)$
    b. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
    c. $SO_4^{2-}(aq) + 2Na^+(aq) \rightarrow NaSO_4^{2-}(aq)$
    d. $H_2SO_4(aq) + 2OH^-(aq) \rightarrow 2H_2O(l) + SO_4^{2-}(aq)$
24. When a given volume of concentrated solution is diluted 100 fold, which one of the following is true?
    a. the dilute solution has more number of moles of solute per unit volume than that of concentrated solution
    b. the number of moles of solute per unit volume of the dilute solution is equal to that of the concentrated solution
    c. the number of moles of solute in a given volume of a solution is the same before and after dilution
    d. all of the above

Part II: Short Answer Questions

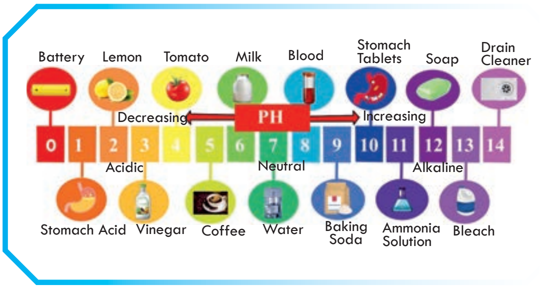
25. Give one example of each: a gaseous solution, a liquid solution, a solid solution.
26. What are the two factors needed to explain the differences in solubilities of ionic solids in water?
27. Explain in terms of intermolecular attractions why octane is immiscible in water.
28. Give the type of colloid (aerosol, foam, emulsion, sol, or gel) that each of the following represents.
    a. rain cloud
    b. milk of magnesia
    c. soapsuds
    d. silt in water
29. Concrete is a mixture of ______ , ______ , ______ , and ______ .
30. Explain on the basis that "like dissolves like" why glycerol, $CH_2OHCHOHCH_2OH$, is miscible in water but benzene, C_6H_6, has very limited solubility in water.
31. Explain why ionic substances show a wide range of solubilities in water.
32. Indicate the type of solute-solvent interaction that is most important in each of the following solutions:
    a. KBr in water
    c. ammonia in water
    b. hexane, C_6H_{14}, in gasoline
33. Describe the characteritics of endothermic and exothermic dissolution processes.
34. Consider the following solutions. In each case, predict whether the solubility of the solute should be high or low. Justify your answer.
    a. KCl in H_2O
    d. H_2O in CH_3OH
    b. HF in H_2O
    e. CCl_4 in H_2O
    c. KCl in CCl_4
35. What is the usual solubility behavior of an ionic compound in water when the temperature is raised? Give an example of an exception to this behavior.
36. Give one example of each: a salt whose heat of solution is exothermic and a salt whose heat of solution is endothermic.

37. What do you expect to happen to a concentration of dissolved gas in a solution as the solution is heated?
38. Explain why a carbonated beverage must be stored in a closed container.
39. Pressure has an effect on the solubility of oxygen in water but a negligible effect on the solubility of sugar in water. Why?
40. Which of the following ions would be expected to have the greater energy of hydration, Mg^{2+} or Al^{3+}?
41. Which of the following ions would be expected to have the greater energy of hydration, F^{-} or Cl^{-}?

Part III: Work Out Problems

42. The solubility of carbon dioxide in water is 0.161 g CO_2 in 100 mL of water at 20oC and 1.00 atm. A soft drink is carbonated with carbon dioxide gas at 5.50 atm pressure. What is the solubility of carbon dioxide in water at this pressure?
43. Calculate the molarity of each of the following solutions:
    a. 10.5 g NaCl in 350.0 mL of solution
    b. 40.7 g $LiClO_4 \cdot 3H_2O$ in 125 mL of solution
44. What mass of solution containing 5.00% potassium iodide, KI, by mass contains 258 mg KI?
45. Caffeine, $C_8H_{10}N_4O_2$, is a stimulant found in tea and coffee. A sample of the substance was dissolved in 45.0 g of chloroform, $CHCl_3$, to give a 0.0946 m solution. How many grams of caffeine were in the sample? Molar mass caffeine = 194.19 g/mol
46. A 100.0-g sample of a brand of rubbing alcohol contains 65.0 g of isopropyl alcohol, C_3H_7OH, and 35.0 g of water. What is the mole fraction of isopropyl alcohol in the solution? What is the mole fraction of water?
47. A bleaching solution contains sodium hypochlorite, NaClO, dissolved in water. The solution is 0.650 m NaClO. What is the mole fraction of sodium hypochlorite?
48. The concentrated sulfuric acid we use in the laboratory is 98.0% H_2SO_4, by mass. Calculate the molality and molarity of the acid solution. The density of the solution is 1.83 g/mL.
49. Calculate the approximate volume of water that must be added to 250 mL of 1.25 N solution to make it 0.500 N.
50. An antiseptic solution contains hydrogen peroxide, H_2O_2, in water. The solution is 0.610 m H_2O_2. What is the mole fraction of hydrogen peroxide?
51. Citric acid, $H_3C_6H_5O_7$, occurs in plants. Lemons contain 5% to 8% citric acid by mass. The acid is added to beverages and candy. An aqueous solution is 0.688 m citric acid. The density is 1.049 g/mL. What is the molar concentration?
52. A solution of vinegar is 0.763 M acetic acid, $HC_2H_3O_2$. The density of the vinegar is 1.004 g/mL. What is the molal concentration of acetic acid?

UNIT 3



IMPORTANT INORGANIC COMPOUNDS

Unit Outcomes

At the end of this unit, you will be able to

- classify inorganic compounds on the basis of their composition and/ or their chemistry;
- discuss types of oxides and their chemical properties;
- explain the Arrhenius concept of acids and bases;
- mention the classification of acids and salts;
- describe the general properties, preparation and uses of common acids, bases and salts;
- distinguish the differences between strong and weak acids/ bases; and concentrated and dilute acids/ bases;
- recognize the corrosive nature of acids and bases, and exercise the necessary precautions in handling and using them;
- develop skills for identifying acidic, basic and neutral compounds;
- demonstrate scientific inquiry skills: observing, classifying, comparing and contrasting, inferring, predicting, communicating, measuring, asking questions, interpreting data, drawing conclusion, applying concepts, relating cause and effect and problem solving.

Start-up Activity

Discuss the following in groups and present your conclusion to the class.

1. Which flavour (salty, sour, bitter or none) each of the following substances does have? Coffee, unripe orange, cheese, outer peel of lemon, kosso, water, common salt, vinegar.
2. Consider the following: carbohydrates, proteins, urea, oils, diamond, salt, water, baking soda. Classify them into organic and inorganic and analyze the source of each substance.
3. Can you identify organic and inorganic compounds using their melting points?

3.1 Introduction

By the end of this section, students will be able to

☞ define inorganic compounds;

☞ classify inorganic compounds as oxides, acids, bases and salts.

The term inorganic compound refers to all compounds that do not contain carbon except simpler compounds of carbon like oxides (eg CO_2), carbonates (eg $CaCO_3$) and carbides (eg SiC). Inorganic compounds are compounds consisting of mineral constituents of the earth or generally found in nonliving things. Compounds found in living things are generally organic compounds (see unit 6 for the details). The majority of metal compounds are inorganic.

Inorganic compounds are mostly found in nature as silicates, oxides, carbonates, sulphides, sulphates, chlorides, nitrates, etc.

Inorganic compounds are generally classified into four groups namely oxides, acids, bases and salts.

We are familiar with the utility and importance of oxides, acids, bases and salts in our daily life. For example, we know that the carbon dioxide (CO_2) gas in CO_2-fire extinguisher is an oxide, the sour taste of vinegar is due to acetic acid (CH_3COOH), sodium hydroxide (NaOH) is a base used in the preparation of soaps, and sodium chloride (NaCl) is common salt which is useful as a food ingredient. This unit deals with the chemical nature and formation of more oxides, acids, bases and salts.

Exercise 3.1

1. List some inorganic compounds with which you are familiar. Write on the board.
2. Lime stone ($CaCO_3$) and common salt (NaCl) are inorganic substances. Why?
3. Which branch of chemistry is concerned with the study of all elements?
4. What is the source of inorganic compounds ?

3.2 Oxides

Activity 3.1

Form a group and discuss the following and present your findings to the class;

1. Why oxygen is the most abundant element on the earth's crust?
2. How oxygen exists in nature?
3. Which elements react with oxygen?
4. What do you call the product when oxygen reacts with an element?

At the end of this section, students will be able to

- define and classify oxides as acidic, basic, neutral, amphoteric and peroxides
- define and give examples of acidic oxides, basic oxides, amphoteric oxides and neutral oxides
- explain the chemical properties of acidic oxides, basic oxides, amphoteric oxides and neutral oxides
- differentiate basic oxides from acidic oxides using experimental results
- compare and contrast acidic and basic oxides
- explain the salt-forming nature of acidic oxide, basic oxide and amphoteric oxide
- define neutral oxides and peroxides and give examples for each of them
- discuss the chemical properties of peroxides
- differentiate peroxides from other oxides by conducting an experiment

Oxygen reacts directly with almost all elements except the noble gases and inactive metals like gold, platinum, and palladium. Such compounds of oxygen are called ***oxides***. Oxides are binary compounds containing oxygen and any other element (metal, non-metal or metalloid). Note that binary compounds are those consisting of only two elements.

Oxygen + Element (Metal, Non-metal or Metalloid) → Oxide

Example 3.1: magnesium oxide, MgO, aluminium oxide, Al_2O_3, carbon dioxide, CO_2, and sulphur dioxide, SO_2 and carbon monoxide, CO are examples of oxides.

Exercise 3.2

Classify the following compounds as oxides or not oxides.

A. Na_2O
B. N_2O_5
C. Na_2CO_3
D. P_4O_{10}
E. KNO_3
F. SO_3
G. H_3PO_4
H. Al_2O_3
I. H_2O
J. H_2SO_4

Oxides, based on their properties are classified as shown in *Figure 3.1*.

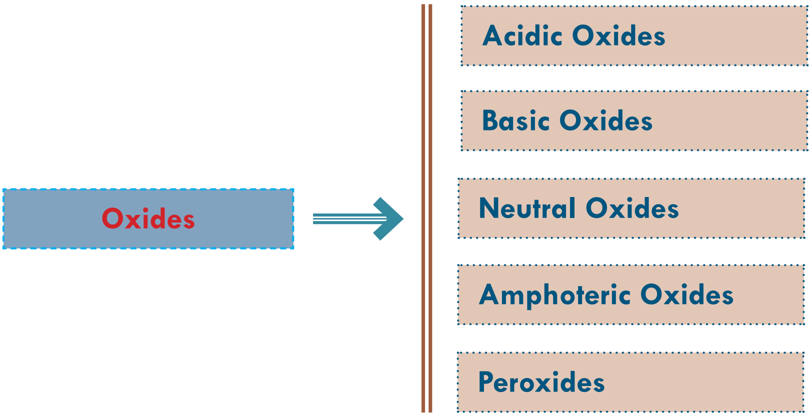


Figure 3.1 Classification of oxides.

A. Acidic Oxides

Activity 3.2

Discuss in your group and present your conclusions to the class.

1. What is acid rain and what are the major causes of acid rain?
2. What would be the pH (decreases, increases or remains unchanged) of solution when carbon dioxide (CO_2), dinitrogen pentoxide (N_2O_5), tetraphosphorous decaoxide (P_4O_{10}), sulfur dioxide (SO_2) and dichlorine heptoxide (Cl_2O_7) are dissolved in water? What can you conclude regarding the acid-base behavior of oxides of nonetallic elements like C, N, S, Cl, and P?

Acidic oxides are the oxides formed by the chemical combination of oxygen with non-metals (Groups 14 -17).

These oxides are also called acid anhydrides, since they form acidic solutions when reacted or dissolved in water. Acid anhydride means acid without water. Generally speaking, acidic oxides are non-metal oxides. Examples of acidic oxides include carbon dioxide, CO_2, nitrogen dioxide, NO_2, and sulphur dioxide, SO_2. However, it is very important to note that all non-metal oxides are not necessarily acidic oxides. For example, carbon monoxide, CO, and di-nitrogen monoxide, N_2O, are non-metal oxides, but they are neutral oxides which will be discussed later (D). There are acidic oxides that do not react directly with water. The acids of such oxides are formed by other methods. For example; SiO_2 is not soluble in water, but it neutralizes basic oxides, thus it is acidic.

$$SiO_2 + Na_2O \rightarrow Na_2SiO_3 \text{ (sodium silicate)}$$

The acid of SiO_2 (H_2SiO_3) is prepared by the reaction of HCl with Na_2SiO_3 as shown below:

$$Na_2SiO_3 \text{ (aq)} + HCl \text{ (aq)} \rightarrow H_2SiO_3 \text{ (aq)} + 2NaCl \text{ (aq)}$$

Chemical properties of acidic oxides

1. Acidic oxides (acid anhydrides) dissolve in water to form acidic solution (acid).

$$\text{Acid anhydride} + \text{Water} \rightarrow \text{Acid}$$
$$CO_2 + H_2O \rightarrow H_2CO_3 \text{ (Carbonic acid)}$$
$$SO_2 + H_2O \rightarrow H_2SO_3 \text{ (Sulphurous acid)}$$

2. Acidic oxides react with basic or metallic oxides to form salt.

$$\text{Acidic oxide} + \text{Basic oxide} \rightarrow \text{Salt}$$
$$CO_2 + Na_2O \rightarrow Na_2CO_3 \text{ (sodium carbonate)}$$
$$SO_3 + CaO \rightarrow CaSO_4 \text{ (calcium sulphate)}$$

3. Acidic oxides react with bases to form salt and water. This reaction is called neutralization reaction.

$$\text{Acidic oxide} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$
$$SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$$
$$CO_2 + 2LiOH \rightarrow Li_2CO_3 + H_2O$$

Exercise 3.3

1. List some examples of acidic oxides.
2. Give two examples of non-metallic oxides which are not acidic oxides.
3. Complete and balance the following equations.
   a. $N_2O_5 + H_2O \rightarrow$
   b. $P_4O_{10} + H_2O \rightarrow$
   c. $MgO + CO_2 \rightarrow$
   d. $Ca(OH)_2 + SO_3 \rightarrow$
   e. $KOH + CO_2 \rightarrow$
   f. $SrO + SO_2 \rightarrow$

B. Basic Oxides

Basic oxides are oxides that are composed of metals and oxygen. Most metals form oxides which exhibit basic properties and dissolve in water to give alkaline solutions. But, all metal oxides are not necessarily basic oxides; for example Al_2O_3 and ZnO are amphoteric oxides, which will be discussed in part (C). Oxides of metals that dissolve in water and react with it to form basic or alkaline solutions are called basic anhydrides. There are metallic oxides which have basic properties but are insoluble in water. These oxides react with acids to give salt and water

Example: $FeO + 2HCl \rightarrow FeCl_2 + H_2O$

The oxides of active metals, group IA and heavier members of group IIA, dissolve in water and readily form bases. The term base is used to describe both soluble and insoluble basic oxides. Some examples of basic oxides are Li_2O, Na_2O, K_2O, MgO, CaO, BaO, and CuO.

Activity 3.3

Discuss the following questions in your group and present your conclusion to the class.

a. What products do basic oxides form in their reactions with water, acidic oxides and acids?
b. What is the type of reaction that takes place between lime (CaO) and carbon dioxide (CO_2)?
c. Write the balanced chemical equation for each of the chemical reactions that occur in (a) and (b).

Chemical properties of basic oxides

1. Basic oxides dissolve in water to form alkaline solutions. As they dissolve, they react with water to form the corresponding metal hydroxides.

Basic oxide + Water → Base (Alkali)

Examples:

$$Li_2O + H_2O \rightarrow 2LiOH$$
$$CaO + H_2O \rightarrow Ca(OH)_2$$

2. Basic oxides react with acidic oxides to form salts.

Basic oxide + Acidic oxide → Salt

Examples: $BaO + SO_3 \rightarrow BaSO_4$
$Na_2O + CO_2 \rightarrow Na_2CO_3$

3. Basic oxides react with acids to form a salt and water.

Basic oxide + Acid → Salt + Water

Examples: $CaO + 2HCl \rightarrow CaCl_2 + H_2O$
$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

Table 3.1 Some common indicators and the colour they develop in acidic and basic solutions

Indicator	Colour in aqueous solution of acidic oxide	Colour in aqueous solution of basic oxide
Universal Indicator	Yellow – orange (in weakly acidic) and red (in strongly acidic)	Blue (in weakly basic) and purple (in strongly basic)
Litmus	Red	Blue
Phenolphthalein	Colourless	Pink (red)
Methyl orange	Red	Yellow

Note that acidic and basic oxides can be identified either by their effects on indicators or chemical interaction they have with one another. Acidic oxides react with bases while basic oxides react with acids. But acidic oxides do not react with acids and basic oxides do not react with bases.

Experiment 3.1
Test for Acidity and Basicity of Oxides

Objective: To identify basic and acidic oxides.
Apparatus: Deflagrating spoon, gas jar and gas jar lid, test tubes.
Chemicals: Sulphur, magnesium or calcium metal, water, universal indicator, litmus paper (blue and red).

Procedure:

1. Take a small amount of powdered sulphur in a deflagrating spoon and heat it.

As soon as sulphur starts burning, introduce the spoon into a gas jar. Cover the gas jar with a lid to ensure that the gas produced does not escape. Remove the spoon after some time. Add 10 mL of water into the gas jar and quickly replace the lid. Shake the gas jar well. Take two test tubes and pour 5 mL of the solution to each test tube. Add a few drops of universal indicator solution to the first test tube and blue litmus paper in the second.

2. Ignite a small amount of magnesium or calcium metal on a deflagrating spoon and insert in to a gas jar. Add 10 mL water to the ash formed and shake. Take two test tubes and pour 5 mL of the solution to each of the test tubes. Add a few drops of universal indicator in the first and red litmus paper to the second test tube.

Questions

a. What compounds are formed by the combustion of sulphur and magnesium or calcium? Write chemical equations to show the reactions.
b. What happens when water is added to the gas jars in which sulphur was burnt?
c. What colours are observed by adding drops of universal indicator and blue or red litmus to the solutions in the test tubes?
d. Why does the change in the colour of indicators occur in the various solutions? Write a laboratory report and submit to your teacher.

Note: Universal indicator and litmus paper serve as indicators. Indicators are substances used to identify whether a given solution is acidic or basic by showing colour changes.

Exercise 3.4

1. Complete and balance the following chemical equations:

a. $K_2O + H_2O \rightarrow$
b. $MgO + H_2O \rightarrow$
c. $Na_2O + CO_2 \rightarrow$
d. $Li_2O + SO_2 \rightarrow$
e. $BaO + H_2SO_4 \rightarrow$
f. $CuO + HCl \rightarrow$

2. Classify the following oxides as basic or acidic:

a. MgO
b. BaO
c. P_4O_{10}
d. N_2O_5
e. Cu_2O
f. Fe_2O_3
g. K_2O
h. SO_2

C. Amphoteric Oxides

Activity 3.4

What trend do you observe in the properties of oxides of the elements in a period as you go from left to right in the periodic table? Which elements form a basic oxide, which elements form acidic oxides? Discuss this in your group and present to the class.

An oxide which shows both acidic and basic behavior is called amphoteric oxide (see *Figure 3.1*).

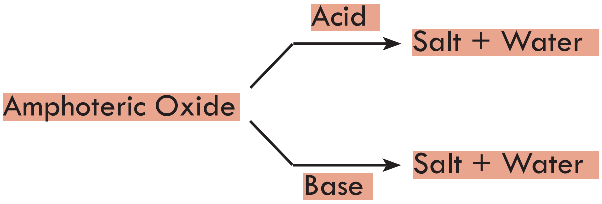
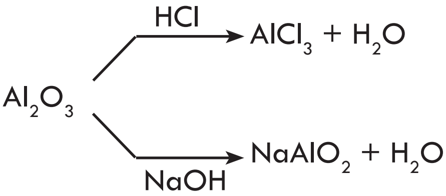


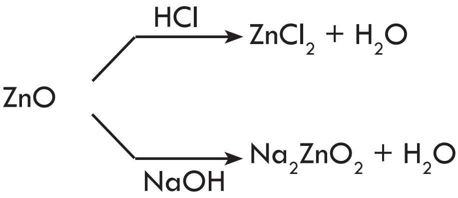
Figure 3.1 The reaction of amphoteric oxide with acids and bases.

For example, oxides of aluminium (Al_2O_3) and zinc (ZnO) are amphoteric. These oxides react with strong bases as well as strong acids.

A. Reactions showing amphoteric nature of aluminium oxide (Al_2O_3)



B. Reactions showing amphoteric nature of zinc oxide (ZnO)



Some other examples of amphoteric oxides are PbO, PbO_2, SnO, and SnO_2. It is also important to realize that hydroxides which react with both acids and bases are described as amphoteric hydroxides. For example, aluminium hydroxide, $Al(OH)_3$, reacts with both acids and bases to form salt and water. So, $Al(OH)_3$, is amphoteric in nature.

What is the common characteristic of acidic, basic and amphoteric oxides? Acidic oxides form salts when reacted with basic oxides and bases. Basic oxides also produce salts in their reactions with acidic oxides and acids. Amphoteric oxides form salts when they react with acids and bases. Thus, acidic oxides, basic oxides and amphoteric oxides are salt-forming oxides.

Exercise 3.5

1. Write the chemical equations to show the amphoteric properties of, Al_2O_3 and PbO when they react with:
   a. NaOH
   b. HNO_3
   c. KOH
   d. H_2SO_4

Experiment 3.2

Investigating Amphoteric Behaviour of Oxides

Objective: To observe the amphoteric behaviour of Al_2O_3.

Apparatus: Spatula, reagent bottles, beakers, and glass rod.

Chemicals: Al_2O_3, HCl, NaOH, Universal indicator and water.

Procedure:

1. Mix 20 mL concentrated HCl and 80 mL water in one reagent bottle;
2. Dissolve 8 g NaOH in 100 mL water in another reagent bottle.
3. Add universal indicator to the acid and base, and observe the colour change.
4. Take two beakers and place a spatula full of Al_2O_3 in each of the beakers.
5. Pour the HCl solution (which you prepared) into one of the beakers and NaOH solution into the other. Stir the mixture with a glass rod.
6. Add universal indicator in the solutions of beakers and observe the colour change.

Questions

1. Does Al_2O_3 react with the solutions in both beakers?
2. What does the change in colour of the indicator in the mixtures indicate?
3. Write the chemical equations to show what has happened?

Write a laboratory report and present to the class.

D. Neutral Oxides

Activity 3.5

1. Explain the acid-base behavior of oxides according to the electropositive character (oxidation state) of the oxide's central atom. Consider all nitrogen oxides
   a. Assign oxidation state for each.
   b. Classify them as soluble and in soluble in water.
   c. Which oxidation state (higher or lower) favors solubility in water?
   d. Which of the nitrogen oxides produce acidic solution in water?
2. In general, oxides of nonmetals are acidic whereas oxides of metals are basic or amphoteric. Can an oxide be neither acidic nor basic?

Discuss in your group and present to the class

Neutral oxides react neither with acids nor with bases to form salt and water. Hence, neutral oxides do not show basic and acidic properties. Examples of neutral oxides are water, H_2O, carbon monoxide, CO, dinitrogen monoxide, N_2O, and nitrogen monoxide, NO. Neutral oxides are very few in number. Neutral oxides are not salt-forming oxides.

E. Peroxides

Activity 3.6

Discuss the following in your group and present your conclusion to the class.

a. Oxides contain O^{2-} ions or the group –O– or =O. What is the oxidation state of oxygen in oxides?
b. What is the oxidation state of oxygen in Al_2O_3, N_2O_5, P_4O_{10}, KO_2 and SO_2?
c. What kind of bonds form oxygen in these oxides?

Compounds which contain oxygen with oxidation number -1 are called peroxides. In peroxides, the two oxygen atoms are linked to each other and with atoms of other elements. Thus, peroxides contain the peroxide, "– O – O –" link. Some examples of peroxides are hydrogen peroxide, H_2O_2, sodium peroxide, Na_2O_2, calcium peroxide, CaO_2, barium peroxide, BaO_2, and strontium peroxide, SrO_2.

Most peroxides of metals are formed by burning the metals in a sufficient amount of oxygen.

Example 3.2:

$$2Na\ (s) + O_2\ (g) \rightarrow Na_2O_2\ (s)$$
$$Ca\ (s) + O_2\ (g) \rightarrow CaO_2\ (s)$$

There are many organic peroxides (ROOR') which are important activators in a process called polymerization, which creates plastic polymers.

Chemical Properties of Peroxides

1. Peroxides are powerful oxidizing agents; they react with different substances by losing oxygen.

Example 3.3:

$$PbS\ (s) + 4H_2O_2\ (aq) \rightarrow PbSO_4\ (s) + 4H_2O\ (l)$$
$$2KI\ (aq) + H_2SO_4\ (aq) + H_2O_2\ (aq) \rightarrow I_2\ (s) + K_2SO_4\ (aq) + 2H_2O\ (l)$$

2. Peroxides react with aqueous acids to form hydrogen peroxide.

Example 3.4:

$$Na_2O_2\ (s) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + H_2O_2\ (aq)$$
$$CaO_2\ (s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (s) + H_2O_2\ (aq)$$

Hydrogen Peroxide

Hydrogen peroxide, H_2O_2, is a colorless liquid whose solutions are usually used as bleach and an antiseptic. Hydrogen peroxide decomposes to release oxygen. This reaction is slow but can be speed up by the addition of catalysts like MnO_2 or Pt.

$$2H_2O_2\ (aq) \rightarrow 2H_2O\ (l) + O_2\ (g)$$

Hydrogen peroxide is a strong oxidizing agent. Its oxidizing power is responsible for its effectiveness as an antiseptic for mouthwash and cleansing wounds. When hydrogen peroxide is added to a colored dye, the molecule responsible for the color will oxidize and so the color will disappear, showing its bleaching action. For example, if hydrogen peroxide is added to a black dye (paint) that contains lead sulphide, PbS, the black colour turns white. This is due to the oxidation of PbS to $PbSO_4$. The equation for this process is:

$$PbS\ (s) + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O\ (l)$$

Experiment 3.3

Distinguishing Peroxides from Other Oxides

Objective: To identify peroxides from other oxides.

Chemicals: CaO or MgO, Al_2O_3 or ZnO or PbO, Na_2O_2 or BaO_2, P_4O_{10}, KI, dilute H_2SO_4, starch, water.

Materials: Reagent bottle, six beakers, dropper, and spatula.

Procedure:

1. Take three beakers. In the first beaker, dissolve 20 g KI in water to prepare a 250 mL solution. Dilute 2 mL concentrated H_2SO_4 by pouring it in 18 mL of water in the second beaker.
2. Prepare starch solution in the third beaker by boiling 1g starch in 100 mL water.
3. Take four beakers and pour 50 mL KI solution to each of them. Acidify the solutions in each of the beakers by adding 5-10 drops of dilute H_2SO_4.
4. Add a spatula full of CaO or MgO in the first, Al_2O_3 or ZnO or PbO in the second, Na_2O_2 or BaO_2 in the third and P_4O_{10} in the fourth beaker and then add about 5 mL starch solution to each of the four beakers.

Questions

1. In which beaker do you see a color change?
2. What is the cause of the colour change?
3. Write a balanced chemical equation for the change?
4. Write a laboratory report and present to the class.

Exercise 3.6

1. Classify the following oxides as acidic, basic, amphoteric, neutral and peroxides

a. Na_2O
b. NO
c. CO_2
d. ZnO
e. MgO
f. N_2O_5
g. CaO
h. CO
i. CaO_2
j. Na_2O_2
k. Li_2O
l. H_3PO_4
m. Al_2O_3
n. SO_2

2. Write an acidic oxide that should react with water to form each of the following acids?
   a. H_2SO_3 c. HNO_3 e. H_3PO_3
   b. H_2SO_4 d. H_2CO_3 f. H_3PO_4
3. Identify the basic anhydrides that react with water to form each of the following bases:
   a. $Ca(OH)_2$ c. $Mg(OH)_2$ e. $Sr(OH)_2$
   b. KOH d. NaOH f. $Ba(OH)_2$
4. Complete and write the balanced equation for each of the following
   a. $MgO + H_2SO_4 \rightarrow$ c. $Al_2O_3 + HNO_3 \rightarrow$ e. $NaOH + CO_2 \rightarrow$
   b. $CaO + H_3PO_4 \rightarrow$ d. $K_2O + H_2SO_4 \rightarrow$ f. $ZnO + H_3PO_4 \rightarrow$
5. How peroxides differ from oxides? Write the structures of H_2O_2, Na_2O_2, CaO_2
6. How can you identify whether an oxide is acidic, basic or neutral?

3.3 Acids

Activity 3.7

Discuss in your group and present it to the class.

a. Describe the reason why unripe citrus fruits or lemon juice give sour taste?
b. List as many sour foods as possible which you have ever tasted.
c. Give some examples of acids that you have encountered in your everyday life?

By the end of this section, students will be able to

- define acids in terms of the concepts of Arrhenius;
- give examples of acids based on Arrhenius;
- categorize acids as monoprotic and poly protic based on the number of ionizable(replaceable) hydrogen atom;
- group acids as binary and ternary based on the number of elements, they contain
- define strong and weak acids;
- differentiate between strong and weak acids;
- define concentrated and dilute acids;
- describe the conceptual difference between strong and concentrated acids;
- explain the conceptual difference between weak and dilute acids;
- apply the necessary precautions while working with acids;
- define pH;
- describe the pH scale;
- identify a given pH-labelled solution as acidic, basic or neutral;

- perform activities to determine the pH of some common substances using universal indicators or pH meter;
- calculate the pH of a given acidic solution;
- calculate the concentration of hydrogen ion from the given information;
- perform activities to investigate some physical properties of acids;
- do activities to investigate some chemical properties of acids;
- discuss the direct combination of elements, the reaction of acidic oxides with water and formation of volatile acids from non -volatile acids as the three methods of preparation of acids;
- carry out simple experiment to prepare acids in laboratory;
- describe the uses of the three common laboratory acids;
- explain the general properties of acids.

Acids play key roles in our bodies, in our homes, and in the laboratories. We seem to like the sour taste of acids; we add them to salad dressings, and spices. Many foods, including citrus fruits and some vegetables, contain acids.

Acetic acid in vinegar, citric acid in lemons and other citrus fruits, are among the acids that we encounter every day. Hydrochloric acid is the acid in gastric juice; it is essential to digestion. Phosphoric acid gives flavour to many carbonated beverages. In Latin the word *acidus* is used for sour. In the study of chemistry the term acid has been used to name a group of compounds which have a sour taste.

Arrhenius Definition of Acids

The simplest definition of acids is suggested by Savante Arrhenius, a Swedish Chemist. Arrhenius defined an acid as a substance that releases hydrogen ion or proton, H^+, or hydronium ion, H_3O^+, in aqueous solution. In general, the ionization of acids takes place as follows.

$$HA(aq) \rightarrow H^+(aq) + A^-(aq) \quad \text{or} \quad HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$$

Example 3.5:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq) \text{ or } HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

$$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq) \text{ or } H_2SO_4(aq) + 2H_2O(l) \rightarrow 2H_3O^+(aq) + SO_4^{2-}(aq)$$

According to this definition, for a substance to be an Arrhenius acid, it should contain ionizable hydrogen. Some examples of Arrhenius acids are HNO_3, H_2SO_4, H_3PO_4, HBr, HI and HF.

Although the Arrhenius concept has limited scope yet, it led to the development of more general theories of acid-base behavior. Broader Acid-base definitions will be dealt in grade 12 chemistry.

Classification of Acids

Activity 3.8

Sulphuric acid, H_2SO_4, has two hydrogen atoms and it is a diprotic acid. Propionic acid, CH_3CH_2COOH, has six hydrogen atoms but is classified as a monoprotic acid.

Discuss the reason in your group and present it to the class.

A. Classification of Acids Based on the Number of Ionizable Hydrogen

Acids are classified as monoprotic, diprotic and polyprotic depending upon the number of replaceable hydrogen ions (H^+) or hydronium ions (H_3O^+ ions) per molecule of the acid.

An acid that contains only one ionisable hydrogen atom per molecule is called monoprotic.

Example 3.6: Hydrochloric acid (HCl) and nitric acid (HNO_3) are monoprotic acids because each give only one H^+ ion upon dissociation in aqueous solution.

$$HCl\ (aq) \longrightarrow H^+\ (aq) + Cl^-\ (aq)$$
$$HNO_3\ (aq) \longrightarrow H^+\ (aq) + NO_3^-\ (aq)$$

Polyprotic acids are those acids containing more than one ionizable (replaceable) hydrogen ion in aqueous solution. Polyprotic acids which contain two ionizable hydrogen atoms such as H_2SO_4, H_2S, and H_2CO_3 are called diprotic acids; those containing three ionizable hydrogen atoms like H_3PO_4 are called triprotic acids. The dissociation of polyprotic acids in aqueous solution is shown in the following chemical equations:

$$H_2SO_4\ (aq) \longrightarrow 2H^+\ (aq) + SO_4^{2-}\ (aq)$$
$$H_2CO_3\ (aq) \rightleftharpoons 2H^+\ (aq) + CO_3^{2-}\ (aq)$$
$$H_3PO_4\ (aq) \rightleftharpoons 3H^+\ (aq) + PO_4^{3-}\ (aq)$$

Activity 3.9

Ammonia (NH_3) and phosphoric acid (H_3PO_4) have three hydrogen atoms but ammonia does not behave as a triprotic acid. It is a base. Why? Discuss in your group and present it to the class.

B. Classification of Acids Based on the Number of Elements they Contain

Acids can also be classified depending on the number of their constituent elements as binary and ternary acids. Binary acids are those acids composed of only two different elements. The two elements are hydrogen and a non-metal.

Examples of binary acids are HCl, HBr, HF, HI, and H_2S.

Ternary acids also called oxy-acids are acids composed of three different elements. The three elements usually are hydrogen, oxygen and a non-metal. Examples are H_2SO_4, H_2CO_3, $HClO_4$, and H_3PO_4.

Activity 3.10

1. Add a drop of lemon juice on your tongue and try to identify its taste. Repeat your observation with orange juice. Try to classify them as acid or base, depending on their taste.
2. Why should curd and acid-containing substances not be kept in brass and copper vessels? (Brass is an alloy of copper and Zinc). Discuss in class.

Acids generally have the following properties:

General Properties of Acids

I. **Acids have a sour taste.** Aqueous solutions of acids have a sour taste. Lemon juice and orange juice taste sour due to the presence of citric acid. Citric acid in lemon juice and orange juice is harmless. However, concentrated acids are corrosive and poisonous. So the test 'taste' is never allowed to identify acids especially in the laboratories.

II. **Acids change the color of certain acid-base indicators.** The common indicators available in high school laboratories are litmus, phenolphthalein, methyl orange or methyl red and universal indicator. Note that natural dyes like red extract of cabbage leaves, petals of some colored flowers and turmeric also act as indicators to test the presence of acid or base.

Experiment 3.4

Effect of Acids on Indicators

Objective: To detect the acidity of a solution using indicators.
Chemicals: Lemon juice, dilute HCl, dilute HNO_3, dilute H_2SO_4 phenolphthalein, litmus, methyl red, universal indicator
Apparatus: Test tubes, test tube rack, test tube holder, and reagent bottles.

Procedure:

Take four clean test tubes and place some lemon juice in the first, dilute HCl in the second, dilute HNO_3 in the third and dilute H_2SO_4 in the fourth. Dip a strip of blue litmus paper into each of the four test tubes and observe. Follow the same procedure and repeat the experiment until each acid has been tested by each indicator. Record your observation.

Questions

1. What colours have you observed when each indicator was added to each of the four acid solutions? Record the colour change of each indicator in each solution (lemon juice, dilute HCl, HNO_3 and H_2SO_4) using the following table.
2. Why do all the acids show similar color changes in the same indicator?
3. What can you say about the acid-base behavior of indicators? Are indicators acidic or basic ?

Table 3.2 Colour of indicators in lemon juice, dilute HCl, HNO_3 and H_2SO_4

Indicator	Colour of indicator in			
	Lemon juice	Dilute HNO_3	Dilute HCl	Dilute H_2SO_4

III. **Acids react with active metals to yield hydrogen gas.** Acids react explosively with metals like sodium, potassium and calcium. The reaction is very dangerous and should not be performed. However, dilute acids (HCl, H_2SO_4) react moderately with reactive metals like: Mg, Zn, Fe and Al to form their respective salts with the evolution of hydrogen gas.

Example 3.7:

$$H_2SO_4 + Zn \rightarrow ZnSO_4 + H_2$$
$$6HCl + 2Al \rightarrow 2AlCl_3 + 3H_2$$

Concentrated nitric acid and hot concentrated sulphuric acid react with copper producing nitrogen dioxide and sulphur dioxide gases, respectively, instead of hydrogen. This is because concentrated HNO_3 and hot concentrated H_2SO_4 are oxidizing acids. The reactions of these acids with copper are given by the following equations:

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$
$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

Experiment 3.5

Investigating the Reactions of Metals with Dilute Acids

Objective: To investigate the reaction between active metals and dilute acids.
Chemicals: Dilute H_2SO_4, dilute HCl, zinc, magnesium and iron
Apparatus: Test tubes, test tube holder, test tube rack, burner, match, cork and spatula.

Procedure:

Take three test tubes and place a spatula-full of powdered zinc in the first, powdered magnesium in the second and iron filings in the third. Pour dilute HCl into each of the test tubes until the metals are completely covered by the acid. To test the type of gas evolved, cover one of the test tubes with a cork which has a glass delivery (jet tube as shown in *Figure 3.2*). Bring a burning match stick close to the glass tube mouth, the gas burns with a pale blue flame with a characteristic sound. Repeat the experiment with dilute H_2SO_4 after placing each of the three metals in three different test tubes.

Questions:

A. What does the formation of bubbles indicate?
B. What sound do you hear when the burning splint is close to the mouth of the test tube?
C. Which gas is evolved during the reaction?
D. Which metal's reaction with dilute HCl or H_2SO_4 is the most violent?
E. Write a laboratory report on your observations and present to the class.

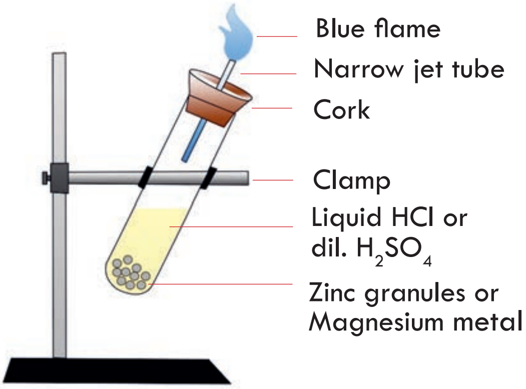


Figure 3.2 Action of dilute acids on relatively active metals.

Acids react with carbonates and hydrogen carbonates to form salt, water and carbon dioxide gas

Acid + Hydrogen carbonate → Salt + Water + Carbon dioxide
Acid + Carbonate → Salt + Water + Carbon dioxide

Example 3.8

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

$$HCl + NaHCO_3 \rightarrow NaCl + H_2O + CO_2$$

Experiment 3.6

Reactions of Acids with Carbonates and Hydrogen Carbonates

Objective: To investigate the reaction between acids and carbonates and bicarbonates.
Chemicals: Dil. H_2SO_4, dil. HCl, Zinc, Magnesium, iron
Apparatus: Test tubes , test tube holder, test tube rack, burner , match, cork and spatula

Procedure:

1. Take three conical flasks and add powdered Na_2CO_3 in the first, powdered $CaCO_3$ to the second and powdered sodium bicarbonate to the third. Pour dilute HCl into each of the three conical flasks until the acid covers the carbonates and bicarbonate. Hold damp blue litmus paper close to the mouth of the conical flasks. Repeat this with damp red litmus paper and record your observations. Bubble the gas through limewater as shown in ***Figure 3.3***.
2. Repeat the experiment using the same carbonates and hydrogencarbonate with dilute H_2SO_4 and dilute HNO_3.

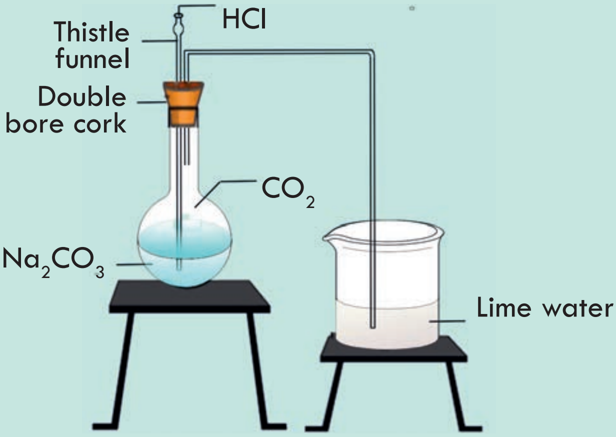


Figure 3.3 Test for Carbon dioxide.

Questions

A. What does the formation of bubbles in the conical flasks indicate?

B. Did the colour of the damp blue litmus paper change when you held it close to the mouth of the conical flasks? What about the colour of damp red litmus?

C. What happened to the colour of lime water when you bubbled the gas through it? If there was any change, what did it prove? Write a balanced chemical equation for the change?

Write a laboratory report on your observation and present to the class.

IV. Concentrated acids react with sulphites to form salts with the liberation of sulphur dioxide gas.

Sulphite + Acid → Sulphur dioxide + Salt + Water

$CaSO_3 + H_2SO_4 \rightarrow SO_2 + CaSO_4 + H_2O$

$NaHSO_3 + HCl \rightarrow SO_2 + NaCl + H_2O$

Experiment 3.7

Neutralization Reaction

Objective: To investigate the reaction between acids and bases

Chemicals: 1 M HCl, 1 M NaOH, 0.5 M $Ca(OH)_2$, red and blue litmus papers

Apparatus: Four 150 mL beakers, two droppers, stirring rod, two watch glasses.

Procedure

1. Make 1 M NaOH solution by dissolving 4.0 g NaOH in enough water to make 100 mL solution.
2. Make a 1 M HCl solution by dissolving 8.3 mL of concentrated HCl in enough distilled water until the volume of the solution is 100 mL.
3. Dissolve 3.7 g of $Ca(OH)_2$ in enough distilled water to make 100 mL solution.
4. To a 150 mL beaker add 10 mL HCl solution and 9.5 mL NaOH solution and stir thoroughly and test with blue and red litmus paper. Continue adding NaOH solution dropwise using a dropper, stirring after each addition and checking with red and blue litmus until the blue remains blue and the red stays red. Put 2 mL of the neutral solution in a watch glass and allow the water to evaporate until the next day.
5. To another 150 mL beaker add 10 mL $Ca(OH)_2$ solution and 9.5 mL HCl solution. Repeat all the steps, which you followed in procedure 4.

Questions

A. Is there any colour change, when you dip blue and red litmus papers into the solution of the acid and the base?

B. Why is it necessary to add NaOH solution dropwise (one drop at a time) in procedure 4?

C. During this experiment, under what conditions does the blue litmus remain blue and the red remain red?

D. What are the products formed in procedures 4 and 5? Write balanced chemical equations for the reactions?

F. Write a laboratory report on your observations and submit to your teacher.

V. Acids react with bases (oxides and hydroxides of metal and ammonium hydroxide) to form salts and water. The reaction of acids with basic oxides or bases to form salt and water is called neutralization reaction

Acid + Base (basic oxides or hydroxides of metal) → Salt + Water

Example 3.9: reaction of acids with basic oxides

$$2HCl + MgO \rightarrow MgCl_2 + H_2O$$
$$H_2SO_4 + CaO \rightarrow CaSO_4 + H_2O$$

Example 3.10: Reaction of acids with hydroxides of metals

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$
$$HNO_3 + KOH \rightarrow KNO_3 + H_2O$$

VI. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great number of ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.

Example 3.11: A solution of hydrochloric acid (HCl) conducts electricity as it contains replaceable H atom which is furnished in solution as H^+ ion and carries the current.

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

HCl is a strong electrolyte because it gives large amount of ions in aqueous solution.

Strong electrolytes give bright light. CH_3COOH gives few ions in aqueous solution; hence it is a weak electrolyte. *Weak electrolytes give dim light.*

Experiment 3.8

Can Aqueous Acid Solution Conduct Electricity?

Objectives: To investigate the conductivity of acids

Chemicals: 1 M HCl, 1 M HNO_3, 1 M CH_3COOH,

Apparatus: Three 150 mL beakers, two graphite electrodes, insulated electric wires, two dry cells, and bulb

Procedure:

Take about 100 mL of 1 M HCl in the first beaker, the same volume and concentration of HNO_3 and CH_3COOH in the second and third beakers respectively.

Arrange the set up as shown in *Figure 3.4*. First test the conduction of electricity through HCl solution by inserting the two graphite electrodes. See whether the bulb glows or not. Is the light bright or dim? Record your observation. Repeat the same activity with HNO_3 and CH_3COOH. Rinse the electrodes after use in each of the acid solutions. Compare the intensity of light produced with HNO_3 and CH_3COOH.

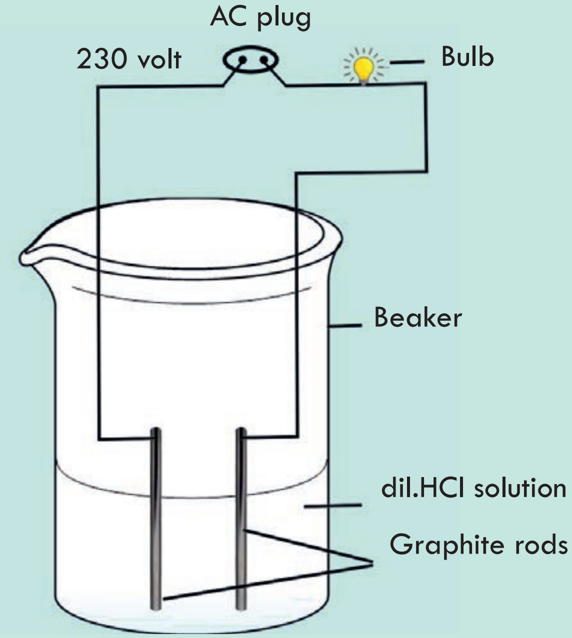


Figure 3.4 Conductivity of acid solutions.

Questions

A. In which of the acid solutions does the bulb produce?
   I) bright light? II) dim light?

B. What conclusions can you draw from your observations?

Write a laboratory report on your observations and present to the class.

Strength of Acids (Strong and Weak Acids)

Activity 3.11

Discuss in your group and present your conclusion to the class.

1. Add a few drops of citric acid solution and hydrochloric acid separately on a piece of waste cotton cloth; What happens? Based on your observations, can you classify these acids as weak acid or strong acid?
2. Why is it safe to put acetic acid-containing vinegar on food yet extremely dangerous to use dilute nitric acid for the same purpose?

Based on their degree of dissociation, acids are divided into strong acids and weak acids. If an acid completely dissociates in an aqueous solution, it is called a strong acid. Strong acids produce large amount of H^+ ions in the aqueous solution as they dissociate completely. Hydrochloric acid (HCl), nitric acid (HNO_3), and sulfuric acid (H_2SO_4) are some examples of acids that dissociate completely in aqueous solution. Their dissociation is indicated by a single arrow (→) as shown below:

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$
$$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$$
$$H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2-}\ (aq)$$

A dilute aqueous solution of strong acids contains predominantly the ions derived from the acids instead of the acid molecules. For example, HCl and HNO_3 are almost completely ionized in water.

An acid is called weak if it dissociates only partially in aqueous solution. Weak acids release few hydrogen ions in aqueous solution. The aqueous solution of a weak acid contains hydronium ions, anions and dissolved molecules of the acid. Examples of weak acids are acetic acid (CH_3COOH), oxalic acid ($H_2C_2O_4$), carbonic acid (H_2CO_3). These acids dissociate only partially as indicated by the double arrow ($\rightleftharpoons$) shown below:

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$
$$H_2C_2O_4 \rightleftharpoons 2H^+ + C_2O_4^{2-}$$
$$H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{2-}$$

Organic acids which contain the acidic carboxyl group – COOH, are generally weak acids.

Exercise 3.7

1.
   a. Give examples of organic acids and inorganic acids.
   b. Classify your examples as strong and weak acids
2. Why do HCl, HNO_3, etc. Show acidic characteristics in aqueous solutions while solutions of compounds like alcohol and glucose do not show acidic property ?

Concentrated and Dilute Acids

Activity 3.12

Discuss in your group and present your conclusion to the class.

a. In a solution, if there is 96% H_2O and the rest is H_2SO_4, what kind of solution is it?
b. In a solution, if there is 96% H_2SO_4 and the rest is H_2O, what kind of solution is it?

The terms "concentrated" and "dilute" are used to describe the relative amount of acid present in a given solution. A concentrated acid has a relatively large amount of solute dissolved in the solvent. A dilute acid has a relatively smaller amount of solute dissolved in the solvent. Both strong acid and a weak acid may be concentrated or diluted depending on the number of moles of acid present.

Precautions in Handling Acid

Concentrated acids are corrosive and may cause visible destruction or irreversible alterations to living tissue by chemical action at the site of contact. Hence, the following safety precautions should be taken when handling acids:

a. Wear goggles, gloves and a laboratory coat
b. If a concentrated acid is spilled or splashed on your body, first wash the affected part with running water and then with 10% Na_2CO_3 solution
c. To dilute concentrated acids, pour the concentrated acid in to water and not water in to the acid.
d. If corrosive acids are swallowed, administer weak bases such as $Mg(OH)_2$ or $Al(OH)_3$.
e. Use bellows to pipette acid instead of sucking using yours lips
f. If an acid enters your eye, wash with water repeatedly and then consult a doctor
g. If concentrated acid is spilled on to cloth, immediately wash it with running water

pH and pH scale

The molar concentrations of H_3O^+ in aqueous solutions are generally very small and expressing acidity as the concentration of H_3O^+ becomes cumbersome. Instead, a

more convenient quantity called pH is used as a measure of the acid strength of a solution. pH scale is used to rank solutions in terms of acidity or basicity (alkalinity). Concentration of hydrogen ion $[H^+]$ in pure water is the basis for the pH scale. Water is a weak electrolyte because it ionizes very slightly into ions in a process called autoionization or self-ionization; $H_2O \rightleftharpoons H^+ + OH^-$
The ionic product constant of water 'Kw' is given as

$$Kw = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25\ ^\circ C$$

It is known that the amount of $[H^+]$ and $[OH^-]$ in pure water are equal

$$[H^+] = [OH^-] \quad \text{or} \quad [H^+]^2 = 1.0 \times 10^{-14}$$
$$[H^+] = 1.0 \times 10^{-7} \quad \text{at } 25\ ^\circ C$$

As it is difficult to deal with such small figures having negative exponents, it is convenient to convert these figures into a positive figure using a numerical system. It is taking the common (base-10) logarithm of the figure and multiplying it with -1. 'p' before a symbol means' negative logarithm of the symbol. Therefore, pH is the negative logarithm of molar concentration of the hydrogen ions.

$$pH = -\log[H^+]$$

With reference to this equation, a scale develops according to the molar concentration of H^+ ions that is called pH scale. It ranges from 0 to 14. According to this scale, pH of water is calculated as: $pH = -\log[H^+] = -\log(1\times10^{-7}) = 7$,

similarly $pOH = -\log[OH^-] = -\log(1\times10^{-7}) = 7$

pH value normally varies from 0 to 14. Therefore:
$pH + pOH = 14$

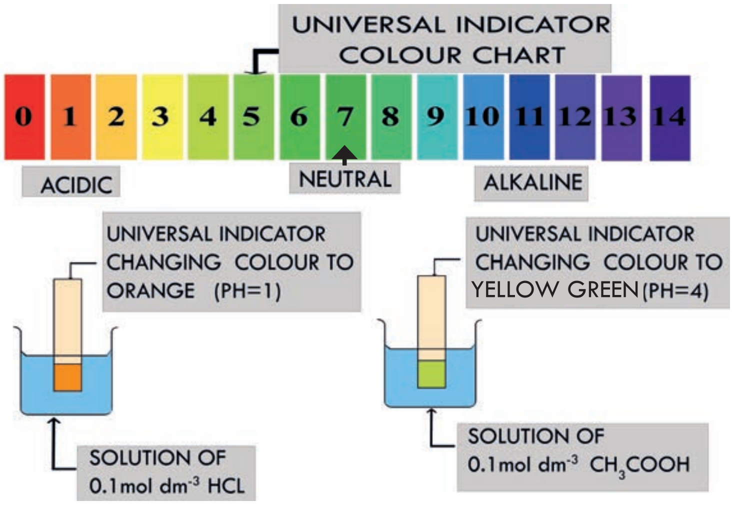


Figure 3.5 Measurement of pH using a universal indicator.

Hence, the sum of the pH and pOH of the solution is always 14 at 25 °C. The pH of a solution can be measured using a universal indicator solution (see ***Figure 3.5***). When universal indicator solution is added to an acid solution, it develops different colors depending on the pH of the solution. In order to identify the pH, we need to compare the color is developed with the standard color chart. A solution of a compound with pH 7 or pOH 7 is considered as a neutral solution. Solutions of pH less than 7 are acidic and those more than 7 are basic. pH values of some common items are given in ***Figure 3.6*** as examples.

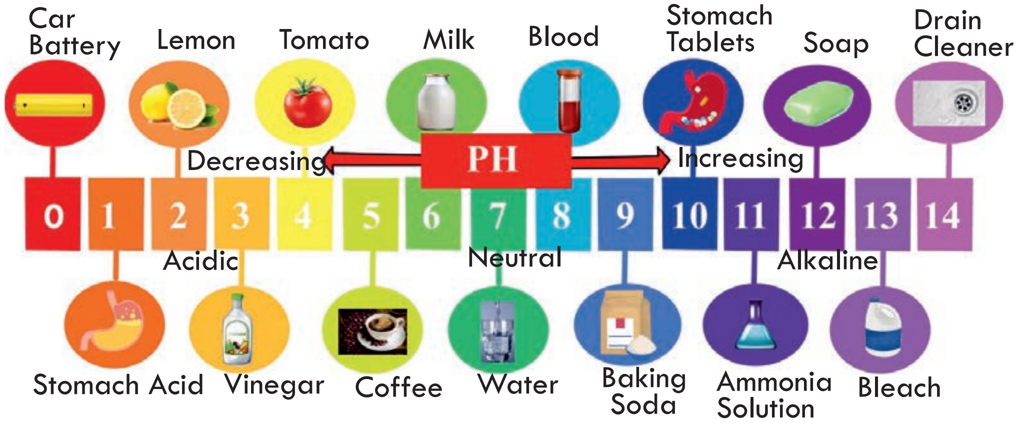


Figure 3.6 pH values of some common items.

As the pH value decreases the acidity of the solution increases. For example, if two solutions A and B have pH values of 4 and 6, respectively, then we can conclude that solution A is more acidic than solution B. The hydrogen ion concentration in solution A is one hundred times greater than that in B.

Example 3.12: Calculate the pH value of a solution when hydrogen ion concentration is 1.0×10^{-5}mo L^{-1}

Solution: pH is related to hydrogen ion concentration by an equation

$pH = -\log[H^+]$

It is given that $[H^+] = 1.0 \times 10^{-5}\ mol\ L^{-1} = 10^{-5}\ mol\ L^{-1}$

$pH = -\log[H^+] = -\log(10^{-5})$

$= -(-5 \log (10) = 5 \times 1 = 5$

Example 3.13: Calculate the hydrogen ion concentration of an aqueous solution

when its pH is 9.

Solution: The concentration of hydrogen ion and the pH value of an aqueous solution are related by $pH = -\log [H^+]$

$-pH = \log [H^+]$

$[H^+] = \text{antilog } (-pH) = 10^{-9} = 1.0 \times 10^{-9}\ mol\ L^{-1}$

The Relationship Between pH and Concentration

Since the pH scale is logarithmic, a solution of pH 1 has 10 times higher concentration of $[H^+]$ than that of a solution of pH 2; 100 times than that of a solution of pH 3 and so on. Hence, low pH value means strong acid while high pH value means a strong base and vice versa.

Example 3.14: A solution of hydrochloric acid is 0.01M. What is its pH value?
Solution: Hydrochloric acid is a strong acid so it ionizes completely.

That is: $HCl(aq) \rightarrow H^+ (aq) + Cl^- (aq)$

0.01 0.01 0.01

$$pH = -\log[H^+]$$

By putting the values of H^+ ions in the above equation:

$$pH = -\log[H^+] = -\log(1.0 \times 10^{-2}) = 2$$

Example 3.15: The hydrogen ion concentration in a dilute solution of citric acid is 1×10^{-6} M. What is the pH of the solution?

Solution: $pH = -\log[H^+] = -\log 1 \times 10^{-6} = 6$

Example 3.16: Calculate the pH value of 0.05 M H_2SO_4. Assume that H_2SO_4 is completely ionized.
Solution: H_2SO_4 is a strong acid and ionizes as

$$H_2SO_4\,(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$$

0.05 2×0.05

Molarity of $H^+ = [H^+] = 0.1 \text{ mol L}^{-1}$

$$pH = -\log(H^+) = -\log(0.1) = 1$$

Exercise 3.8

1. Find the pH of 0.01M sulphuric acid?
2. HCl and H_2SO_4 are strong acids. Why do the equimolar solutions of the two acids have different pH?
3. Two solutions A and B have pH values of 2 and 6 respectively. How many times greater is the hydrogen ion concentration in solution A than that of solution B?

Experiment 3.9

pH of Solutions of Common Substances

Objective: To determine the pH of different substances

Chemicals: Lemon juice, vinegar, tonic water, tomato juice

Materials: Beakers, and universal indicator solution or pH indicator paper

Procedure:

Take four beakers and place lemon juice in the first, vinegar solution in the second, tonic water in the third and filtered tomato juice solution in the fourth. Then add a few drops of universal indicator solution or dip a piece of pH indicator paper into each of the solutions. Compare the colour developed with standard colour chart to decide the pH of each solution.

Questions:

a. What is your conclusion based on your observations?

b. Were the substances used in this experiment acidic or neutral? Why?

c. Record your observations using the following Table:

Table 3.3 Colour and corresponding pH.

Substance	Color developed	pH
Lemon juice		
Vinegar solution		
Tomato Juice		
Tonic water		

Write a laboratory report on your observations and present to the class.

Preparation of Acids

Acids can be prepared by:

1. The reaction of oxides of non-metals (acidic oxides) and water:

Acidic oxide + Water → Acid

Example 3.17: $N_2O_5\ (s) + H_2O\ (l) \rightarrow 2HNO_3\ (aq)$

$P_4O_{10}\ (s) + 6H_2O\ (l) \rightarrow 4H_3PO_4\ (aq)$

2. Direct combination of some non-metals like S and Cl with hydrogen:

This method is mostly used to prepare binary acids (acids consisting only two elements

Nonmetal + Hydrogen → Binary acid

Example 3.18: $H_2 (g) + Cl_2 (g) \rightarrow 2HCl (g)$

$H_2(g) + S (s)) \rightarrow H_2S (g)$

When gaseous hydrogen chloride and hydrogen sulphide dissolve in water, they form hydrochloric acid, and hydrosulphuric acid respectively.

3. Using a non-volatile acid: Volatile acids can be prepared by heating their salts with a non-volatile acid .For example, concentrated sulphuric acid (H_2SO_4) is a nonvolatile acid and it is used for the preparation of volatile acids Hydrochloric acid (HCl) and nitric acid (HNO_3) according to the following equations:

$NaCl (s) + H_2SO_4 (l) \rightarrow NaHSO_4 (s) + HCl (l)$

Nonvolatile acid Volatile acid

$NaNO_3 (s) + H_2SO_4 (l) \rightarrow NaHSO_4 (s) + HNO_3(l)$

Nonvolatile acid Volatile acid

Experiment 3.10

Preparation of Chlorous Acid

Objective: To investigate the product formed from $Ba(ClO_2)_2$ and H_2SO_4

Chemicals: $Ba(ClO_2)_2$, concentrated H_2SO_4, water, blue, red litmus papers, methyl red.

Materials: Two 250 mL beakers, glass rod, dropper, test tubes, test tube rack.

Procedure:

1. Dissolve 12 g $Ba(ClO_2)_2$ to prepare 100 mL solution in one beaker. Dilute 2 mL concentrated H_2SO_4 with water to prepare 50 mL of dilute H_2SO_4 solution in the second beaker.
2. Add 80 mL of $Ba(ClO_2)_2$ solution to the dilute solution of H_2SO_4 in the second beaker. Is there any formation of the white precipitate? What do you think is the precipitate formed? Wait for some time till the precipitate settles. Continue adding $Ba(ClO_2)_2$ solution using a dropper, one drop at a time until formation of white precipitate stops. After all the precipitate settles, take 5 mL of the clear liquid in a test tube and test with litmus or methyl red.

Questions:

a. Is the final solution acidic or basic?
b. Which acid is formed?
c. Write a balanced chemical equation for the reaction.

Write a laboratory report as a group and present to the class

Use of Some Important Acid

Activity 3.13

1. List uses of some common acids in your daily life.
2. High consumption of sulphuric acid in a country indicates the economic growth of the country. What is the reason? Discuss in groups and present to the class.

Table 3.4 Uses of the common acids H_2SO_4, HCl and HNO_3

Acid	Use
Sulphuric acid, H_2SO_4	H_2SO_4 is the leading industrial chemical. It is used: * In production of sulphate and phosphate fertilizers, synthetic, fibers, paints, drugs, detergents, paper and dyes * In petroleum refining * In production of metals * As electrolyte in car batteries
Hydrochloric Acid, HCl	* It is present naturally in the gastric juice of our body and helps in the digestion of food * Industrially, HCl is used for pickling of iron and steel (to remove surface impurities) before galvanizing and tin plating * Used to produce aniline dyes, drugs, photographic films, plastics like polyvinyl chloride (PVC) * Used to recover magnesium from sea water
Nitric Acid, HNO_3	Used industrially in the manufacture of * Explosives such as trinitrotoluene (TNT) and trinitroglycerine * Fertilizers such as KNO_3 and NH_4NO_3 * Rubber, chemicals, plastics, dyes and drugs.

Reading Assignment

Read from reliable sources the uses of common acids H_3PO_4, H_2CO_3, CH_3COOH and discuss in class with your group

Exercise 3.9

1. Which one of the following substances are Arrhenius acids:
   a. $HBrO_4$
   b. BCl_3
   c. H_2S
   d. HF
   e. HBr
   f. PCl_5

2. Classify the following acids as monoprotic, diprotic or triprotic, binary or ternary, strong or weak acids. (give three answers for each type):
   a. $HClO_4$
   b. H_2SO_4
   c. H_2CO_3
   d. H_2SO_3
   e. HF
   f. H_2S
   g. HCl
   h. H_3PO_4
   i. HNO_2
   j. HCN
   k. HNO_3
   l. CH_3COOH
3. A reagent bottle (labelled as A) is filled with HCl solution and the other (labelledas B) is filled with water. Both liquids in the bottles are colourless. What method do you recommend to identify the acid and water?
4. What is the basis for the classification of acids as strong and weak?
5. What is the pH of a solution having the following hydrogen ion concentrations?
   a. 5×10^{-3} M
   b. 0.003 M
   c. 2.0×10^{-6} M
6. Calculate the hydrogen ion concentrations in solutions having the following pH values:
   a. 4
   b. 2
   c. 5
7. How many moles of H_2SO_4 are present in 0.500 L of a 0.150 M H_2SO_4 solution?
8. Identify the following solutions as acidic or basic, estimate $[H_3O^+]$ values for each, and rank them in order of increasing acidity:
   a. Saliva, pH =6.5
   b. Orange juice, pH = 3.7
   c. Pancreatic juice, pH = 7.9
   d. Wine, pH = 3.5

3.4 Bases

Activity 3.14

Discuss the following questions in your group and present your conclusions to the class.

1. What is the main ingredient in stomach antacid ?
2. Make a list of some bitter-tasting foods and beverages.
3. Have you ever wondered about the origin of the saying “it is a bitter pill (kosso) to swallow”? Discuss and give a meaning to the saying.
4. Have you ever had soap in your mouth? How does it taste?

After completing this section, you should be able to

- define bases in terms of the concepts of Arrhenius;
- give examples of bases based on Arrhenius;
- discuss the general properties of bases;

- define strong and weak bases;
- distinguish between strong and weak alkalis (soluble bases);
- define concentrated and dilute alkalis;
- distinguish between concentrated and dilute alkalis (soluble bases);
- use the necessary precautions while working with bases;
- define pOH;
- show the mathematical relationship between pH and pOH;
- calculate the pOH of a given basic solution;
- calculate the concentration of hydroxide ion from the given information;
- carry out activities to investigate some chemical properties of bases;
- discuss the reaction of active metals with water, the reaction of basic oxides with water and double displacement reactions as the three methods of preparation of bases;
- carry out simple experiments to prepare bases in laboratory;
- explain the uses of the two common laboratory bases.

Bases are either oxides or hydroxides of metals, which are therefore called basic oxides and hydroxides respectively. They are of great importance in chemical industries and in our daily lives, directly or indirectly. For example, sodium hydroxide, NaOH is used in the production of soap, paper, textile etc.

Potassium hydroxide, KOH is used to produce soft soap, fertilizers etc. Calcium hydroxide, $Ca(OH)_2$, is used to manufacture mortar and bleaching powder, to remove soil acidity etc.

Arrhenius Definition of Bases

According to the Arrhenius theory, a base is a substance that produces a hydroxide ion (OH^-) when it is dissolved in water. Soluble bases are called **alkalis**.

For example, sodium hydroxide and potassium hydroxide are bases because they produce OH- ion when dissolved in water.

$$NaOH\ (aq) \rightarrow Na\ (aq)^+ + OH^-(aq)$$
$$KOH\ (aq) \rightarrow K^+\ (aq) + OH^-\ (aq)$$

Note that the ions produced by the base can come from either of two sources. Metal hydroxides, such as NaOH, KOH, and $Ba(OH)_2$, are ionic compounds that already contain ions and merely release those ions when they dissolve in water. Ammonia, however, is not ionic and contains no ions in its structure. Nonetheless, ammonia is a base because it undergoes a reaction with water when it dissolves, producing NH_4^+ and OH^- ions:

$$NH_3\ (aq) + H_2O\ (l) \rightleftharpoons NH_4OH(aq) \rightarrow NH_4^+\ (aq) + OH^-(aq)$$

The reaction of ammonia with water is a reversible process whose equilibrium strongly favors unreacted ammonia.

General Properties of Bases

Activity 3.15

1. What is the taste of carbonated beverage, coffee and tonic water? How can it be differ from the taste of lemon juice or vinegar?
2. Explain the reason why a dilute solution of sodium bicarbonate can be used to treat bee stings, whereas a dilute solution of acetic acid can be used to treat wasp stings.

Discuss with your group and present your conclusion to the class.

1. Alkalis have a soapy feeling and a bitter taste: Bases have a bitter taste; feel slippery like soap to the skin in dilute aqueous solutions. Strong bases such as NaOH and KOH are very corrosive and poisonous. So they should be neither brought into contact with the skin nor tasted.

2. Alkalis change the color of indicators: Aqueous solutions of bases turn the color of red litmus to blue, phenolphthalein to pink (red), methyl red to yellow and universal indicator blue (purple). This property is because of the OH^- ion given by the base.

3. Alkalis release hydroxide ion in aqueous solution: The characteristic properties of bases in aqueous solutions are due to the presence of the hydroxide ion, OH^- which they release on dissolution.

Example 3.19: $NaOH\ (aq) \rightarrow Na\ (aq)^+ + OH^-(aq)$
$KOH\ (aq) \rightarrow K^+\ (aq) + OH^-(aq)$

4. Bases neutralize acids or acidic oxides to form salt and water:

Example 3.20: Base + Acid (or Acidic Oxide) → Salt + Water
$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$
$2NaOH\ (aq) + CO_2\ (aq) \rightarrow Na_2CO_3\ (aq) + H_2O(l)$

5. Aqueous solutions of bases conduct electricity: Soluble bases are electrolytes. Solutions of strong bases are good conductors while solutions of weak bases are poor conductors

Activity 3.16

Discuss with your group and present your conclusion to the class.

1. What is the difference between concentrated base and strong base, dilute base and weak base? Can a strong base be a dilute base and a weak base concentrated base? Discuss in class.
2. Why solutions of weak bases give dim light and solutions of strong bases give bright light? Discuss in class.

Strength of Bases (Strong and Weak Bases)

A strong base is almost completely ionized in aqueous solution to give its constituent cations and anions. That is the extent of ionization of a strong base is very high. Thus, the amount of hydroxide ions (OH^-) in the solution of a strong base is large.

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

Hydroxides of alkali (Group IA) metals and lower members of alkaline earth metals such as LiOH, NaOH, KOH, and $Ba(OH)_2$ are examples of strong bases. Strong bases are strong electrolytes as they produce large amount of ions and conduct high current. A weak base is only partially ionized into its cations and anions in aqueous solution. That is, the extent of ionization of a weak base is small. Thus, the amount of hydroxide ions (OH^-) in the solution of a strong base is small.

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

The double arrow shows that the dissociation of a weak base does not proceed to completion. This means an aqueous solution of a weak base contains only a small amount of ions derived from the dissociation of the base and a large amount of the non-ionized base. For example, a solution of ammonia containing 0.1 mole NH_3 per litre of solution ionizes only to the extent of 1.3%. Weak bases are weak electrolytes as they allow low current. Examples of weak bases include $Mg(OH)_2$ and $Ca(OH)_2$.

Concentrated and Dilute Bases

The concentration of an alkali is a measure of the number of moles of the alkali dissolved in 1 litre of solution and is therefore expressed in mol per liter. Concentrated bases contain relatively large amounts of a base in a given volume of solution while dilute base solutions contain only a small amount of base. The concentration is expressed in terms of mole per litre (Molarity).

The greater the number of moles of the base per liter of the solution, the more concentrated is the solution. For example, a solution containing ten moles of NaOH per litre is more concentrated than the solution containing two moles of NaOH per liter. The latter solution is more dilute than the former. Thus, both a strong base and a weak base may be concentrated or dilute depending on the number of moles of the base present per liter.

Precautions in Handling Bases

Strong bases like NaOH and KOH are known as caustic alkalis. The word "caustic" refers to a substance that can cause burning. Thus, it is very important to avoid contact of these bases with any part of our body or clothing. Not only strong bases but also weak bases are corrosive. For example, concentrated ammonia solution can cause blindness if splashed into the eye.

The following safety precautions are useful in handling bases in school laboratories or anywhere while working with them

a. Wear eye goggles, gloves and a laboratory coat.
b. If bases are spilled on your working tables wipe the spillages immediately.
c. Whenever bases are splashed on your cloth, wash the affected part with running water.
d. If a base enters your eyes, wash with water repeatedly as first aid treatment and seek medical advice.
e. If a base is swallowed by accident, drink 1-2% acetic acid or lemon juice immediately.
f. Whenever bases come into contact with your skin, wash the affected part with plenty of water and then wash the affected part with a very dilute solution (about 1%) of a weak acid such as acetic acid.

pOH (note that 'p' is small, O and H are capitals!)

Activity 3.17

1. Since water, H_2O, has H^+ and OH^-, it is both acidic and basic. Do you agree?
2. Why pure water doesn't change the color of either red litmus or blue litmus?
3. If $[OH^-]$ increases in a solution what happens to its $[H^+]$? Share your comments with your group.

pOH is a measure of the concentration of hydroxide ions in an acidic or a basic solution. It is defined as the negative logarithm of hydroxide, OH^- ion concentration.

$$pOH = -\log_{10}[OH^-]$$

Relationship Between pH and pOH

The pH and pOH are related mathematically as pH + pOH = 14

This equation can be derived using the ionic product of water, Kw, as follows:

The ionic product of water, Kw is the product of the concentration of $[H^+]$ or $[H_3O^+]$ and $[OH^-]$.

Electrical conductivity measurements of pure water such as distilled water show that concentrations of H_3O and OH are 1.0×10^{-7} M at 25^0. Therefore, the ionic product of pure water is $Kw = [H^+] \text{ or } [H_3O^+] \times [OH^-]$

$$= (1.0 \times 10^{-7}) \times (1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$

Taking the logarithm of both sides of the equation $\log10\{[H^+][OH^-]\} = \log1.0 \times 10^{-14}$

We get $\{-\log_{10}[H^+]\} + \{-\log_{10}[OH^-]\} = 14$

Hence, pH + pOH = 14

Example 3.21: Find out the pH and pOH of a 0.0001M solution of NaOH?

Solution: Potassium hydroxide solution is a strong base. It ionizes completely such that one mole of NaOH gives one mole of OH ions.

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

0.0001M 0.0001M 0.0001M

Therefore, 0.0001M solution of NaOH produces 0.0001M OH ions

$$[OH^-] = 0.0001M = 1 \times 10^{-4}$$

$$pOH = -\log[OH^-] = -\log1 \times 10^{-4} = 4$$

$$pH = 14 - 4 = 10$$

Exercise 3.10

1. Which solution has a higher $[H_3O^+]$ concentration, one with pH =5, or with pH =9? Which solution has the higher OH^- concentration?
2. Give the pH and pOH for each solution with the following concentrations:
   a. $[H_3O^+] = 1.0 \times 10^{-5}$ b. $[OH^-] = 1.0 \times 10^{-9}$ M
3. Give the hydroxide ion concentrations of solutions with the following values of pH. Which of the solutions is most acidic? Which is most basic?
   a. pH 13.0 b. pH 3.0 c. pH 8.0

Preparation of Bases

Bases containing hydroxide ion (hydroxide bases) can be prepared by the following methods:

I. By the reaction of highly reactive metals from Group IA and Group IIA (below magnesium) with water. This reaction produces the metal hydroxide with the liberation of hydrogen gas

$$2Li\ (s) + 2H_2O\ (l) \rightarrow 2LiOH\ (aq) + H_2\ (g)$$

$2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$
$Ca\ (s) + 2H_2O\ (l) \rightarrow Ca(OH)_2\ (aq) + H_2\ (g)$

II. By the reaction of Group IA or Group IIA metal oxides with water, which gives the metal hydroxides?

Metal oxide + Water → Metal hydroxide
$Na_2O\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq)$
$BaO\ (s) + 2H_2O\ (l) \rightarrow Ba(OH)_2\ (aq)$

III. By double displacement reaction
The reaction of an aqueous solution of a soluble base and a soluble salt, gives another soluble base and an insoluble salt as products

$Ba(OH)_2\ (aq) + K_2SO_4\ (aq) \rightarrow 2KOH\ (aq) + BaSO_4\ (s)$
$Ca(OH)_2\ (aq) + Na_2CO_3\ (aq) \rightarrow 2NaOH\ (aq) + CaCO_3\ (s)$

Common uses of NaOH and $Ca(OH)_2$
Sodium hydroxide, NaOH (lye or caustic soda) is used in the manufacture of soaps and detergents, degreasers and as the main ingredient in oven and drain cleaners. Calcium hydroxide, $Ca(OH)_2$ (slaked lime) is used in the manufacture of cement and lime water. It is usually added to neutralize acidic soil. $Ca(OH)_2$ is used also for lime water test for carbon dioxide.

Experiment 3.11

Chemical Behavior of Bases

Objectives: To investigate the thermal stability and reaction of bases with acids.
Chemicals: NaOH or KOH, $Ca(OH)_2$, HNO_3, water, blue and red litmus papers, cobalt chloride
Apparatus: Test-tube, test-tube rack, test-tube holder, Bunsen burner, three beakers, dropper, measuring cylinder, glass rod and watch glass.

Procedure:
1. Place 4.5 g NaOH or KOH in one test-tube and the same amount of $Ca(OH)_2$ in another test-tube. Heat the test-tube containing NaOH or KOH gently using a Bunsen burner, by holding the test tube with a test tube holder. Hold the cobalt chloride paper partly inserted in the test tube. See whether the cobalt chloride paper shows a colour change or not. Repeat the same experiment with the second test tube that contains $Ca(OH)_2$.

2. Dissolve 3.6 g KOH in distilled water to prepare 100 mL solution in one beaker. Dilute 2 mL concentrated HNO_3 to make a 50 mL solution in another beaker. Add 10 mL KOH solution and 9.5 mL HNO_3 to the third beaker, stir thoroughly and test with blue and red litmus papers. Continue adding HNO_3 using a dropper, one drop at a time stirring after each addition and checking with red and blue litmus until the blue remains blue and the red remains red.
3. Put 5 mL of the neutral solution on a watch glass and allow the water to evaporate until the next day.

Questions:

1. What type of reaction occurred between the KOH and HNO_3? Write the balanced chemical equation for the reaction.
2. What was left on the watch glass?
3. Which hydroxide was melted on heating? Which base was decomposed on heating to give metal oxide and water? How do you know this? Write the balanced chemical equation for the reaction. Write a laboratory report and present it to the class.

Experiment 3.12

Preparation of Bases

Objective: To prepare bases from metal oxides and using double displacement reaction

Chemicals: Water, MgO, K_2SO_4, $Ba(OH)_2$ solution

Apparatus: Spatula, balance, three beakers, measuring cylinder, dropper, glass rod.

Procedure:

1. Add a half a spatula measure of MgO to a test tube containing clean water. Shake well and then test whether the solution is acidic or basic using red litmus paper. Does the colour of the red litmus paper change?
2. Add 7 g K_2SO_4 to the first and 7 g $Ba(OH)_2$ to the second beaker. Add water to each beaker and stir with a glass rod to prepare a 100 mL solution of each sample. After dissolution is complete, take 20 mL $Ba(OH)_2$ solution and pour it into the third beaker. Measure 20 mL K_2SO_4 solution, mix it with the solution in the third beaker and shake well. Observe what is happening. Continue adding K_2SO_4 solution drop by drop using a dropper while shaking the solution after the addition of each drop.

When the formation of a white precipitate does not occur anymore stop adding K_2SO_4 solution and test whether the solution in the third beaker is acidic or basic using litmus paper.

Questions:

1. Did the colour of red litmus paper change when dipped into the solutions obtained in procedure 1. Is the solutions acidic or basic?
2. Why do you think is the white precipitate formed in procedure 2?
3. What is the nature of the final solution in the third beaker?
4. Write balanced chemical equations for the reactions 1 and 2.

Write a laboratory report on your observations and present to the class.

3.5 Salts

Activity 3.18

Discuss the following in class and present your conclusion to the class.

1. List as many salts as you can that are used in home, medicine or agriculture.
2. What is the natural source of salts? How can you define salts?
3. Make a list of coloured salts available in your school laboratory.
4. What is the most common salt in seawater such as in Red Sea?

By the end of this section, students will be able to

- define salts;
- give examples of salts;
- classify salts as acidic and normal salts;
- discuss the direct combination of elements, the reaction of acids with bases, neutralization and the reaction between acids and metals as the methods of salt preparation;
- carry out simple experiment to prepare a salt by neutralization;
- list some important salts;
- explain the uses of some important salts;
- discuss the properties of salts;
- explain the chemical tests of some salts by conducting activities.

Salts are ionic compounds generally formed by the neutralization of an acid with a base. A salt gets its name from the names of the metal ion (derived from the base) and the acid radical derived from the acid. For example, NaCl is formed from Na^+ derived from NaOH and Cl^- is derived from HCl. Hence, NaCl (sodium cloride) is a salt of NaOH and HCl. Similarly, KNO_3 (potasium nitrate) is a salt of KOH and HNO_3.

Salt formation

$$\underset{\text{Base}}{Na^+OH^-} + \underset{\text{Acid}}{H^+Cl^-} \rightarrow \underset{\text{Salt}}{Na^+Cl^-} + \underset{\text{Water}}{H_2O}$$

Salt formation

$$\underset{\text{Base}}{K^+OH^-} + \underset{\text{Acid}}{H^+NO_3^-} \rightarrow \underset{\text{Salt}}{KNO_3} + \underset{\text{Water}}{H_2O}$$

During neutralization reaction, the cation of the base combines with the anion of the acid and a salt is formed (while H^+ and OH^- combine to form H_2O) as shown above. Salts are also defined as ionic compounds formed when the ionizable hydrogen of acids are partly or completely replaced by metal ions or ammonium ion.

Classification of Salts

Activity 3.19

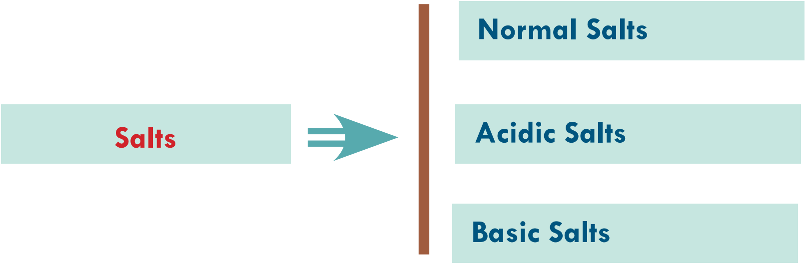
Discuss in your group and present the conclusion to the class.

Recall the concept of the neutralization reaction between an acid and a base to give a salt and water.

a. What would be the nature of the salt, if all the acid hydrogen ions are not replaced by metal ions or ammonium ions?

b. What would be the nature of the salt, if all the base hydroxide ions are not replaced by the anions of the acid?

Salts can be classified into three main groups as shown below in *Figure 3.7*:



Salts

Normal Salts

Acidic Salts

Basic Salts

Figure 3.7 Classification of Salts

I. Normal Salts

A salt formed by the complete replacement of ionizable H^+ ions of an acid by a positive metal ion or NH_4^+ ions is called normal or neutral salt. These salts are neutral to litmus.

$HCl\ (aq) + KOH\ (aq) \rightarrow KCl\ (aq) + H_2O\ (l)$
$H_2SO_4\ (aq) + ZnO\ (aq) \rightarrow ZnSO_4(aq) + H_2O\ (l)$
$HNO_3(aq) + NH_4OH\ (aq) \rightarrow NH_4NO_3(\ aq) + H_2O\ (l)$

II. Acidic salts

These salts are formed by the partial replacement of replaceable H^+ ions of an acid by a positive metal ion.

$H_2SO_4\ (aq) + KOH(aq) \rightarrow KHSO_4(aq) + H_2O(l)$
$H_3PO_4(aq) + NaOH(aq) \rightarrow NaH_2PO_4(aq) + H_2O(l)$

These salts $KHSO_4$ and NaH_2PO_4 turn blue litmus red. For example, when $KHSO_4$ dissolves in aqueous solution, it releases the ions K^+, H^+, SO_4^{2-}. It releases hydrogen ions hence it is acidic.
Acidic salts react with bases to form normal salts, K_2SO_4 and Na_3PO_4

$KHSO_4\ (aq) + KOH\ (aq) \rightarrow K_2SO_4\ (aq) + H_2O\ (l)$
$NaH_2PO_4(aq) + 2NaOH\ (aq) \rightarrow Na_3PO_4(aq) + H_2O\ (l)$

III. Basic Salts

These are salts in which not all of the hydroxide ions in a base have been replaced by the anions of the acid. Basic salts are formed by the incomplete neutralization of a polyhydroxy base by an acid.

$Al(OH)_3\ (aq) + HCl\ (aq) \rightarrow Al(OH)_2Cl\ (aq) + H_2O(l)$
$Zn(OH)_2\ (aq) + HNO_3\ (aq) \rightarrow Zn(OH)NO_3(aq) + H_2O(l)$

The above basic salts further react with acids to form normal salts.

$Al(OH)Cl\ (aq) + HCl\ (aq) \rightarrow AlCl_3\ (aq) + H_2O(l)$
$Zn(OH)\ (aq) + 2HNO_3\ (aq) \rightarrow Zn(NO_3)_2\ (aq) + 2H_2O(l)$

General Methods for the Preparation of Salts

Activity 3.20

How salts could be prepared? List a few methods to prepare sodium chloride in the laboratory? Discuss in your group and present it to the class.

Usually soluble salts are prepared by methods that involve crystallization, while insoluble salts are prepared by methods that involve precipitation.

1. Preparation of Soluble Salts

A. The reaction of an acid and a metal (Direct Displacement method)
This is direct displacement method in which hydrogen ion of acid is replaced by a reactive metal. Such as calcium, magnesium, zinc and iron, e.g.

Active metal + Acid → Salt + Hydrogen

$Mg(s) + 2HCl\ (aq) \rightarrow MgCl_2(aq) + H_2(g)$

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

Note that the reactions of Group IA metals like sodium and potassium are very vigorous and it is not advisable to use the metals of Group IA for the preparation of salts by this method.

B. The reaction of an acid and a base (Neutralization method)

It is a neutralization reaction in which acid and base react to produce a salt and water

Acid + Base → Salt + Water

$HCl\ (aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(g)$

Exercise 3.11

Identify the base and the acid used for the formation of each of the following salts

a. Zinc sulphate ($ZnSO_4$)

b. Calcium phosphate $Ca_3(PO_4)_2$

c. Silver acetate (CH_3COOAg)

d. Sodium carbonate (Na_2CO_3)

C. By the reaction of an acid and metallic oxide

Mostly the insoluble metallic oxides react with dilute acids to form salt and water

Metallic oxide + Dilute Acid → Salt + Water

$CuO(s) + H_2SO_4\ (aq) \rightarrow CuSO_4(aq) + H_2O(aq)$

$ZnO\ (s) + 2HNO_3\ (aq) \rightarrow Zn(NO_3)_2\ (aq) + H_2O\ (l)$

D. The reaction of an acid and a carbonate or bicarbonate:

Dilute acids react with metallic carbonates to produce salts, water and carbon dioxide gas.

Metal carbonates or Bicarbonates + Dilute acid → Salt + Water + Carbon dioxide

$Na_2CO_3\ (aq) + 2HNO_3\ (aq) \rightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$

$Ca(HCO_3)_2\ (s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + 2H_2O\ (l) + 2CO_2\ (g)$

2. Preparation of Insoluble Salts by Double Decomposition (Precipitation)

In this method, usually solutions of soluble salts are mixed. During the reaction, exchange of ionic radicals (i.e., metallic radicals exchange with acidic radicals) takes place to produce two new salts. One of the salts is insoluble and the other is soluble. The insoluble salt precipitates (solidify in solution).

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

$Na_2CO_3(aq) + CuSO_4(aq) \rightarrow CuCO_3\ (s) + Na_2SO_4\ (aq)$

Experiment 3.13

Preparation of Sodium Chloride

Objective: to prepare sodium chloride from sodium bicarbonate

Chemicals: Sodium bicarbonate, concentrated hydrochloric acid.

Apparatus: Evaporating dish, graduated cylinder, Burner.

Procedure

1. Add to the dish about 5 g of $NaHCO_3$ and weigh again. Record the mass in blank (a) in the table below.
2. Add 5 to 6 mL of distilled water to the dish to wet the bicarbonate. Cover the dish with a watch glass.
3. Move the watch glass aside slightly and add, in small portions, about 6 mL of concentrated hydrochloric acid from a 10 mL graduated cylinder. These small portions of acid should be added so that the acid runs down the inside wall of the evaporating dish. After the addition of 6 mL of acid, continue adding acid only as long as CO_2 (gas) continues to be evolved. Remove the watch glass and evaporate to dryness over a water bath (the evaporating dish is placed on top of a beaker containing boiling water).
4. Next, heat the dish on wire gauze with the burner for about 3 minutes. Allow the dish to cool and weigh accurately. Again, heat the dish, cool, and weigh. Continue heating and weighing until the dish reaches constant mass. Record this constant mass in blank (d) in the table below.

Questions:

1. What is the product formed? Write the balanced chemical equation for the reaction.
2. What is the change as the reaction occurs? What does this show?

Write a laboratory report and present to the class.

Some of the most important bases are listed corresponding to their uses in the following table.

Table 3.5 Some important salts and their uses.

Salts	**Uses**
Sodium chloride (common salt, NaCl)	1. Preparation and preservation of food 2. Raw material for the manufacture of sodium, chlorine, and sodium hydroxide 3. Component of Oral Rehydration Salt (ORS) 4. Manufacture of baking soda ($NaHCO_3$) and Na_2CO_3 5. In making a freezing mixture which is used by ice cream vendors
Ammonium nitrate, (NH_4NO_3	Nitrogenous fertilizer and in explosives
Copper (II) sulphate, ($CuSO_4$)	1. Used to make Bordeaux mixture (mixture of $CuSO_4$ and $Ca(OH)_2$ and other fungicides. Bordeaux mixture is used to prevent fungal attack of leaves and vines 2. Useful in electroplating and dyeing
Iron (III) chloride, ($FeCl_3$)	1. Used in the treatment of waste water 2. Used for etching printed circuits
Potassium nitrate, (KNO_3)	1. Used in making gun powder ((mixture of KNO_3, carbon and sulphur) and other explosives 2. Used as fertilizer in agriculture
Calcium sulphate, ($CaSO_4.2H_2O$, Gypsum)	Used for plastering walls and supporting fractured bones.
Barium sulphate, ($BaSO_4$)	1. Given to patients as a "barium meal" before gastrointestinal x-ray photography 2. Used as a white pigment
Iron (II) sulphate ($FeSO_4$)	Given as iron tablets to patients who suffer from anemia

Properties of Salts

Salts can be classified depending on the anion (negative ion) they possess, because the anion is partly responsible for the solubility of the salt.

1. Solubility of salts

When salts are added to water or water is added to salt, different types of actions take place. Some salts are soluble in water; some are insoluble while some salts are only slightly soluble.

2. Tendency to absorb water from the atmosphere or release water to the atmosphere.

   Salts can be classified as hygroscopic, deliquescent and efflorescent depending on their tendency to absorb water from or release water to the atmosphere.

Hygroscopic salts are those which absorb water from the atmosphere but remain solid.

Example: Anhydrous copper (II) sulphate, $CuSO_4$

Deliquescent salts absorb water from the atmosphere to form a solution. The process of absorbing water from the atmosphere by a solid to form a solution is called deliquescence.

Example: Calcium chloride($CaCl_2$), Sodium nitrate ($NaNO_3$), Iron(III) chloride ($FeCl_3$)

Efflorescent salts lose their water of crystallization to the atmosphere. The loss of water of crystallization by solid crystals to the atmosphere is known as efflorescence. It is very important to note that all deliquescent substances are hygroscopic, but all hygroscopic substances are not necessarily deliquescent.

Example 3.22: Hydrated sodium carbonate ($Na_2CO_3 \cdot 10H_2O$), Hydrated sodium sulphate ($Na_2SO_4 \cdot 10H_2O$)

3. **Aqueous solutions of soluble salts** are good conductors of electricity, because they release mobile positive and negative ions in solution.

$$NaNO_3\,(aq) \rightarrow Na^+\,(aq) + NO_3^-\,(aq)$$
$$CaCl_2\,(aq) \rightarrow Ca_2^{+}(aq) + 2Cl^-\,(aq)$$

4. **Thermal stability of salts:** Thermal stability of a salt is the property of a salt to resist irreversible change in its chemical or physical structure, often by resisting decomposition at a high relative temperature. When different salts containing the same anion are heated, they may not show similar behavior. Some salts are thermally stable while others undergo decomposition. The following examples illustrate this fact.

A. Thermal decomposition of carbonates

All the carbonates of Group IIA undergo thermal decomposition to the metal oxide and carbon dioxide gas. The term "thermal decomposition" describes splitting up a compound by heating it. Carbonates are salts containing a carbonate, CO_3^{2-}, as an anion.

If "M" represents any one of the elements, the following describes this decomposition:

$$MCO_3(s) \rightarrow MO(s) + CO_2$$

The carbonates of Group IIA are: $MgCO_3$, $CaCO_3$, $SrCO_3$, $BaCO_3$

The carbonates become more thermally stable down the group. In Group IA, lithium carbonate behaves in the same way, producing lithium oxide and carbon dioxide:

$$Li_2CO_3(s) \rightarrow Li_2O(s) + CO_2$$

The rest of the Group IA carbonates do not decompose at laboratory temperatures, although at higher temperatures this becomes possible. The decomposition temperatures again increase down the group.

Most transition metals decompose on heating to give the metal oxides and carbon dioxide:

Example 3.23:

$$CuCO_3\ (s) \rightarrow CuO\ (s) + CO_2\ (g)$$

B. Thermal decomposition of nitrates

Nitrates are salts containing a nitrate, NO_3^-, as an anion. The Group IIA and most transition metal nitrates undergo thermal decomposition to the metal oxide, nitrogen dioxide and oxygen gas. .

Nitrate → Metal oxide + Nitrogen dioxide + Oxygen

Example 3.24: $2Mg(NO_3)_2\ (s) \rightarrow 2MgO\ (s) + 4NO_2\ (g) + O_2\ (g)$

$$2Pb(NO_3)_2\ (s) \rightarrow 2PbO\ (s) + 4NO_2\ (g) + O_2\ (g)$$

Nitrates of sodium and potassium decompose on heating to give nitrites (instead of the oxides) and oxygen gas;

$$2NaNO_3\ (s) \rightarrow 2NaNO_2\ (s) + O_2\ (g)$$
$$2KNO_3\ (s) \rightarrow 2KNO_2\ (s) + O_2\ (g)$$

The carbonate and nitrate of lithium differ from those of sodium and potassium; they decompose on heating in the following manner:

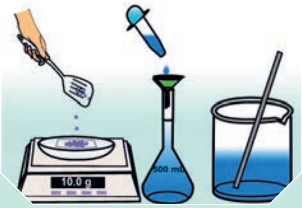
$$4LiNO_3\ (s) \rightarrow 2Li_2O\ (s) + 4NO_2\ (s) + O_2\ (g)$$
$$Li_2CO_3\ (s) \rightarrow Li_2O\ (s) + CO_2\ (g)$$

Note that both the nitrates and carbonates of lithium are thermally unstable similar to Group IIA. This reveals that lithium behaves more like Group IIA rather than Group IA.

Chemical Tests of Some Ions in Salt

Activity 3.21

1. Do you know why fireworks give different colors? Consult books in your library, and discuss in your group. Explain to the class.
2. How do chemists know of the presence of certain ions in a compound?



Experiment 3.14

Test for Cations-I

Objective: To identify the presence of Li^+, Na^+, K^+, Ca^{2+}, Sr^{2+} and Ba^{2+} in salts by flame tests.

Chemicals: Distilled water, salts containing Li^+, Na^+, K^+, Ca^{2+}, Sr^{2+} and Ba^{2+}.

Apparatus: Platinum or nichrome wire, watch glass, Bunsen burner.

Procedure:

1. Place a small amount of the salt containing a Li^+ ion on a watch glass, moisten it with pure concentrated HCl, dip the tip of the platinum or Nichrome wire into the moist salt and then bring to the Bunsen flame.
2. Observe the colour produced.
3. Repeat the same step for salts containing Na^+, K^+, Ca^{2+}, Sr^{2+} and Ba^{2+} and record your observations. Rinse the platinum or nichrome wire with distilled water after each test.

Question:

a. Write the colour of the flame produced in the following table.

Table 3.6 Metal ion and corresponding flame colour

Metal ion in the salt	Colour of flame produced
Lithium (Li^+)	
Sodium (Na^+)	
Potassium (K^+)	
Calcium (Ca^{2+})	
Strontium (Sr^{2+})	
Barium (Ba^{2+})	

b. Write a laboratory report and present to the class.

Flame tests are used to identify the presence of a relatively small number of metal ions in a compound. Certain metals give a characteristic color to a Bunsen flame when their solid salts or moist salts are heated directly in the flame. A flame test is commonly used to identify the presence of lithium, sodium, potassium, calcium, strontium and barium ions in salts.

Experiment 3.15
Test for Cations - II

Objectives: To identify the presence of Cu^{2+}, Fe^{2+} and Fe^{3+} in salts.

Chemicals: Ammonia solution, sodium hydroxide solution, salts containing each of Cu^{2+}, Fe^{2+} and Fe^{3+} ions

Apparatus: Beakers, test tubes, test tube rack, glass rod.

Procedure:

1. Dissolve a salt containing Cu^{2+} in the first beaker, Fe^{2+} in the second and Fe^{3+} in the third. Take three test tubes and add a 2 mL solution of (each ion) Cu^{2+} salt to the first, Fe^{2+} salt solution to the second and Fe^{3+} salt solution to the third.
2. Add aqueous ammonia in small quantities until it is in excess, to the first test tube containing a copper (II) salt solution and record your observations.
3. Add sodium hydroxide solution in small quantities, until it is present in excess, to the second and third test tubes, and record your observations.

Questions:

1. What does the formation of a blue precipitate in the first indicate? Write the chemical equation for the blue precipitate formation.
2. Observe the colours of the precipitates formed and complete the following

Table 3.7 Colour observed and its indication.

Colour observed	Confirms the presence of

Write balanced chemical equations for the reactions. Write a laboratory report in group and present to the class.

Experiment 3.16
Test for Sulphates

Objective: To identify for the presence of sulphate using barium salts.

Chemicals: Any soluble sulphate salt (such as sodium sulphate), barium chloride or barium nitrate solution, and dilute HCl.

Apparatus: Beakers, test tubes, and test tube rack,

Procedure: Add some sodium sulphate solution to a test tube and acidify the solution by adding a few drops of dilute HCl. Then add $BaCl_2$ or $Ba(NO_3)_2$ solution and note if a white precipitate is formed.

Questions:

1. Name the white precipitate formed?
2. Why is it necessary to add a few drops of dilute HCl?
3. Write a balanced chemical equation for the reaction.

Write a laboratory report and present to the class

Experiment 3.17

Test for Carbonates and Hydrogen Carbonates

Objective: To distinguish between carbonates and hydrogen carbonates
Chemicals: Na_2CO_3, $NaHCO_3$, dilute HCl, lime water
Apparatus: Conical flasks, Beaker

Procedure:

1. Take 20 mL solution of Na_2CO_3 and add it to one conical flask and 20 mL $NaHCO_3$ solution to another. Add the same amount of dilute HCl to each of the conical flasks; Fit a rubber stopper to which a delivery tube is inserted to each conical flask. Allow the gas produced to pass through lime water and observe the changes.
2. Again take Na_2CO_3 solution in one conical flask and $NaHCO_3$ solution in the other. Add the same amount of $CaCl_2$ solution to each of the conical flasks.

Questions:

1 What happened to the lime water in each case?
2. Which solution formed a white precipitate upon the addition of $CaCl_2$ solution?
3. Write balanced chemical equations for all the reactions.

Write a laboratory report and present it to the class.

Experiment 3.18

Test for Nitrates (Brown Ring Test)

Objectives: To identify the presence of nitrate in a solution.
Chemicals: Nitrate solution, iron (II) sulphate solution, concentrated H_2SO_4
Apparatus: Test tube, test tube rack, beaker,

Procedure:

Take 2 mL of nitrate solution in a test tube and add an equal volume of freshly prepared iron (II) sulphate solution. Hold the test tube in an inclined position and carefully pour concentrated sulphuric acid down along the inclined side of the test tube. The acid sinks to the bottom. Carefully observe the changes in test tube.

Questions:

1. Where is the brown ring formed?
2. Write the chemical equation for the formation of the brown ring.
3. What does the formation of brown ring in the solution indicate? Write a laboratory report

Exercise 3.12

1. You are provided separate solutions of NaCl, HCl, NaOH and Phenolphthalein (Marked 1). All the solutions are colorless. How would you recognize the solutions of acid, base and salt?
2. Suggest at least three methods for the preparation of salts.
3. Classify the following salts as soluble or insoluble in water:
   a. NH_4Cl
   b. K_3PO_4
   c. $FeCO_3$
   d. $AgCl$
   e. $AgNO_3$
   f. $CaCl_2$
   g. $PbSO_4$
   h. Na_2S
   i. $CaSO_4$
   j. $BaCO_3$
4. Why do aqueous solutions of soluble salts conduct electricity?
5. Which carbonates do not decompose on heating?
6. Nitrates mostly decompose by heat to give metal oxide, nitrogen dioxide and oxygen. Which nitrates do not give these products on heating?
7. What reagents do you use to identify the presence of the following ions in salts?
   a. Halide ions
   b. Fe^{2+} and Fe^{3+} ion
   c. Sulphate, SO_4^{2-} ions?

Plant Nutrients

By the end of this topic, students will be able to

- explain the applications of some salts as fertilizers, pesticides and herbicides;
- give examples of salts used for fertilizers, herbicides and pesticides.

Activity 3.22

Form groups, discuss the following questions in class and present it to the class.

1. Plants need nitrogen for their growth, and there is plenty of nitrogen in the air. If this is so, why do we apply nitrogenous fertilizers to the soil to support plants growth?
2. Some people prefer to eat organic food that has been grown without fertilizers and pesticides. Does eating organic food overcome the problems caused by artificial fertilizers?

Essential plant nutrients could be either macronutrients or micronutrients, depending on the amount required by plants.

Table 3.8 Essential plant nutrients

Macronutrients	Micronutrients
These are the elements required in relatively large amounts by plants. These elements are nitrogen, phosphorus, potassium, calcium, magnesium, sulphur, carbon, hydrogen and oxygen.	These are the elements required by plants in relatively small (trace) amounts. These elements are iron, cobalt, zinc, manganese, molybdenum, copper, boron and chlorine. The elements carbon, hydrogen and oxygen are obtained by plants from air and water. These are not considered as mineral nutrients.

Nitrogen, phosphorous and potassium are said to be primary mineral nutrients. What are the roles of nitrogen, phosphorus and potassium in plant growth?

Nitrogen (N)

- is a very important plant nutrient and absorbed by plants in the form of nitrate ions (NO_3^-).
- is utilized in the synthesis of amino acids, proteins, coenzymes and nucleic acids which are taking place in the growth and development of plants
- is also involved in the synthesis of chlorophyll to produce a deep green color.

Phosphorus (P)

- is absorbed by plants mainly in the form of $H_2PO_4^-$ and HPO_3^{2-} in small amounts.
- facilitates early growth and root formation, quick maturity and promotes seed or fruit production.
- has a role in the formation of some amino acids and proteins, coenzymes, nucleic acids and high energy phosphate compounds like Adenosine Triphosphate (ATP).

Potassium (K)

- is absorbed by plants in the form of the K^+ ion.
- is a component of enzymes that facilitate the process of photosynthesis , protein synthesis, and adjusts stomata movement.
- is the most important ion in controlling the turgidity of plant cells.
- increases in the solute potential (sap concentration) within the cells which leads to an increase in the amount of water that enters the cells osmotically.

Exercise 3.13

1. What is the importance of the following elements for the growth of food crops?

A. Nitrogen B. Phosphorus C. Potassium

2. In what form do plants absorb these elements?
3. Why nitrogen is not absorbed by plants in the form of N_2? And phosphorous not in the form of P_4?

Fertilizers

Fertilizers are materials that are added to soils to increase the growth, yield or nutritional value of crops. There are organic(natural) and synthetic (artificial) fertilizers.

Organic (natural) fertilizers are fertilizers derived from animals and plants.

- Organic fertilizers provide readily available nutrients to plants after some period of decay and decomposition and can supply the elements nitrogen, phosphorus, potassium and various trace elements to the soil.
- Synthetic (artificial) fertilizers contain nitrogen, phosphorous and potassium, and are known as complete fertilizers.
- A mixed artificial fertilizer may contain nitrogen, phosphorus and potassium. Such type of fertilizer is called a complete fertilizer.
- The term NPK is used to describe fertilizers containing the elements nitrogen, phosphorus and potassium.

Examples of artificial fertilizers are ammonium sulphate, $(NH_4)_2SO_4$, potassium nitrate (KNO_3), ammonium nitrate (NH_4NO_3), and diammonium hydrogen phosphate, DAP $(NH_4)_2HPO_4$).The most important mineral nutrients that need to be added to the soil are usually nitrogen, phosphorus and potassium.

- All three elements are needed in large quantities. Commercial fertilizers are normally given a "grade", which reflects the percentages they contain of N, P and K by dry weight. For example, a NPK fertilizer may be described as 10:10:10 indicating that there is 10% of each element in the fertilizer.
- The suitable proportions are best determined in relation to the tested fertility of the soil and the requirements of the particular crop that is being grown on it. This means fertilizers of the same composition are not used for different purposes.
- In nitrogen-deficient soil, a fertilizer containing a higher percentage of nitrogen than phosphorus and potassium must be used. In addition to this, all fertilizers are not used for the treatment of all kinds of soil. For example, ammonium sulphate is more suitable for use in basic soils than in acidic or neutral soils.

Exercise 3.14

1. Explain the terms mixed fertilizer , complete fertilizer and fertilizer grade
2. Classify the following as organic or synthetic fertilizers:
   A. NH_4NO_3 B. Urea C. $(NH_4)_2SO_4$
3. Can the same fertilizer be applied to all types of soil? Explain.

Pesticides

- Pesticides are chemicals that can be applied to crops to kill pests that affect plant growth and development.

Herbicides

- Herbicides (weed killers) are chemicals used to manipulate or control undesirable plants. Some examples of herbicidal salts are: Common salt ($NaCl$), copper sulphate ($CuSO_4$), sodium arsenite ($NaAsO_2$), Ammonium sulphamate ($NH_4SO_3NH_2$), sodium chlorate ($NaClO_3$).

Reading Assignment

Read about the use of more salts as pesticides and herbicides from a reliable source and present it to your class.

Check List for Reference

Key Terms of the Unit

- Alkalis
- Acid salt
- Acidic oxide
- Acid, basic and normal salt
- Amphoteric oxide
- Arrhenius acid
- Arrhenius base
- Basic salt
- Basic oxide
- Binary acid
- Concentrated acid
- Deliquescent
- Dissociation constant for water (kw)
- Dilute acid
- Electrolyte
- Efflorescent
- Fertilizer
- Flame test
- Hygroscopic
- Litmus
- Macronurient
- Molarity
- Micronutrient
- Monoprotic acid
- Neutral oxide
- Neutralization
- Non-volatile acid
- Normal salt
- pH
- Peroxide
- pOH
- Pesticide
- Polyprotic acid
- Salt
- Strong acid
- Strong base
- Ternary acid
- Universal indicator
- Weak acid
- Weak base

Unit Summary

- Inorganic compounds are those compounds that originate from mineral constituents of the earth's crust. Inorganic compounds may be classified as oxides, acids, bases and salts.

- Oxides are binary compounds consisting of oxygen and any other element. Most common oxides are classified as acidic, basic, amphoteric, neutral oxides and peroxides. Acidic oxides are oxides of non-metallic elements. Basic oxides are oxides of metals. These metal oxides which dissolve in water are also called basic anhydrides. Amphoteric oxides are those oxides which show the properties of both acids and bases. Neutral oxides are those oxides which do not show basic or acidic properties. Peroxides are oxides containing a peroxide (-O-O-) link and the oxidation state of oxygen is -1.

- Arrhenius acids are substances that release hydrogen ions or protons in aqueous solution. Arrhenius bases are substances that release hydroxide (OH–) ions in aqueous solution. Strong acids and strong bases ionize almost completely in aqueous solution. Weak acids and weak bases ionize only slightly in aqueous solution.

- pH is the negative logarithm of hydrogen ion concentration $\text{pH} = -\log [\text{H}^+]$.
  pOH is the negative logarithm of hydroxide ion concentration $\text{pOH} = -\log[\text{OH}^-]$.
  $\text{KW} = [\text{H}^+][\text{OH}^-] = 1 \times 10^{-14}$ at 25°C.
  $\text{pKw} = \text{pH}^+\ \text{pOH} = 14$ at 25°C

- Salt is an ionic compound containing a cation derived from a base and anion derived from an acid. Salts may be classified as acidic, normal and basic salts. An acid salt is formed when ionizable hydrogen atoms of an acid are replaced partly by a metal ion or ammonium ion. A normal salt is formed when all ionizable hydrogen atoms of an acid are completely replaced by a metal or ammonium ion. A basic salt is a salt containing ionizable hydroxide ion.

- Hygroscopic salt absorb water from the atmosphere. Deliquescent salt absorb water from the atmosphere and dissolve in the water absorbed to form solutions. Deliquescent salt absorb water from the atmosphere and dissolve in the water absorbed to form solutions. Efflorescent salt lose their water of crystallization to the atmosphere.

- Plant nutrients are minerals required by plants for their growth and development. The elements plants need for their growth and development are classified as Macronutrients and micronutrients depending on the amount utilized by them. Fertilizers are natural products or synthetic chemicals that are added to the soil to increase its crop producing potential. Synthetic fertilizers are classified as nitrogen, phosphorus and potassium fertilizers. Pesticides are chemicals used in agriculture to kill pests that reduce crop yields.

Review Exercises

Part I: Match the substances in column X with those properties in column Y

	Column X		Column Y
1.	Apple (pH = 3.3)	A.	Strongly Acidic
2.	Tomato (pH =4.4)	B.	Strongly Alkaline
3.	Household ammonia (pH = 11.9)	C.	Neutral
4.	Distilled water (pH =7)	D.	Weakly acidic
5.	Hydrochloric acid (pH = 0.0)	E.	Weakly alkaline
6.	Grapes (pH = 4.5)		
7.	Milk of magnesia (pH = 10.5)		
8.	Dish-washing powder (pH = 12)		

Part II: Match the ions or molecules in column X with the correct test reagent in column Y

	Column X		Column Y
9.	Iron (II) ion	a.	Silver nitrate
10.	Iodide ion	b.	Acidified barium nitrate
11.	Sulphate ion	c.	Heat with aluminium
12.	Carbon dioxide	d.	Iron (II) sulphate & sulphuric acid
13.	Nitrate ion	e.	Sodium hydroxide
		f.	Limewater

Part III: Fill in the blanks by using the correct words /terms given in the brackets

14. The pH value of an acid is _____ than 7. (greater/less)
15. A solution is said to have _____ character when its pH value is greater than seven (basic/acidic)
16. The pH value of a neutral solution is _____ at 298K. (Zero/7)
17. The pH of 10^{-6} is ____ than the pH of 10^{-4} HCl. (greater/less)
18. The pH of water decreases when _____ is dissolved in it. (NH_3/SO_2)
19. The hydrogen ion concentration of a solution with pH = 3 is ____ than the solution with pH = 6. (greater/less)
20. An acid produces _____ ions in aqueous solution. (Hydrogen/hydroxide)
21. The concentration of H^+ ions given by HCl is _____ than that given by equal concentration of CH_3COOH (greater/less).

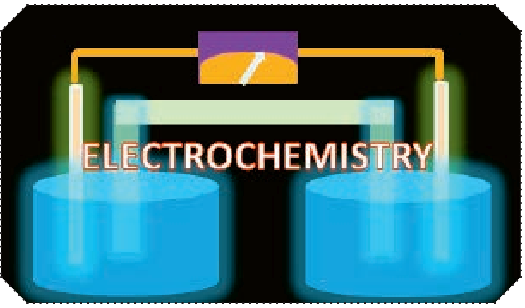
Part IV: Select the correct answer from the given choices

22. Which one of the following is a diprotic acid?
    - a. Hydrochloric acid
    - b. Phosphoric acid
    - c. Sulfuric acid
    - d. Acetic acid
23. Which of the following is not the characteristic of an acid:
    - a. An acid changes the color of an indicator
    - b. An acid has a bitter taste
    - c. An acid ionizes in water
    - d. An acid produces hydronium ions in water
24. When an acid reacts with an active metal:
    - a. Hydronium ion concentration increases
    - b. Hydrogen gas is produced
    - c. Metal forms anions
    - d. Carbon dioxide gas is produced
25. The pH of a solution is 8. What is the pOH of the solution?
    - a. 8
    - b. 8×10^{-7}
    - c. 6
    - d. 2.1×10^{-6}
26. Which of the following solutions would have a pH value greater than 7?
    - a. $[OH^-] = 2.4 \times 10^{-2}$ M
    - b. 0.0001 M HCl
    - c. $[H_3O^+] = 1.53 \times 10^{-2}$ M
    - d. $[OH^-] = 4.4 \times 10^{-9}$ M

Part V: Give Short Answers

27. The hydrogen ion concentration in an acidic solution is 10^{-5} ? What is the pH value of the solution?
28. An element combines with oxygen to form an oxide. This oxide dissolves in water. This aqueous solution changes blue litmus to red. Write: A. the nature of the aqueous solution (alkaline or acidic) B. the nature of the element (metal or nonmetal)
29. What do you understand by the statement 'acetic acid' CH_3COOH, is a monoprotic acid?
30. Write the balanced equations for the reaction of zinc oxide with:
    a. hydrochloric acid
    b. aqueous sodium hydroxide
31. How many times will the hydrogen ion concentration change when the pH value of an aqueous solution is changed from 6 to 5?
32. Write the balanced chemical equations for the preparation of the following salts:
    a. A soluble sulphate by the action of an acid on a metal
    b. A soluble sulphate by the action of an acid on an insoluble metal oxide
    c. An insoluble sulphate by the action of an acid on another salt
33. Calcium oxide reacts with hydrochloric acid to form calcium chloride and water.
    a. Write a balanced equation for this reaction.
    b. Farmers often add calcium oxide to the soil. Explain why they do this.
34. Fertilizers are spread on fields by farmers.
    a. Why do farmers use fertilizers?
    b. State the names of three elements most commonly found in fertilizers.
    c. Ammonium sulphate is a fertilizer. Describe how you can make ammonium sulphate in the laboratory from aqueous ammonia and sulphuric acid.
    d. Ammonium nitrate is also a fertilizer. Write a word equation to show how ammonium nitrate can be produced.
35. State the names of a suitable acid and alkali you can use to make each of the following fertilizers:
    a. Ammonium sulphate
    b. Potassium phosphate
    c. Ammonium nitrate
36. A classmate states, "All compounds containing H atoms are acids, and all compounds containing OH groups are bases." Do you agree? Give examples and explanations.

UNIT 4



ENERGY CHANGES AND ELECTROCHEMISTRY

Unit Outcomes

At the end of this unit, you will be able to

- describe energy changes in chemical reactions;
- describe how a chemical reaction produces electric current and how electricity brings about a chemical reaction in electrochemical cells;
- distinguish the difference between metallic conduction and electrolytic conduction;
- develop skills in writing the oxidation half-reaction, reduction half-reaction and cell reaction for the electrolysis of molten electrolytes that occur in electrolytic cells;
- describe the three types of Voltaic cells;
- explain the difference between electrolytic cells and voltaic cells and
- demonstrate scientific inquiry skills: observing, classifying, comparing and contrasting, inferring, predicting, communicating, measuring, asking questions, interpreting data, drawing conclusion, applying concepts, relating cause and effect and problem solving.

Start-up Activity

1. Describe the different types of chemical reactions and their characteristics.
2. Argue for or against this idea "without energy, there is no development!"
3. Describe the energy sources of our country Ethiopia? Compare their relative importance to the country's development. Which one is the principal source? Which one is the most important source that is underutilized? What are the different forms of energy contained in each of these sources?
4. Discuss the relevance of hydroelectric dams such as the GERD (The Grand Ethiopian Renaissance Dam) to the country's development vis-à-vis climate change? What energy changes are involved? What about batteries? How do they generate electrical energy and how light bulbs convert it into light? Can you make a battery? Some people do not want to think about these types of questions. They just spend money to buy one and gratify their needs. But others are very curious to learn new things and want to create their own device named in honor of their names (Made by Mr/Miss "X"). Which category are you from?

4.1 Introduction

At the end of this section, students will be able to describe energy changes in chemical reactions.

Activity 4.1

1. We feel warm when we stand in the sunlight. Discuss the reason.
2. Your body sweats when you take a hot shower after an intensive physical exercise?
3. Why do you burn wood, or charcoal at your home? What energy change is involved?
4. The major component of dynamite is nitroglycerin, a very unstable material, mixed with diatomaceous earth. What type of energy change occurs when dynamite explodes?

You have studied in unit 1 that chemical reactions almost always involve energy change. This means, the energy of the reaction mixture is different before and after the reaction. For example, coke is mainly carbon, and the combustion of carbon in oxygen, which can be represented by the chemical equation

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

results not only in the formation of gaseous carbon dioxide, but also in the release of energy in the form of heat and light. During this process, the chemical energy contained in fossil fuels is converted into heat and light energy. Similarly, natural gas (principally methane, which has the formula CH_4) bums in oxygen

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

gives not only the gaseous products carbon dioxide and water, but energy in the form of heat and light.

To study the energy change, (ΔE), involved in chemical reactions like the combustion reaction we have just cited, scientists conceptually divide the universe into two: the system (the part being studied) and the surroundings (everything else).

$$\text{Universe} = \text{System} + \text{Surrounding}$$

A system is part of the universe that is being studied. Surrounding is anything else than the system in the universe. Each particle in a system under investigation, such as a reaction mixture in a test tube, has potential energy and kinetic energy, and the sum of all these energies is the internal energy, E of the system.

Internal energy of a system is the sum of potential energy and kinetic energies of the components of the system.

$$(\text{Internal Energy})_{\text{system}} = (\text{potential energy} + \text{kinetic energy})_{\text{system}}$$

Potential energy is stored energy. It is stored in chemical bonds. Kinetic energy is the energy due to motion. It causes work to be done through movement.

When the reactants in a chemical system change to products, the system's internal energy is changed. This change, ΔE, is the difference between the internal energy after the change (E_{final}) and before the change (E_{initial}):

$$\Delta E = E_{\text{final}} - E_{\text{initial}} = E_{\text{products}} - E_{\text{reactants}}$$

Where Δ (Greek delta) means "change (or difference) in" and refers to the final state minus the initial state. Thus, ΔE is the final quantity of energy of the system minus the initial quantity.

Note that the total energy of the universe remains constant. When the system loses a certain amount of energy, the surrounding gains the same amount of energy. This means, energy is simply being exchanged between the two components of the universe keeping the total energy of the universe constant (*Figure 4.1*).

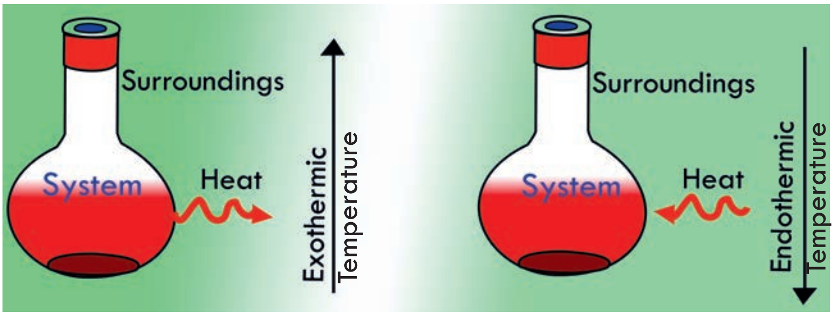


Figure 4.1 Energy exchange between system and surrounding.

Energy transferred from system to surroundings or vice versa appears in two forms: heat and work. For instance, the various chemicals that make up gasoline contain a large amount of chemical (potential) energy that is released when the gasoline is burned in a controlled way in the engine of the car. The release of that energy does two things: some of the potential energy is transformed into work, which is used to move a car; at the same time, some of the potential energy is converted to heat and makes the car's engine very hot. Therefore, the energy changes of a system occur as either heat (q) or work (w), or some combination of both.

$$\Delta E = q + w$$

4.1.1 Exothermic and Endothermic Chemical Reactions

At the end of this section, students will be able to
- define endothermic reaction;
- describe endothermic reaction;
- define exothermic reaction;
- describe exothermic reaction;
- elucidate endothermic and exothermic reactions using diagrams;
- do simple experiment to demonstrate exothermic and endothermic reactions.

Activity 4.2

1. From your discussion on ***Activity 4.2***, you can conclude that reactions that cause the surrounding to heat up are called heat releasing reactions and the process is exothermic. Reactions that take up energy from the surrounding such as the reaction flask are called heat absorbing and the process is endothermic. Based on this information, discuss whether energy (in the form of heat) is released or absorbed during
   a. baking of bread (consider the heat of stove before and after the cold dough is loaded)
   b. cooking food
   c. burning of wood and then classify each of them as exothermic or endothermic.

   What other processes can you cite from your daily life experience that either releases heat or absorbs heat?
2. Is photosynthesis an endothermic or exothermic reaction? Write the balanced equation and draw an energy diagram indicating all the necessary information. Repeat this activity for respiration.

All of the combustion reactions mentioned in the previous section release energy in the form of heat. Such reactions are described as being exothermic (from the Greek exo meaning outside and therme meaning heat). One of the examples cited earlier is the combustion of coke in oxygen, which obviously results in the release of a considerable amount of energy. Similarly, although on a much more modest scale, when NaOH dissolves in water, the test-tube becomes warm. In other words, the process results in a temperature rise. This observation indicates that the dissolving process for NaOH is also exothermic; that is, it is accompanied by the release of energy, which heats up the solution.

$$NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq)$$

The other is the reaction between calcium chloride and water:

$$CaCl_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + 2HCl(g) + heat$$

In this case too, heat is released during the reaction, elevating the temperature of the reaction mixture, and thus the reaction flask becomes hot. But, when potassium nitrate, KNO_3, dissolves in water, the mixture gets colder. This process can be represented by the following equation:

$$KNO_3(s) + Heat \xrightarrow{H_2O} K^{+}(aq) + NO_3^{-}(aq)$$

Reactions like this are described as being endothermic (from the Greek endo meaning within).

To study whether the energy being released or absorbed during a chemical reaction, we need to set the temperature and pressure constant. Thus, if a reaction takes place at constant temperature and pressure, the heat transferred is called the enthalpy of reaction, and this is denoted by the symbol ΔH:

$$\Delta H = H_{Products} - H_{reactants}$$

For an exothermic reaction: ΔH is negative, $\Delta H < 0$.
For an endothermic reaction: ΔH is positive, $\Delta H > 0$.

The relative energies of reactants and products in exothermic and endothermic reactions can be represented as shown schematically in ***Figure 4.1 a*** and ***b***, respectively.

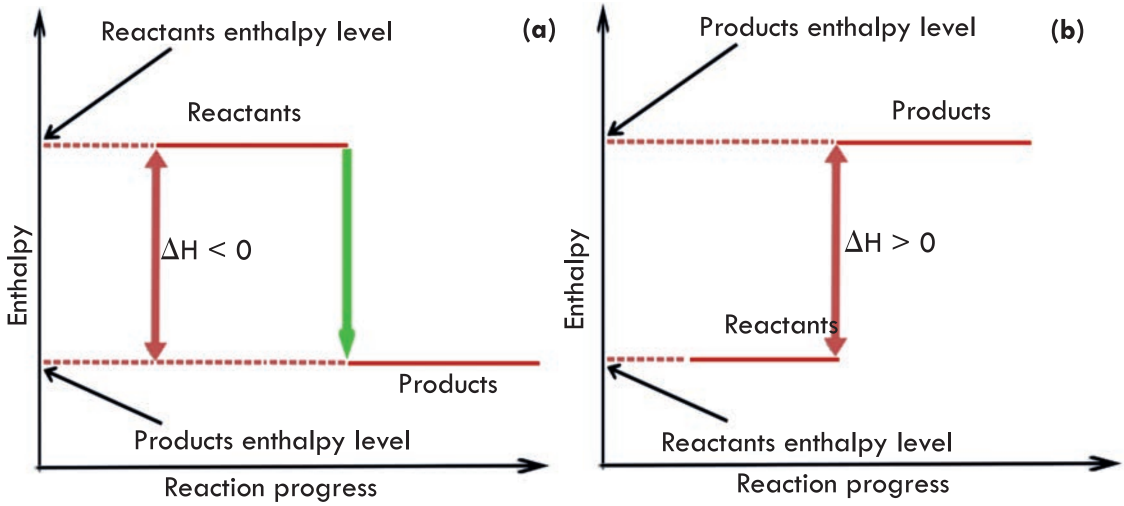


Figure 4.1 Enthalpy diagram for (a) exothermic reactions (b) endothermic reactions.

The device used to measure the heat released (or absorbed) by a chemical reaction is called calorimeter. To find the energy change during chemical reaction, we measure the change in temperature and determine the quantity of heat released or absorbed using the following relation:

$$q \propto \Delta T \text{ or } \quad q = \text{constant} \times \Delta T \quad \text{or} \quad q/\Delta T = \text{constant}$$

The proportionality constant in the above equation is called heat capacity. Every object has its own heat capacity. Heat Capacity of a substance is defined as the quantity of heat required to change its temperature by 1 K.

$$\text{Heat capacity} = q/\Delta T$$

The unit of heat capacity is Joule per Kelvin (J/K). Heat capacity is related to a substance's ability to retain heat and the rate at which it will heat up or cool down. For example, a substance with a low heat capacity, such as iron, will heat and cool quickly, while a substance with a high heat capacity, such as water, heats and cools slowly.

A related property is specific heat capacity (c). Specific heat capacity (c) of a substance is defined as the quantity of heat required to change the temperature of 1 gram of the object by 1 K:

$$\text{Specific heat capacity (c)} = \text{(Heat capacity)/mass} = q/(\text{mass} \times \Delta T)$$

If we know c of the object being heated (or cooled), we can measure the mass and the temperature change and calculate the heat absorbed (or released):

$$q = c \times \text{mass} \times \Delta T$$

Notice that metals have relatively low values of c and water has a very high value: for instance, it takes over 30 times as much energy to increase the temperature of a gram of water by 1 K as it does a gram of gold. This is one of the properties that make water unique and responsible for its use as a coolant in car radiators and in industries

Exercise 4.1

1. Where does the energy released during chemical reaction come from?
2. How is the energy that you get from food stored in your body? How is it supplied to the cellular process in your body when needed?
3. Explain the energy changes involved when
   a. quicklime or calcium oxide (CaO) reacts with water to form slaked lime [$Ca(OH)_2$]?

   $$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2\,(aq)$$

   b. limestone ($CaCO_3$) decomposes to give quicklime (CaO)
   c. nitroglycerine (major component of dynamite) decomposes

   $$4C_3H_5\,(NO_3)_3(l) \rightarrow 12CO_2(g) + 10H_2O(g) + O_2\,(g) + 6N_2\,(g)$$

   Is it exothermic or endothermic? Explain! Draw an enthalpy diagram for each of these three reactions.
4. On a hot summer day the water in a lake stays cool even though the air above it heats quickly, and the water stays warm at night after the air has cooled. Can you explain this?
5. Have you encountered an earthen stove "*Injera Mitad*" that takes longer time to feel hot? Explain why. Which stove stays warm for longer time after the heat source is disconnected, the one heats up quickly or slowly?
6. What property of substance is responsible for the observations in Q1 and Q2?

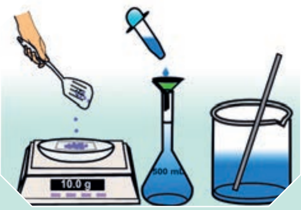
4.1.2 Importance of Chemical Changes

At the end of this section, students will be able to discuss the importance of chemical changes in the production of new substances and energy.

Activity 4.3

1. Discuss the importance of chemical changes in your life process. Consider, for instance, oxidation of glucose in your body cells. The photosynthesis and combustion reactions, the digestion, etc. Do you think life is possible without chemical reactions or chemical changes?
2. Can you mention some pharmaceuticals and their use? How are they made?
3. What two major uses of chemical reactions or changes can you suggest from the above discussion?

Chemical reactions are the most important types of events in the universe. It's through chemical reactions that plants grow, produce fruit, and become compost for new plants. It's because of chemical reactions that human beings (and all other animals) reproduce, digest, grow, hear, and think. Chemical changes are involved during digestion and respiration. You get energy because of respiration. You get food because of the chemical change called photosynthesis. Therefore, the two principal advantages of chemical changes are production energy and useful substances.



Experiment 4.1

Energy Changes in Chemical Reactions

Objective: To explore the energy changes involved in the reaction between acetic acid and baking soda.

Materials needed: Plastic cup or beaker, vinegar (acetic acid), baking soda ($NaHCO_3$), spatula/teaspoon, thermometer.

Procedure:

Pour about 10 mL of vinegar into a small plastic cup. Then, place a thermometer into the vinegar. Record the initial temperature (T_i).

While the thermometer is in the cup, add about ½ teaspoon of baking soda to the cup.

Watch the thermometer for any change in temperature. After it has stopped changing, record the final temperature (T_f).

1. Based on your observations of the baking soda and vinegar reaction, is the reaction exothermic or endothermic?
2. Use energy level diagram to visualize the energy changes
3. Repeat the above steps using calcium chloride instead of vinegar. Briefly, dissolve 1/2 a spoonful of calcium chloride in 20 mL water in a test tube. Record your observation. In a second test tube, dissolve equal amount of baking soda and record your observation. Allow both solutions to cool to equal (room) temperature and record this temperature as Ti. Mix the two solutions in a plastic cup/beaker and record the final temperature.
4. Discuss the endothermic and exothermic processes observed in the above procedure.
5. Write the reaction equation and use energy level diagram to visualize the energy changes during reaction of calcium chloride and baking soda.

4.2 Energy Changes in Electrochemistry

At the end of this section, students will be able to

- describe electrochemistry;
- define electrical conductivity;
- explain metallic conductivity;
- explain electrolytic conductivity;
- differentiate between metallic conduction and electrolytic conduction.

4.2.1 Electrochemistry

Activity 4.4

1. You know that metals are good conductors of heat and electricity. But what is electricity? What are the sources? What is the analogy between electricity and water flow via tap to your home? How do metals conduct electricity? Can you suggest other materials that conduct electricity?
2. What do substances need to possess in order to conduct electricity?
3. What about batteries such as hand batteries and car batteries? What is the energy change involved?
4. Can you suggest devices that generate electrical energy? Discuss the uses of electrical energy?

Electrochemistry is a field of chemistry that deals with the interconversion of electrical energy and chemical energy. It is concerned with the use of electricity to cause chemical reactions (changes) to happen or generating electricity from chemical reactions. Thus, electrical energy and chemical energy are inter-convertible.

4.2.2 Electrical Conductivity

Why do metals conduct electricity? Do you know any non-metal which conducts electricity? Electrical conductivity is the ability of a substance to transmit electricity. The materials that allow the passage of electricity through them are called electrical conductors. Electricity is constituted by movement of charged particles or charge carriers. Electrons carry negative charge whereas ions may carry positive (cations) or negative (anions) charges. Depending on the nature of the particles responsible for the flow of electric charges through conductors, electrical conductivity can be classified as metallic conductivity and electrolytic conductivity.

a) Metallic conductivity

Electrical conductivity in metals is a result of the movement of electrically charged particles called electrons. The atoms of metal elements are characterized by the presence of valence electrons; which are electrons in the outer shell of an atom that are free to move about. It is these "free electrons" that allow metals to conduct an electric current. Because valence electrons are free to move, they can travel through the metallic lattice that forms the physical structure of a metal. Silver is the best electrical conductor.

Metallic lattice can be described as an atmosphere of positive ions in a sea of mobile electrons (*Figure 4.3*). The electrons entering the metal displace (repel) the freely moving electrons at the point of entry. The displaced electrons occupy new positions by pushing neighboring electrons ahead. This will continue until electrons are forced out of the wire at the opposite end. The charge carriers in metallic conduction are electrons. Hence, metallic conductivity is also called electronic conductivity.

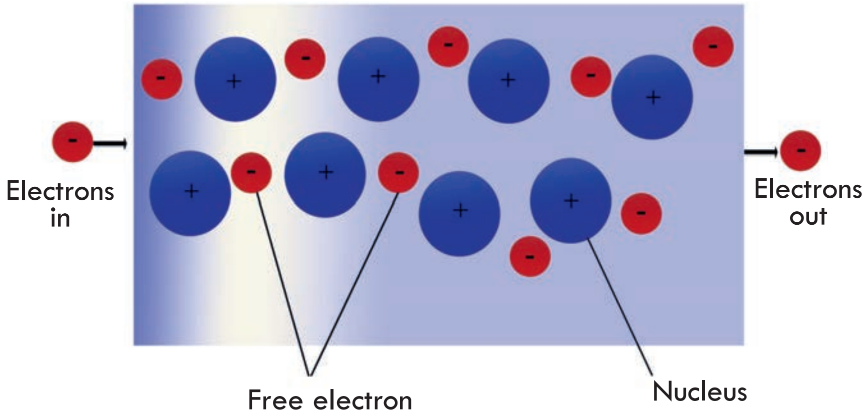


Figure 4.3 Electrical conductivity in metals.

Non-metals are generally non-conductors of electricity, because they do not have freely moving electrons. Graphite is a form of carbon in which the carbon atoms are bonded in trigonal planar fashion to the three other carbon atoms, to form inter-connected hexagonal rings, as shown in ***Figure 4.4***. Electrons move freely through the hexagonal layers, making graphite a good conductor of electricity.

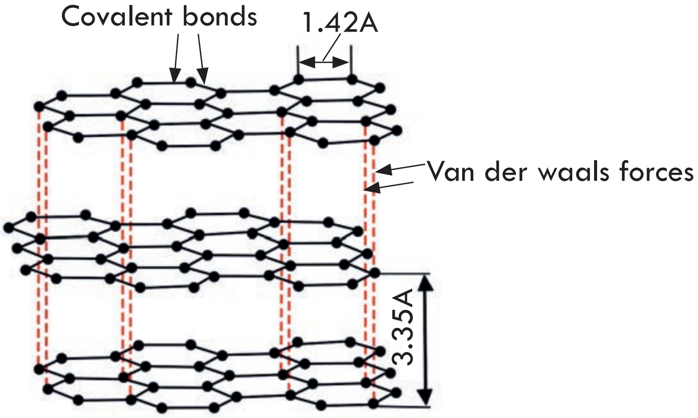


Figure 4.4 Structure of graphite.

b) Electrolytic conductivity

Electrolytes are substances that transmit electricity in a molten state or in aqueous solution. Unlike metallic conductivity, the conductivity of electrolytic solutions depends on the type and concentration of ions in solution. Based on their degree of ionization or the extent to which they produce anions and cations, electrolytes can be classified as strong electrolytes or weak electrolytes. Recall from unit 2 that many ionic compounds such as NaCl dissociate completely in water and are therefore called strong electrolytes.

$$NaCl\ (aq) \rightarrow Na^+ (aq) + Cl^- (aq)$$

When electrical potential is applied through an electrolyte solution, the positive ions (cations) move in one direction and the negative ions (anions) move in the opposite direction. This movement of ions through the electrolyte, brought about by the application of electricity, is called electrolytic conductivity. Hence, the charge-carriers in electrolytic conductivity are ions (anions and cations). That is electrolytic conductors are also called ionic conductors.

Difference between metallic conduction and electrolytic conduction

Metallic Conduction	Electrolytic Conduction
Conduction through metals is due to the movement of electrons.	Conduction through electrolytes is due to the movement of ions.
No change in the chemical properties of the conductors is observed.	It may cause chemical decomposition of the electrolytes.

It does not involve any transfer of matter.	It involves the transfer of matter as ions.
The resistance of metals is due to obstacles of vibrating kernels of metal atoms.	The resistance of electrolytic solutions is due to factors like concentration inter-ionic attractions, the viscosity of solvent etc.
It decreases with the increase in temperature.	It increases with the increase in temperature.

Note that metallic conduction is a physical process, but electrolytic conduction is accompanied by chemical changes that occur on the surface of the electrodes.

Experiment 4.2

Conductivity of Materials

Objectives: To test the conductivity of substances and classify them as strong, weak or non-conductors of electricity.

Apparatus: Copper wire, light bulb, power supply (DC) or dry cells, scotch tape, distilled water, 1M NaCl solution, 1 M HCl, 1 M acetic acid solution, 1 M sugar solution, pencil graphite, wood, magnesium ribbon and an iron nail.

Procedure:

1. Arrange the conductivity apparatus as shown in *Figure 4.5*.

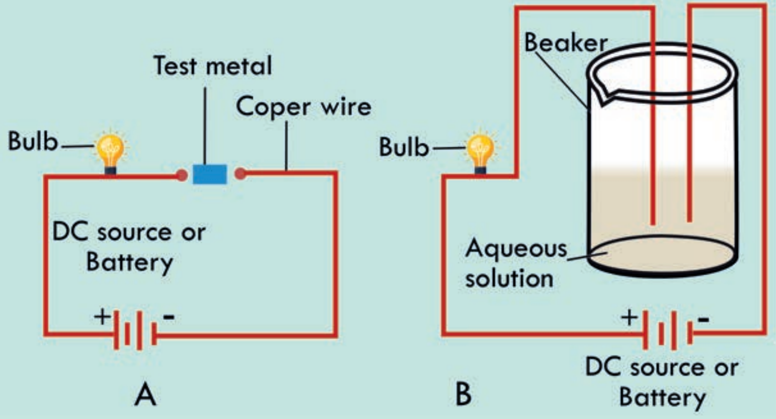


Figure 4.5 Conductivity apparatus for (A) metals and (B) electrolytic solution

2. Pour some distilled water in the beaker (as illustrated in *Figure 4.5B*), dip the electrodes into the water and turn the switch on. Set-up the experiments to test the conductivity of pencil graphite, magnesium ribbon, wood or an iron nail as illustrated in *Figure 4.5A*.

3. Repeat the experiment with separate solutions of table salt, copper sulphate, hydrochloric acid, sodium hydroxide, acetic acid, ammonia solution, molten lead bromide and sugar solution. (use 1.0 M solutions of each).

Observations and Analysis:

a. Does the bulb glow when the switch is turned on?

b. Solutions of which substances make the bulb glow and not glow, when you turn on the switch?

c. Solutions of which substances make the bulb to glow with a: i) bright light? ii) dim light?

d. Classify the substances used in this experiment as strong conductors, weak conductors and non-conductors by completing the table below:

Strong Conductors	Weak Conductors	Non-conductors

e. Which substances in the experiment are used as: i strong electrolytes ii weak electrolytes

Repeat the experiment with lemon juice, and orange juice and answer the questions.

Note that if the solution contains ions, current flows through the circuit and the light bulb glows. The brightness of the bulb increases with current strength, which increases with the number of ions in the solution. Strong electrolytes dissociate completely and produce more ions in solutions brighter light than weak electrolytes that dissociate slightly.

f. In your experiment using set-up 4.5 A, what observations did you make? In which case did the bulb glow and in which case did not? Pencil graphite, magnesium ribbon, wood, or iron nail?

Write a laboratory report on what you have observed and submit to your teacher.

Exercise 4.2

1. Na and Ca metals conduct electricity in solid state but NaCl and $CaCl_2$ conduct electricity only when they are dissolved in water or when they are in the molten form, but not in the solid state. Why?
2. Why are solutions of strong electrolytes better conductors of electricity than weak electrolytes?
3. Which of the following substances are capable of conducting electricity? Give reason for your answer.
   - a. Iron
   - b. Sulphur
   - c. Solid sodium chloride
   - d. Molten calcium chloride

Electrolytes and Nonelectrolytes

At the end of this section, students will be able to
- distinguish between strong and weak electrolytes;
- use conductivity apparatus to test conductivity of substances.

All solutes that dissolve in water fall into one of two categories: electrolytes and nonelectrolytes. An electrolyte is a substance that, when dissolved in water, results in a solution that can conduct electricity. In order to conduct a current, a substance must contain mobile ions. All ionic compounds are electrolytes. When ionic compounds dissolve, they dissociate (break apart) into ions, which are then able to conduct a current. Thus, we can represent sodium chloride dissolving in water as

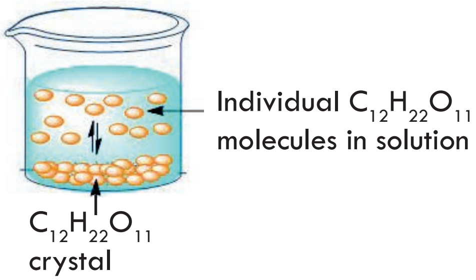
$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

What this equation says is that all the sodium chloride that enters the aqueous solution ends up as Na^+ and Cl^- ions; there are no undissociated NaCl units in solution.

Even insoluble ionic compounds, such as $CaCO_3$, are considered electrolytes because they can conduct a current in the molten (melted) state. Some polar covalent compounds such as acetic acid (CH_3COOH) that can, at least partially, dissociate into ions when dissolved in water are also classified as electrolytes.

A nonelectrolyte is a substance that does not conduct an electric current in either aqueous solution or in the molten state. Most polar covalent compounds, such as table sugar ($C_{12}H_{22}O_{11}$), and all nonpolar covalent compounds are good examples of nonelectrolytes. When these compounds dissolve in water, they do not produce ions. An equation can still be written that simply shows the solute going into solution without dissociating into ions. For example, the process of dissolving sucrose in water can be written as follows:

$$C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$$



Strong and Weak Electrolytes

A strong electrolyte is a substance that when dissolved in water dissociates completely into its ions. All ionic compounds are strong electrolytes. NaCl is a good example. A weak electrolyte is a substance that when dissolved in water dissociates partially into its ions. For example, acetic acid (CH_3COOH)- the compound in vinegar and nitrous acid (HNO_2) only partially ionizes into respective ions (H^+ and CH_3COO^- or H^+ and

NO_2^-) when dissolved in water. Aqueous nitrous acid is composed of only about 5% ions and 95% intact (undissociated) nitrous acid molecules. *Table 4.1* shows some strong and weak electrolytes.

$$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$$
$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^- (aq)$$

Table 4.1: Classification of Solutes in Aqueous Solutions.

Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
HCl	CH_3COOH	$(NH_2)_2CO$ (urea)
HNO_3	HF	CH_3OH (methanol)
$HClO_4$	HNO_2	C_2H_5OH (ethanol)
H_2SO_4	NH_3	$C_6H_{12}O_6$ (glucose)
NaOH	H_2O	$C_{12}H_{22}O_{11}$ (sucrose)
$Ba(OH)_2$		
Ionic compounds		

Experiment 4.3

The Distinction Between Electrolyte and Nonelectrolyte Solutions

Objective: To distinguish between electrolyte and nonelectrolyte solutions
Chemicals: Water, Table salt
Apparatus: power source, wires, beaker (or water glass) and light bulb

Procedure

1. Construct the light bulb experiment as shown in *Figure 4.6*.
2. Fill the glass with water as shown in the Figure. Write your observation
3. Add a tea spoonful of fine NaCl in the water and let it dissolve. Make sure the circuit is complete. Write your observation.

Caution: Bare wires are a hazard! Skin will conduct an electric current and will cause a shock. Do not touch the bare part of the wires when the light-bulb apparatus is plugged in.

Question

In which case does the light bulb glow? With just water or salt dissolved in it? Explain your observations. Try the same experiment using table sugar (sucrose) instead of salt. Apply the concept to test whether the drinking water in your locality is salty or potable. Note that the brightness of the light bulb depends on the concentration of the salt.

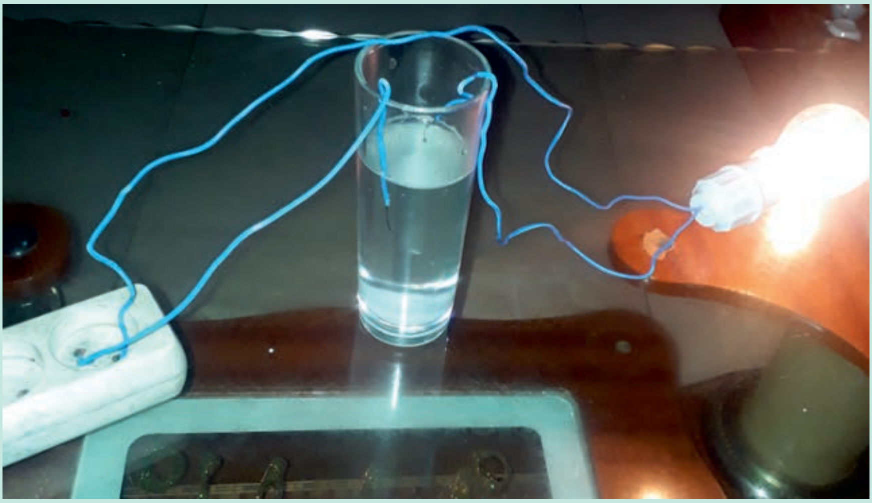
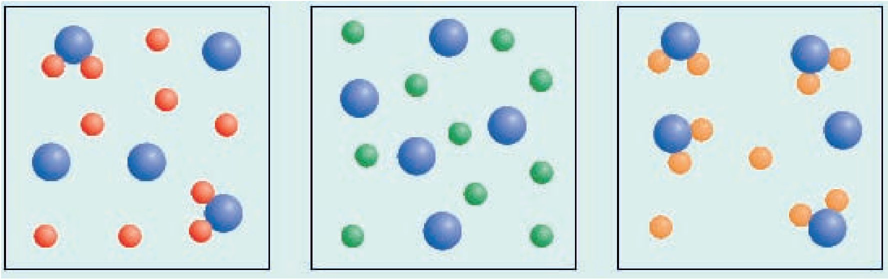


Figure 4.6 Setup for electrolytic conductivity test.

Exercise 4.3

1. Classify each of the following as electrolytes or nonelectrolytes:
   - a. air
   - b. pure water
   - c. saline solution
   - d. lemon water
   - e. sulphuric acid
   - f. nitric acid
   - g. ammonia solution
2. The diagrams show three compounds (a) AB_2, (b) AC_2, and (c) AD_2 dissolved in water. Which is the strongest electrolyte and which is the weakest? (For simplicity, water molecules are not shown.)



(a) (b) (c)

4.3 Electrochemical Cells

At the end of this section, students will be able to describe the importance of electrochemical cells in daily life (Votic cell and car battery).

Activity 4.5

Consider the catabolic biochemical reactions such as breakdown of glucose during respiration which releases energy (exothermic) and anabolic reactions such as joining of amino acids to form proteins (endothermic). Where do you think these reactions take place in your body? What about photosynthesis in plants? Where does it take place? Can you draw analogy between the places (setup) where biochemical reactions take place in your body to the setup for redox reactions to produce electrical energy or produce useful compounds?

An electrochemical cell is a system that incorporates a redox reaction to produce or use electrical energy. It consists of two or more electrodes that are dipped into an electrolyte (see ***Figure 4.7 & 4.10***) in which a chemical reaction either uses or generates an electric current, and the electronic circuitry for controlling and measuring the current and the potential. Electrodes are strips of metal or graphite that allow electrons to leave or enter the electrolytes. They can be chemically active or inert. Active electrodes directly take part in reactions. Examples include zinc and magnesium. Inert electrodes do not directly take part in chemical reactions. They only serve to transfer electrons. Examples include platinum and graphite.

Depending on the purpose or the types of reactions taking place, electrochemical cells are divided into two. These are Galvanic (Voltaic) cells and Electrolytic cells.

4.3.1 Galvanic (Voltaic) Cells

What is a galvanic cell?

A Galvanic cell or voltaic cell is an electrochemical cell in which chemical energy contained in reactants is converted into electrical energy via a spontaneous redox reaction. Galvanic cell or voltaic cell was so named after the Italian scientists Luigi Galvani and Alessandro Volta, who constructed early versions of the device.

Examples include the cells we use in our hand batteries (dry cells), flashlight batteries, wrist watches, cameras and car batteries. The redox reactions between the chemicals in these cells are responsible for the generation of electricity. The reaction proceeds on its own and without any external influence. This type of chemical reaction is called spontaneous redox reaction. Therefore, Galvanic cells are electrochemical cells in which spontaneous redox reactions generate electricity. Galvanic cells are classified into primary cells, secondary cells, and fuel cells. Here, we shall be concerned with the primary and secondary cells only. Fuel cells will be discussed in grade 11. Both primary and secondary cells are batteries that store electrical energy in the form of chemical energy. The chemical energy is converted into electrical energy when we the battery is in use.

Primary Cells (Battery Cells)

Primary cells or also known as battery cells are galvanic cells that are not rechargeable. This is because once the chemicals in the cells that serve as reactants are completely used up, it is not possible to recover them by charging the cells. Examples of primary cells include Daniell's cell and zinc-carbon (Leclanche) dry cells. The common feature of all Galvanic cells is that they contain two electrodes in contact with an electrolyte. The electrolyte in a Galvanic cell can be in the form of a solution (Daniell cell) or a paste (Leclanche cell). The cells containing electrolytes in the form of solution are called wet cells, and those containing electrolytes in the form of paste are called dry cells.

An example of a wet primary cell is the Daniell cell. It consists of a zinc strip placed in $ZnSO_4$ solution in one compartment and a copper strip placed in copper sulphate, $CuSO_4$, solution in another compartment (*Figure 4.7*).

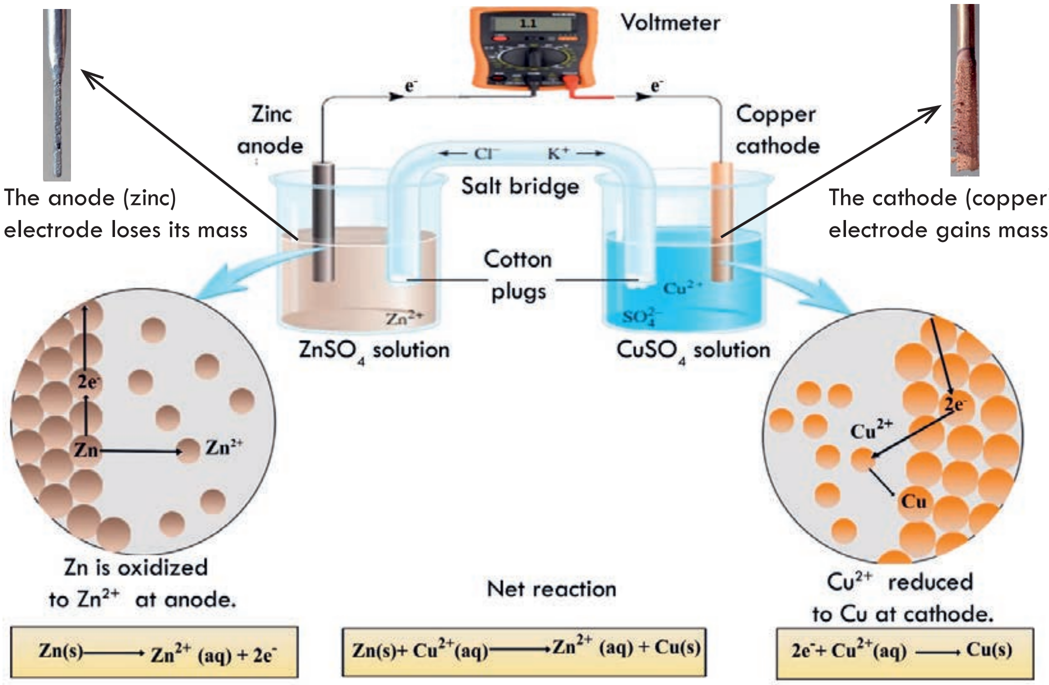


Figure 4.7 The Daniell Cell.

Each compartment is called a half-cell, and the reactions occurring in each compartment are called half-cell reactions. The solutions in the two compartments are linked by a salt bridge. The salt bridge consists of a delivery tube filled with warm mixture of conc. KCl solution and agar solution, which is then allowed to cool so that it sets in the form of a gel.

In the anode compartment, the zinc atoms from the zinc electrode lose two electrons each and become zinc ions, Zn^{2+}. The ions enter into the solution, and the electrons remain on the electrode and flow through the external wire to the copper electrode (see anode half reaction below). This situation causes the zinc electrode to be negative and the solution to have an overall positive charge. In the process the anode (zinc) electrode loses its mass. That is why the anode of the galvanic cell is sometimes called sacrificial electrode. On the other hand, in the compartment containing the copper electrode, copper ions, Cu^{2+}, from the solution move to the cathode and gain two electrons each, to become copper atoms and deposit on the surface of the copper electrode (see cathode half reaction below). This condition causes the electrode to be positive and the solution to have a negative charge. In addition, the cathode continuous to grow in mass as more and more copper deposits on it. The electron flows from anode to cathode in the external conductor.

The half-cell reactions and the cell reaction in Daniell cells are represented as

Anode reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e-$
Cathode reaction: $Cu^{2+}(aq) + 2e- \rightarrow Cu(s)$
Overall (cell) reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Due to the oxidation-reduction reaction in the cell, the Daniell cell generates electricity. What is the purpose of the salt bridge in ***Figure 4.7***? From the preceding discussion, it is clear that the solution in which the zinc electrode is placed has an overall positive charge while the solution in the copper compartment has a negative charge. Unless the two solutions are neutral, the cell cannot produce electricity. Thus, the purpose of the salt bridge is to maintain electrical neutrality between the two solutions.

Although, wet cells like the Daniell cell can serve as a source of electricity, they are not portable since they contain solutions. Due to this practical problem of using wet cells, dry cells were developed. In a dry cell, a moist electrolyte paste is used instead of solutions. This cell was invented by Georges Leclanche, a French chemist.

We commonly use Leclanche cells as convenient, portable sources of energy. Flashlights and radios are examples of devices that are often powered by the zinc–carbon, or Leclanché, dry cell (***Figure 4.8***). This voltaic cell has a zinc can as the anode; a graphite rod in the center, surrounded by a paste of manganese dioxide, ammonium and zinc chlorides, and carbon black, as the cathode. It produces electricity as a result of a spontaneous redox reactions:

Anode reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e-$
Cathode reaction: $2NH_4^+(aq) + 2MnO_2(s) + 2e \rightarrow Mn_2O_3(s) + H_2O(l) + 2NH_3(aq)$

Overall (cell) reaction:

$$Zn(s) + 2NH_4^+(aq) + 2MnO_2(s) \longrightarrow Zn^{2+}(aq) + Mn_2O_3(s) + H_2O(l) + 2NH_3(aq)$$

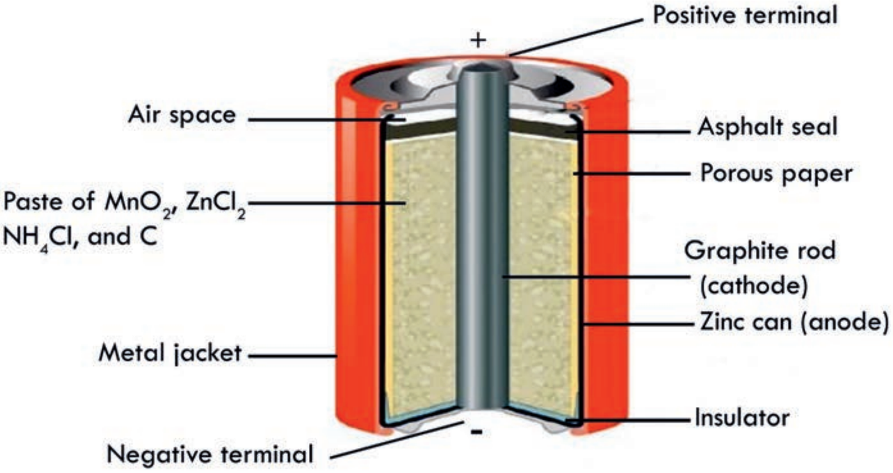


Figure 4.8 Zinc-carbon dry cell (Leclanche cell).

A buildup of ammonia gas around the cathode may disrupt the current. However, this is prevented by the reaction between Zn^{2+} and NH_3 to form a complex ion, $[Zn(NH_3)_2]^{2+}$ which crystallizes as a chloride salt through reaction with chloride (spectator ion).

$$Zn^{2+}(aq) + 2NH_3(aq) + 2Cl^-(aq) \longrightarrow Zn(NH_3)2Cl_2(s)$$

Therefore, the overall reaction occurring in a Leclanche cell is often written as:

$$Zn(s) + 2NH_4Cl(aq) + 2MnO_2(s) \longrightarrow Zn(NH_3)2Cl_2(s) + Mn_2O_3(s) + H_2O(l)$$

The voltage of this dry cell is initially about 1.5 V, but it decreases as current is drawn off. The voltage also deteriorates rapidly in cold weather. There are other examples of dry cells, such as the alkaline dry cell, silver oxide cell, and copper oxide cell which are left for interested students to read.

Exercise 4.4

A voltaic cell is constructed from a half-cell in which a cadmium rod is dipped into a solution of cadmium nitrate, $Cd(NO_3)_2$, and another half-cell in which a silver rod is dipped into a solution of silver nitrate, $AgNO_3$. The two half-cells are connected by a salt bridge. Silver ion is reduced during operation of the voltaic cell.

- ☞ Draw a sketch of the cell.
- ☞ Label the anode and cathode, showing the corresponding half reactions at these electrodes.
- ☞ Indicate the electron flow in the external circuit (with a lightbulb) and the signs of the electrodes.

Project 4.1

Constructing Simple Galvanic Cells

Objectives: To construct a simple Galvanic cell.

Direction: Construct a modified zinc-carbon dry cell (Leclanche cell) using your own materials.

Write a project report and submit to your teacher. Demonstrate the functioning of the developed cell in the class. Your report should describe the basic features of the developed cell. The equations for the redox reactions involved. The half-reaction reactions and where (the electrode) they occur. Compare the performance (light intensity) of your cell with the commercially available cells.

b. Secondary (Rechargeable) Cells

Unlike primary cells, secondary cells are rechargeable. The electrode reactions can be reversed, and the original reactants can be regenerated. This can be achieved by passing a direct current through the cell. The process is called charging or recharging. A secondary cell needs to be recharged when it stops producing electricity. A lead storage battery is an example of a secondary cell. A lead storage battery is the common automobile battery that usually delivers either 6 or 12 volts, depending on the number of cells used in its construction. The inside of this Voltaic cell consists of electrodes of lead alloy grids; one electrode is packed with a spongy lead to form the anode, and the other electrode is packed with lead (IV) oxide to form the cathode (see ***Figure 4.9***). Both are bathed in an aqueous solution of 35% sulfuric acid, H_2SO_4. A single lead-storage cell delivers 2 volts. Therefore, a 12 V battery contains six cells connected in series (***Figure 4.9***).

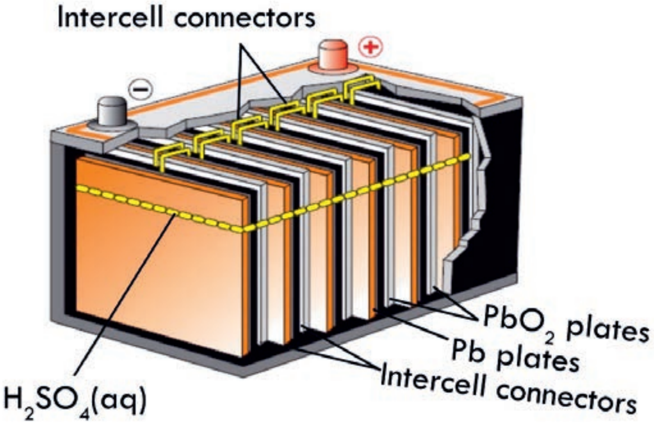


Figure 4.9 A Lead - Storage Battery. Each cell delivers about 2 V, and a battery consisting of six cells in series gives about 12 V

When a lead-storage battery is in operation (on discharge), the following reactions occur at the electrodes:

Anode reaction: $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e-$

Cathode reaction: $PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e- \rightarrow PbSO_4(s) + 2H_2O(l)$

Overall (cell) reaction: $Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$

From the electrode reactions it can be noticed that $PbSO_4$ is produced at both electrodes. Lead-storage battery, as a secondary cell, is rechargeable when it runs down. The electrode reactions can be reversed by applying a potential across the electrodes that is slightly larger than that which the battery can deliver. The reaction that takes place on recharging a lead storage battery is given by the following equation.

$$2PbSO_4(s) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(s) + 2H_2SO_4(aq)$$

Note: The electrode that served as anode during discharging (use) is forced to behave as cathode during recharging. Consequently, the direction of the electron flow that was from left (Pb = anode) to right (PbO_2 = cathode) is changed to right (PbO_2 = anode) to left (Pb = cathode). Oxidation always takes place at anode and reduction at cathode. The direction of electron flow is always from anode to cathode.

Project 4.2

Do you know how car batteries are maintained? Your teacher will arrange a visit to the nearby car battery maintenance center. Before the trip, prepare the following questions for interview.

Why car battery fails?

What are the components in the car battery?

How are they assembled to a self-contained device?

What is the advantage of car battery over other types of batteries such as dry cells? Each dry cell can generate 1.5 V. If we assemble 8 of them in series, we can generate 12.0 V. So, why can't we use this option?

How the health or the degree of discharge is tested?

How is the battery recharged?

After getting answers for the above questions, and reaching conclusion after discussion in group, write a report about what you have learnt during your visit and present it to the class.

4.4 Electrolysis

At the end of this section students will be able to

- define the term electrolysis;
- define the terms electrode, anode, cathode, electrolyte, anion, and cation;
- describe electrolytic cell;
- draw labelled diagram of an electrolytic cell;
- define the terms' half-reaction and cell reaction;

- write the oxidation half-reaction, reduction half reaction and cell reaction for the electrolysis of molten or fused electrolytes;
- perform an activity to show electrolysis of molten electrolytes.

Activity 4.6

1. In the voltaic cells such as primary or secondary you have just learnt, chemical reaction generates electricity or chemical energy is converted into electrical energy. What about the reverse? Which of the two types of voltaic cells can be adapted to serve as electrolytic cell; the primary or secondary?
2. Can you infer the meaning of the word "electrolysis" by breaking it into 'electro' and 'lysis'? Remember that you applied an electrical potential across the electrodes to cause the reaction:

$2PbSO_4(s) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(s) + 2H_2SO_4(aq)$ in secondary cells.

Now, based on the information from the above analysis can you define electrolysis?

In contrast to spontaneous redox reactions, which result in the conversion of chemical energy into electrical energy, electrolysis is a process by which electrical energy is used to drive an otherwise nonspontaneous reaction. Electrolysis stands for "electro" means electricity and "lysis" means to separate or break. So in terms, electrolysis would mean "breakdown via electricity". An electrochemical cell in which electrolysis occurs is called electrolytic cell (***Figure 4.10***). It has two components, electrodes (anode and cathode) and electrolyte.

In electrolytic cell, the electrode connected to the positive terminal of the source (e.g. battery) is positively charged and is called the anode. It is the electrode through which electrons leave the cell. The electrode connected to the negative terminal of the source is negatively charged and is called the cathode. It is the electrode through which electrons enter the cell (see ***Figure 4.10***).

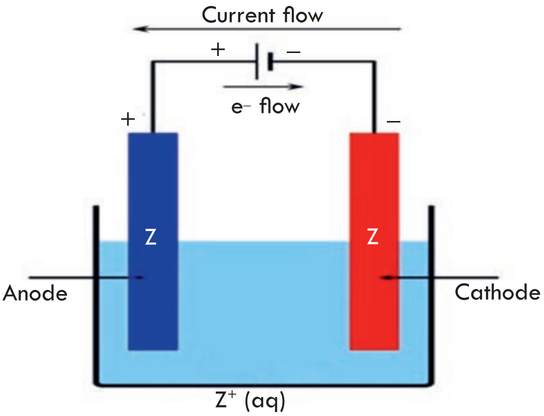


Figure 4.10 A sketch of an electrolytic cell.

During electrolysis, the ions of the electrolyte migrate to the electrodes of the opposite charge. The positive ions are attracted to the cathode and are called cations. Since the cathode has excess electrons, the cations will discharge by gaining electrons. This process of gaining electrons is called reduction. The negative ions are attracted by the positive electrode (anode) and are called anions. These ions are discharged by losing electrons (get oxidized) at the anode. This process of losing electrons is called oxidation. Thus, the cathode is the electrode at which reduction occurs and the anode is the electrode at which oxidation takes place.

4.4.1 Electrolysis of Molten (Fused) Electrolytes

Ionic solids can melt at elevated temperature. NaCl, for instance, melts at 801 °C. The process is called melting or fusion. When ionic solids melt, they dissociate into positive and negative ions that are free to move. To understand the chemical reactions that occur during electrolysis of molten (fused) electrolytes, consider electrolysis of molten NaCl. *Figure 4.11* shows a simple electrolytic cell. Wires from a battery are connected to electrodes that are dipped into molten sodium chloride. At the electrode connected to the negative pole of the battery, globules of sodium metal form; chlorine gas evolves from the other electrode. The half-reactions are

$$Na^+(l) + e^- \rightarrow Na(l)$$
$$Cl^-(l) \rightarrow 1/2Cl_2(g) + e^-$$

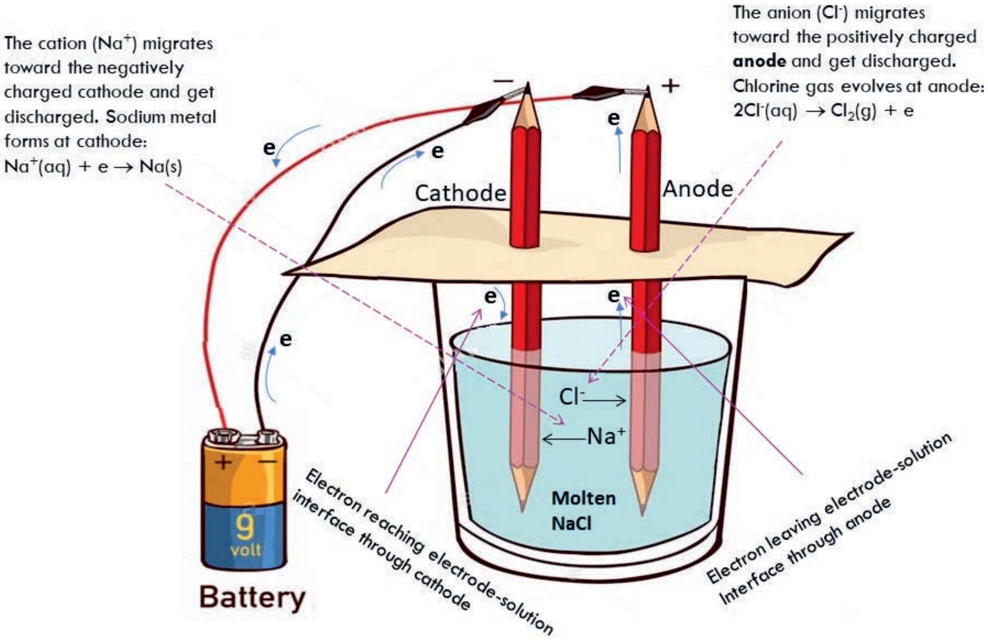


Figure 4.11 Electrolysis of molten sodium chloride. Sodium metal is produced commercially this way.

The reduction of Na^+ to Na occurs at the cathode, and the oxidation of Cl^- to Cl_2 occurs at the anode (note the labeling of electrodes in ***Figure 4.11***.
You obtain the cell reaction by adding the half-reactions.

$$\begin{array}{l} Na^+(l) + e^- \rightarrow Na(l) \\ Cl^-(l) \rightarrow 1/2Cl_2(g) + e^- \\ \hline Na^+(l) + Cl^-(l) \rightarrow Na(l) + \frac{1}{2} Cl_2(g) \end{array}$$

As you can see from the cell reaction, the electrolysis of molten NaCl is used commercially to obtain sodium metal from sodium chloride. A number of other reactive metals are obtained by the electrolysis of a molten salt or ionic compound. Lithium, magnesium, and calcium metals are all obtained by the electrolysis of the chlorides.

Exercise 4.5

1. A lead-storage buttery is an example of _____
   - a. an electrolytic cell called secondary cell
   - b. galvanic cell called secondary cell
   - c. rechargeable electrolytic cell called secondary cell
   - d. only b and c
2. How do you know if an electrode is a cathode or an anode in voltaic cells and in electrolytic cells?
3. Can you suggest why positive and negative ions are named as cations and anions respectively?
4. Distinguish between Voltaic cells and electrolytic cells.
5. During electrolysis of fused $PbBr_2$,
   - a. Which ions are responsible for the conduction of electricity through the molten salt?
   - b. What half-cell reactions occur at the anode and cathode?

Experiment 4.4

Electrolysis of Molten (Fused) Lead Bromide

Objectives: To observe substances produced at the electrodes during electrolysis.

Materials required: Stand and clamp, two graphite electrodes, wires, switch, light bulb, test tube (bigger in size), Bunsen burner, lead bromide crystals (melting point: 373°C).

Precaution: Bromine causes very severe burn on the skin.

Procedure:

1. Assemble the materials as shown in *Figure 4.12*.
2. Place small amount of lead bromide crystals in a beaker. Insert the two electrodes, as shown in *Figure 4.12*, until they are in contact with the lead bromide crystals. Then turn on the switch. Does the bulb glow?
3. Heat the lead bromide in a beaker gently, using a Bunsen burner. When the lead bromide melts, turn on the switch.

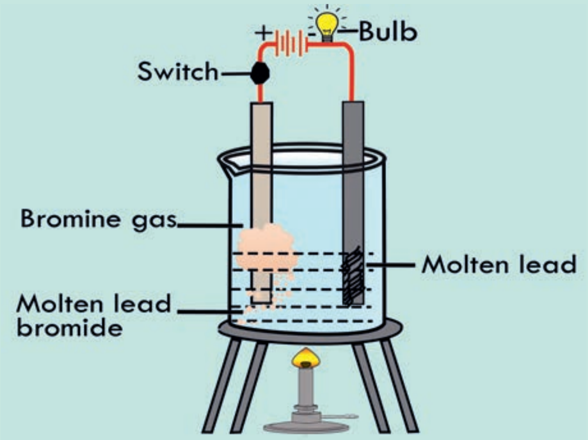


Figure 4.12 Electrolysis of Fused Lead Bromide.

Observations and Analysis:

1. Does the bulb glow? If yes, what is the reason?
2. Write the dissociation reaction for $PbBr_2$.
3. Identify the ions which migrate to the respective electrodes
4. Write the products formed at each electrode.
5. Identify the half reactions as oxidation and reduction.
6. Write the cell reaction

Write a laboratory report on your observation and present to the class.

Activity 4.7

Consider the text entitled '**Energy and Climate**' below and answer the following questions!

1. Discuss the impact of nation's energy demand on climate.
2. Explain the role of youth in controlling climate change.
3. Which type of nations do you think contribute largely to climate change, developed or developing?
4. Discuss how Ethiopia should benefit from large-scale tree planting.
5. Explain the relevance of tree planting in the carbon capture and sequestration.

Energy and Climate

Today, no scientific challenge is greater than reversing the climatic effects of our increasing dependence on the combustion of fossil fuels-coal, petroleum, and natural gas. Because these fuels form so much more slowly than we consume them, they are nonrenewable. In contrast, wood and other fuels derived from plant and animal matter are renewable.

All carbon-based fuels release CO_2 when burned. The ability of CO_2 to absorb heat plays a vital temperature-regulating role in the atmosphere. Much of the sunlight that shines on Earth is absorbed by the land and oceans and converted to heat. Like the glass of a greenhouse, CO_2 traps some of the heat radiating back from Earth's surface and, thus, warms the atmosphere. This process is called the natural greenhouse effect which changes Earth's climate.

Scientists are trying to develop alternative energy sources that do not release greenhouse gases. In addition to developing alternative energy sources to reduce fossil-fuel consumption, some researchers are studying carbon capture and sequestration, accomplished by large-scale tree planting and by liquefying CO_2 released from coal fired power plants and burying it underground or injecting it deep into the oceans.

The climate change is a concern to the global nations too. In 1997, the United Nations Conference on Climate Change in Kyoto, Japan, created an international treaty that set legally binding limits on release of greenhouse gases. It was ratified by 189 countries, but not by the largest emitter of CO_2, the United States. The 2005 conference in Montreal, Canada, presented overwhelming scientific evidence that confirmed the human impact on climate change, but the 2009 conference in Copenhagen, Denmark, which was attended by representatives from 192 nations, failed to produce a legally binding agreement on emission targets. In late 2010, however, the conference in Cancun, Mexico, resulted in some financial commitments to help developing countries rely on alternative energy sources.

Exercise 4.6

1. Write the half-reactions for the electrolysis of the following molten compounds: a. KCl; b. KOH.
2. Is the polarity of the electrodes in Galvanic cells similar to that of the electrodes in electrolytic cells? Explain your answer.

Key Terms of the Unit

- anode
- cathode
- conductivity apparatus
- Daniell cell
- electrical conductivity
- electrochemical cell
- electrochemistry
- electrode
- electrolysis
- electrolyte
- electrolytic cell
- endothermic reaction
- energy
- energy change
- enthalpy
- enthalpy change
- exothermic reaction
- galvanic cell
- half - cell
- half - reaction
- heat
- heat of reaction
- heat capacity
- internal energy
- kinetic energy
- Leclanche cell
- potential energy
- primary cell
- salt bridge
- secondary cell
- specific heat capacity
- voltaic (galvanic) cell
- work

Unit Summary

- All chemical reactions accompany energy change.
- There are many forms of energy and they are all interconvertible.
- The internal energy of a substance is the sum of its potential energy and kinetic energy.
- Potential energy is stored energy.
- Potential energy is stored in chemical bonds and as intermolecular forces.
- Chemical energy is a form of potential energy.
- The energy change occurs when energy transfer takes place between system and surrounding.
- Energy transfer takes place in the form of heat or work or both.
- Any chemical reaction that gives off heat to the surroundings is called an exothermic reactions; any Chemical reaction that absorbs heat from the surroundings is called an endothermic reaction.
- Enthalpy is a measure of energy change of the system of reactant and products.
- Enthalpy change is calculated as enthalpy of products minus enthalpy of reactants.
- Enthalpy change is positive for endothermic reactions and negative for exothermic reactions.
- Like any other reaction, redox reactions also involve energy change.
- The energy change in redox reactions is dealt in electrochemistry.
- Electrochemistry is a field of chemistry that studies how chemical reactions produce electricity and how electricity is used to bring about chemical reactions in electrochemical cells.
- Electrical conductivity is the ability of substances to conduct electricity.
- Metallic conductivity is the flow of electricity through metals, and the conduction of electricity through metals is due to the presence of freely moving (delocalized) valence electrons.
- Electrolytes are substances that conduct electricity in solution or in a molten state.
- The conduction of electricity through electrolytes is due to the movement of

anions and cations towards electrodes of opposite charge, and the charge carriers in electrolytic conduction are ions.

- Cations carry positive charge whereas anions carry negative charge
- Electrochemical cells include Galvanic (Voltaic) cells and electrolytic cells.
- Galvanic cells convert chemical energy into electrical energy.
- Electrolytic cells convert electrical energy into chemical energy.
- Electrolysis is a process in which electricity is used to bring about an oxidation-reduction reaction in an electrolytic cell and produces useful substances.
- Electrolysis is also the process of decomposition of an electrolyte, using electrical energy.
- Electrodes are either metal strips or graphite rods.
- Primary cells are voltaic cells that are not rechargeable, and the reactions taking place in them are irreversible.
- Secondary cells are voltaic cells that are rechargeable since the reactions taking place in them are reversible.
- Unlike electrolytic cells, the anode is negative and the cathode is positive in galvanic/voltaic cells.
- In voltaic cells, the polarity (charge) of the electrodes is determined by the redox reactions.
- In electrolytic cells the polarity of the electrodes is determined by the external power source.
- An anode is the electrode attached to the positive terminal of a direct current source in electrolytic cell, at which oxidation (loss of electrons) by anions occurs, and electrons leave the cell.
- A cathode is the electrode attached to the negative terminal of a DC source in electrolytic cell, the negative electrode at which reduction of cations occurs and electrons enter the cell.
- Oxidation always takes place at anode and accompanied by reduction at cathode.
- During electrolysis, anions move to the anode and cations move to the cathode.
- The reaction taking place at each electrode (cathode or anode) is said to be a half-cell reaction.
- Cell reaction is the reaction that takes place in the entire cell and obtained by adding half reactions.

Review Exercise

Part I: Choose the correct answer from the suggested options

1. The energy released during chemical reactions comes from
   a. the bond being formed
   b. the bond being broken
   c. the heat supplied from the surrounding
   d. all

2. Which of the following is not correct about voltaic cells?
   a. the anode is negative
   b. the cathode is positive
   c. oxidation takes place at the cathode
   d. redox reactions produce electricity in the cell
3. Which one of the following is used to produce useful compounds?
   a. Voltaic cells
   b. Electrolytic cells
   c. Daniel's cell
   d. Batteries
4. Electrolysis is not used for the:
   a. purification of metals
   b. production of metals
   c. electroplating of metals
   d. production of electricity
5. Which substance is not used in the Leclanche cell?
   a. H_2SO_4
   b. NH_4Cl
   c. MnO_2
   d. powdered carbon
6. Which one of the following is correct about automobile batteries?
   a. the electrodes are graphite electrodes
   b. each cell in the battery delivers 1.5 volts
   c. the electrolyte is aqueous H_2SO_4 solution
   d. lead (IV) oxide is used as anode
7. Which of the following occurs during electrolysis of the molten binary salt of a metal:
   a. the metal will deposit on the cathode
   b. reduction will take place at the anode
   c. oxidation will take place at the cathode
   d. no current will flow through the molten salt
8. Two copper electrodes dipped in copper sulphate solution are connected to a12 volt battery. The electrode connected to the end of the battery marked with a "–" is:
   a. anion
   c. anode
   b. cathode
   d. cation
9. The charge-carriers in electrolytic conduction are:
   a. anions only
   c. cations and anions
   b. cations only
   d. delocalized electrons
10. During the electrolysis of fused lead bromide, the anode half reaction involves:
    a. oxidation of lead atoms to ions
    b. reduction of bromine atoms to give bromide ions
    c. reduction of lead ions to form free metal
    d. oxidation of bromide ions to elemental bromine

11. Voltaic cells and electrolytic cells are similar in that:
    a. the anode is positive and cathode is negative in both types of cells
    b. oxidation half-reaction occurs at the cathode in both types of cells
    c. both types of cells contain two electrodes in contact with electrolytes
    d. reduction half-reaction occurs at the anode in both types of cells
12. Four different solutions of equal volume (1L) were prepared by dissolving one mole of each of the following substances. The conduction of electricity is least in the solution containing:
    a. HCl
    c. CH_3COOH
    b. HNO_3
    d. KCl
13. Which of the following is a wet voltaic cell:
    a. Leclanche cell
    b. cells used in mobile telephones
    c. lead-storage cell

Part II: Give short answers to each of the following questions

14. Under what condition is the heat of a reaction equal to the enthalpy change of the same reaction?
15. Describe at least two exothermic processes and two endothermic processes.
16. Decomposition reactions are usually endothermic, whereas combination reactions are usually exothermic. Give a qualitative explanation for these trends.
17. Draw an enthalpy diagram for a general exothermic reaction; label the axis, reactants, products, and ΔH with its sign.
18. Draw an enthalpy diagram for a general endothermic reaction; label the axis, reactants, products, and ΔH with its sign.
19. Consider these changes.
    a. $Hg(l) \rightarrow Hg(g)$
    b. $3O_2(g) \rightarrow 2O_3(g)$
    c. $CuSO_4.5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(g)$
    d. $H_2(g) + F_2(g) \rightarrow 2HF(g)$

At constant pressure, which of the reactions is exothermic and which is endothermic?

20. Why are ionic compounds like NaCl, KCl, $CaCl_2$, $PbBr_2$, etc. non-conductors in the solid state but conductors in aqueous solutions?
21. Describe the basic features of a galvanic cell. Why are the two components of the cell separated from each other?
22. Explain the differences between a primary galvanic cell-one that is not rechargeable-and a storage cell (for example, the lead storage battery), which is rechargeable.
23. During electrolysis of fused $CaCl_2$,
    a. Which ions are responsible for the conduction of electricity through the molten salt?
    b. Which ion migrates towards anode?
    c. Which ion migrates towards cathode?
    d. What half-cell reactions occur at the anode and cathode?

UNIT 5

METALS AND NONMETALS

Unit Outcomes

At the end of this unit, you will be able to

- describe the general properties of metals and nonmetals;
- describe the extraction, chemical properties and uses of aluminium, iron, copper, nitrogen, phosphorus, oxygen, sulphur, and chlorine;
- demonstrate scientific inquiry skills: observing, classifying, communicating, and asking questions, applying concepts and problem solving.

Start-up Activity

Discuss in your group and explain your ideas to the class.

1. What properties did you consider when classifying elements as metalic or non metalic? How are this properties related to the uses of the elements?
2. Which elements, metals or non metals, are
   a. mostly used in electrical wires, cooking utensils, and boilers?
   b. used for breathing, purifying water, and fertilizers?
3. List at least eight metals and eight non metals used in our daily lives?
4. How do metals and nonmetals exist in nature?

5.1 Introduction

By the end of this unit, you will be able to

☞ describe the source of metals and non-metals;

☞ explain the natural occurrence of metals and nonmetals.

Natural resources are resources that exist without any actions of human kind such as sunlight, land, fossil fuels (crude oil, coal, and natural gas), crops, water and air. Natural resources can be classified as renewable and non-renewable resources. Renewable resources are resources that can be replenished (replaced) by natural processes. Examples are water, air and plant materials (biomass). Non-renewable resources are resources that are found in a fixed amount in nature and cannot be replenished. Fossil fuels, nuclear energy, metallic and non metallic minerals are examples of non-renewable resources. Metallic minerals are minerals which contain one or more metallic elements. Non-metallic minerals are the minerals that do not contain metals. Metallic minerals occur in rare, naturally formed concentrations known as mineral deposits.

The principal source of metallic and nonmetallic minerals is the planet Earth. The majority of metals and non-metals are extracted from their minerals. Minerals such as iron haematite (Fe_2O_3) or aluminum bauxite ($Al_2O_3.2H_2O$), sulphides such as zinc blende (ZnS), or carbonates such as malachite ($Cu(OH)_2.CuCO_3$) are examples of metallic minerals. Diamond, gravel, and mica are examples of nonmetallic minerals. Metals and non-metals can also be found in seawater. Some non-metallic elements, such as oxygen and nitrogen, are found in the atmosphere.

5.2 General Properties of Metals and Production of Some Metals

At the end of this section, students should be able to

- mention general properties of metals;
- describe the uses of some common metals.

In this section, you will study the general characteristics and extraction of metals, uses of some common metals, alloys and production of Aluminium, Copper and Iron.

5.2.1 Properties and Extraction of Metals

Activity 5.1

Discuss in your group the following metal properties using an example and present your conclusion to the class.

a. A metal which is liquid at room temperature. Is there any metal that exists in the gaseous state at room temperature?
b. A metal which is the best conductor of heat
c. A metal which is the poorest conductor of heat
d. A metal which can be cut with a knife

About 80 % of the known elements are metals. These metals have some common properties

A. Some physical properties of metals

- They have a shining appearance(lustrous) and can be polished. Gold is shining yellow and copper is brown. Iron, aluminum, zinc and silver are lustrous grey or silvery.
- They are malleable, (can be beaten into thin sheets). Aluminum foils, which are used to wrap cigarettes, medicines, etc., gold and silver ornaments of different designs, are due to the malleability of these metals.
- They are ductile, i.e. they can be drawn into wires. For example copper, gold, iron and silver are ductile.
- They are hard and have tensile strength except lithium, potassium, sodium.
- They generally have a high density except lithium, potassium, sodium.
- They are usually sonorous-produce metallic sound when struck (e.g. school bell)

B. Chemical properties of metals

- Metals possess positive valency and have a tendency to lose electrons.

$$M(g) \longrightarrow M^{n+}(g) + ne^-$$

- They have 1, 2, or 3 valence electrons.
- They are oxidized by losing electrons and act as reducing agents
- They mostly form basic oxides and some amphoteric oxides (refer to Unit 3)
- They form chlorides that are true salts and electrovalent.
- They form hydrides which are ionic, unstable and reactive.
- They usually replace hydrogen from dilute non-oxidising acids like HCl and H_2SO_4. Exceptions are copper, silver and gold.

C. Reactivity series of metals

If we take solution of different metals and place other metals in these solutions, we can compare the reactivity of metals with respect to each other. The arrangement of metals in the decreasing order of their activity is known as activity or reactivity series. The most active metal is at the top and the least active metal is at the bottom of the reactivity series. The hydrogen, a nonmetal, has been included in the series only for the comparison of the activity of the metals. The metals above hydrogen like potassium, sodium, calcium and magnesium in the reactivity series liberate hydrogen gas when treated with dilute acids. The metals below hydrogen in the reactivity series like copper, silver, and gold do not liberate hydrogen when treated with dilute acids. A less reactive metal (low in reactivity series) can be displaced from its compound or from its salt solution by a more reactive metal (high in reactivity series). Thus, potassium can displace all the metals from their salt solution

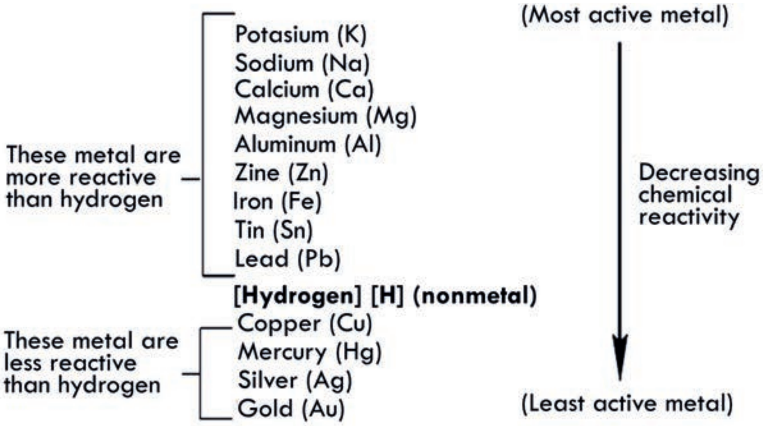


Figure 5.1 Reactivity series of metals.

Metals at the top of the reactivity series are strong reducing agents. Metals at the bottom in the reactivity series are weak reducing agents. Hence, potassium is the strongest reducing agent whereas gold is the weakest reducing agent in the above reactivity series.

D. Natural occurrence and extraction of metals

Noble metals Ag, Au, Bi, Cu, Pd, Pt exist in nature as uncombined or free state. More active metals like alkali and alkaline earth metals never exist in uncombined state. They always exist in compounds. Active metals may exist in the form of carbonates, halides, oxides, phosphates, silicates, sulphides and sulphates.

The constituents of the earth's crust which contain these metals or their compounds are known as minerals. For example, sodium exist as the mineral halite ($NaCl$), potasium as sylvite (KCl), magnesium as magnesite ($MgCO_3$), calcium as lime stone ($CaCO_3$). At some places minerals contain a high percentage of particular metals and the metal can be profitably extracted from it, such minerals are called ores. An ore taken out from the earth contains a lot of impurities in the form of sand and other undesirable materials called gangue. In fact, metals exist in these ores in the form of compounds. The science and technology of extracting metals from their ores and of compounding alloys is called metallurgy.

The three principal steps in the extraction of a metal from its ore are (1) preparation (concentration) of the ore , example : oil floatation, magnetic separation, (2) production of the metal. Example: roasting, calcination, and (3) purification of the metal (last step), example: chemical reduction, electrolytic reduction, as shown in ***Figure 5.2***.

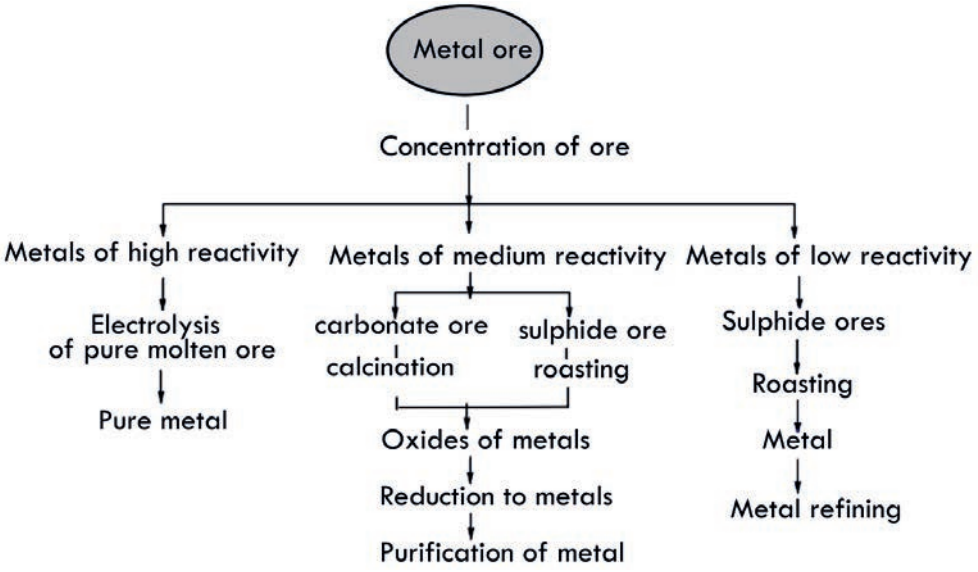


Figure 5.2 General metallurgical processes of metals.

The most active metals such as K, Na, Ca. Mg are extracted from their compounds only by electrolysis because there is no economic reducing agent that can reduce them from their compounds. It is only using electricity (electrolysis) that these metals are extracted or reduced.

Project

Read from a reliable source about the uses of some common metals (sodium, potasium, magnsium, calcium and tantalum) and their common compounds. Submit to your teacher and discuss in class.

5.2.2 Alloys

Alloys are mixtures of two or more metals or metals and nonmetals when molten and do not separate when solidified. The constituent elements are melted together and then allowed to cool to form a solid material called alloy.

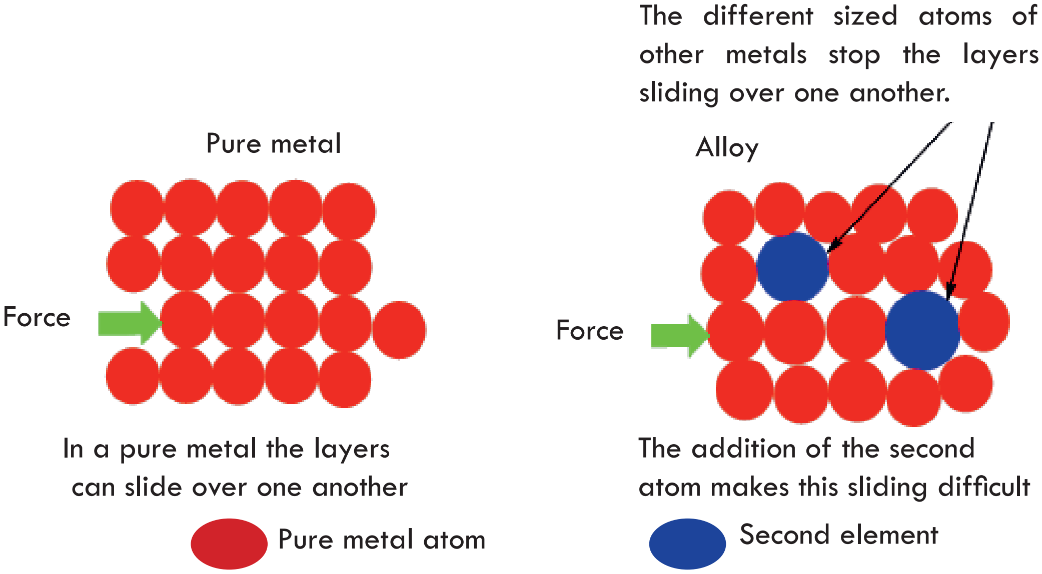


Figure 5.3 Arrangement of metal atoms in pure metal and alloy.

Alloying of metals increases the hardness and strength, modify the color and melting point, decreases the electrical conductivity and increase the resistance to corrosion of metals. For example, amalgam is an alloy of mercury and another metal. Gun metal is an alloy of copper (87 %), tin (10 %) and zinc (3 %).

Solder is called fuse metal. Its composition is 67 % tin and 33 % lead. It melts at 183 ^{0}C which is lower than the melting points of tin (Sn melts at 232 ^{0}C). Solder is used to join wires and electrical resistances, etc.

Gold alloys: Gold is a very soft metal and therefore hardened by alloying it with copper and silver. The quantity of gold in an alloy is called its gold content. The gold content of the alloy is expressed either in carats or as mass percent. A carat is equal to the mass unit of gold in 24 mass units of the alloy. Thus, 24 carat represents pure gold. It doesn't contain any other metal.

22 carat gold contains 22 parts of pure gold by mass in 24 carats of the alloy.

Nowadays 22 carat gold is used to make jewelry. 18 carat contains 18 parts of pure gold by mass in 24 parts of alloy. Thus, the mass percent of gold in different types of alloys can be calculated as follows:

Mass % of gold in 24 carat gold = 24 x100/24 = 100.00 (pure gold, no Cu or Ag)
Mass % of gold in 22 carat gold = 22 x100/24 = 91.67 (alloy, 8.33 % Cu or Ag)

5.2.3 Production of Aluminum, Iron and Copper

At the end of this section, students should be able to

- explain Occurrence and extraction of Aluminum, Iron and Copper;
- describe the the applications of Aluminum, Iron and Copper.

A. Aluminum

At the end of this section, students should be able to

- explain properties, Occurrence and extraction of Aluminum;
- describe the the applications of Aluminum.

Occurrence and extraction

Aluminum is the most abundant metal in the earth's crust and the third abundant element after oxygen and silicon. Approximately eight percent of the earth's crust is aluminum. In terms of consumption, it is the second most important metal after iron. Aluminum does not occur in uncombined or free metal state in nature. Its main mineral is bauxite ($Al_2O_3.2H_2O$). Other minerals containing aluminum are orthoclase ($KAlSi_3O_8$), cryolite (Na_3AlF_6), corundum (Al_2O_3), beryl ($Be_3Al_2Si_6O_8$) and china clay ($Al_2Si_2O_7.2H_2O$).

Activity 5.2

1. What properties of aluminum make it useful for packaging cigarettes and medicines?
2. Steel-cored aluminum cables are used for conducting high-voltage electricity over long distances.
   a. give two reasons why aluminum is used for these cables.
   b. what is the purpose of the steel core?
3. Can you mention materials, in your home or outside, that are made of aluminum? What are their properties for example weight, rusting, shape etc.?

Discuss in groups and give presentation to the class

Industrially, aluminum is extracted from bauxite, $Al_2O_3•2H_2O$, mainly by the Hall–Héroult process or simply the Hall process. In this process, the bauxite must be purified first, because it is often contaminated by silicon dioxide (SiO_2), iron oxide, and titanium (IV) oxide. To separate the pure Al_2O_3 from the bauxite, the powdered ore is first heated with a sodium hydroxide solution to convert the silica to soluble silicate.

$$\underset{\text{Insoluble}}{SiO_2\,(s)} + 2NaOH\,(aq) \rightarrow \underset{\text{soluble}}{Na_2SiO_3\,(aq)} + H_2O\,(l)$$

At the same time aluminum oxide is converted to soluble sodium aluminate.

$$\underset{\text{Insoluble}}{Al_2O_3\,(s)} + 2NaOH\,(aq) \rightarrow \underset{\text{soluble}}{NaAlO_2\,(aq)} + H_2O\,(l)$$

The impurities like iron oxides and titanium (IV) oxide remain unaffected by the base and are filtered off. Why they are not affected?
Then the solution is treated with acid to precipitate aluminium hydroxide.

$$AlO^{2-}\,(aq) + H_3O^+\,(aq) \rightarrow Al(OH)_3\,(s)$$

Aluminum hydroxide is collected by filtration, washed, dried, and then heated strongly in a furnace to get Al_2O_3.

$$2Al(OH)_3\,(s) \xrightarrow{\text{Heat}} Al_2O_3\,(s) + 3H_2O\,(g)$$

Pure aluminum oxide is mixed with some cryolite, Na_3AlF_6, and then melted. Cryolite is added to Al_2O_3 to reduce its melting point from 2045°C to 1000°C. The molten cryolite also provides a good conducting medium for the electrolysis process. The molten mixture is then electrolyzed to obtain pure aluminum. The electrolytic cell, in which the molten mass is electrolyzed, contains graphite electrodes as both anode and cathode (See ***Figure 5.4***).

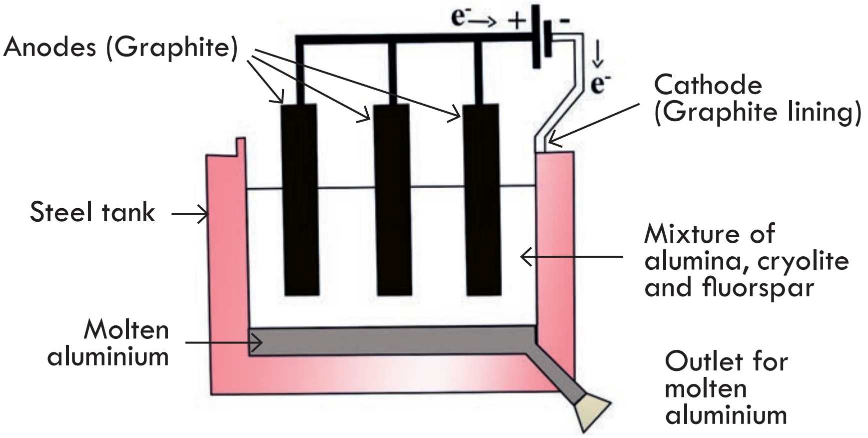


Figure 5.4 Electrolytic reduction of aluminum based on the Hall's process.

Al_2O_3 is converted to Al, by electrolysis. The reactions at the graphite (carbon) electrodes are:

Anode reaction: $3C(s)\ (\text{graphite electrode}) + 6O^{2-} \rightarrow 3CO_2(g) + 12e$

Cathode reaction: $4Al^{3+}(l) + 12e \rightarrow 4Al(l)$

Over all reaction: $4Al^{3+}(l) + 6O^{2-}(l) + 3C(s) \rightarrow 4Al(l) + 3CO_2(g)$

Or $2Al_2O_3(l) + 3C(s) \rightarrow 4Al(l) + 3CO_2(g)$

Note that under the prevailing conditions in the electrolytic cell, the anode (graphite electrode) will gradually oxidize to carbon dioxide and must be replaced regularly. Molten aluminum is siphoned or drawn from the bottom of the electrolytic cell. We do not use reducing agents for reduction of aluminum from aluminum oxide; why do you think?

Physical Properties: Aluminum is a soft silvery-white and light metal having a density of 2.7 g/cm^3. It melts at 660°C. It can be shaped into wires, rolled, pressed or cast into different shapes. It is a good conductor of heat and electricity.

Chemical Properties: Aluminum is a reactive metal, even though its reactions are not vigorous in comparison with those of sodium, potassium, calcium and other metals of group IA and IIA. It undergoes the following reactions:

i. Reaction with Oxygen

Aluminum has affinity for oxygen and reacts with atmospheric oxygen to form a thin film of aluminum oxide on its surface.

$$4Al\ (s) + 3O_2\ (g) \rightarrow 2Al_2O_3\ (s)$$

This thin film of oxide inhibits further reaction with oxygen. The thin film can be removed by rubbing with mercury or mercury (II) chloride solution. What type of oxide is Al_2O_3?

ii. Aluminum burns in Nitrogen gas to form Aluminum Nitride.

$$2Al\ (s) + N_2\ (g) \rightarrow 2AlN\ (s)$$

iii. Reaction with dilute acids

Aluminum reacts with dilute acids like HCl and H_2SO_4, forming salts and liberating hydrogen gas.

$$2Al\ (s) + 3H_2SO_4\ (aq) \rightarrow Al_2(SO_4)_3\ (aq) + 3H_2\ (g)$$
$$2Al\ (s) + 6HCl\ (aq) \rightarrow 2AlCl_3\ (s) + 3H_2\ (g)$$

Aluminum does not react with dilute or concentrated HNO_3 due to the formation of a protective oxide layer on its surface. Therefore, materials made of Al may be used to transport HNO_3.

iv. Aluminum burns in chlorine gas to form aluminum chloride

$$2Al\ (s) + 3Cl_2\ (g) \rightarrow 2AlCl_3\ (s)$$

v. Aluminum reacts with sodium hydroxide solution

$$2Al\ (s) + 2NaOH\ (aq) + 6H_2O\ (l) \rightarrow 2NaAl(OH)_4\ (aq) + 3H_2\ (g)$$

Uses of Aluminum

Activity 5.3

Discuss in your group and present your conclusion to class.

1. Why should the food containing iron salts are not cooked in aluminum utensils?
2. Aluminum is a highly reactive metal, however, still it is used to make utensils for cooking.

Aluminum is a lightweight metal and resistant to corrosion, which is why it is used to make the light alloys, such as duralumin (a mixture of Al, Cu, and Mg). Aluminum alloy is widely used in the transportation industry to make airplanes, ships, and automobiles. Due to its high thermal conductivity and corrosion resistance, it is used to make household cookwares. It is also used in the manufacture of door and window frames and building roofs, as a packaging material in the food industry, and for power transmission lines. Al is also used in the thermite welding process.

Thermite Welding

Thermite welding is useful in welding rails, propeller shaft and other steel parts of machinery that are difficult to repair by any other method. The process is explained as follows:

Aluminum is more reactive than iron, therefore it displaces iron from iron oxide. In the thermite welding process, powdered aluminum when mixed with iron (III) oxide and ignited, produces a temperature of about 3000°C. This temperature is quite sufficient for welding metals. The reaction in this process is as follows:

$$\underset{\text{Aluminum}}{2Al\ (s)} + \underset{\text{Iron oxide}}{Fe_2O_3\ (s)} \rightarrow \underset{\text{Iron}}{2Fe\ (l)} + \underset{\text{Aluminum oxide}}{Al_2O_3\ (s)}$$

The mixture of powdered aluminum and iron oxide is called thermite. Similar reactions are used to insulate other metals such as Chromium:

$$Al(s) + Cr_2O_3 \rightarrow 2\ Cr\ (l) + Al_2O_3\ (s)$$

Exercise 5.1

1. Aluminum is the most abundant metal but it is of the secondary importance next to iron, in terms of consumption. Why?
2. Cryolite, Na_3AlF_6 is added to Al_2O_3 to reduce its melting point from 2045°C to 1000°C . Explain why it is required to reduce the melting point of Al_2O_3?

3. Write the balanced chemical equation for the dissociation of Al_2O_3 in the molten state.
4. Identify the ions that move to the cathode and the anode during the electrolysis of molten aluminum oxide? Write the half-reactions at the anode, cathode and overall cell reactions
5. At which electrode will aluminum be produced during the process?
6. Why aluminum is preferred to copper for high voltage wires such as the line from Koka Dam to Addis Ababa?

Project

Perform the following project activities referring to reliable resources and submit to your teacher.

1. Make a list of materials made of aluminum you encounter at home, and in the city or town where you live.
2. What property of aluminum makes it useful in the construction of cook wares?
3. Regarding the manufacturing of aluminum by electrolysis, answer the following questions.

A. The electrolyte does not need heating to keep it in its molten state; why?

B. The graphite anode gradually disappears during the electrolysis. How? (Think about the product that is given off at the anode).

Present your findings to the class

B. Iron

Activity 5.4

1. Iron is the cheapest among all metals. Give reasons for this.
2. What is the difference between stainless steel and wrought iron?
3. What do you understand by the term alloy? Is stainless steel an alloy?

Discuss in groups and give a presentation to the class.

Occurrence and Extraction

At the end of this section, students should be able to

☞ explain properties ,Occurrence and extraction of Iron;

☞ describe the the applications of Iron.

Iron is the second most abundant metal in the earth's crust after aluminum, and the fourth most abundant element. It accounts for about 4.7% of the weight of the earth's crust. It has never been found as a free metal in nature. It exists only in the form of compounds in nature, such as oxides, carbonates and sulphides. The chief ores of iron are hematite (Fe_2O_3), limonite ($Fe_2O_3.H_2O$), magnetite (Fe_3O_4), and siderite ($FeCO_3$). It is also found in the form of iron pyrites (FeS_2) which is commonly called fool's gold. Iron is generally extracted from hematite (Fe_2O_3), magnetite (Fe_3O_4) and siderite ($FeCO_3$).

Producing pig iron (impure iron)

The extraction of iron from its ores is carried out in a blast furnace. The raw materials for the extraction of iron are iron ore, coke, limestone and hot air. The furnace is charged with a mixture of iron ore, limestone and coke at the top and a blast of hot air is blown at the bottom. The coke burns to form carbon dioxide and the temperature approaches 2000°C near the bottom. As the carbon dioxide rises up, the coke reduces it to carbon monoxide. The carbon monoxide then reduces the iron oxide to iron metal. The reactions taking place in the blast furnace can be summarized as follows:

i. Oxidation of coke to carbon dioxide: $C\ (s) + O_2\ (g) \rightarrow CO_2\ (g) + \text{Heat}$

ii. Reduction of carbon dioxide to carbon monoxide: $CO_2\ (g) + C\ (s) \rightarrow 2CO\ (g)$

iii. Reduction of iron oxides to metallic iron by carbon monoxide:

$$3Fe_2O_3\ (s) + CO\ (g) \rightarrow 2Fe_3O_4\ (s) + CO_2(g)$$
$$Fe_3O_4\ (s) + CO\ (g) \rightarrow 3FeO\ (s) + CO_2\ (g)$$
$$FeO\ (s) + CO\ (g) \rightarrow Fe\ (l) + CO_2(g)$$

iv. Decomposition of the limestone by heat in the blast furnace serves as a flux to remove impurities: $CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$

The Blast Furnace

Charge: Iron ore, coke, limestone

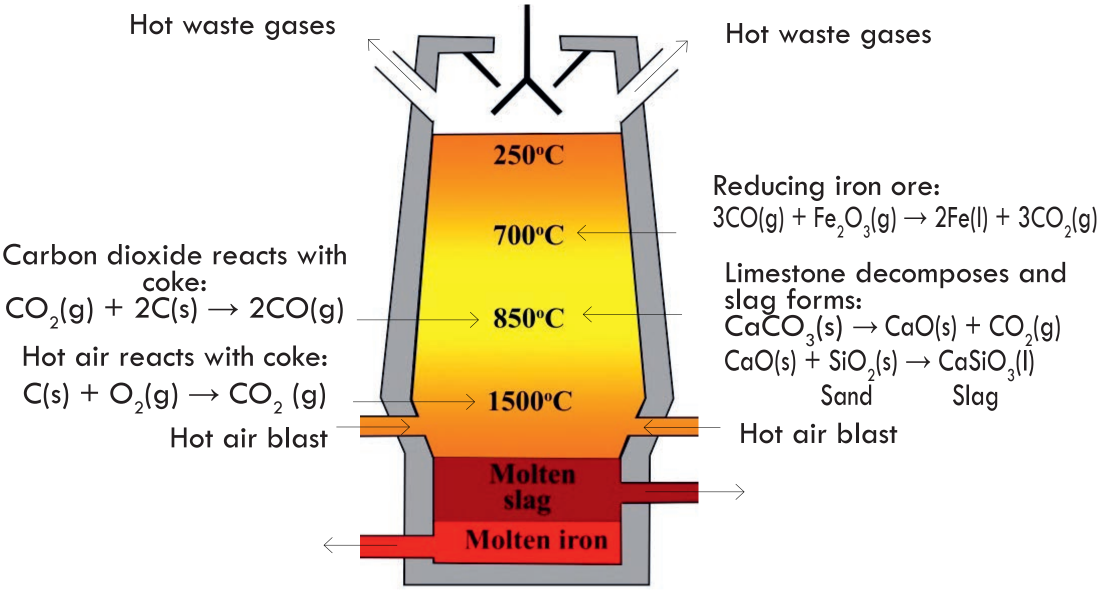


Figure 5.5 Production of pig iron from iron ore in the blast furnace.

v. Calcium oxide (CaO) removes the impurities like sand or silica (SiO_2) to form calcium silicate which is a glass-like material called slag.

CaO	+	SiO_2	$\rightarrow$	$CaSiO_3$
Lime	+	sand		calcium silicate
(Flux)		(Impurity)		(Slag)

The molten slag is lighter than the molten metal. Therefore, it floats over the molten metal and thus the two can be easily separated. The slag is used mostly for the manufacture of cement.

The iron obtained directly from the blast furnace is called pig iron. It is impure and contains about 2 % silicon, up to 1% phosphorus and manganese, and traces of sulphur. These impurities make pig iron brittle. Pig iron, also known as crude iron, is an intermediate product of the iron industry in the production of steel. When pig iron is re-melted with scrap iron and cast in to moulds, it forms cast iron. Pig iron contains at least 92% Fe and has very high carbon content, typically 3.5 - 4.5%. Along with silica and other constituents of dross, which makes it very brittle and not useful directly as a material except for limited applications.

Activity 5.5

1. Explain the principle of blast furnace
2. Why calcium carbonate is used in the blast furnace in the production of pig iron ?
3. What is the composition and uses of slag?

Discuss in your group and present your conclusions to the class.

The purest form of commercial iron is called wrought iron. It is obtained by removing most of the impurities from pig iron. It is manufactured by heating impure iron with hematite and limestone in a furnace. This increases the purity of the iron to 99.5%. Wrought iron is a tough, malleable and ductile form of iron.

Steel making from pig iron (purification of pig iron)

The relatively high carbon-content of the iron recovered from the blast furnace (pig iron) makes it very hard and brittle. The conversion of pig iron to steel is essentially a purification process in which impurities are eliminated through oxidation. This purification process is carried out in a furnace at high temperatures. Three distinctive techniques are used to convert pig iron into steel. These are the Bessemer converter, Open-hearth Furnace and the Basic Oxygen Process.

i. **The Bessemer Converter:** In the process of converting pig iron to steel, molten pig iron from the blast furnace is transferred to a cylindrical vessel containing a refractory lining of $MgCO_3$ and $CaCO_3$. A blast of hot air is blown through the molten metal from a set of small holes at the bottom of the vessel. The oxygen passing through the molten metal converts silicon, phosphorus and sulphur to their oxides, which then react with the lining to form a slag. The carbon in the pig iron is also oxidized to carbon monoxide, so its concentration is also reduced. The conversion of the pig iron to steel by this process requires only 15 minutes. However, the reaction is difficult to control and the quality of the steel produced can be quite variable.

ii. **The Open-hearth Furnace**: It has a large, shallow hearth, which is usually lined with a basic oxide refractory (MgO and CaO). The furnace is charged with a mixture of pig iron, Fe_2O_3, scrap iron and limestone. A blast of hot air and burning fuel is directed over the surface of the charge to maintain it in the molten state. Impurities in the pig iron are oxidized by the Fe_2O_3 and air. Carbon dioxide, and sulphur dioxide formed by the oxidation of the carbon and sulphur, respectively, in the pig iron bubble out of the mixture as waste gases.

The SiO_2 and other acidic oxides like P_4O_{10} combine with CaO (from limestone) and the refractory lining to form a slag.
The impurities in pig iron are oxidized in the following way:

$$C + O_2 \rightarrow CO_2$$
$$S + O_2 \rightarrow SO_2$$
$$12P + 10Fe_2O_3 \rightarrow 3P_4O_{10} + 20Fe$$
$$3Si + 2Fe_2O_3 \rightarrow 3SiO_2 + 4Fe$$

Calcium oxide, formed by the decomposition of limestone, reacts with oxides of silicon and phosphorus to form slag.

$$P_4O_{10} + 6CaO \rightarrow 2Ca_3(PO_4)_2$$
$$SiO_2 + CaO \rightarrow CaSiO_3$$

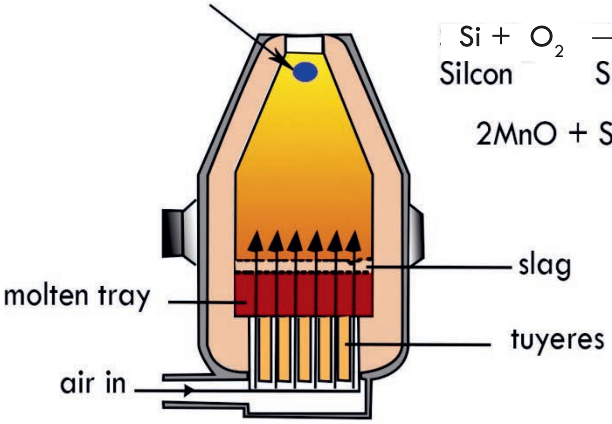
Impurities manganese and silicon form respective oxides and removed as slag.

$$\underset{\text{Carbon}}{2C} + O_2 \rightarrow \underset{\text{Carbon monoxide}}{2CO}$$

$$\underset{\text{Manganese}}{2Mn} + O_2 \rightarrow \underset{\text{Manganese dioxide}}{2MnO}$$

$$\underset{\text{Silcon}}{Si} + O_2 \rightarrow \underset{\text{Silcon dioxide}}{SiO_2}$$

$$2MnO + SiO_2 \longrightarrow \underset{\text{(slag)}}{MnSiO_3}$$



(a)

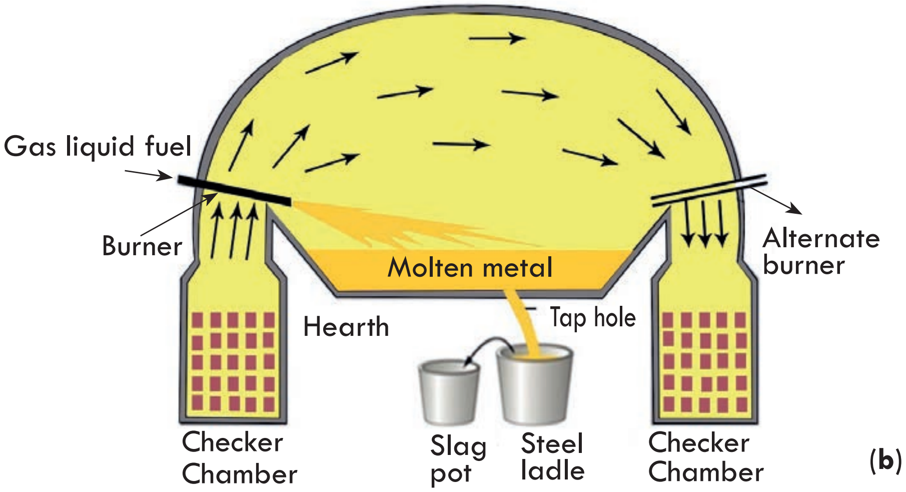


Figure 5.6 Bessemer converter (a) and Open-hearth furnace (b)

This entire process takes 8 to 10 hours to complete and the quality of the steel is much more easily controlled. A calculated amount of various elements can be added to the steel to form steel alloys with desired properties

iii. **Basic Oxygen Process:** This newer procedure has largely replaced the open hearth, because of its high speed. It involves forcing a mixture of powdered calcium oxide, CaO, and oxygen gas directly into the surface of the molten pig iron. The oxygen reacts exothermically with carbon, sulphur, silicon, phosphorus and impurity metals. Carbon and sulphur are oxidized to CO_2 and SO_2, respectively, and are released as exhaust gases. The oxides of silicon (SiO_2), phosphorus, (P_4O_{10}) and impurity metals combine with lime (CaO), forming slag. The charge in the converter is maintained in the molten state by the energy released in these oxidation reactions. The converter is tilted to remove the slag. After a second blow of oxygen, the converter is tilted once again and the molten steel is poured out. This process takes only about 20 to 25 minutes.

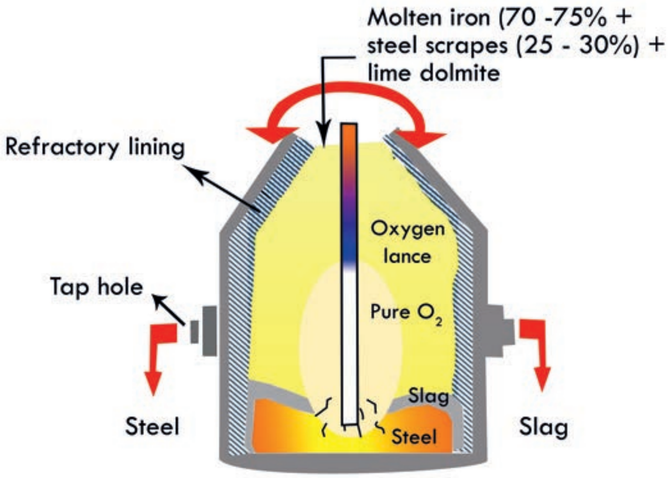


Figure 5.7 The basic oxygen converter

Tempering of steel: is a process by which steel is conditioned to a desired hardness by heating and controlled rate of cooling. Some of the carbon, in steel is present as a compound called cementite (FeC_3). It reduces the hardness of steel. Cementite can be removed by tempering which is difficult to remove by the above methods.

Activity 5.6

1. Which form of iron is the purest: wrought iron, pig iron, steel?
2. Prepare a list of materials which are made of steel and describe their uses. Discuss in groups and give a presentation to the class.

Physical Properties of Iron

Do you remember physical properties of iron? Can you mention some of them?
Iron is a gray lustrous, malleable and ductile metal. It is a good conductor of heat and electricity. It has a high melting point (1580°C) and a high density (7.87 g/cm^3). It is a ferromagnetic metal, which means it can be permanently magnetized.

Chemical Properties of Iron

Have you ever noticed the formation of a reddish thin film on the surface of an iron object? Do you know the chemical nature of this thin film? From your observations, what do you conclude about the reactivity of iron?

Iron is a reactive metal. Its reactivity is very slow, compared to group IA and IIA metals. Iron rusts in the presence of air and moisture to form hydrated iron (III) oxide.

$$4Fe(s) \quad + \quad 3O_2(g) \xrightarrow{\text{Moisture}} 2Fe_2O_3.xH_2O(s)(rust)$$

Iron reacts with dilute acids, like HCl and H_2SO_4, forming iron (II) salts and liberating hydrogen gas

$$Fe\ (s) + 2HCl\ (aq) \longrightarrow FeCl_2\ (aq) + H_2\ (g)$$
$$Fe\ (s) + H_2SO_4\ (aq) \longrightarrow FeSO_4\ (aq) + H_2\ (g)$$

Like other transition metals, iron also exhibits different oxidation states. It commonly exists as the ferrous (Fe^{2+}) and ferric (Fe^{+3}) ions. Iron (II) and iron (III) compounds are coloured.

Aqueous solutions of iron (II) compounds are pale-green and solutions of iron (III) compounds are yellowish brown.

Aqueous solutions of some iron (II) compounds are rapidly oxidized to the corresponding iron (III) compounds. For example, iron (II) hydroxide, $Fe(OH)_2$, on exposure to air, oxidizes to iron (III) hydroxide

$$4Fe(OH)_2\ (s) + O_2\ (g) + 2H_2O\ (l) \rightarrow 4Fe(OH)_3\ (s)$$

Heated iron reacts with hydrogen chloride gas, forming iron (II) chloride and hydrogen gas.

$$Fe\ (s) + 2HCl\ (g) \rightarrow FeCl_2\ (g) + H_2\ (g)$$

Heated iron reacts with chlorine and sulphur to form the chloride and sulphide, respectively

$$2Fe\ (s) + 3Cl_2\ (g) \rightarrow 2FeCl_3\ (s)$$
$$Fe\ (s) + S\ (s) \rightarrow FeS\ (s)$$

Iron has the ability to displace less active metals from solutions of their salts. For example, when a piece of iron is placed in a solution of $CuSO_4$, it is oxidized to Fe^{2+} and a reddish brown deposit of copper metal forms on the surface of the iron.

$$Fe\ (s) + Cu^{2+}\ (aq) \rightarrow Fe^{2+}\ (aq) + Cu\ (s)$$

Uses of Iron

Iron is the most widely used metal and takes first position in worldwide consumption. It is used as a structural metal in the construction of buildings and bridges. In the form of pig iron, it is used to make domestic boilers, hot-water radiators, railings, water pipes, castings, and moldings. As wrought iron, it is used in making nails, sheets, horseshoes, ornamental gates, door knockers, farm machinery etc. Iron is also used in the manufacture of alloys such as carbon steels and alloy steels.

Table 5.1 Types and uses of carbon steels.

Type	% of carbon	Use
Mild steel	Up to 0.2	Making screws, motors, car batteries, railway lines, ships, nuts and bolts
Medium steel	0.3 - 0.6	Making springs and chains
High carbon steel	0.6 - 1.5	Making drill bits, knives, hammers and chisels
Stainless steel	Fe = 73, C =1, Cr = 18, Ni = 8	For making knife blades, kitchen pots, pans, food machinery, automobile parts, laundry machines

Alloy Steel: contains iron and a relatively higher percentage of other metals. For example, stainless steel contains chromium (14 –18%) and nickel (7 – 9%). Tool steels include tungsten steel and manganese steel. These metals give the alloy hardness, toughness and heat resistance. They are used in rock drills, cutting edges and parts of machinery that are subjected to heavy wear.

C. Copper

Activity 5.7

Discuss in a group the subsequent questions and present your response to the class.

1. What common items are made of copper?
2. Copper is used in electrical appliances. Give two reasons based on its physical properties
3. A solution of copper sulphate cannot be stored in a vessel made of iron. Give a reason and justify it by a balanced ionic equation.
4. Why is copper so important in our technological society?

Occurrence and Extraction

At the end of this section, students should be able to

- explain properties, occurrence and extraction of Copper;
- describe the the applications of Copper.

Copper is occasionally found as native copper. However, it is found mainly in compounds such as sulphides, oxides and the carbonates. The most important sulphide ores are chalcopyrite ($CuFeS_2$), chalcocite (Cu_2S), covellite (CuS) and bornite (Cu_5FeS_4). The principal oxide ores are cuprite (Cu_2O) and tenorite (CuO). In the carbonate form, it exists as malachite ($CuCO_3.Cu(OH)_2$).

Copper is principally extracted from chalcopyrite. Since the amount of copper in the ore is very small, the crushed and ground sulphide ore is first concentrated by froth flotation. This treatment changes the concentration of the ore from 2 % copper to as high as 30% copper. The concentrated ore is then roasted with a limited supply of air (oxygen).

$$2CuFeS_2\ (s) + 4O_2\ (g) \rightarrow Cu_2S\ (s) + 2FeO(s) + 3SO_2\ (g)$$

The roasted mixture is smelted by adding limestone and sand to form a molten slag that removes many of the impurities. FeO present in the roasted mixture can also be removed as slag in the form of iron silicate ($FeSiO_3$), and silica in the form of calcium silicate ($CaSiO_3$).

$$CaCO_3\ (s) + SiO_2\ (s) \rightarrow CaSiO_3\ (l) + CO_2\ (g)$$
$$FeO\ (s) + SiO_2\ (s) \rightarrow FeSiO_3\ (l)$$

The Cu_2S obtained by roasting chalcopyrite is then reduced by heating it in a limited supply of oxygen.

$$Cu_2S\ (s) + O_2\ (g) \rightarrow 2Cu\ (l) + SO_2\ (g)$$

The copper produced by the above process is called blister copper and it has 98.5 – 99.5 % purity. Blister copper contains iron, silver, gold and sometimes zinc as impurities. So it is refined further by electrolysis as shown in *Figure 5.8*.

Electrolytic refining of copper

The main parts of the electrolytic cell are:

Anode: A thick block of impure copper
Cathode: A thin strip of pure copper
Electrolyte: An aqueous solution of copper sulphate, Small quantity of dilute sulphuric acid is also added to the salt solution to prevent hydrolysis.

When electric current is passed, the copper of the anode is oxidized to Cu^{2+} ions. The Cu^{2+} ions pass in to solution.

$$\text{Cu (impure metal)} \rightarrow Cu^{2+}(aq) + 2e \qquad \text{oxidation at anode}$$

At the same time an equal number of Cu^{2+} ions of the solution move towards cathode where they are reduced and deposited as metallic copper.

$$Cu^{2+}(aq) + 2e \rightarrow \text{Cu (metal)} \qquad \text{reduction at cathode}$$

The process of oxidation at anode and reduction at cathode continues as long as current is passed. Thus, the cathode becomes thicker and the anode becomes thinner. Pure copper deposited at cathode is removed and used. Copper obtained by electrolytic refining is 99.99% pure.

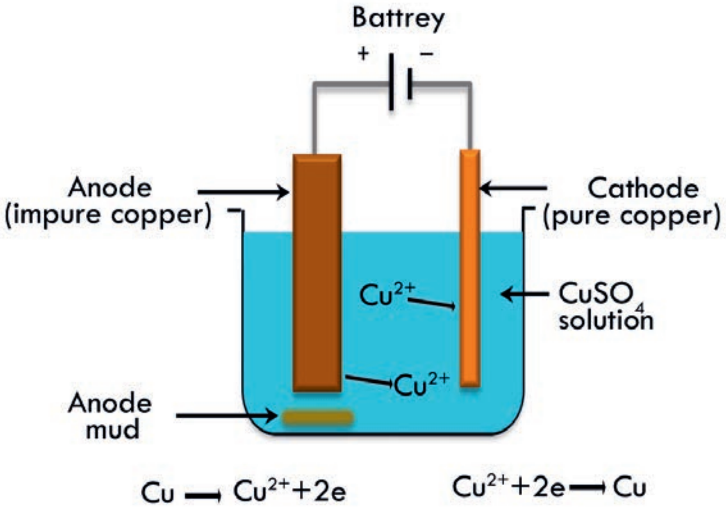


Figure 5.8 Electrolytic refining /purification/ of copper.

Physical Properties of Copper

Copper is a soft, ductile, malleable, reddish-brown metal with a density of 8.96 g/cm³. It is second to silver in electrical conductivity. It melts at 1086°C and boils at 2310°C.

Chemical Properties of Copper

Copper is less reactive metal that is why it is found in the native state. Although it is not very reactive, it undergoes the following reactions:

Powdered copper, when heated in air forms a black powder of copper (II) oxide, CuO.

$$2Cu\ (s) + O_2\ (g) \rightarrow 2CuO\ (s)$$

Copper does not react with dilute acids like HCl and H_2SO_4. It cannot displace hydrogen from acids. However, it can be oxidized by oxidizing acids such as dilute and concentrated nitric acid and hot concentrated sulphuric acid, H_2SO_4.

$$3Cu(s) + 8HNO_3\ (aq) \xrightarrow{\text{dilute}} 3Cu(NO)_3)_2\ (aq) + 2NO\ (g) + 4H_2O\ (l)$$

$$Cu(s) + 8HNO_3\ (aq) \xrightarrow{\text{concentrated}} Cu(NO)_3)_2\ (aq) + 2NO_2\ (g) + 2H_2O\ (l)$$

$$Cu(s) + 2H_2SO_4\ (aq) \xrightarrow{\text{Hot and concentrated}} CuSO_4(aq) + SO_2\ (g) + 2H_2O\ (l)$$

Copper corrodes in moist air over a long period of time as a result of oxidation caused by a mixture of water, oxygen and carbon dioxide. It turns green, due to the formation of verdigris: a basic copper carbonate ($CuCO_3.Cu(OH)_2$) or $Cu_2(OH)_2CO_3$.

$$2Cu\ (s) + H_2O\ (l) + O_2\ (g) + CO_2\ (g) \rightarrow CuCO_3.Cu(OH)_2$$

Basic copper carbonate is responsible for the green layer that forms on the surface of copper and bronze objects. This layer adheres to the surface and protects the metal from further corrosion. Copper is a transition metal and exhibits different oxidation states. It exists as cuprous (Cu^+) and cupric (Cu^{2+}) ions. Compounds containing copper in the +2 oxidation state are stable. Cu^{1+} is unstable and cannot exist in aqueous solution since it reduces and oxidizes (disproportionates) itself as follows:

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

Soluble compounds containing Cu^+, such as CuCl, CuI and CuBr, are all colourless solids. The most stable oxidation state of copper is +2. Aqueous solutions of many copper (II) compounds are blue, thus copper forms coloured compounds in the oxidation state of +2. For example, hydrated copper sulphate, $CuSO_4.5H_2O$, consists of blue crystals, and its aqueous solution is also blue.

Uses of Copper

Copper is used to manufacture alloys. For example, bronze is an alloy of copper and tin, and it is used to make coins, medals, bells, machinery parts, etc. Brass is an alloy of copper and zinc, and it is used for hardware tops, terminals and pipes. Copper is widely used in the electrical industry in the form of electric wires, cables etc. Copper compounds, such as copper chloride, copper carbonate and copper hydroxide, are used as pesticides

Exercise 5.2

1. Classify the following as naturally available and made by humans: water, sunlight, land, minerals, statues, concrete, plastics, buildings, fossil fuels (crude oil, coal, and natural gas)
2. Categorize the following as renewable and nonrenewable resources: rocks, metal ores, soil, wood, air, water, crops, fossil fuels, biofuels
3. Describe the occurrence and extraction of the following metals; aluminium, iron, and copper.
4. Which of the three metals, Al, Fe and Cu is produced by electrochemical reduction?
5. Explain how Al, Fe and Cu behave in dry air, moisture and acidic solutions.
6. Why is limestone used in steel?
7. Discuss the main areas of applications of Al, Fe and Cu in home and industry.
8. The kitchen utensils used for cooking foods are made of metals like aluminum, iron, copper or the alloys of some metals. Give a reason.
9. Why are wires employed in electrical appliances made of copper and (or) aluminum?

5.3 Production of Some Important Nonmetals

5.3.1 General Properties of Nonmetals and Common Uses of Some Nonmetallic Compounds

At the end of this section, students should be able to

- mention the general properties of non-metals and their uses;
- describe some common uses of compounds of nonmetals such as CO_2, Na_2CO_3, NH_3, HNO_3, H_3PO_4, $Ca_3(PO_4)_2$, SO_2 & H_2SO_4 ;
- describe the occurrence, extraction and uses of nitrogen, phosphorous, oxygen, sulphur and chlorine.

Nonmetals have opposite characteristics to that of metals.

A. Physical properties

Nonmetals are

- Solids, liquids, gases
- Non-lustrous

- Nonmalleable and non-ductile
- Varying hardness and have low density
- Low melting and boiling points
- do not show any metallic luster.
- are soft in comparison to metals
- are non-sonorous
- Poor conductors of heat and electricity

B. Chemical properties of non-metals

- Nonmetals react with oxygen on heating or burning to form their oxides
- Do not displace hydrogen on reaction with dilute acids
- React with oxygen to form acidic or neutral oxides
- Combine with hydrogen to form stable hydrides
- Do not react with water
- Electronegative i.e for negative ions by gaining electrons
- Oxidizing agents

5.3.2 Production of Nitrogen, Phosphorous, Oxygen, Sulphur and Chlorine

A. Nitrogen

At the end of this section, students should be able to

- explain properties ,Occurrence and extraction of Nitrogen;
- describe the the applications of Nitrogen.

Activity 5.8

Discuss in group and give a presentation to the class.

1. Draw the structure of a nitrogen molecule. How many bonds are present in it?
2. Why does nitrogen form triple bond?
3. Which form of nitrogen is absorbed by green plants?

Occurrence and Production

Nitrogen occurs in nature in the elemental form. It also exists in the form of compounds. In the elemental state, it exists as a diatomic molecule, N_2, in atmospheric air. It constitutes about 80% by volume of the atmosphere. In the form of compounds, it exists as sodium nitrate (Chile salt peter, $NaNO_3$) and potassium nitrate (KNO_3) also called saltpetre. It is also found in DNA molecules and proteins of all living things.

In the industrial production of nitrogen, the first step is to remove impurities like dust and other particles from air. The air is then compressed under high pressure and low temperature to remove carbon dioxide and water vapor since they solidify at a relatively high temperature compared to nitrogen and oxygen. As the temperature continues to decrease, the air containing mainly nitrogen and oxygen condenses to give a pale-blue liquid. Nitrogen is thus separated on an industrial scale by fractional distillation of liquid air. When liquid air is fractionally distilled, nitrogen is collected and stored in steel cylinders under pressure.

The second fraction consists of argon, which distills off the mixture at –186°C, leaving behind a blue liquid of oxygen that boils at –183°C. In the laboratory, nitrogen is prepared by warming an aqueous solution containing ammonium chloride and sodium nitrite

$$NH_4Cl\ (aq) + NaNO_2\ (aq) \rightarrow NaCl\ (aq) + N_2\ (g) + 2H_2O\ (l)$$

Physical Properties of Nitrogen

Nitrogen is colorless, odorless and tasteless gas. It is inert under ordinary conditions, and that is why it is found in atmospheric air as N_2. The inertness of nitrogen at low temperatures is directly related to the strength of the triple bond, which requires high energy to break.

Chemical Properties of Nitrogen

Activity 5.9

Discuss in group and give a presentation to the class.

1. Why nitrogen is the most abundant gas in the atmosphere?
2. Why nitrogen exists mostly in the elemental state?
3. What is the purpose of running a reaction under an inert atmosphere such as nitrogen gas?

Though nitrogen is inert at low temperatures, it reacts with metals of group IA and IIA as well as oxygen at higher temperatures. When heated with reactive metals like lithium, calcium and magnesium, it forms compounds, known as nitrides.

$$6Li\ (s) + N_2\ (g) \rightarrow 2Li_3N\ (s)$$
$$3Ca\ (s) + N_2\ (g) \rightarrow Ca_3N_2\ (s)$$
$$3Mg\ (s) + N_2\ (g) \rightarrow Mg_3N_2\ (s)$$

Nitrogen combines with oxygen at elevated temperatures or in an electric arc to form oxides.

$$N_2\ (g) + O_2\ (g) \rightarrow 2NO\ (g)$$
$$N_2\ (g) + 2O_2\ (g) \rightarrow 2NO_2\ (g)$$

Nitric oxide, also called nitrogen monoxide, NO, is a colorless and reactive gas. It forms nitrogen dioxide (NO_2), a reddish brown gas, when it comes into contact with oxygen. This reddish brown gas dimerizes at low temperatures to give a colorless gas of dinitrogen tetraoxide, N_2O_4.

$$2NO_2\ (g) \rightarrow N_2O4\ (g)$$

Nitrogen also forms oxides, like dinitrogen monoxide, N_2O, (commonly called laughing gas), dinitrogen trioxide (N_2O_3) and dinitrogen pentoxide (N_2O_5). Nitrogen reacts directly with hydrogen in the Haber process to form ammonia. In this process, a mixture of N_2 and H_2 gas, at a pressure of 200 - 300 atm and a temperature of 400 - 600°C, is passed over a catalyst of finely divided iron.

$$N_2(g) + 3H_2(g) \underset{400\text{-}600^0C}{\overset{Fe/200\ -\ 300atm}{\rightleftharpoons}} 2NH_3(g)$$

The ammonia produced by this method is primarily used in the manufacture of fertilizers. It is also used for the production of nitric acid (HNO_3).

Uses of Nitrogen

Nitrogen is used in food packaging to prevent oxidation, and to create an inert atmosphere in the production of semiconductors. Liquid nitrogen is used as a refrigerant to preserve bulls' semen and blood. Its major use is in the production of ammonia.

B. Phosphorus

Activity 5.10

Discuss in group and give a presentation to the class. Have you ever observed glowing worms? Why do they glow? Anyone who has observed them should explain to the class.

Occurrence and Extraction

At the end of this section, students should be able to

- explain properties ,Occurrence and extraction of Phosphorous;
- describe the the applications of Phosphorous.

Does phosphorus exist in the pure form (elemental state) as does nitrogen? If not, why? Phosphorus is a relatively abundant element, ranking 12th in the earth's crust. It exists naturally only in the combined state, such as in rock phosphate, $Ca_3(PO_4)_2$, fluoroapatite, $Ca_{10}(PO_4)_6F_2$ or $3Ca_3(PO_4)_2.CaF_2$. It is also found in teeth, bones and DNA.

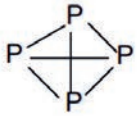
What is allotropy? What are the two common allotropes of phosphorus?

Allotropes are different forms of the same element with the same physical state but with different physical properties. This is due to the difference in the arrangement of atoms. The two common allotropic forms of phosphorus are white phosphorus and red phosphorus.

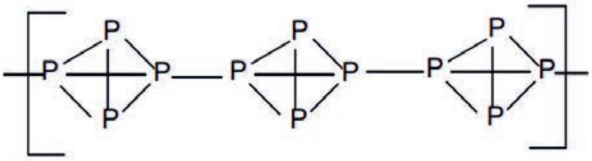
Physical Properties of Phosphorus

White phosphorus is a very poisonous, white waxy-looking substance that melts at 44.1°C and boils at 287°C. Its density is 1.8 g/cm^3. It consists of individual tetra-atomic (P_4) molecules and is an unstable form of phosphorus

Red phosphorus is denser (2.16 g/cm^3) and is much less reactive than white phosphorus at normal temperatures. It consists of P_4 molecules linked together to form a polymer (large molecule) as shown below.



The white phosphorus molecule



The arrangement of atoms in red phosphorus

Industrially, white phosphorus is manufactured by heating a mixture of crushed rock phosphate, $Ca_3(PO_4)_2$, silica, SiO_2, and coke in an electric furnace. The reaction in the furnace is:

$$2Ca_3(PO_4)_2\,(s) + 6SiO_2\,(s) + 10C\,(s) \rightarrow 6CaSiO_3(l) + P_4(g) + 10CO\,(g)$$

The vaporized phosphorus (P_4) is condensed, collected and stored under water. Red phosphorus is prepared by heating white phosphorus in sunlight for several days. White phosphorus is stored under water because it spontaneously ignites (bursts into flames) in the presence of oxygen. Red phosphorus is much less reactive than white phosphorus and does not need to be stored under water. Although red phosphorous is insoluble in water, it is very soluble in carbon disulphide. Solutions of P_4 in CS_2 are reasonably stable. As soon as CS_2 evaporates, however, the phosphorous bursts in to flame.

Activity 5.11

Discuss in group and give a presentation to the class.

Use reference books available in your school library and write down the main properties of white and red phosphorus.

Chemical Properties of Phosphorus

Phosphorus exhibits the following common chemical reactions. Phosphorus reacts with limited and excess supplies of oxygen to form tetraphosphorus hexoxide (P_4O_6) and tetraphosphorus decoxide (P_4O_{10}) respectively

$$P_4\,(s) + 3O_2\,(g) \rightarrow P_4O_6(s)$$
$$P_4\,(s) + 5O_2\,(g) \rightarrow P_4O_{10}(s)$$

P_4O_6 and P_4O_{10} dissolve in water to form phosphorous acid, H_3PO_3, an orthophsophoric acid, H_3PO_4 respectively.

$$P_4O_6\ (s) + 6H_2O\ (l) \rightarrow 4H_3PO_3\ (aq)$$
$$P_4O_{10}\ (s) + 6H_2O\ (l) \rightarrow 4H_3PO_4\ (aq)$$

Phosphorus also reacts with limited and excess supplies of chlorine to form phosphorus (III) chloride (PCl_3) and phosphorus (V) chloride (PCl_5) respectively

$$P_4\ (s) + 6Cl_2\ (g) \rightarrow 4PCl_3\ (s)$$
$$P_4\ (s) + 10Cl_2\ (g) \rightarrow 4PCl_5\ (s)$$

Uses of Phosphorus

Red phosphorous is used to make matches. Most of the white phosphorous produced is used to make phosphoric acid or other phosphorous compounds. Phosphorous is also used in making fireworks, smoke bombs, rat-poisons, and tracer bullets. Phosphorous is essential for plant growth. Large quantities of phosphorous compounds are converted into acids and salts to be used as fertilizers, in baking powder in the chemical industries.

C. Oxygen

Activity 5.12

Discuss the following questions in groups and give a presentation to the class.

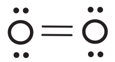
1. Write the chemical reactions for photosynthesis and respiration. What are the products formed in photosynthesis and respiration?
2. What are the uses of oxygen in hospitals?
3. How do astronauts breathe while in space?

Occurrence and Production

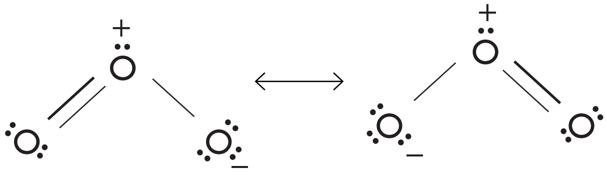
At the end of this section, students should be able to

☞ explain properties ,Occurrence and extraction of Oxygen;
☞ describe the the applications of Oxygen.

Oxygen is the most abundant element on earth. It constitutes about 46.6% by weight of the earth's crust in the form of compounds such as oxides, silicates, carbonates, phosphates etc. It exists in nature in the elemental state in atmospheric air constituting about 20% by volume. Oxygen is manufactured industrially by the fractional distillation of liquid air in a way similar to that of nitrogen. Oxygen has two allotropic forms, diatomic (O_2) and triatomic (O_3), ozone.



Structure of O_2 molecule



Resonance structures of O_3 molecule

Ozone is naturally produced in the stratosphere by a two- step reactive process. In the first step, solar ultraviolet radiation (sunlight) breaks apart an oxygen molecule to form two separate oxygen atoms. In the overall process, three oxygen molecules plus sunlight react to form two ozone molecules according to the equation:

$$3O_2 \xrightarrow{\text{UV - light}} 2O_3$$

Occurrence of this process in the upper atmosphere removes radiation that would be harmful to life on the surface of the Earth.

Physical Properties

Oxygen is a colorless, odorless, tasteless gas. It changes from a gas to a liquid at a temperature of -182.96°C. The liquid formed has a slightly bluish color to it. Liquid oxygen can then be solidified or frozen at a temperature of -218.4°C. The density of oxygen is 1.429 grams per liter. By comparison, It is denser than air. The density of air is about 1.29 grams per liter

Chemical Properties of Oxygen

Oxygen is relatively reactive and combines directly with most elements to form oxides. Oxygen combines with metal to form metal oxides (*basic oxides*). For example,

$$2Mg\ (s) + O_2\ (g) \rightarrow MgO\ (s)$$
$$2Ca\ (s) + O_2\ (s) \rightarrow 2CaO\ (s)$$

Oxygen combines with non-metals to form acidic oxides; for example:

$$S_8\ (s) + 8O_2\ (g) \rightarrow 8SO_2\ (g)$$
$$P_4\ (s) + 5O_2\ (g) \rightarrow P_4O_{10}\ (g)$$

Oxygen supports combustion and is necessary for the burning of substances like charcoal, hydrocarbon fuel, etc.

$$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$$
$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (g)$$

D. Sulphur

Activity 5.13

Discuss the following in groups and give a presentation to the class.

Have you observed the color of sulphur? What is the use of sulphur in your area?

Occurrence and Extraction

At the end of this section, students should be able to

- explain properties, occurrence and extraction of Sulphur;
- describe the the applications of Sulphur.

Sulphur is found in nature and in the form of compounds such as galena (PbS), pyrites (FeS_2) (fool's gold), cinnabar (HgS), sphalerite (ZnS), gypsum ($CaSO_4.2H_2O$), barite ($BaSO_4$), and as hydrogen sulphide (H_2S) in natural gas and crude oil.

Sulphur is extracted from underground deposits of elemental sulphur by the Frasch process (*Figure 5.9*). In this process, three concentric pipes are sent down to the sulphur deposit.

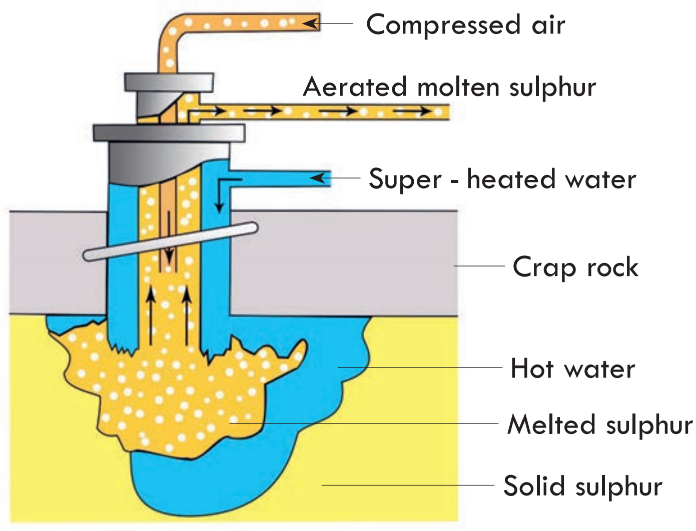


Figure 5.9 Frasch process- extraction of sulphur.

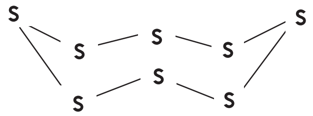
Superheated water at about 170°C is pumped through the outermost pipes. Hot air is compressed in the innermost tube. A froth of sulphur, air and water come out to the surface of the earth forced by hot compressed air in the middle tube. The molten sulphur is then cooled and solidified.

Uses of Sulphur

Nearly one-half of the sulphur needed by chemical industries is obtained as a waste product of other industrial processes. For example, sulphur can be obtained in the form of hydrogen sulphide (H_2S) from natural gas and during the purification of crude oil. It can also be obtained in the form of sulphur dioxide from the roasting of sulphide ores during the manufacture of metals. The use of sulphur obtained from waste products of other industries reduces the demand for the element from natural deposits. This also reduces the pollution of atmospheric air and the formation of acid rain.

Allotropic Forms of Sulphur

Sulphur exhibits allotropy. The most important allotropes of sulphur are rhombic and monoclinic sulphur. Rhombic sulphur is the most stable form of sulphur consisting of S_8 molecules.



Crown shape of S_8 molecules

Rhombic Sulphur (α-sulphur): rhombic sulphur is crystalline in nature and has octahedral shape. On heating the solution of roll sulphur in CS_2 we get rhombic sulphur. It is yellow with a melting point of 385.8K and specific gravity 2.06. Rhombic sulphur cannot be dissolved in water but can be dissolved in benzene, ether, alcohol etc. Rhombic sulphur is the most stable form of sulphur consisting of S_8 molecules.

Monoclinic Sulphur (β-sulphur): β-Sulphur is a yellow solid with a monoclinic crystal form and is less dense than α-sulphur. Like the α- form it contains puckered S_8 rings and only differs from it in the way the rings are packed in the crystal. It is unusual because it is only stable above 95.3 °C; below this temperature it converts to α-sulphur.

Physical Properties of Sulphur

Pure sulphur is a tasteless, odourless, brittle solid that is pale yellow in colour, a poor conductor of electricity, and insoluble in water

Chemical Properties of Sulphur

Sulphur is relatively stable and unreactive at room temperature. However, it reacts with metals and non-metals when heated.

Sulphur combines with metals when heated to form sulphides.

$$8Fe\ (s) + S_8\ (s) \rightarrow 8FeS\ (s)$$

Sulphur burns in oxygen to form oxides.

$$S_8\ (s) + 8O_2\ (g) \rightarrow 8SO_2\ (g)$$

Sulphur is the raw material for the production of sulphuric acid (H_2SO_4) by the Contact Process. This process of sulphuric acid production involves the following steps

Step 1: Sulphur is first oxidized to produce sulphur dioxide:

$$S_8\,(s) + 8O_2\,(g) \rightarrow 8SO_2\,(g)$$

Step 2: Sulphur dioxide is converted to sulphur trioxide at high temperatures in the presence of a catalyst:

$$2SO_2\,(g) + O_2 \rightarrow 2SO_3\,(g)$$

Step 3: Sulphur trioxide is absorbed into concentrated sulphuric acid to produce oleum ($H_2S_2O_7$).

$$SO_3\,(g) + H_2SO_4\,(l) \rightarrow H_2S_2O_7\,(l)$$

Step 4: The oleum is then diluted with water to produce the desired concentration of sulphuric acid. Commercially used concentrated sulphuric acid is 98% H_2SO_4.

$$H_2S_2O_7\,(l) + H_2O\,(l) \rightarrow 2H_2SO_4\,(aq)$$

E. Chlorine

Activity 5.14

Discuss the following in groups and give a presentation to the class.

1. When we turn on the water tap, the water coming out of the tap sometimes appears white and turns normal after a few minutes. Many people say that the white colour we see is due to chlorine. Do you agree? Discuss in group and present your opinion to the class.
2. What is the purpose of adding chlorine to drinking water?

Occurrence and Extraction

At the end of this section, students should be able to

- explain properties, occurrence and extraction of Chlorine;
- describe the the applications of Chlorine.

Chlorine belongs to group VIIA, known as the halogens. It is the most abundant element among the halogens. It is found in nature in the form of compounds only. Chlorine is found in nature chiefly in the form of chlorides of sodium, potassium, calcium and magnesium. Sodium chloride is the chief source of chlorine, and it is obtained from seawater or as deposits of rock salt. Commercially, chlorine is manufactured by the electrolysis of a concentrated aqueous solution of sodium chloride. During electrolysis of the concentrated solution, chlorine is formed at the anode and the reduction of water occurs at the cathode, producing hydrogen gas and hydroxide ions (OH^-). The electrode and the overall cell reactions are as follows;

Oxidation at anode: $2Cl^- (aq) \rightarrow Cl_2 (g) + 2e^-$

Reduction at cathode: $2H_2O (l) + 2e^- \rightarrow H_2 (g) + 2OH^- (aq)$

Cell reaction: $2NaCl (aq) + 2H_2O (l) \rightarrow 2NaOH (aq) + Cl_2 (g) + H_2 (g)$

Although NaOH forms at the cathode compartment and Cl_2 at the anode, the two substances must be kept apart from one another. This is because, if they come into contact with each other, they can react according to the following equation.

$$NaOH (aq) + Cl_2 (g) \rightarrow NaOCl (aq) + HCl (aq)$$

What is the oxidation number of chlorine in NaOCl?

Physical Properties of Chlorine

Activity 5.15

Discuss the following in groups and give a presentation to the class.

In urban area, tap water possesses a peculiar smell, which disappears after a few minutes. Many people say that the smell is due to chlorine. Do you agree?

Some of its physical properties are the following:

- Chlorine is a greenish-yellow gas at room temperature.
- It melts at –102°C and boils at –34°C.
- It is fairly soluble in water.
- It is extremely poisonous, causing inflammation of the lungs and mucous membranes if inhaled, even in very small amounts. It is fatal if inhaled in higher amounts.

Chemical Properties of Chlorine

Chlorine is a highly reactive non-metal. It reacts directly with almost all elements except the noble gases, carbon and nitrogen.

It is a powerful oxidizing agent and oxidizes most of the elements with which it reacts.

The following are the main reactions of chlorine.

Chlorine reacts with heated metals to form chloride salts

$$2Fe (s) + 3Cl_2 (g) \rightarrow 2FeCl_3 (s)$$
$$2Al (s) + 3Cl_2 (g) \rightarrow 2AlCl_3 (s)$$

Chlorine reacts smoothly with hydrogen, in the presence of a charcoal as a catalyst, at room temperature to form hydrogen chloride.

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

Chlorine displaces less reactive halogens (Br_2 and I_2) from aqueous solutions of their compounds.

$$Cl_2\ (g) + 2KBr\ (aq) \rightarrow 2KCl\ (aq) + Br_2\ (l)$$
$$Cl_2\ (g) + 2KI\ (aq) \rightarrow 2KCl\ (aq) + I_2\ (aq)$$

Can bromine displace chlorine in a chemical reaction?
Chlorine dissolves in water and reacts with it, forming an acidic solution containing hydrochloric acid (HCl) and hypochlorous acid (HOCl).

$$Cl_2\ (g) + H_2O\ (l) \rightarrow HCl\ (aq) + HOCl\ (aq)$$

This type of reaction is called a disproportionation reaction, where chlorine is itself reduced and oxidized.

Uses of Chlorine

Chlorine is commercially used as a bleaching agent and disinfectant (kills germs). As elemental chlorine or as chlorine water, it is strong enough to oxidize dyes and to bleach coloured materials. For example, it can be used to remove yellow or brown colors from wood pulp, paper and cotton.

Exercise 5.3

1. What are the main natural sources of N, P, O, S, and Cl?
2. Describe the main processes involved in the production (extraction) of N, P, O, S and Cl.
3. How do N, P, and S behave when heated in air?
4. Write an equation for the reaction of chlorine with water.
5. Describe and discuss the common uses of CO_2, Na_2CO_3, NH_3, HNO_3, H_3PO_4, $Ca_3(PO_4)$, SO_2, H_2SO_4

A Checklist for Review

- Allotrope
- Alloy
- Basic – oxygen process
- Bauxite
- Bessmer converter
- Blast furnace
- Blister coppe
- Cryolite
- Chlorine
- Cacination
- Chemical reduction
- Contact process
- Electrolytic reduction
- Fractional distillation
- Flotation
- Fluoroapatite
- Frasch process
- Haber process
- Hall process
- Mineral
- Metallurgy
- Nitrogen
- Ore
- Open – hearth fumace
- Red phosphorus
- Roasting
- Rock phosphate
- Sulphur
- Slag
- Steel
- Thermite process
- White phosphorus

Unit Summary

- About 80% of the elements are metals. They have common physical and chemical properties. Metals are ductile, malleable, sonorous, conductors of heat and electricity. Metals are reducing agents and mostly form basic oxides. The specific properties of metals vary widely from one metal to the other. Active metals always occur in compounds in the form of minerals. The most active metals are extracted from their compounds only by electrolysis because there is no stronger reducing agent that can reduce them from their compounds. The reactivity series is a useful guide to predict the relative reactivity of metals.

- Aluminum is the most abundant metal in the earth's crust. It is manufactured by the electrolysis of molten aluminum oxide mixed with cryolite using the Hall's process. Aluminum is used to make light-weight alloys that can be used for aircraft and similar constructions. Aluminum resists corrosion due to the formation of a thin film of an oxide on its surface.

- Iron is the second-most abundant metal in the earth's crust. Iron is extracted in a blast furnace. The furnace is charged with iron ore, coke and limestone, and hot air is blown up from the bottom. Limestone is used in the blast furnace to remove impurities like sand (silica) in the form of slag. The type of iron obtained directly from the blast furnace is called pig iron and it is impure. Wrought iron is the purest form of iron obtained from pig-iron by removing most of the impurities. Steel is an alloy of iron containing definite amounts of carbon and other metals like chromium, nickel, tungsten, vanadium, manganese etc. Steel is manufactured from the pig iron by the Bessemer converter, Open-Hearth Furnace and Basic Oxygen Processes.

- Copper is extracted by roasting chalcopyrites ($CuFeS_2$). Copper that is obtained by roasting chalcopyrites is called blister copper. It is impure, and it is refined further by electrolysis. Copper is malleable and ductile, a good conductor of heat and electricity and is used to make electric wires and in the manufacture of alloys.

- The properties of nonmetals are opposite to that of metals. Some examples of nonmetals are nitrogen, oxygen, phosphorous and chlorine.

- Nitrogen constitutes about 80% by volume of atmospheric air. Nitrogen is manufactured by the fractional distillation of liquefied air. Nitrogen is relatively inert. However, its reactivity increases when heated and combined with metals of group IA and IIA to form nitrides and with oxygen to form oxides like NO and NO_2.

- Phosphorus is manufactured by heating a mixture of rock phosphate, coke (carbon) and silica in an electric furnace. Phosphorus has two common allotropes, white and red phosphorus. White phosphorus is highly reactive, while red phosphorous is relatively stable. White phosphorus is stored under water. Phosphorus reacts with oxygen and chlorine.

- Oxygen makes up about 20% by volume of air. It is manufactured by the fractional distillation of liquid air. Oxygen is the most abundant element in the earth's crust.

- Sulphur exists in the elemental state in underground deposits from which it is extracted by the Frasch process. Sulphur reacts with oxygen to form sulphur dioxide and sulphur trioxide. Sulphur is used for the manufacture of sulphuric acid by the Contact Process.

- Chlorine is manufactured by the electrolysis of brine. Chlorine is a powerful oxidizing agent and reacts with heated metals to form chlorides. Chlorine and chlorine water are used to bleach colored materials and as disinfectant.

Review Exercises

Part I: Multiple choice questions. Choose the best answer from the given alternatives

1. Which one of the following metals is the most ductile?
   a. Aluminum b. Copper c. Gold d. Silver
2. Which one of the following is the least dense?
   a. Mercury b. Gold c. Sodium d. Aluminum
3. Which one of the following metals will not liberate hydrogen gas when treated with dilute hydrochloric acid ?
   a. Magnesium b. Zinc c. Copper d. Iron
4. Which one of the following metallic elements is most likely to be found as the free metal in nature ?
   a. Ca b. Au c. Al d. Fe
5. What is the purpose of a converter in steel production?
   a. to reduce the iron in the ore to elemental iron
   b. to allow the addition of nitrogen for increased strength
   c. to allow slow solidification of the molten metal so it will purify as it crystallizes
   d. to remove impurity elements by oxidation
6. Which of the following is a non-renewable resource:
   a. soil c. petroleum
   b. water d. plants
7. Which of the following elements possess allotropes?
   a. carbon b. phosphorus c. sulphur d. all
8. Which one is used as anode in the electrolytic refining of copper?
   a. Pure copper c. Impure copper
   b. Graphite rod d. Pure iron rod

9. Which of the following is not a use of nitrogen:
   a. making matches
   b. production of ammonia
   c. making fertilizers
   d. as a refrigerant
10. The most abundant element in the earth's crust is:
   a. Silicon
   b. Iron
   c. Oxygen
   d. Aluminium
11. Which of the following elements is not found as a free element in nature:
   a. Sulphur
   b. Oxygen
   c. Copper
   d. Chlorine
12. Which of the statements about the extraction of iron in the blast furnace is true?
   a. limestone is added to combine with excess carbon dioxide
   b. a slag of iron oxide forms at the bottom of the furnace
   c. hot air is blown in at the top of the furnace
   d. carbon monoxide reduces iron (III) oxide to iron
13. Calcium oxide is added to a steel making furnace to:
   a. oxidize basic oxides
   b. oxidize carbon to carbon dioxide
   c. react with acid impurities
   d. react with silicon to form silicon dioxide

Part II: Short answer questions.

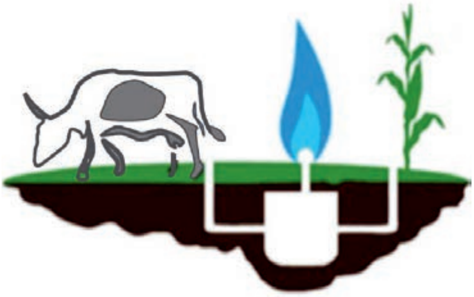
Answer the following appropriately

14. You are given, a hammer, a battery, wires, bulb and a switch.
   a. How will you use them to distinguish between samples of metals and non-metals
   b. Assess the usefulness of these tests to distinguish between metals and nonmetals
15. In the electrolytic refining of a metal M what would you take as cathode, as the anode and as the electrolyte ?
16. State two ways to prevent the rusting of iron.
17. Give reason for the following:
   a. Aluminum is a highly reactive metal, yet it is used for making cooking utensils
   b. You must have seen tarnished copper vessels being cleaned with lemon or tarmarind juice
18. Explain why I_2 is solid, Br_2 is liquid, Cl_2 and F_2 are both gases at room temperature?
19. Which one of Cu, Ag and Au is most malleable?
20. Which one of Fe, Ag and Pb is the best conductor of heat?
21. Name a metal and a nonmetal existing in liquid state at room temperature
22. Why is limestone added to the blast furnace?
23. Compare and contrast the Bessemer and open hearth processes.
24. Describe the advantages of using steel alloys instead of pure iron.

25. Write balanced chemical equations for the given reactions, which take place in the blast furnace during the extraction of iron:
    a. The reaction of iron (III) oxide with carbon monoxide.
    b. The reaction of calcium oxide with silicon (IV) oxide

Part III: Fill in the blank spaces

26. Gold can be hammered in to sheets. Therefore, it is ________. (malleable/ductile)
27. Copper is used in electrical appliances because it is good conductor of electricity and _________. (ductile/brittle)
28. Copper can displace ______ from its salt solution. (Zinc/silver/iron)
29. The purest form of iron is known as ____________
30. Aluminum metal is extracted by electrolysis from the mineral ____________



UNIT 6

HYDROCARBONS AND THEIR NATURAL SOURCES

Unit Outcomes

At the end of this unit, you will be able to

- discuss the historical development of organic chemistry;
- classify organic compounds;
- write the general formula of alkanes, alkenes and alkynes;
- name simple alkanes, branched-chain alkanes, simple alkenes, branched-chain alkenes and simple alkynes;
- write the molecular and structural formula of simple alkanes, branched chain alkanes, simple alkenes, branched chain alkenes and simple alkynes;
- explain isomerism;
- draw the possible isomers of alkanes, alkenes, and alkynes;
- explain the physical and chemical properties; and the general methods of the preparation of alkanes; alkenes, alkynes, and benzene;
- list the major natural sources of hydrocarbons;
- demonstrate scientific inquiry skills: observing, classifying, communicating, measuring, asking questions, interpreting data, drawing conclusions, applying concepts, predicting and problem-solving.

6.1 Introduction

Start-up Activity

1. Consider the common say,-"The Kingdom of Heaven Runs on Righteousness, but the Kingdom of Earth Runs on Alkanes". Argue for and/or against this say!
2. What sort of materials do you use as a source of energy to cook your food or for transport? Do you know the chemical components of gasoline or diesel oil ('naphtha') that people use for automobiles, 'bajaj' or a motorbike? How is it obtained?
3. Have you encountered a biogas plant in your area? What is its use? How does it work?
4. Discuss the socio-economic impacts of plants as a source of chemicals used in traditional medicine and doping.
5. Can you define organic chemistry or organic compounds? What about the hydrocarbons?

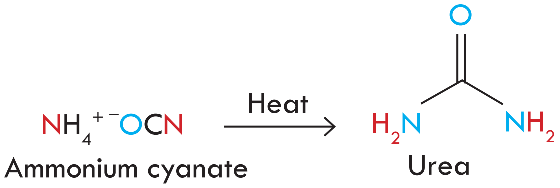
The hydrocarbons that you are going to study in this unit are a class of organic compounds. But, what are organic compounds and organic chemistry? The answers to these questions are all around you.

Most of the foodstuff that we consume every day such as fruits, vegetables, sugar, fats, starch, and vinegar are basically organic compounds. Similarly, substances such as pesticides, herbicides, fuels, dry cleaning chemicals, soaps and detergents, pharmaceuticals, alcoholic beverages, etc. are also organic compounds. Every living organism is made of organic compounds. A large number of organic compounds have been isolated and identified from plants and animals. The organic compounds that are synthesised by living organisms are called natural products. Although plants have principally been used as a source of useful organic compounds such as those in traditional medicine, they also provide some organic compounds that are considered drugs of abuse because they lead to addiction and have negative socio-economic effects. Psychotropic plant chemicals such as nicotine from tobacco plant, cathinone from Chat (khat); from *Catha edulis*; opium (morphine) from poppy plant and cocaine from coca leaf are some examples. Some of these drugs are also used in doping. Doping is the misuse of certain substances and methods to enhance sporting performance and stamina. Although these compounds are perceived as useful to boost zealous, increase energy, and decrease fatigue, they have serious side effects such as personality disorder, low productivity, increased accidents, reduction of competitiveness, inability to cope with tasks that require higher-level judgment and fine motor skills.

6.1.1 History of Organic Chemistry

At the end of this section, students will be able to narrate the historical development of organic chemistry.

The science of organic chemistry began to flourish with the demise of a nineteenth-century theory called Vitalism. According to Vitalism theory, organic compounds were only those that came from living organisms, and only living things could synthesize organic compounds through the intervention of a vital force. Inorganic compounds were considered those compounds that came from non-living sources. The German chemist Friedrich Wöhler, however, discovered in 1828 that an organic compound called urea (a constituent of urine) could be synthesized in the laboratory by evaporating an aqueous solution of the inorganic compound ammonium cyanate. With this discovery, the synthesis of an organic compound began the evolution of organic chemistry as a scientific discipline. Since Wohler's discovery, chemists have developed a remarkably sophisticated ability to design and synthesize new organic compounds in the laboratory-medicines, dyes, polymers, and a host of other substances that are helpful to society.



To date, over 50 million organic compounds are known. The list included organic compounds isolated from natural sources and also those which were synthesized in the laboratories. All of the organic compounds whether of synthetic or natural origin contain the element carbon. Therefore, organic chemistry, today, is defined as the study of carbon compounds except the oxides of carbon, carbonates, hydrogen carbonates, cyanides and cyanates. But why is carbon special? If you have a look at the periodic table, you will see over a hundred elements listed there. The question that comes to mind is this: why should an entire field of chemistry be based on the chemistry of compounds that contain this specific element, carbon?

The reason is linked to the unique property of carbon called catenation. Catenation is the ability of atoms of the same element to join together to form short or long chains and rings. Carbon atoms can form strong covalent bonds to one another to form rings and chains of carbon atoms. And also they can form strong bonds to elements such as hydrogen, nitrogen, oxygen, and sulfur. Because of these bond-forming properties, carbon can be the basis for the huge diversity of compounds. A few other elements like sulphur, silicon and boron show this behavior, but to a much lesser extent.

Activity 6.1

Form a group, discuss on the answers for the questions under this activity, and present your conclusions to the class.

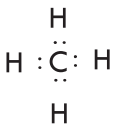
1. Despite the demise of vitalism in science, the word 'organic' is still used today by some people as 'organic vitamins', 'organic fertilizers', or 'organic food.' What do you understand from these terms? Is pure 'natural vitamin C from orange juice', for example, healthier than pure 'synthetic vitamin C'; or both are important? Discuss!

6.1.2 Structural Representations of Organic Molecules

Before we proceed to the main topic of this unit-the hydrocarbons, we shall briefly review the bonding concept. Recall from your Grade 9 that carbon has four valence electrons and it can form four covalent bonds with other atoms. You also studied Lewis structure- a representation of covalent bonding in which shared electron pairs of covalent bonds are shown as pairs of dots or lines between two atoms and lone pairs are shown as pairs of dots on individual atoms. However, as the size of the compound increases, it becomes difficult to portray the structure of molecules with the use of Lewis structure. This is why scientists have developed other structural representations including structural formula, condensed structural formula, bond-line representation, and polygon formula. Let us take a look at these structures.

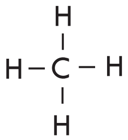
a. Complete Structural Formula

Complete Structural formula is obtained by representing the two-electron covalent bond by a single dash. A single dash represents a single bond, a double dash represents a double bond and a triple dash represents a triple bond. Lone pairs of electrons on heteroatoms such as oxygen, halogens and more, may or may not be shown. Methane (CH_4), for instance, contains one carbon atom and four hydrogen atoms. The carbon atom contributes its 4 valence electrons and the four hydrogen atoms contribute four electrons (each H atom contributes one electron). Therefore, the Lewis structures for a methane molecule having 4 + 4 = 8 electrons is drawn following the rule (see your grade 9 lesson) as:



$$\begin{array}{c} H \\ H:\underset{..}{\ddot{C}}:H \\ H \end{array}$$

Methane
Lewis Formula

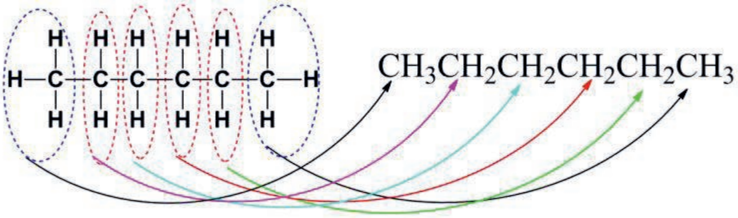


$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

Methane
Structural Formula

b. Condensed Structural Formula

The condensed structural formula is an abbreviated form of structural formula. It is obtained by omitting some or all the dashes that represent covalent bonds. Hexane (C_6H_{14}) can be represented by the structural formula or the condensed structural formula as shown below.



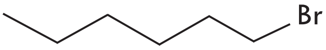
Structural Formula Condesed Structural Formula

The condensed structural formula can further be condensed by putting identical repetitive units in parenthesis and indicating their repetition using subscripts. For example, we can further condense $CH_3CH_2CH_2CH_2CH_2CH_3$ to $CH_3(CH_2)_4CH_3$.

c. Bond-Line Structural Representation

In a bond-line structural representation, carbon and hydrogen atoms are not shown but are understood to be there based on the fact that each carbon forms 4 covalent bonds. The lines representing carbon-carbon bonds are drawn in a zig-zag fashion as exemplified by the bond-line structure for hexane and hexyl bromide:

C_6H_{14}, hexane

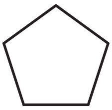


$C_6H_{13}Br$, hexyl bromide

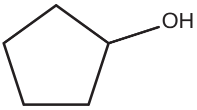
Only heteroatoms need to be indicated in bond line representation. The terminals in the bond-line structures denote methyl (CH_3) groups unless indicated otherwise by a functional group, while the line junctions denote carbon atoms bonded to an appropriate number of hydrogen atoms required to satisfy the valency of the carbon atoms. Note that in hexyl bromide one of the hydrogens in the terminal CH_3 on one side of the structure is replaced Br leaving a CH_2Br group.

d. Polygon Formula

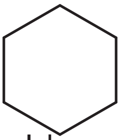
There are many organic compounds, in which carbon atoms are not joined in a chain, but rather are joined in a ring. These cyclic compounds are usually represented by polygon using bond-lines without showing carbon and hydrogen atoms. The corners of a polygon represent a carbon atom and the sides of a polygon denote a carbon-carbon bond. Similarly, if an atom or a group of atoms other than hydrogen is attached to carbon, then that atom or a group of atoms manifest in this structure. As example, cyclopentane (C_5H_{10}), cyclopentanol ($C_5H_{10}O$), cyclohexane (C_6H_{12}), and bromocyclohexane ($C_6H_{11}Br$) are depicted below:



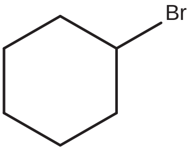
cyclopentane



cyclopentanol



cyclohexane



bromocyclohexane

6.1.3 Classes of Organic Compounds

At the end of this section, students will be able to

- classify organic compounds;
- define the term functional group.

Activity 6.2

Form a group, discuss on the following activity and present your responses to the class.

You are familiar with a classification in chemistry as well as biology. For instance, you have studied that the elements in periodic table are classified as metals, non-metals, and metalloids. What is the purpose of classification? Why don't we deal with individual entities? What is the basis of classification? Why do you think we should classify organic compounds and how?

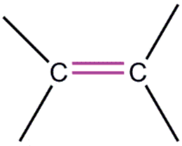
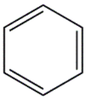
We have noted in section 6.1 about the unique properties of carbon giving rise to a very vast number of organic compounds known. Indeed, more and more are being added to the list each day. Each of these compounds has its own physical properties, such as melting point and boiling point, and each has its own chemical reactivity. Chemists have learned through years of experience that organic compounds can be classified into families according to their structural features that are responsible for their properties so that members of a given family often have similar chemical behavior.

Thus, instead of having 50 millions of compounds with random reactivity, it is better to have a few dozen families of organic compounds whose chemistry is reasonably predictable. The structural features that make it possible to classify compounds into families are called functional groups.

Functional Groups

A functional group is a group of atoms and/or bonds within a molecule that has a characteristic chemical behavior. Chemically, a given functional group behaves in nearly the same way in every molecule it's a part of. ***Table 6.1*** lists the common functional groups of hydrocarbons, name ending, and specific examples. It also gives simple examples of their occurrence. The functional group of an alkene, for example, is its carbon-carbon double bond. The functional group of an alkyne is its carbon–carbon triple bond. Alkanes do not have a functional group. Their molecules have carbon–carbon and carbon–hydrogen single bonds only, but these bonds do present in molecules of almost all organic compounds. In addition, C-C and C-H bonds are, in general, much less reactive than common functional groups.

Table 6.1 Some classes of organic compounds, their functional groups, name ending, and examples.

Organic compound	**Functional group**	**Name ending**	**Example**
Alkane	-	-ane	CH_3-CH_3 Ethane
Alkene		-ene	$H_2C=CH_2$ Ethene
Alkyne	$—C\equiv C—$	-yne	$H—C\equiv C—H$ Ethyne
Aromatic		None	 benzene

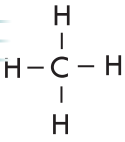
6.2 Saturated Hydrocarbons: Alkanes (C_nH_{2n+2})

At the end of this section, students will be able to

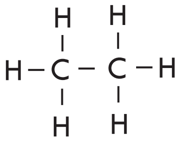
- define hydrocarbon;
- define saturated hydrocarbon;
- define homologous series;
- drive the general formula of alkanes from based on the number of hydrogen and carbon;
- write the first ten members of alkanes homologous series;
- write the molecular formulas of alkanes from the given numbers of carbon atom.

Activity 6.3

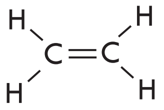
In unit 2, you have classified a solution as saturated, unsaturated, or supersaturated based on the amount of solute it contains relative to the maximum dissolving capacity of the solvent. In grade 9, you also studied that the valence of carbon is 4; meaning the maximum number of atoms that carbon can form bonds to is 4. Based on this analysis, can you classify the following compounds as saturated and unsaturated?



I.



II.



III.

$H-C\equiv C-H$

IV.

Compounds that contain only carbon and hydrogen atoms are called hydrocarbons. Hydrocarbons may be classified as aliphatic ("fat-like") or aromatic (***Figure 6.1A***). Aliphatic hydrocarbons exhibit a "fat-like" behavior. This includes alkane, alkene, alkyne, or their derivatives. Aromatic hydrocarbons are those hydrocarbons that contain a low hydrogen-to-carbon ratio. Benzene (C_6H_6), for instance, is a typical aromatic hydrocarbon having a ratio of number of hydrogen atoms to number of carbon atoms only 1.

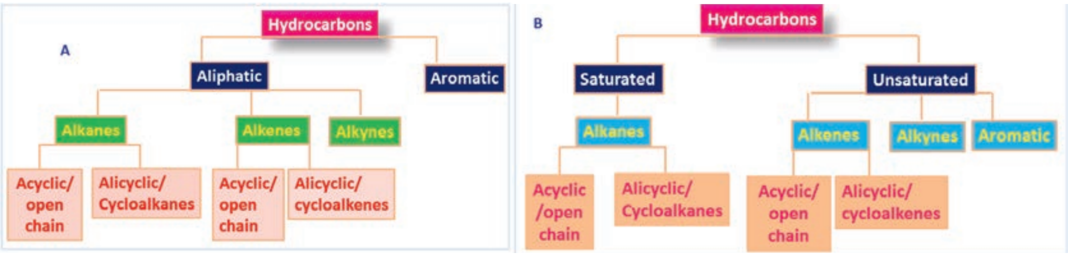


Figure 6.1 Classification of hydrocarbons, A as aliphatic and aromatic and B as saturated and unsuturated.

Hydrocarbons may also be classified as suturated and unsaturated. Generally speaking, compounds such as alkanes (e.g. methane, CH_4), whose molecules contain only single bonds are referred to as saturated compounds. Because they contain the maximum number of hydrogen atoms per carbon atoms that the carbon compound can possess. Remember from Grade 9 that the valency of carbon is 4. This means C can form a maximum of 4 covalent bonds with 4 hydrogen atoms because each hydrogen can form only one (valency of H is 1). Compounds with multiple bonds, such as alkenes (e.g. ethylene), alkynes (e.g. acetylene), and aromatic hydrocarbons (e.g. benzene), are called unsaturated compounds because they possess fewer than the maximum number of hydrogen atoms per carbon, and they are capable of reacting with hydrogen under proper conditions. Let's begin with saturated compounds!

Alkanes

Structures I and II in ***Activity 6.3*** represent methane and ethane, respectively, the typical alkanes. Alkanes are often described as saturated hydrocarbons: hydrocarbons because they contain only carbon and hydrogen; saturated because they contain the maximum possible number of hydrogens per carbon. Alkanes are also occasionally called aliphatic compounds, a name derived from the Greek aleiphas, meaning "fat." Alkanes can be unbranched (normal) or branched. ***Table 6.2*** lists the common names and formulae of some members of the unbranched alkane family.

Table 6.2 Some members of the unbranched alkane family.

Name	No of C atoms	Structure	Molecular formula	Name	No of C atoms	Molecular formula
Methane	1	CH_4	CH_4	Undecane	11	______
Ethane	2	CH_3CH_3	C_2H_6	Dodecane	12	______
Propane	3	$CH_3CH_2CH_3$	C_3H_8	Tridecane	13	______
Butane	4	$CH_3(CH_2)_2CH_3$	C_4H_{10}	Tetradecane	14	______
Pentane	5	$CH_3(CH_2)_3CH_3$	C_5H_{12}	Pentadecane	15	______
Hexane	6	$CH_3(CH_2)_4CH_3$	C_6H_{14}	Hexadecane	16	______
Heptane	7	$CH_3(CH_2)_5CH_3$	C_7H_{16}	Heptadecane	17	______
Octane	8	$CH_3(CH_2)_6CH_3$	C_8H_{18}	Octadecane	18	______
Nonane	9	$CH_3(CH_2)_7CH_3$	C_9H_{20}	Nonadecane	19	______
Decane	10	$CH_3(CH_2)_8CH_3$	$C_{10}H_{22}$	Eicosane	20	______

Activity 6.4

1. Consider ***Table 6.2*** and answer the following questions
   - i. What trends do you observe from the Table?
   - ii. What does family name and specific name for a member hold in common?
   - iii. Based on the trend you have observed from the Table, write the general formula for an alkane having "n" carbon atoms.
   - iv. Using the general formula, complete the missing molecular formula in the Table for the members undecane through eicosane.
   - v. Draw condensed structures and bond line representations for the first ten alkanes.
   - vi. Discuss the analogy in the usage of the terms "saturated" and "unsaturated" in this section and in unit 2.

6.2.1 Alkane Homologous Series

If you examine the alkanes in ***Table 6.2***, you can notice that each alkane differs from the preceding alkane by one $-CH_2-$ group. Butane, for example, is $CH_3(CH_2)_2CH_3$ and pentane is $CH_3(CH_2)_3CH_3$. A series of compounds like this, where each member differs from the next member by a constant unit, is called a homologous series. A homologous series has the following characteristics:

- has a constant unit between two consecutive members;
- has a general formula;
- shows trends in physical properties and
- exhibit the same reactivity (chemical property). Members of a homologous series are called homologues.

Exercise 6.1

1. Define homologous series.
   Describe characteristics of a homologous series.
2. Determine if each statement about alkane homologous series is true or false.
   a. It has general formula and a name.
   b. Consecutive members differ by a constant ($-CH_2-$) unit.
   c. All members contain the same functional group.
   d. Consecutive members differ by a constant molecular mass of 14.

6.2.2 Physical Properties of Alkanes

At the end of this section, students will be able to explain the physical properties of alkanes.

Form a group, discuss on the following questions and present your responses to the class.

You have learnt in unit 2, that particles in liquid and solid state are hold together by intermolecular forces while the gaseous state is characterized by absence of such forces. You have also studied that among the various intermolecular forces the dispersion forces are the ones possessed by nonpolar molecules and increases as molecular weight increases.

Activity 6.5

You also learnt that polarity is the result of difference in electronegativity of bonded atoms. Given the fact that carbon and hydrogen atoms have similar electronegativity, answer the following questions:

a. Are alkanes polar or nonpolar molecules?

b. What is the intermolecular forces that could exist in alkanes that are liquids and solids at room temperature

c. Which member of alkanes do you think are gases, lower mass or higher mass? Why?

d. What happens to the boiling point and melting point of alkanes as number of carbon atoms or molecular masses increases? Can you explain your reasoning?

e. Are alkanes soluble in water? Explain based on the like dissolves like rule you learnt in unit 2.

At room temperature (25°C and 1 atm pressure), the first four members of the homologous series of unbranched alkanes are gases (*Table 6.3*), the C_5-C_{17} unbranched alkanes (pentane to heptadecane) are liquids, and the unbranched alkanes with 18 and more carbon atoms are solids. As a class, the alkanes are the least dense of all groups of organic compounds. All alkanes have densities considerably less than 1.00 g mL^{-1} (the density of water at 4 °C). As a result, petroleum (a mixture of hydrocarbons rich in alkanes) floats on water.

Table 6.3 Physical properties and isomers of some members of alkane family.

Name	No of C atoms	Physical state (at room temp)	M.P (°C)	B.P (°C)	Density	No of isomers
Methane	1	gas	-183	-162	0.717 g/L	0
Ethane	2	Gas	-172	-89	1.35 g/L	0
Propane	3	Gas	-187	-42	2.02 g/L	0
Butane	4	Gas	-135	-0.5	2.48 g/L	2
Pentane	5	Liquid	-130	36	0.63 g/mL	3
Hexane	6	Liquid	-94	68	0.66 g/mL	5
Heptane	7	Liquid	-91	98	0.68 g/mL	9
Octane	8	Liquid	-57	126	0.70 g/mL	18
Nonane	9	Liquid	-54	151	0.71 g/mL	35
Decane	10	Liquid	-30	174	0.73 g/mL	75

Exercise 6.2

Explain each of the following facts about alkanes.

1. Alkanes show regular increase in boiling point as their molecular weight increases.
2. Branching decreases the boiling point of alkanes.
3. Alkanes are insoluble in water.

6.2.3 Nomenclature of Alkanes

At the end of this section, students will be able to

- apply IUPAC rules to name straight and branched chain alkanes;
- write the structural formulas of the first ten alkanes.

Activity 6.6

Form a group, discuss on the following and present your responses to the class.

1. There might be many students in your class. Each student has his/her name? Can you explain the purpose of naming? Is the way girls are named different from boys? If yes, why do you think? Is there any pattern of systemic doing on the way girls and boys are named? Why do you think we have such a systematic naming for girls and boys? Discuss!

How do your parents call you? Like you are called in school or different? If you have a different name commonly used by your parents, do you mind telling it to your classmate as if it were your common name "my common name is ____?" How do you think parents select a common name to their children? Any criteria?

As the science of organic chemistry slowly grew in the 19th century, so too did the number of known compounds and the need for a systematic method of naming them grow. Alkanes can be named in two ways, common names and IUPAC names.

Common Names of Alkanes

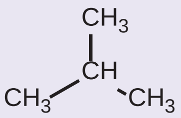
Lower members of the alkane homologous series have common names. The prefixes used in the common names are n- (normal), 'iso-' and 'neo-'. The prefix n- is used when all the carbon atoms form a continuous chain.

Example 6.1 $CH_3CH_2CH_2CH_2CH_3$ $CH_3CH_2CH_2CH_2CH_2CH_3$

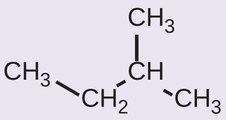
n-pentane n-hexane

Iso- is used when all of the carbon atoms form a continuous chain, except for the one next to the last carbon.

Example 6.2:



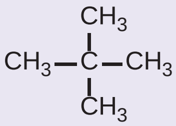
isobutane



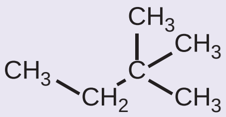
isopentane

Neo-is used when the central carbon is bonded to four other carbon atoms.

Example 6.3:



neopentane



neohexane

The System of IUPAC Nomenclature

The system of nomenclature we'll use in here is the one that was devised by the International Union of Pure and Applied Chemistry (IUPAC, usually read as eye-you-pac). The IUPAC nomenclature can be applied to both branched and unbranched alkanes.

IUPAC Rules to Name Unbranched Alkanes

Compounds like butane and pentane, whose carbons are all connected in a row, are called straight-chain alkanes, or normal alkanes. Compounds like 2-methylpropane (isobutane), 2-methylbutane, and 2, 2-dimethylpropane, whose carbon chains branch, are called branched-chain alkanes. In unbranched alkanes, each carbon atom within the chain is bonded to no more than two other carbon atoms and that unbranched alkanes contain only primary and secondary carbon atoms. Primary (connect directly to one carbon), secondary (connect directly to two carbons), and tertiary (connect directly to three carbons) carbon atoms, quaternary (connected to four other carbons). The names for several of the unbranched alkanes are listed in ***Table 6.2***. The ending (suffix) for all of the names of alkanes is -ane. The stems of the names of most of the alkanes (above C_4) are of Greek and Latin origin. Thus the parent names for one, two, three, four, and five become meth-, eth-, prop-, but-, and pent-.

IUPAC Rules to Name branched-chain Alkanes

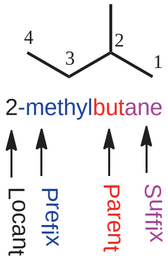
A chemical name for branched alkanes typically has four parts in the IUPAC system of nomenclature: locant, prefix, parent, and suffix.

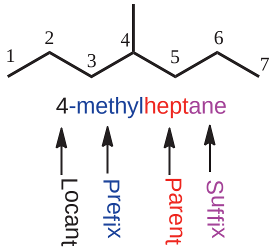
Locant — Prefix — Parent — Suffix

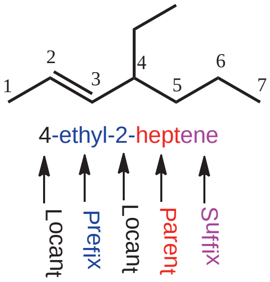
Locant	Prefix	Parent	Suffix
Where are the substituents and functional groups?	What are the substituents	How many carbons?	What is primary functional groups?

The prefix identifies the various substituent groups in the molecule, the parent selects the main part (longest chain) of the molecule and tells how many carbon atoms are in that part, the locants give the positions of the functional groups and substituents, and the suffix identifies the primary functional group.

The examples below illustrate the basic components of the IUPAC nomenclature. The methyl and ethyl groups in, for example, 2-methylbutane and 4-ethyl-2-heptene are substituents and are called alkyl groups. What are alkyl groups?





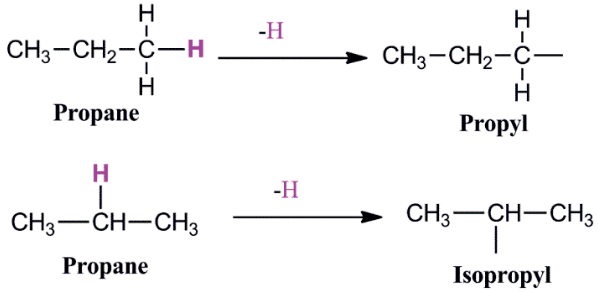


Alkyl Groups and the Symbol R

If we remove one hydrogen atom from an alkane, we obtain what is called an alkyl group or alkyl radicals. These alkyl groups have names that end in -yl. When the alkane is unbranched, and the hydrogen atom that is removed is a terminal hydrogen atom, the names are straightforward as illustrated below:

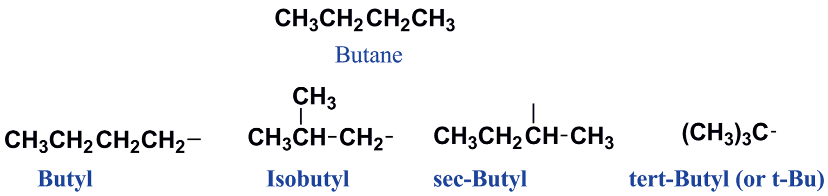
Alkane	Alkyl Group	Abbreviation
CH_3-H methane	CH_3- Methyl	Me-
CH_3-CH_2-H ethane	CH_3-CH_2- Ethyl	Et-
$CH_3-CH_2-CH_2-H$ propane	$CH_3-CH_2-CH_2-$ Propyl	Pr-
$CH_3-CH_2-CH_2-CH_2-H$ butane	$CH_3-CH_2-CH_2-CH_2-$ Butyl	Bu-

While only one alkyl group can be derived from methane or ethane (the methyl and ethyl groups, respectively), two groups can be derived from propane. Removal of a hydrogen atom from one of the end carbon atoms gives a group that is called the propyl group; removal of a hydrogen atom from the middle carbon atom gives a group that is called the isopropyl group.

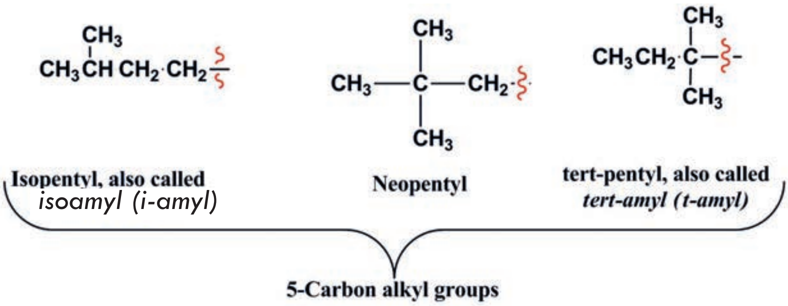


There are four C_4 alkyl groups.

Four-Carbon Alkyl Groups



There is one five-carbon group with an IUPAC approved common name that you should also know: the 2,2-dimethylpropyl group, commonly called the neopentyl group. Other common names for five carbon alkyl groups include isopentyl and tert-pentyl groups.

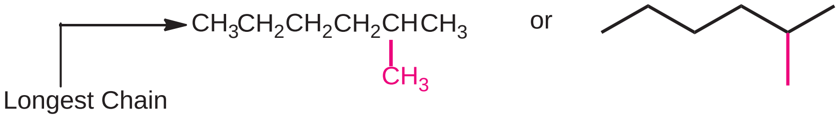


The common names isopropyl, isobutyl, *sec*-butyl, *tert*-butyl, isopentyl, neopentyl, and *tert*-pentyl are approved by IUPAC for the unsubstituted groups, and they are still very frequently used. You should learn these groups so well that you can recognize them any way they are written.

The symbol R is used as a general symbol to represent any alkyl group. For example, R might be a methyl group, an ethyl group, a propyl group, an isopropyl group, etc. in the general formula for alkane R-H. Similarly, in R-COOH which is the general formula for carboxylic acid, R can be any alkyl group (methyl, ethyl, isopentyl, etc.).

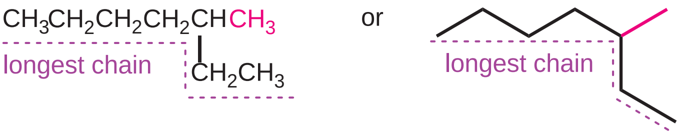
Branched-chain alkanes are named according to the following rules:

1. Locate the longest continuous chain of carbon atoms; this chain determines the parent name for the alkane. The parent name for the following compound, for example, is hexane because the longest continuous chain contains six carbon atoms:



$$\underset{\textstyle CH_3}{CH_3CH_2CH_2CH_2CHCH_3}$$

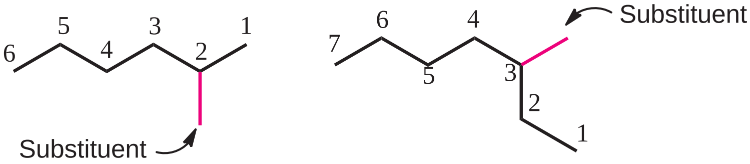
The longest continuous chain may not always be obvious from the way the formula is written. Notice, for example, that parent name of the following alkane is designated as a heptane because the longest chain contains seven carbon atoms:



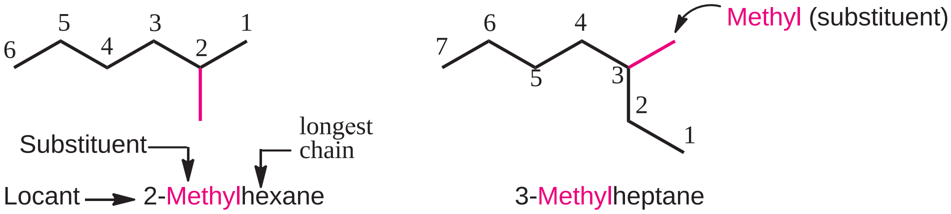
$$\underset{\textstyle CH_2CH_3}{CH_3CH_2CH_2CH_2CHCH_3}$$

2. Number the longest chain beginning with the end of the chain nearer to the substituent.

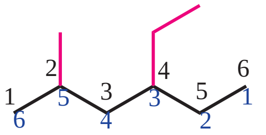
Applying this rule, we number the two alkanes that was illustrated previously in the following way:



3. Use the numbers obtained by the application of rule 2 to designate the location of the substituent group. The parent name is placed last, and the substituent group, preceded by the number designating its location on the chain, is placed first. Numbers are separated from words by a hyphen. Our two examples are 2-methylhexane and 3-methylheptane, respectively:

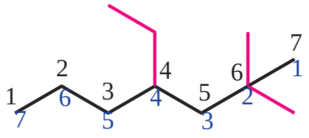


4. When two or more substituents are present, give each substituent a number corresponding to its location on the longest chain. For example, we designate the following compound as 4-ethyl-2-methylhexane:



4-Ethyl-2-methylhexane

Not 3-Ethyl-5-methylhexane



4-Ethyl-2, 2-dimethylheptane

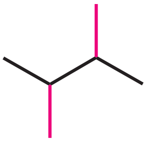
Not 4-Ethyl-6,6-dimethylheptane

To check which direction of numbering is correct in this case; add the locants (numbers assigned to the carbon atoms to which substituents are attached). The sum of the locants in the first case is 2 + 4 = 6 and in the second case is, 3 + 5 = 8. So numbering should be done from left to right in this case. Thus, the first option is correct. The substituent groups should be listed alphabetically (i.e., ethyl before methyl). In deciding on alphabetical order, disregard multiplying prefixes such as 'di', 'tri', and 'tetra'. Ethyl comes first compared to methyl when 'di is ignored, as illustrated in 4-ethyl-2,2-dimethylheptane. In addition, disregard structure-defining prefixes that are written in italics and separated from the name by a hyphen. Thus tert-butyl precedes ethyl (tert is ignored), but ethyl precedes isobutyl ('iso' not ignored.

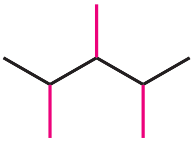
5. When two substituents are present on the same carbon atom, use that number twice:

3-Ethyl-3-methylhexane

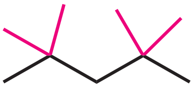
6. When two or more substituents are identical, indicate this by the use of the prefixes di-, tri-, tetra-, and so on. Then make certain that each substituent has a number. Commas are used to separate numbers from each other:



2,3-Dimethylbutane



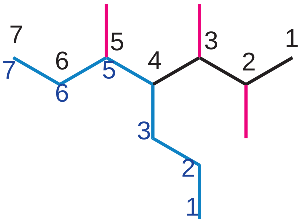
2,3,4-Trimethylpentane



2,2,4,4-Tetramethylpentane

The application of these six rules allows us to name most of the alkanes that we shall encounter. Two other rules, however, may be required occasionally:

7. When two chains of equal length compete for selection as the parent chain, choose the chain with the greater number of substituents:



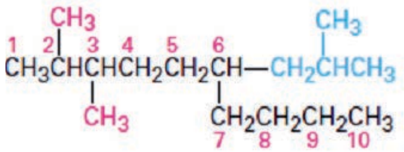
2,3,5-trimethyl-4-propylheptane

Not 4-(1,2-dimethylpropyl)-5-methylheptane

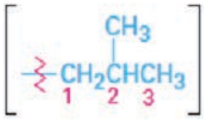
8. When branching first occurs at an equal distance from either end of the longest chain, choose the name that gives the lower number at the first point of difference:

2,3,5-trimethylhexane
Not 2,4,5-trimethylhexane

9. Name a complex substituent as though it were itself a compound. In some particularly complex cases, a fifth step is necessary. It occasionally happens that a substituent on the main chain has sub-branching. In the following case, for instance, the substituent at C_6 is a three-carbon chain with a methyl sub-branch. To name the compound fully, the complex substituent must first be named.



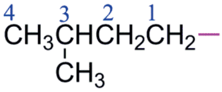
$CH_3CH(CH_3)CH(CH_3)CH_2CH_2CH(CH_2CH(CH_3)CH_3)CH_2CH_2CH_2CH_3$



$-CH_2CH(CH_3)CH_3$

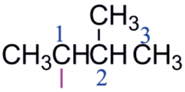
2,3-Dimethyl-6-(2-methylpropyl)decane A 2-methylpropyl group
or 2,3-Dimethyl-6-isobutyldecane (isobutyl group)

Note that, for alkyl group, numbering begins at the carbon directly connected to the parent chain and continues to the longest direction. Further examples:



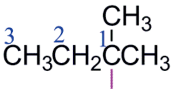
$CH_3CH(CH_3)CH_2CH_2-$

3-methylbutyl



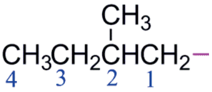
$CH_3CH(-)CH(CH_3)CH_3$

1,2-dimethylpropyl



$CH_3CH_2C(CH_3)(-)CH_3$

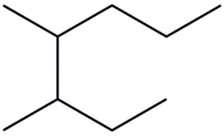
1,1-dimethylpropyl



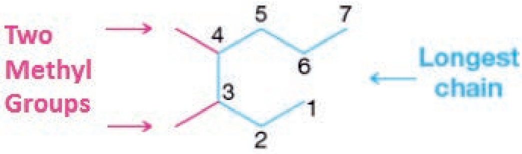
$CH_3CH_2CH(CH_3)CH_2-$

2-methylbutyl

Example: Write IUPAC name for:

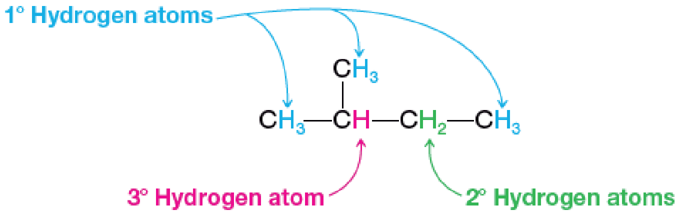


Strategy and solution: We find the longest chain (shown in blue) to be seven carbons; therefore the parent name is heptane. There are two methyl substituents (shown in red). We number the chain so as to give the first methyl group the lower number. The correct name, therefore, is 3,4-dimethylheptane. Numbering the chain from the other end to give 4,5-dimethylheptane would have been incorrect.

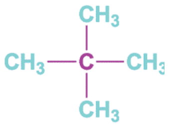


How to Classify Hydrogen Atoms

We have just seen how to classify carbon in alkanes as primary, secondary, tertiary, and quaternary. The hydrogen atoms of an alkane are classified based on the carbon atom to which they are attached. A hydrogen atom attached to a primary carbon atom is a primary (1°) hydrogen atom, and so forth. The following compound, 2-methylbutane, has primary, secondary (2°), and tertiary (3°) hydrogen atoms:


$\mathrm{CH_3-CH(CH_3)-CH_2-CH_3}$

Note that 2,2-dimethylpropane, a compound that is often called neopentane, has only primary hydrogen atoms:

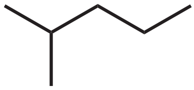

$\mathrm{CH_3-C(CH_3)_2-CH_3}$

Neopentane

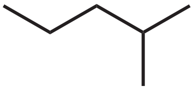
Exercise 6.3

1. Which structure does not represent 2-methylpentane?

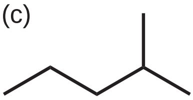
(a)



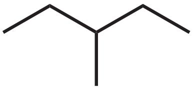
(b)



(c)



(d)



2. Write the structure and give the IUPAC name for an alkane with formula C_6H_{14} that has only primary ($-CH_3$) and secondary ($-CH_2-$) carbon atoms.

6.2.4 Isomerism in Alkanes

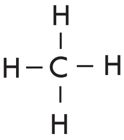
At the end of this section, students will be able to

- define isomerism;
- define structural isomerism;
- draw the possible structural isomers for C_4H_{10}, C_5H_{12} and C_6H_{14}.

Activity 6.7

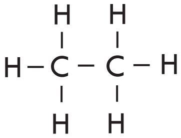
Among animals, it is very common to encounter some dogs that appear pretty the same in terms of size, and colour. How would you identify them apart if you encounter two such dogs? Assume that one of the dogs is your friendly neighbour. What about chemicals? Can you name two different compounds that have the same molecular formula and molecular weight? What do you call them?

Unlike the first three members of the alkane homologous series, i.e. methane, ethane, and propane,



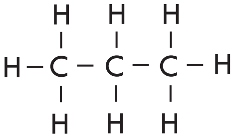
$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H \\ & & | & & \\ & & H & & \end{array}$$

Methane



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H \\ & & | & & | & & \\ & & H & & H & & \end{array}$$

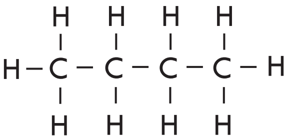
Ethane



$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

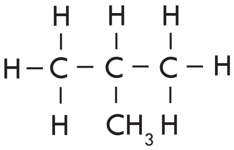
Propane

in molecules having four and above carbon atoms the carbons can be arranged in two or more ways. A molecule having the formula C_4H_{10} you know as butane can have two structures.



$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & \end{array}$$

Butane



$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & H \\ & & | & & | & & | & & \\ & & H & & CH_3 & & H & & \end{array}$$

Isobutane

The two structures are drawn using the possibilities of connecting the four carbon atoms in two different ways. In the first case, all the four carbon atoms in butane are connected in raw (linearly). In the second case, the four carbon atoms are connected in a branched manner to give isobutane. Butane and isobutane have the same molecular formula, C_4H_{10}, but have different structures and properties. Similarly, the 5 C atoms in pentane (C_5H_{12}) can be arranged in three ways. Compounds like the two C_4H_{10} molecules and the three C_5H_{12} molecules, which have the same molecular formula but different structures, are called structural isomers, from the Greek isos + meros, meaning "made of the same parts." A formula to determine the number of isomers of alkanes containing 4 – 7 carbon atoms is $2^{n-4} + 1$.

Definition: Isomers are compounds that have the same numbers and kinds of atoms but differ in the way the atoms are arranged.

Definition: Structural isomers are two or more different compounds that have the same molecular formula but different connectivity or structures. As ***Table 6.3*** shows, the number of possible alkane isomers increases dramatically as the number of carbon atoms increases.

Note: Structural isomers have different physical properties. They are always found to have different melting points, boiling points, densities, indexes of refraction, and so forth.

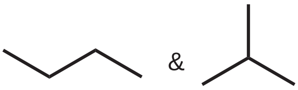
Activity 6.8

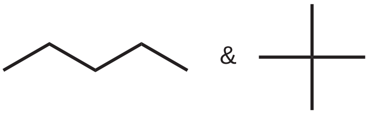
Form a group, discuss on the answers for the questions under this activity, and present your conclusions to the class.

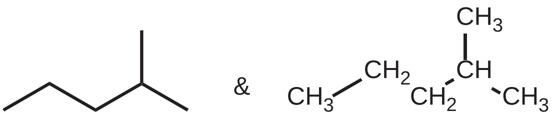
1. Consider two hypothetical compounds A and B having the same molecular formula, C_4H_{10}. Both are colorless and odorless gases at room temperature. Compound A has a boiling point of -1.0 °C whereas B has a boiling point of -11.7 °C. Draw structure of compound A and compound B. What is the relationship between the two compounds? What conclusions can you make about the physical properties of such pairs of compounds?
2. Is the statement "two compounds that have the same molecular formula have the same properties" true or false? Explain your reasoning.

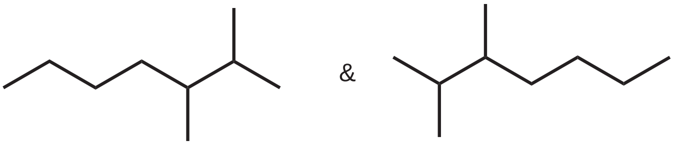
Exercise 6.4

1. Determine whether the pairs of structures in each of the following boxes are isomers or simply different structural representation of the same compound. Hint: If IUPAC names are different, the structures are also different.

(a)


(b)


(c)


(d)


6.2.5 Preparation of Alkanes

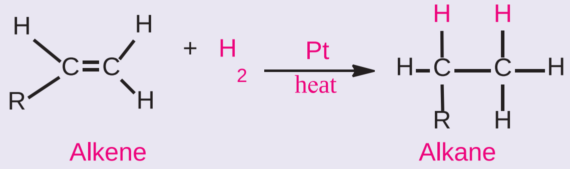
At the end of this section, students will be able to

- describe the general methods for preparation of alkanes in a laboratory;
- synthesize methane in a laboratory by decarboxylation method;
- carry out a project work to produce biogas from cow dung.

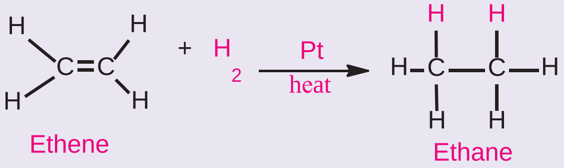
Alkanes are the major constituents of petroleum and natural gas. They are mainly obtained by fractional distillation of petroleum. Alkanes can also be prepared in the laboratory. Some methods of their preparation are as follows:

1. Hydrogenation of alkenes with a metal catalyst.

General Reaction:



Example 6.4:



2. Wűrtz Synthesis (reaction). This method involves the reaction of a halogenated alkane with sodium and the reaction is somewhat exothermic.

General reaction: $2R\text{-}X + 2Na \longrightarrow R\text{-}R + 2NaX$

2R-X: Halogenated alkane/ Alkyl halide
R-R: Alkane
2NaX: Sodium halide

Example 6.5:

$2CH_3Cl + 2Na \rightarrow CH_3 - CH_3 + 2NaCl$

Methyl chloride, Ethane, Sodium chloride

This reaction is named as Wűrtz reaction after the French chemist Charles-Adolphe Wűrtz (1817–1884).

3. Heating sodium salt of an organic acid with soda lime (mixture of sodium hydroxide and calcium oxide); the reaction brings about the removal of the carboxylate group from the sodium salt of the carboxylic acid. This type of reaction is called decarboxylation.

General reaction: $RCOONa + NaOH \xrightarrow[\text{heat}]{CaO} R\text{-}H + Na_2CO_3$

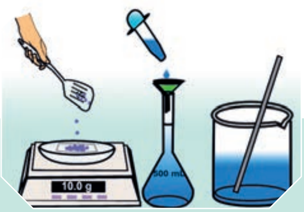
RCOONa: Sodium salt of carboxylic acid
R-H: Alkane
Na_2CO_3: Sodium carbonate

Example 6.6: (Preparation of methane by decarboxylation reaction)

$$\underset{\text{Sodium acetate}}{CH_3COONa} + NaOH \xrightarrow[\text{heat}]{CaO} \underset{\text{Methane}}{CH_3\text{-}H} + \underset{\text{Sodium Carbonate}}{Na_2CO_3}$$

Methane is the main constituent (about 90%) of natural gas. It is obtained during fractional distillation of petroleum. Methane is also formed by the decay of animal and plant remains in swampy or marshy lands.

Note: The materials that can be used to produce biogas include dung from cattle, pigs, chickens, chopped green plants and plant wastes. The advantages of biogas technology include saving on fuel such as kerosene, wood and charcoal, and decrease in local deforestation. The residue obtained after preparing the biogas can be used as fertilizer.



Experiment 6.1

Laboratory Preparation of Methane

Objective: To prepare methane and study its properties.

Materials required: Sodium acetate (CH_3COONa) and soda lime (CaO, NaOH), test tubes, delivery tube, stopper, gas jar and gas jar lid, pneumatic trough, stand, clamp, beehive shelf, Bunsen burner and balance.

Procedure:

Arrange the assembly as shown in ***Figure 6.2***. Mix thoroughly 5 g powdered sodium acetate with 10 g of soda lime. Place the mixture in the test tube and heat it. Collect the gas by the downward displacement of water. Collect several jars of the gas.

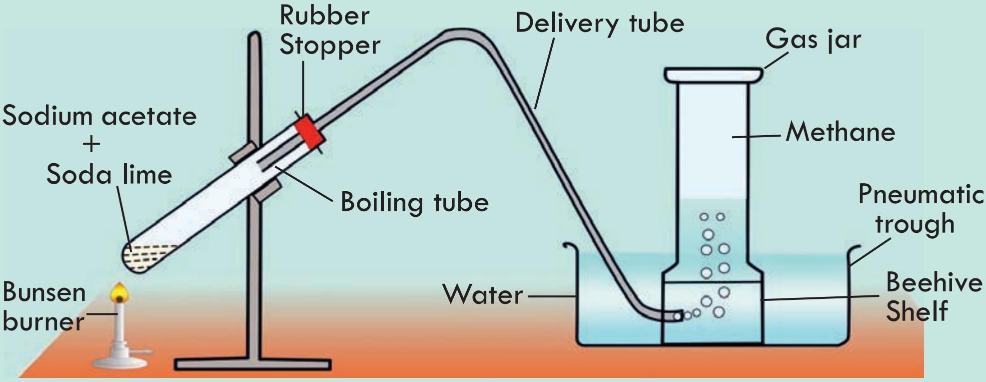


Figure 6.2 Laboratory preparation of methane.

Carry out the following activities with the collected gas:
Insert a burning splint into the jar full of the gas. Observe what happens in the jar.
Add $Ca(OH)_2$ solution to the jar and observe.
Add a few drops of bromine water to any one of the gas jars filled with methane. Cover the gas jar and shake well.

Observations and analysis:

a. What is the colour of the gas? Is it soluble in water?
b. Is the gas combustible? Write a balanced equation for the change.
c. What change did you observe upon the addition of $Ca(OH)_2$?
d. Is there any change when bromine water is added?

Write a laboratory report of your observations, and submit to your teacher.

Exercise 6.5

1. Write the products when:
   a. $CH_3CH=CH_2$ reacts with hydrogen in the presence of Pt as catalyst
   b. Ethyl bromide reacts with sodium metal
   c. Sodium propanoate (CH_3CH_2COONa) reacts with soda lime
2. The reaction of a halogenated alkane with sodium is named in honor of_______________

Experiment 6.2

Production of Methane from Cow Dung

Objective: To prepare biogas from locally available materials.

Materials required: Conical flask, cow dung, water, delivery tube with tap and stopper with one hole.

Procedure:

1. Mix some cow dung with water and pour it into a conical flask, *Figure 6.3*.
2. Fit the conical flask with a stopper in which a delivery tube with a tap is inserted.
3. Cover the conical flask with a ball of cotton wool and place it near a window, and leave it there for 3 to 4 days.
4. Check the formation of methane after 4 days. (Bring a lighted splint closer to the outlet of the delivery tube and open the tap). See what happens.

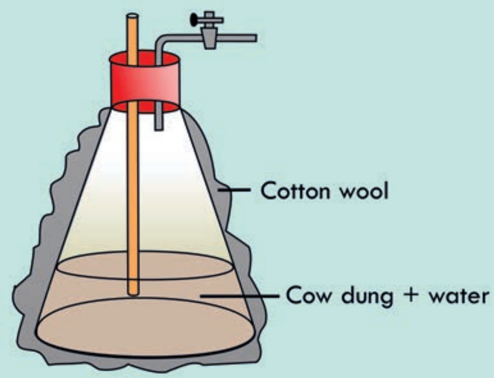


Figure 6.3 Preparation of methane gas from cow dung.

Observations and analysis:

a. What is the importance of covering the conical flask with cotton wool and placing it near the window?
b. What is your observation when you bring a burning splint close to the outlet of the tube?
c. In your opinion, what change did occur in the conical flask that led to the formation of methane?
d. Write a laboratory report in a group and present it to the class.

Project 6.1

1. In consultation with your classmates or friends, identify a place (Kebele or neighboring Kebeles) where a biogas plant is at work..
2. Prepare data collection instrument (e.g. questionnaire)
3. Make a trip to visit the biogas plant and collect the information you require
4. Describe the method involved and depict it graphically
5. Develop a similar set-up and produce your own biogas. Collect it using balloons and bring it to class. Take photos of your set-up and videos while you are constructing and doing, and show it to your teacher.
6. With your sample, in balloon, conduct tests to answer the following questions:
   a. What is the colour of the gas?
   b. Is the gas combustible? Write a balanced equation for the change.
   c. Is there any change when bromine water is added?

6.2.6 Chemical Properties of Alkanes

At the end of this section, students will be able to

☞ describe the general methods for preparation of alkanes in a laboratory;
☞ synthesize methane in a laboratory by decarboxylation method;
☞ carryout a project work to produce biogas from cow dung.

Activity 6.9

1. You have learnt in grade 9 that atoms achieve their valency by forming bonds to other atoms. Now, you have just learnt that alkanes are saturated compounds. Do you think they have the affinity to react with other atoms or molecules?
2. Based on your experiences from the startup activity of this unit and unit 4, can you suggest the most important reaction of alkanes?

Alkanes are sometimes referred to as paraffins, a word derived from the Latin parum affinis, meaning 'little affinity.' This term aptly describes their behavior, for alkanes show a little chemical affinity for other substances and are chemically inert to most laboratory reagents like acids, bases, oxidizing and reducing agents. They are also relatively inert biologically and are not often involved in the chemistry of living organisms.

Alkanes do, however, react with oxygen, halogens, and a few other substances under appropriate conditions.

1. Combustion Reaction:

Combustion is the most important reaction of alkanes. Reaction with oxygen occurs during combustion in an engine or furnace when the alkane is used as a fuel. Carbon dioxide and water are formed as products, and a large amount of heat is released. The general equation for the combustion reaction of alkanes is:

$$C_nH_{2n+2} + \left[\frac{3n+1}{2}\right]O_2 \rightarrow nCO_2 + (n+1)H_2O$$

where n is number of C atoms

For example, methane (natural gas) reacts with oxygen according to the equation:

Example 6.7: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + 890\ \text{kJ/mol}\ (213\ \text{kcal/mol})$

$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$

The heat energy liberated is used to generate electricity, to move motor vehicles or to cook our food.

2. **Substitution Reaction:** This is a reaction that involves the replacement of one atom or a group of atoms by another atom or group of atoms. Halogenation of alkanes is a very good example of substitution reaction. This reaction involves reacting alkanes with chlorine and bromine in presence of heat or sunlight. This type of reaction is called a photochemical reaction. The reaction of alkanes with chlorine and bromine proceeds in a sequence of steps. For example, let us consider the photochemical reaction of methane with chlorine (chlorination):

i. **Chain Initiating Step:** This step involves the absorption of energy by the halogen molecule to generate reactive particles known as free radicals. A free radical is an atom or a group of atoms possessing an unpaired electron. A free radical is electrically neutral. A chlorine molecule absorbs light and decomposes into two chlorine atoms:

$$Cl\text{-}Cl \xrightarrow{\text{light}} 2Cl\bullet$$

ii. **Chain Propagating Step:** This is a step involves reaction of the free radical from step i with the alkane and that consumes a reactive particle (free radical) produced in the chain initiation step and generates another free radical.

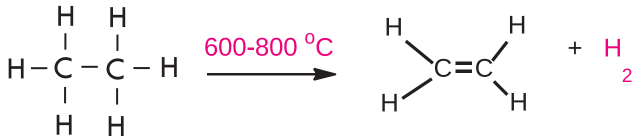
$$Cl\bullet + CH_4 \xrightarrow{\text{light}} HCl + CH_3^{\bullet}$$

iii. **Chain Terminating Step:** In the chain-terminating step reactive particles (free radicals) are consumed but not generated.

$$CH_3^{\bullet} + Cl\bullet \xrightarrow{\text{light}} CH_3Cl$$

Such a detailed step by step description of a chemical reaction is called a reaction mechanism.

3. **Elimination reaction:** This type of reaction involves the removal of smaller molecules from a compound and leads to the formation of a compound containing a multiple bond (unsaturated compound). It can also involve the removal of hydrogen atoms from adjacent carbon atoms at relatively high temperatures.



Ethane

Ethene

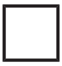
Exercise 6.6

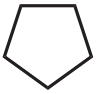
1. Alkanes are sometimes referred to as paraffins. Explain
2. Write the balanced equation for the combustion reaction of hexane
3. Write the major products of reaction of:
   a. methane with bromine in the presence of light.
   b. ethane is heated at 700 °C
4. Explain why alkanes do not undergo addition reaction

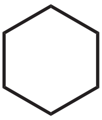
6.2.7 Cycloalkanes (Alicyclic Hydrocarbons)

Cycloalkanes are alkanes in which all or some of the carbon atoms are arranged in a ring. Alkanes have the general formula C_nH_{2n+2}; cycloalkanes containing a single ring have two fewer hydrogen atoms than alkanes and thus have the general formula C_nH_{2n}, where $n \geq 3$. Since cycloalkanes are alkanes themselves they have the same properties as the alkanes we have described in the preceding sections. They are also called alicyclic hydrocarbons ("ali" stands for aliphatic). Examples include cyclopropane, cyclobutane, cyclopentane, cyclohexane, etc. Note that the simplest alicyclic compound is cyclopropane. You will learn more about this in upper grades.


cyclopropane


cyclobutane


cyclopentane


cyclohexane

6.2.8 Uses of Alkanes

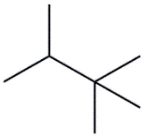
Activity 6.10

Have a look at the icon to the left? What comes to your mind is the gasoline or diesel supplied at fuel stations. You will soon learn that the liquid you see is principally a mixture of alkanes. You have also learnt that they are nonpolar molecules that have very little affinity to react with other substances. Based on this information, discuss on the most important uses of alkanes and present your conclusions to the class.

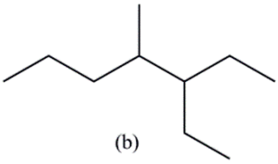
Alkanes are primarily used as fuels. They are also used as solvents and raw materials for making alkenes, alcohols, soaps, detergents and plastics.

Exercise 6.7

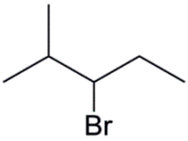
1. Define each of the following terms:
   - a. Homologous series
   - b. substitution reaction
   - c. Alkyl radical
   - d. combustion reaction
   - e. Isomerism
   - f. functional groups
2. Write the balanced chemical equation for the combustion of octane
3. How many chain isomers are there for an alkane that contains 7 carbon atoms?
4. Write IUPAC names for the alkanes represented by the following bond-line structures.



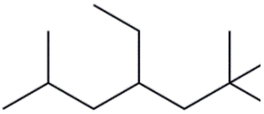
(a)



(b)



(c)



(d)

6.3 Unsaturated Hydrocarbons: Alkenes, Alkynes, and Aromatic Hydrocarbons

At the end of this section, students will be able to

- define unsaturated hydrocarbons;
- define alkenes;
- write the general formula of alkenes;
- write the molecular formula first nine homologous series of alkenes.

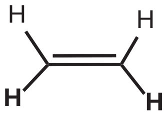
Alkenes, alkynes, and arenes (aromatic compounds) all contain carbon–carbon multiple bonds and are unsaturated hydrocarbons.

6.3.1 Alkenes or Olefins

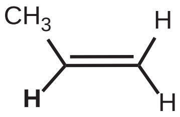
Alkenes are unsaturated hydrocarbons that contain at least one carbon–carbon double bond and have an -ene ending in their names. They are also known as olefins. They form a homologous series represented by the general formula C_nH_{2n}, where n = 2, 3, 4, …

A. Homologous series of Alkenes

Ethene is the simplest alkene. Ethene and propene are among the most important industrial Chemicals. Ethene is used as a starting material for the synthesis of many industrial compounds, including ethanol, ethylene oxide, ethanal, and the polymer polyethylene. Propene is used in making the polymer polypropylene, and, in addition to other uses, propene is the starting material for the synthesis of acetone and cumene. Ethene also occurs in nature as a plant hormone. It is produced naturally by fruits such as tomatoes and bananas and is involved in the ripening process of these fruits.



Ethene



Propene

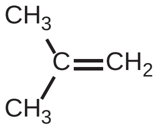
Table 6.4 The homologous series of alkenes and their physical constants.

IUPAC Name	Molecular formula	Condensed Formula	M.P	B.P	Density
Ethene	C_2H_4	$CH_2=CH_2$	-169	-102	0.61g/L
Propene	C_3H_6	$CH_2=CH_2CH_3$	-185	-48	0.61 g/L
1-butene	C_4H_8	$CH_2=CH_2CH_2CH_3$	-130	-6.5	0.63 g/L
1-pentene	C_5H_{10}	$CH2=CH_2(CH_2)_2CH_3$	-130.5	30	0.64 g/L
1-hexene	C_6H_{12}	$CH_2=CH_2(CH_2)_3CH_3$	-138	63	0.67 g/L
1-heptene	C_7H_{14}	$CH_2=CH_2(CH_2)_4CH_3$	-119	93	0.69 g/L
1-octene	C_8H_{16}	$CH_2=CH_2(CH_2)_5CH_3$	-104	122	0.72 g/L
1-nonene	C_9H_{18}	$CH_2=CH_2(CH_2)_6CH_3$	-95	146	0.73 g/L
1-decene	$C_{10}H_{20}$	$CH_2=CH_2\ (CH_2)_7CH_3$	-87	171	0.74 g/L

Exercise 6.8

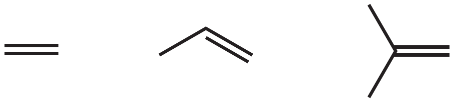
1. Define alkene and explain how it differs from alkanes
2. Describe the general characteristics of the homologous series of alkenes
3. The molecular formulae for the first three members of an alkene series are C_2H_4, C_3H_6 and C_4H_8. Derive the general formula for alkenes and compare it with that of alkanes. Is there a difference in the number of hydrogen atoms?
4. By referring to ***Table 6.4***, what relationship do you observe between the number of carbon atoms and the melting point, boiling point and density? Discuss with your group and present it to the class.
5. Write the molecular formula of the first 10 members of the alkene homologous series
6. Draw their condensed structures and bond line representations following the examples provided hereunder.

Condensed structures: $CH_2{=}CH_2$ $CH_3{-}CH{=}CH_2$



$$(CH_3)_2C{=}CH_2$$

Line-bond formula:

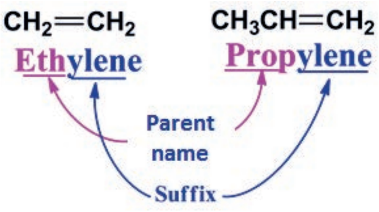


IUPAC name:	Ethene	Propene	2-Methylpropene
Common name:	Ethylene	Propylene	Isobutylene

B. Nomenclature of Alkenes

At the end of this section, students will be able to use IUPAC rules to name straight and branched chain alkenes.

Alkenes can have common names as well as IUPAC names. The common names of alkenes are obtained by replacing the – ane of the corresponding alkane with -ylene' as illustrated below:



The IUPAC rules for naming alkenes are similar in many respects to those for naming alkanes:

1. Determine the parent name by selecting the longest chain that contains the double bond and change the ending of the name of the alkane of identical length from –ane to -ene. Thus, if the longest chain contains five carbon atoms, the parent name for the alkene is pentene; if it contains six carbon atoms, the parent name is hexene, and so on.

2. Number the chain so as to include both carbon atoms of the double bond, and begin numbering at the end of the chain nearer the double bond. Designate the location of the double bond by using the number of the first atom of the double bond as a locant. The locant for the alkene suffix may precede the parent name or be placed immediately before the suffix.

Example 6.8: $\overset{1}{CH_2}=\overset{2}{CH}\overset{3}{CH_2}\overset{4}{CH_3}$

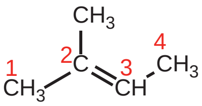
1-Butene
(not 3-butene)

$CH_3CH=CHCH_2CH_2CH_3$

2-Hexene
(not 4-hexene)

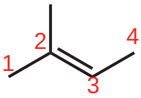
3. Indicate the locations of the substituent groups by the numbers of the carbon atoms to which they are attached:

Structural



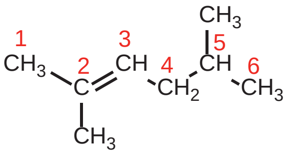
2-methyl-2-butene

Bond-line



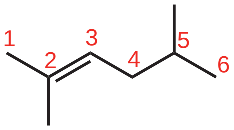
2-methyl-2-butene

Structural

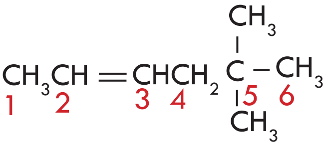


2,5-dimethyl-2-hexene

Bond-line



2,5-dimethyl-2-hexene



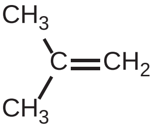
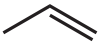
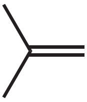
$$CH_3CH=CHCH_2\overset{CH_3}{\underset{CH_3}{C}}-CH_3$$

5,5-Dimethyl-2-hexene
or 5,5-dimethylhex-2-ene

$\overset{4}{CH_3}\overset{3}{CH}=\overset{2}{CH}\overset{1}{CH_2}Cl$

1-chloro-2-butene
or 1-chlorobut-2-ene

Many older names for alkenes are still in common use. Ethene is often called ethylene, Propene is often called propylene, and 2-methylpropene is often called isobutylene.

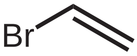
Condensed structures:	$CH_2{=}CH_2$	$CH_3-CH{=}CH_2$	 $(CH_3)_2C{=}CH_2$
Line-bond formula:			
IUPAC name:	Ethene	Propene	2-Methylpropene
Common name:	Ethylene	Propylene	Isobutylene

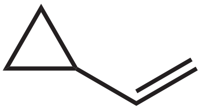
4. Two frequently encountered alkenyl groups are the vinyl group and the allyl group:

$CH_2{=}CH-$
vinyl

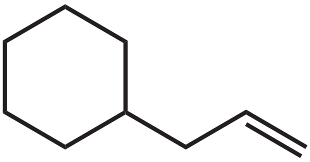
$CH_2{=}CH-CH_2-$
allyl

Using substitutive nomenclature, the vinyl and allyl groups are called ethenyl and prop-2-en-1-yl, respectively. The following examples illustrate how these names are employed:


Bromoethene
or
vinyl bromide
(common)

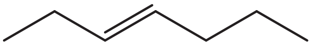

Ethenylcyclopropane
or
vinylcyclopropane

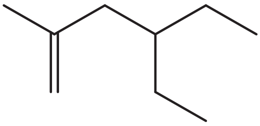

3-chloropropene
or
allyl chloride
(common)

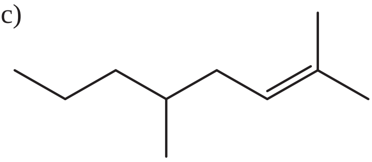

3-cyclohexylpropene
or
allylcyclohexane

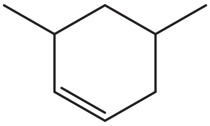
Exercise 6.9

Give IUPAC names for the following alkenes:

(a)


(b)


(c)


(d)


C. Physical Properties of Alkenes

At the end of this section, students will be able to use IUPAC rules to name straight and branched chain alkenes

Activity 6.11

Discuss on the following questions in a group and present your conclusions to the class.

1. Are alkenes polar or nonpolar molecules?
2. Do you think the physical properties of alkenes are different from alkanes?
3. Are alkenes soluble in water? What about in carbon tetrachloride?
4. Can you explain the trends in boiling points, melting points, and density of alkenes as molecular weight increases?
5. Which alkenes are gases, which liquids, and which ones are solids?

D. Isomerism in Alkenes

At the end of this section, students will be able to

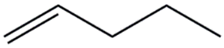
- write possible structural isomers for C_4H_8 and C_5H_{10}
- define geometric (cis-trans) isomerism;
- give examples of molecules that show geometric isomerism;
- construct models that show cis-trans isomerism.

Alkenes containing two and three carbon atoms have only one possible structure. Alkenes containing four or more carbon atoms exhibit isomerism. In the preceding section, you learnt that alkanes exhibit chain isomerism. In addition to chain isomerism, however, alkenes show two more types of isomerism:

- Chain isomerism: This is due to the difference in the arrangement of the carbon atoms in the longest continuous carbon chain.

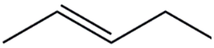
  For example, in the two compounds 1-pentene and 2-methyl-1-butene, the position of the double bond is the same. However, in 1-pentene the carbon chain is linear whereas in 2-methyl-1-butene it is branched. Thus, 1-pentene and 2-methyl-1-butene are chain isomers (See structures below).
- Position isomerism: This is due to the difference in the position of the double bond in the carbon chain.

a



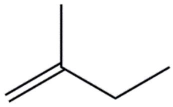
1-pentene

c



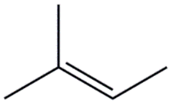
2-pentene

b



2-methyl-1-butene

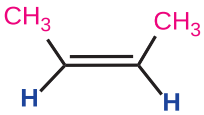
d



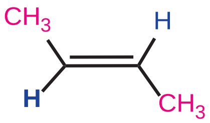
2-methyl-2-butene

Compounds (a) and (c) represent the position isomers of pentene, (b) and (d) are position isomers of methylbutene, (a), and (b) or (a) and (d) are the chain isomers, while all these compounds are the isomers of C_5H_{12}.

- **Geometrical isomerism.** This results from the difference in the relative spatial arrangement of atoms or groups about the double bond. This isomerism exists because free rotation about the double bond is not possible. To differentiate geometrical isomers, we use the prefix 'cis' if two similar groups are on the same side of the double bond and 'trans' when the two similar groups are on opposite sides of it; 'cis' means the same and 'trans' means across. The geometrical isomers of 2-butene are:



cis-2-butene or cis-but-2-ene



trans-2-butene or trans-but-2-ene

Note: Not all alkenes have geometrical (cis-trans) isomers; alkenes cannot have geometrical isomers if two identical groups are attached to any one of the carbon atoms linked by a double bond. E.g. 1-pentene cannot exhibit cis-trans isomerism.

Exercise 6.10

1. Explain the type of isomerism (chain or position) in:
   a. 2-pentene and 2-methyl-2-butene
   b. 1-hexene and 2-methyl-1-pentene
2. Explain why the boiling points and other physical properties of alkenes increase with increasing molecular weight
3. Write three position isomers of hexene.

4. Write structural formulas for the following:
   a. cis-3-Octene
   b. 1-Bromo-2-methyl-1-(prop-2-en-1-yl) cyclopentane
   c. trans-2-Hexene
   d. 3,4-Dimethylcyclopentene
   e. 2,4-Dimethyl-2-pentene
   f. Vinylcyclopentane
   g. trans-1-Chlorobut-2-ene
   h. 1,2-Dichlorocyclohexene
   i. 4,5-Dibromo-1-pentene
   j. trans-1,4-Dichloro-2-pentene

Project 6.2

Prepare a model from locally available materials, to show the cis-trans isomers of 2-butene.

E. Preparation of Alkenes

At the end of this section, students will be able to

☞ explain the general method for preparation of alkenes in a laboratory;

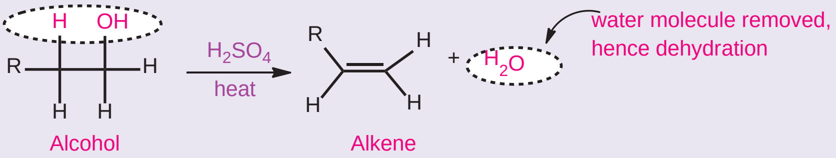
☞ produce ethylene in a laboratory by dehydration of ethanol.

Alkenes are mainly obtained during fractional distillation of petroleum when the process called cracking is carried out.

In the laboratory, alkenes can be prepared by:

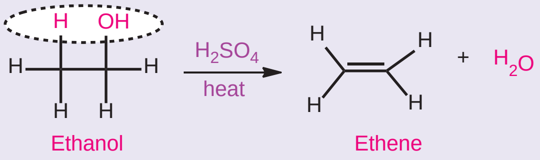
☞ Dehydration of alcohols with concentrated sulphuric acid or alumina (Al_2O_3). Dehydration means removal of water.

General Reaction:



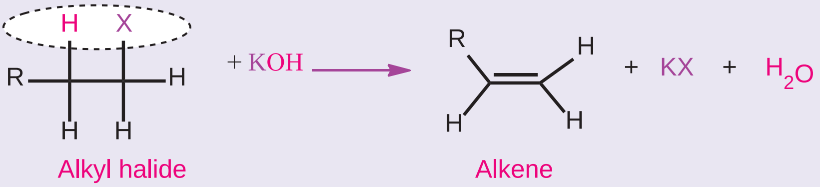
water molecule removed, hence dehydration

Example 6.9:

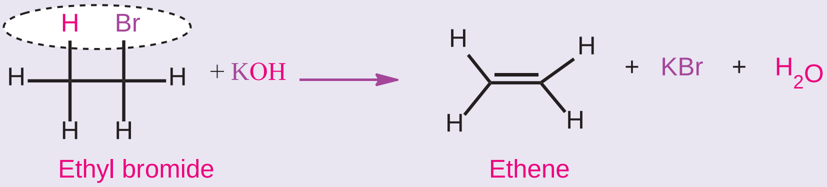


☞ Dehydrohalogenation of alkyl halides with a base (KOH). Dehydrohalogenation means the removal of hydrogen and a halogen atom. Alkyl halides are compounds of an alkyl group and a halogen such as CH_3Cl and CH_3-CH_2-Br.

General Reaction:



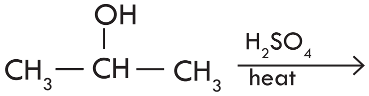
Example 6.10:



Exercise 6.11

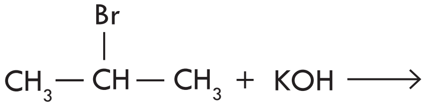
1. In the above two methods of preparations of alkenes, (a) what difference do you observe between the structures of the major reactant and the product?
2. Write the complete and balanced chemical equations for each of the following reactions:

a)



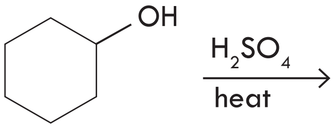
$CH_3-\overset{\overset{OH}{|}}{CH}-CH_3 \xrightarrow[heat]{H_2SO_4}$

b)



$CH_3-\overset{\overset{Br}{|}}{CH}-CH_3 + KOH \longrightarrow$

c)



Experiment 6.3

Laboratory Preparation of Ethene

Objective: To prepare ethene and study some of its properties

Materials required: Ethanol, concentrated H_2SO_4, water, bromine water, basic $KMnO_4$. Round-bottomed flask, stopper, delivery tube, gas jar, pneumatic trough, thermometer, tripod, Bunsen burner, stand and clamp, beehive shelf and wire gauze.

Procedure:

1. Arrange the set-up as shown in ***Figure 6.4***
2. Take ethanol in the flask and add concentrated H_2SO_4 through the thistle funnel; heat carefully until the temperature reaches about 170°C. Discard the initially formed gas as it might contain air; collect the ethene gas in three gas jars by downward displacement of water; perform the following activities.

a. Insert a lighted splint into the jar containing the gas and see what happens.
b. Add two drops of bromine water to the other gas jar, cover the jar and shake well.
c. Add a few drops of alkaline $KMnO_4$ solution to another jar containing the gas and observe what happens.

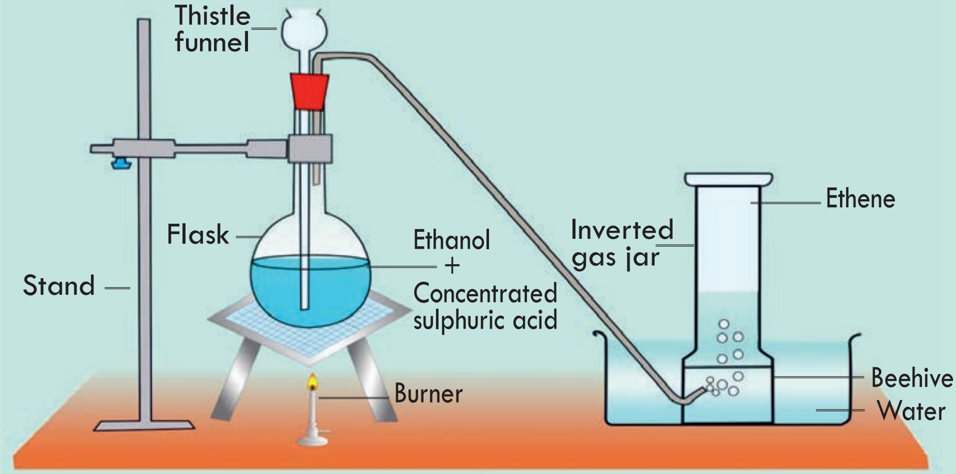


Figure 6.4 Preparation of Ethene.

Observations and analysis:

a. What is the colour and odour of the gas?
b. What is the substance left after combustion of the gas in the jar?
c. What change is observed when the gas is shaken with bromine water or alkaline potassium permanganate? Write equations for the reactions?

d. Name the dehydrating agent used in the experiment?
e. Which method of preparation of an alkene is used in this experiment?

Write a laboratory report about the experiment and submit to your teacher. Write a laboratory report in a group and present it to the class.

F. Chemical Properties of Alkenes

At the end of this section, students will be able to explain chemical properties of alkenes.

Activity 6.12

Do you think alkenes are more or less reactive than alkanes? Explain
Can alkenes undergo combustion reaction? What do you think is the most important reaction of alkenes?

Alkenes are unsaturated hydrocarbons containing double bond (a strong bond and a weak bond) between the two carbon atoms. They are more reactive than alkanes because the weak bond can be used for further reaction. They undergo several types of reactions.

☞ Combustion reaction. Alkenes burn in oxygen with a luminous flame to form carbon dioxide and water.

General reaction:

$$C_nH_{2n} + (3n/2)O_2 \rightarrow nCO_2 + nH_2O + Heat$$

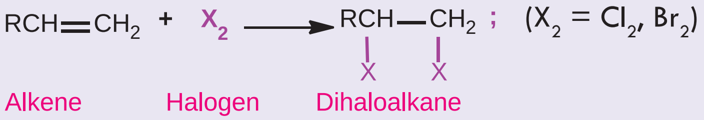
Example 6.11:

$$C_4H_8 + 6O_2 \rightarrow 4CO_2 + 4H_2O + Heat$$

☞ Addition reaction. Alkenes undergo mainly addition reactions. The addition occurs at the carbon-carbon double bond.

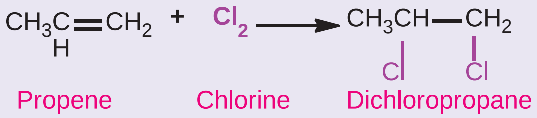
i. **Addition of halogens (halogenation)**: When a halogen molecule, X_2, (where, X_2 = Cl_2 or Br_2) is added to alkenes, the product is a dihaloalkane.

General Reaction:



$$RCH{=}CH_2 + X_2 \longrightarrow \underset{X}{\underset{|}{RCH}}-\underset{X}{\underset{|}{CH_2}} \;;\; (X_2 = Cl_2, Br_2)$$

Example 6.12:

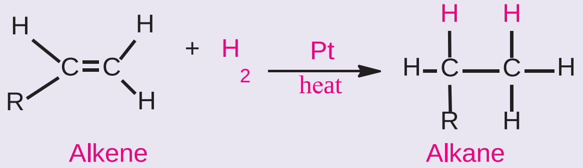


$$CH_3\underset{H}{C}{=}CH_2 + Cl_2 \longrightarrow CH_3\underset{Cl}{\underset{|}{CH}}-\underset{Cl}{\underset{|}{CH_2}}$$

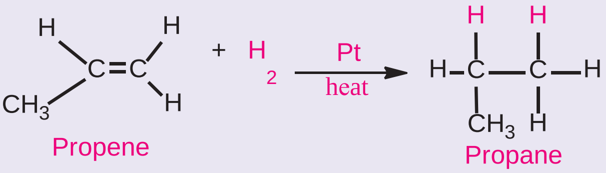
If bromine water (bromine in CCl_4) is added to alkenes, the reddish-brown colour of Br_2 in CCl_4 will disappear. This is due to the addition of bromine (Br_2) across the double bond. So, Br_2 in CCl_4 is used to detect unsaturation (presence of multiple bonds) in a compound.

ii. **Addition of hydrogen (Hydrogenation)**: The addition of hydrogen molecule to alkenes in the presence of a nickel or platinum catalyst yields alkanes.

General Reaction:



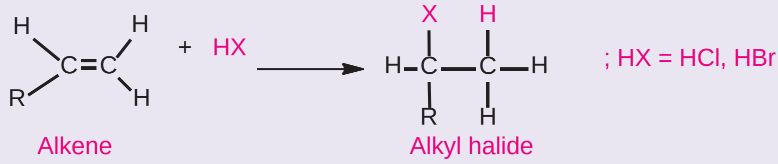
Example 6.13:



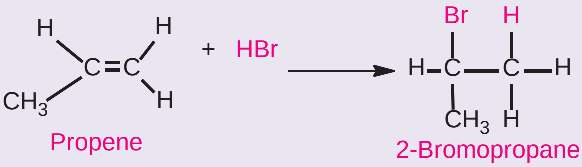
iii. **Addition of hydrogen halides, HX (hydrohalogenation)**: The addition of hydrogen halide (HCl, HBr or HI) to alkenes leads to the formation of alkyl halides. The product of the reaction can be predicted by Markovnikov's rule; which states that "when an alkene reacts with a hydrogen halide to give an alkyl halide, the hydrogen adds to the carbon atom of the double bond that has the greater number of hydrogen atoms, and the halogen to the carbon that has the fewer number of hydrogen atoms". "The Rich Get Richer!!"

The same is true when an alkene reacts with water in an addition reaction to form an alcohol.

General Reaction:



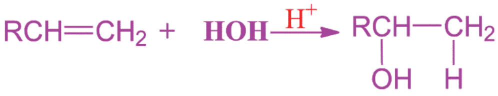
Example 6.14:



Have you noticed the location where H and Br have formed bonds?

iv. Addition of water (hydration): When alkenes are hydrated in the presence of an acid catalyst, they produce alcohols following Markovnikov's rule.

General reaction:



$$RCH{=}CH_2 + HOH \xrightarrow{H^+} \underset{OH}{RCH}{-}\underset{H}{CH_2}$$

Alkene Alcohol

Example 6.15:

$$CH_2{=}CH_2 \xrightarrow{H_3PO_{4(aq)}} CH_3{-}CH_2OH$$

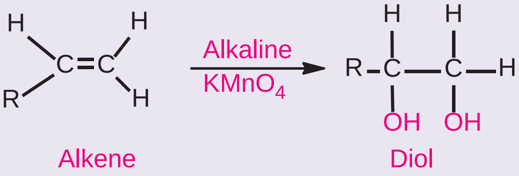
Ethene Ethanol

$$CH_2{=}CHCH_3 \xrightarrow{H_3PO_{4(aq)}} CH_3{-}CH(OH)CH_3$$

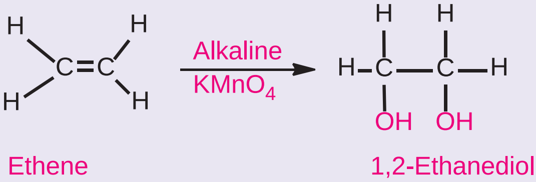
Propene 2-propanol

Oxidation of alkenes: Oxidation of alkenes with cold alkaline potassium permanganate solution ($KMnO_4$) forms diols (compounds containing two hydroxyl groups).

General Reaction:



Example 6.16:



Alkaline $KMnO_4$ solution is also used as a qualitative test for the identification of unsaturation in a compound. The solution is called Baeyer's reagent. In the presence of unsaturated hydrocarbons, the purple colour of alkaline $KMnO_4$ solution fades and a brown precipitate is formed.

v. **Polymerization (self-addition) of alkenes:** Polymerization is the union of small molecules called monomers to form a large molecule called a polymer.

Example 6.17:

$$CH_2{=}CH_2 \xrightarrow[O_2\text{ (catalyst)}]{1000\text{ atm, }473\text{ K}} (-CH_2-CH_2-)n$$

Ethene Polyethene

Exercise 6.12

1. Write the balanced chemical equation for the combustion reaction of hexene.
2. Write the structures of the major products when $CH_3CH{=}CH\text{-}CH_3$ reacts with:
(a) Br_2 b) H_2 c) HCl d) H_2O (in presence of dilute H_2SO_4) e) alkaline $KMnO_4$.
3. Alkenes are more reactive than alkanes. Why do you think?
4. Alkene reacts readily with Br_2 even in the dark but alkanes react only in the presence of sunlight. Why? Are the reaction products the same? What about the type of reaction? Discuss!

G. Uses of Ethene (Ethylene)

At the end of this section, students will be able to explain the uses of ethylene.

Ethene is used as a starting material for the synthesis of many industrial compounds, including ethanol, ethylene oxide, ethanal, and the polymers polyethylene and polyvinylchloride, polystyrene and teflon. It is also used in the production of 1,2-ethanediol, which is used as an antifreeze material; and for the preservation as well as artificial ripening of fruits.

Project 6.3

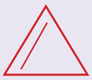
Put some green tomatoes in the first basket, and place a ripe banana in between them. In the second basket, put only green tomatoes (but not a banana). Compare the tomatoes in the two baskets. Have the tomatoes ripen at the same rate in both baskets? If not, why? Share your opinion with your classmates.

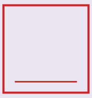
H. Cycloalkenes

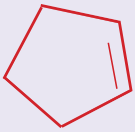
Cycloalkenes are unsaturated cyclic hydrocarbons. The carbon atoms are linked in such a manner as to form a closed chain or a ring structure. They contain a double bond between carbon atoms in the ring and are represented by the general formula C_nH_{2n-2}; where n should be equal to or greater than 3. They are isomeric with alkynes. They are named by prefixing "cyclo" to the name of alkenes containing the same number of carbon atoms. The given structures show some examples of cycloalkenes.

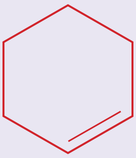
Unsubstituted cycloalkenes are named by adding the prefix "cyclo" to the name of the parent alkene. The simplest cycloalkene is cyclopropene.

Example 6.18:

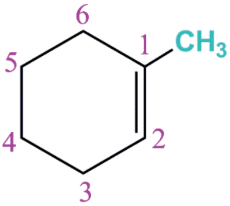

cyclopropene


cyclobutene

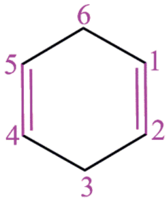

cyclopentene


cyclohexene

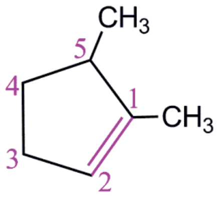
Substituted cycloalkenes are named similarly as alkenes, but because there is no chain end to begin from, we number the cycloalkene so that the double bond is between C_1 and C_2 and the first substituent has as low a number as possible. It's not necessary to indicate the position of the double bond in the name because it's always between C_1 and C_2. As with open-chain alkenes, newer but not yet widely accepted naming rules place the locant immediately before the suffix in a diene.



1-Methylcyclohexene



1,4-cyclohexadiene
(or cyclohexa-1,4-diene)



1, 5-dimethylcyclopentene

Exercise 6.13

Complete each of the following equations and name the products:

a. $CH_3CH_2CH=CH_2 + HCl \rightarrow$

b. $CH_3CH_2CH=CH_2 + Br_2 \rightarrow$

c. $CH_3CH_2CH=CH_2 + H_2O \xrightarrow{H^+}$

d. $CH_3CH_2CH=CH_2 \xrightarrow[KMnO_4]{Alkaline}$

6.3.2 Alkynes and Their Physical Properties

At the end of this section, students will be able to

- define alkynes;
- write the general formula of alkynes;
- write the molecular formula first nine homologous series of alkynes;
- write the molecular formulas of alkynes from the given number of carbon atoms;
- describe the physical properties of alkynes.

Alkynes are unsaturated hydrocarbons that contain at least one carbon–carbon triple bond as a functional group, and have a -yne ending in their names. The simplest alkyne is ethyne (also called acetylene). Alkynes occur in nature and can be synthesized in the laboratory.

The homologous series of alkynes is represented by the general formula C_nH_{2n-2}, where $n \geq 2$. The first member of the group is commonly known as acetylene. The homologous series of alkynes is also called acetylene series.

Table 6.5 Homologous series of alkynes and their physical constants.

Formula of Alkyne	Condensed Structure	IUPAC Name	Melting Point (°C)	Boiling Point (°C)
C_2H_2	$CH\equiv CH$	Ethyne	-82	-75
C_3H_4	$CH\equiv C-CH_3$	Propyne	-105.5	-23
C_4H_6	$CH\equiv C-CH_2-CH_3$	1-butyne	-122	9
C_5H_8	$CH\equiv C-(CH_2)_2-CH_3$	1-pentyne	-98	40
C_6H_{10}	$CH\equiv C-(CH_2)_3-CH_3$	1-hexyne	-124	72
C_7H_{12}	$CH\equiv C-(CH_2)_4-CH_3$	1-heptyne	-80	100
C_8H_{14}	$CH\equiv C-(CH_2)_5-CH_3$	1-octyne	-70	126
C_9H_{16}	$CH\equiv C-(CH_2)_6-CH_3$	1-nonyne	-65	151
$C_{10}H_{18}$	$CH\equiv C-(CH_2)_7-CH_3$	1-decyne	-36	182

Exercise 6.14

1. By looking at the trends for alkynes in Table 6.5, write the molecular formula and the structures of the alkynes containing 11 and 12 carbon atoms.
2. Describe the general characteristics of the alkyne homologous series
3. Compare the physical state, melting points and boiling points of alkynes, alkenes and alkanes as the number of carbon atoms increases. Discuss with your group and present to the class.

A. Nomenclature of Alkynes

At the end of this section, students will be able to

- use IUPAC rules to name straight and branched chain alkynes;
- write the structural formulas alkynes up to nine carbon atoms;
- alkynes are commonly named as a derivative of acetylene.

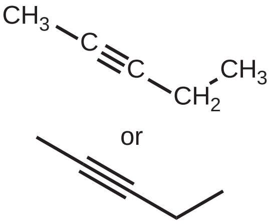
$CH\equiv CH$	$CH\equiv C-CH_3$	$CH\equiv C-CH_2CH_3$
acetylene	methyl acetylene	ethyl acetylene

Substituted Alkynes are named in much the same way as alkenes using the IUPAC system. Unbranched alkynes, for example, are named by replacing the -ane of the name of the corresponding alkane with the ending -yne. The chain is numbered to give the carbon atoms of the triple bond the lower possible numbers.

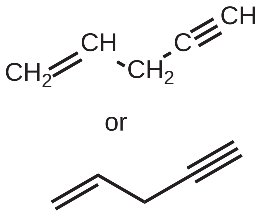
The lower number of the two carbon atoms of the triple bond is used to designate the location of the triple bond. When double and triple bonds are present, the direction of numbering is chosen so as to give the lowest overall set of locants. In the face of equivalent options, then preference is given to assigning the lowest numbers to the double bonds. The IUPAC names of three unbranched alkynes are shown here:

$H-C\equiv C-H$

Ethyne or
acetylene (common name)

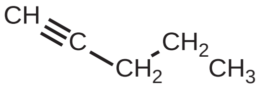


2-pentyne

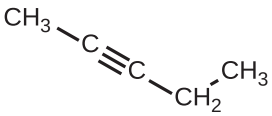


1-pentene-4-yne

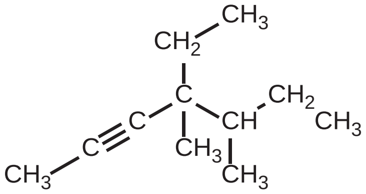
The locations of substituent groups of branched alkynes and substituted alkynes are also indicated with numbers.



1-pentyne



2-pentyne



4-ethyl-4,5-dimethyl-2-heptyne

Exercise 6.15

Give the structures and IUPAC names for all the alkynes with the formula C_6H_{10}, and identify those that are position and chain isomers.

B. Isomerism in Alkynes

At the end of this section, students will be able to write the structural formulas of alkynes up to nine carbon atoms.

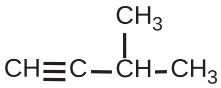
☞ Alkynes show both chain and position isomerism, but not geometrical or *cis-trans* isomerism Consider:

$CH\equiv C-CH_2CH_3$

1-butyne or but-1-yne

$CH_3-C\equiv C-CH_3$

2-butyne or but-2-yne



$CH\equiv C-CH(CH_3)-CH_3$

3-methyl-1-butyne or 3-methylbut-1-yne

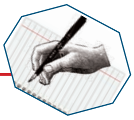
$CH\equiv C-CH_2CH_2CH_3$

1-pentyne or pent-1-yne

1-Butyne and 2-butyne are position isomers whereas 3-methyl-1-butyne and 1-pentyne are chain isomers.

Activity 6.13

Write the structural formula for pentane, 1-pentene, and 1-pentyne. Why are not these three hydrocarbons considered as isomers? Tell your idea to the class.



Exercise 6.16

1. Write the structures of all isomers of 1-pentyne (C_5H_8) and
(a) name them b) identify chain isomers? c) identify position isomers?

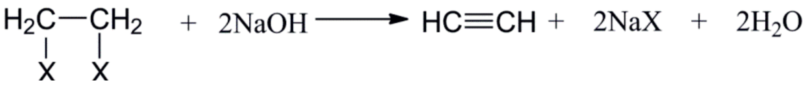
C. Preparation of Alkynes

At the end of this section, students will be able to

- describe the general method for preparation of alkynes in a laboratory;
- prepare acetylene in a laboratory by the reaction of CaC_2 with water.

Alkynes can be prepared by several methods. Some of the general methods of preparation of alkynes are:

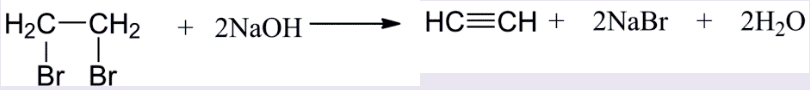
☞ Dehydrohalogenation of vicinal (adjacent) dihalides with a base NaOH or KOH or $NaNH_2$.



$$H_2C(X)-CH_2(X) + 2NaOH \longrightarrow HC\equiv CH + 2NaX + 2H_2O$$

Vicinal dihalide Ethyne

Example 6.19:



$$H_2C(Br)-CH_2(Br) + 2NaOH \longrightarrow HC\equiv CH + 2NaBr + 2H_2O$$

1,2-dibromoethane Ethyne

☞ Alkylation of sodium acetylide (dicarbide) with a primary alkyl halide.

General reaction:

$$HC{\equiv}\bar{C}Na^+ \quad + \quad R\text{-}X \longrightarrow HC{\equiv}C{-}R \quad + \quad NaX$$

Sodium carbide	Primary Alkyl halide	Alkyne

Example 6.20

$$HC{\equiv}\bar{C}\,Na^+ \; + \; CH_2CH_2Br \longrightarrow HC{\equiv}C{-}CH_2CH_3$$

Sodium carbide	Ethyl bromide	But-1-yne (1-butyne)

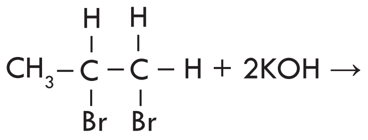
Reaction of calcium carbide with water:

$$CaC_2(s) \; + \; 2H_2O(l) \rightarrow Ca(OH)_2(aq) \; + \; C_2H_2(g)$$

Calcium Carbide	Water	Calcium hydroxide	Acetylene/ ethyne

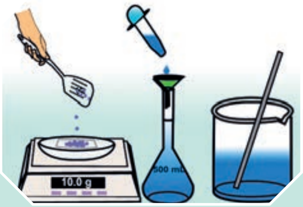
Exercise 6.17

Write the major product of the following reactions:

a.


$$CH_3{-}\underset{Br}{\overset{H}{C}}{-}\underset{Br}{\overset{H}{C}}{-}H + 2KOH \rightarrow$$

b. $H-C{\equiv}\bar{C}Na^+ + CH_3Br \rightarrow$



Experiment 6.4

Laboratory Preparation of Ethyne

Objective: To prepare ethyne from calcium carbide and water.

Materials required: CaC_2, H_2O, bromine water, round-bottomed flask, separating funnel, delivery tube, trough, beehive shelf and gas jar.

Procedure:

1. Arrange the set up as shown in *Figure 6.5*. Put a layer of sand in a conical flask and place calcium carbide on the sand.

2. Add water drop by drop from the separating funnel onto the calcium carbide. Watch carefully and record your observation. Touch the flask with the tip of your finger.
3. Collect several jars of the gas over water and carry out the following tests:

a. Burn the gas as you did with methane and ethene.
b. Add a few drops of bromine water to another jar filled with ethyne and note the changes.

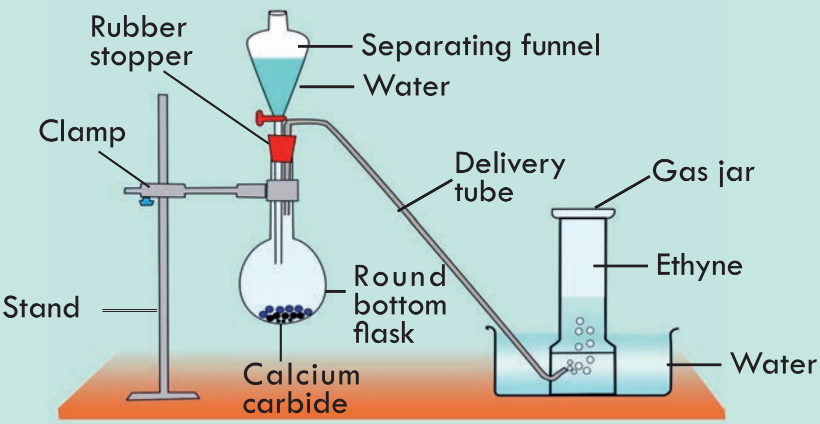


Figure 6.5 Laboratory preparation of Ethyne.

Observations and analysis:

a. What do you feel when you touch the flask? Is the reaction exothermic or endothermic?
b. How do you compare the colour of the flame produced with that of methane and ethene?

Write a complete laboratory report and submit to your teacher. Write a laboratory report about the experiment and submit to your teacher.
Write a laboratory report in a group and present it to the class.

D. Chemical Properties of Alkynes

At the end of this section, students will be able to explain chemical properties of alkynes.

Alkynes are more unsaturated than alkanes and alkenes due to the presence of a carbon-carbon triple bond. They can undergo combustion and addition reactions.
Some of the common reactions of alkynes are:

i. **Combustion reaction:** Alkynes burn with a smoky luminous flame, forming CO_2 and water. Smoky luminous flames result from the combustion of alkynes due to their relatively high carbon content.

$$C_nH_{2n-2} + (\frac{3n-1}{2})O_2 \rightarrow nCO_2 + (n-1)H_2O + \text{Heat}$$

Example 6.21:

$$C_3H_4(g) + 4O_2 \rightarrow 3CO_2 + 2H_2O + \text{Heat}$$

Addition reaction: Alkynes undergo an addition reaction at the carbon-carbon triple bond. Some of the addition reactions of alkynes are the following: **(a) Addition of hydrogen (hydrogenation):** In the presence of nickel or palladium catalyst, alkynes produce alkanes.

General equation:

$$C_nH_{2n-2} + H_2 \xrightarrow[\text{heat}]{\text{Ni}} C_nH_{2n} \xrightarrow[\text{heat}]{H_2,\ \text{Ni}} C_nH_{2n+2}$$

Alkyne Alkene Alkane

Example 6.22:

$$CH{\equiv}CH + H_2 \xrightarrow[\text{heat}]{\text{Ni}} CH_2{=}CH_2 \xrightarrow[\text{heat}]{H_2,\ \text{Ni}} CH_3{-}CH_3$$

Ethyne Ethene Ethane

Partial hydrogenation of alkynes in the presence of Lindlar's catalyst gives alkenes. Lindlar's catalyst is powdered palladium partially deactivated with lead acetate that reduce its activity such that it can only reduce alkynes, not alkenes. It always gives the cis-alkene.

General reaction:

$$C_nH_{2n-2} + H_2 \xrightarrow[\text{Catalyst}]{\text{Lindlar's}} C_nH_{2n}$$

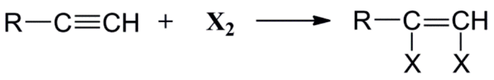
Alkyne Alkene

Example 6.23:

$$CH_3{-}C{\equiv}C{-}CH_3 + H_2 \xrightarrow[\text{Catalyst}]{\text{Lindlar's}} CH_3\text{-}CH{=}CH\text{-}CH_3$$

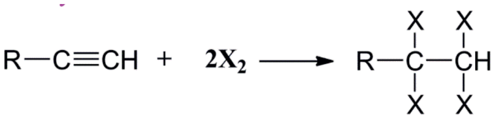
2-butyne *cis*-2-butene

(b) Addition of halogens: When a molecule of an alkyne reacts with one and two molecules of halogens, it gives a dihaloalkene and a tetrahaloalkane, respectively.



$$R{-}C{\equiv}CH + X_2 \longrightarrow R{-}C(X){=}CH(X)$$

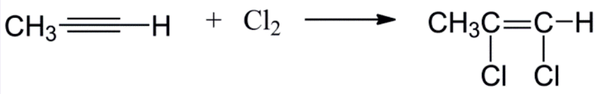
Alkyne Dihaloalkene



$$R{-}C{\equiv}CH + 2X_2 \longrightarrow R{-}CX_2{-}CHX_2$$

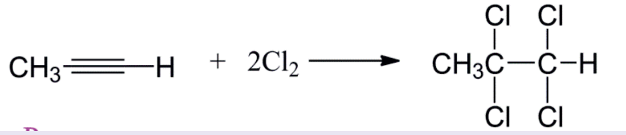
Alkyne Tetrahaloalkane

Example 6.24:



$$CH_3C\equiv C{-}H + Cl_2 \longrightarrow CH_3\underset{Cl}{C}{=}\underset{Cl}{C}{-}H$$

Propyne 1,2-dichloropropene



$$CH_3C\equiv C{-}H + 2Cl_2 \longrightarrow CH_3\overset{Cl}{\underset{Cl}{C}}{-}\overset{Cl}{\underset{Cl}{C}}{-}H$$

Propyne 1,1,2,2-tetrachloropropane

Activity 6.14

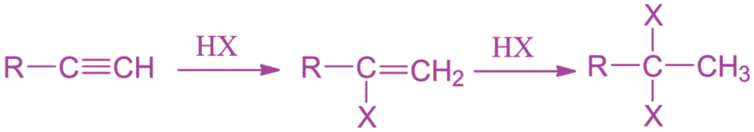
Discuss on the following questions in a group and present your responses to the class.

Do you expect alkynes to decolorize bromine water? If so, why?

Write the balanced chemical equation for the reaction between C_2H_2 and Br_2.

(c) Addition of Hydrogen Halides: Alkynes react with hydrogen halides to form a monohaloalkene and a dihaloalkane. The addition reaction occurs according to Markovnikov's rule.

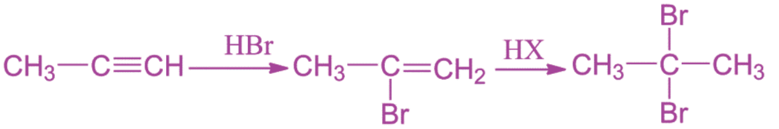
General reaction:



$$R{-}C\equiv CH \xrightarrow{HX} R{-}\underset{X}{C}{=}CH_2 \xrightarrow{HX} R{-}\overset{X}{\underset{X}{C}}{-}CH_3$$

Alkyne Monohaloalkene Dihaloalkane

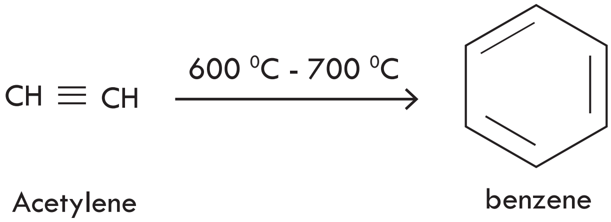
Example 6.25:



$$CH_3{-}C\equiv CH \xrightarrow{HBr} CH_3{-}\underset{Br}{C}{=}CH_2 \xrightarrow{HX} CH_3{-}\overset{Br}{\underset{Br}{C}}{-}CH_3$$

Propyne 2-Bromopropene 2,2-Dibromopropane

(d) Trimerization of Acetylene: Acetylene, on prolonged heating at 600 – 700 °C, yields benzene.



Activity 6.15

Discuss on the following questions in a group and present your responses to the class.

1. Is Markonikov's rule applied in the reaction between propyne and hydrobromic acid? How?
2. The addition reaction of alkynes proceeds in two steps. Explain what changes occur during the first and the second steps.
3. Compare the reaction products when 1-butyne and 2-butyne react with HBr.

E. Properties and uses of Acetylene or Ethyne

At the end of this section, students will be able to

- explain the uses of ethylene and acetylene;
- compare and contrast the properties of ethane, ethene and ethyne;
- test for unsaturation of ethylene and ethyne.

Acetylene is a colourless, sweet-smelling gas in pure form and is insoluble in water. It is usually stored as a solution of acetone in steel cylinders. Combustion of acetylene with oxygen produces an intensely hot flame of about 3000 °C. Thus, a large quantity of acetylene is used as a fuel in oxy-acetylene torches for cutting and welding metals. It is a two-stage reaction where the primary chemical reaction involves the acetylene disassociating in the presence of oxygen to produce heat, carbon monoxide, and hydrogen gas: $C_2H_2 + O_2 \rightarrow 2CO + H_2$

The second reaction is oxidation of carbon monoxide into carbon dioxide while the hydrogen formed in the above reaction forms water:

$CO + H_2 + O_2 \rightarrow CO_2 + H_2O$

Acetylene is also used to prepare acrylonitrile which is a starting material for producing polyacrylonitrile, a raw material for textile fiber. It is also used for making vinyl chloride which polymerizes to give polyvinyl chloride (PVC) commonly used for making floor tiles, electrical insulators, shoe soles, water pipes etc. Acetylene is used to produce 1,1,2,2-tetrachloroethane that serves as a solvent for wax, grease, rubber etc. See *Table 6.6*.

Table 6.6 Comparison of Some Properties of Ethane, Ethene and Ethyne

Property	Ethane	Ethene	Ethyne
Nature of flame	Non-luminous	Luminous	Smoky luminous
Effect on colour of Br_2 in CCl_4	No effect	Decolorizes	Decolorize
Effect on colour of alkaline $KMnO_4$	No effect	Decolorizes	Decolorizes

6.4 Aromatic Hydrocarbons: Benzene

At the end of this section, students will be able to

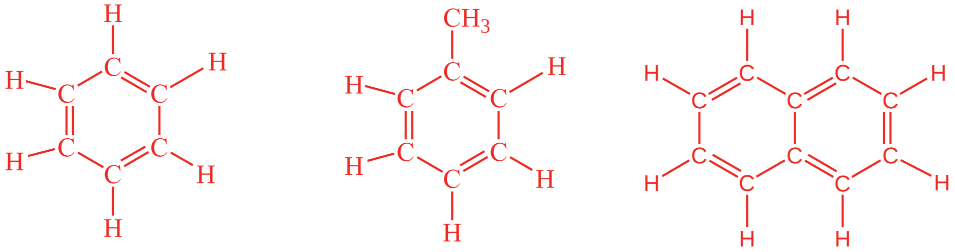
- define aromatic hydrocarbons;
- distinguish aromatic compounds from other hydrocarbons;
- draw and explain the structure of benzene;
- describe the main physical properties and chemical reactions of benzene;
- perform simple chemical tests to identify alkanes, alkenes/alkynes, and aromatic compounds.

Activity 6.16

Discuss the following in your group and present your responses to the whole class:

1. What comes to your mind when you hear the word 'benzene'?
2. What does the word "aroma" mean? Which substances have an aroma smell?
3. Are aromatic hydrocarbons saturated or unsaturated?
4. What are the main sources of aromatic compounds?

The term aromatic is derived from the word 'aroma' meaning pleasant smell, which was originally used to describe naturally occurring compounds with pleasant smells. At present, the term is used without its original significance. Aromatic hydrocarbons are generally obtained from petroleum and coal tar. They are a group of hydrocarbons characterized by the presence of a benzene ring or related structures. ***Figure 6.6*** gives some examples of aromatic hydrocarbon structures.

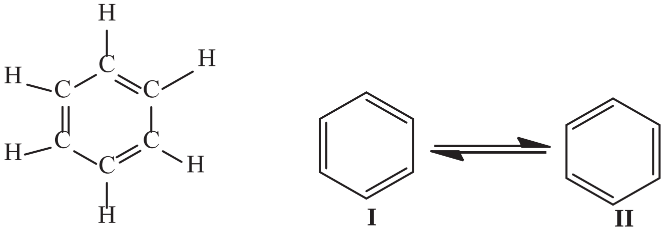


Benzene   Methyl benzene (Toluene)   Naphthalene

Figure 6.6 Structures of some Aromatic Hydrocarbons.

6.4.1 Benzene

Benzene is the simplest aromatic hydrocarbon. Its molecular formula is C_6H_6. The first structure for benzene was proposed by August Kekulé in 1872. It consisted of a six-membered ring with alternating single and double bonds and with one hydrogen bonded to each carbon. He assumed that the ring contains three double bonds that shift back and forth so rapidly that the two forms (***Figure 6.7***) cannot be separated. Later on, Kekulé assumption was found to be incorrect. Each structure has become known as Kekulé structure.

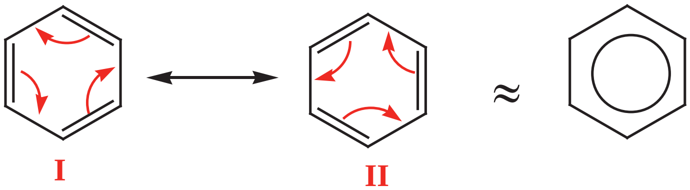


A Kekulé structure
showing all atoms

A Kekulé structure
as line-bond formulas

Figure 6.7 A Kekulé structure of benzene.

Today, we represent benzene as a hybrid of two equivalent contributing Kekulé structures. Each Kekulé structure makes an equal contribution to the hybrid:



Benzene as a hybrid of two equivalent
contributing structures

Resonance hybrid

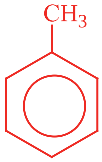
The bonds in benzene are neither single nor double, but have an intermediate character between those of single and double bonds. All the carbon-carbon bonds in the molecule are the same in length and nature. Since structures I or II given above are not the true structures of the benzene molecule, the benzene ring is written in its resonance hybrid form. However, chemists continue to use a single contributing structure to represent benzene.

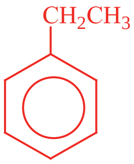
From structures I and II in ***Figure 6.7***, one may think that benzene has three double bonds and has the same chemistry as that of alkenes. But, this is not true. Benzene and other aromatic hydrocarbons contain high degrees of unsaturation, yet fail to undergo characteristic alkene addition and oxidation–reduction reactions. They are more stable than alkenes.

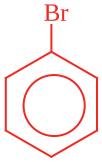
6.4.2 Nomenclature of Substituted Benzenes

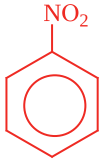
i. Monosubstituted Benzenes

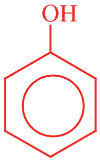
Generally, mono-substituted benzenes are named in a similar manner as hydrocarbons with -benzene as the parent name:

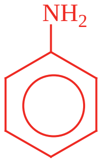

Methylbenzene
(Toluene)


Ethylbenzene


Bromobenzene


Nitrobenzene

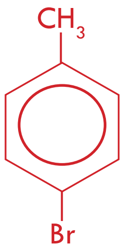

Hydroxybenzene
(Phenol)

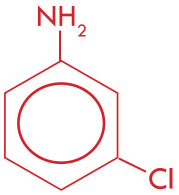

Aminobenzene
(Aniline)

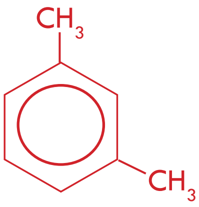
The IUPAC system retains certain common names for several of the simpler mono-substituted benzenes. Examples are toluene (rather than methylbenzene) and phenol (rather than hydroxybenzene).

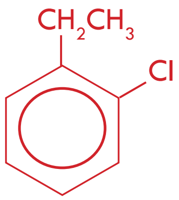
ii. Disubstituted Benzenes

When two substituents occur on a benzene ring, three constitutional isomers are possible. We locate substituents either by numbering the atoms of the ring or by using the locators 1,2- or *ortho (o)*, 1,3- or *meta (m)*, and 1,4- or *para (p)*.


4-Bromotoluene
(p-Bromotoluene)


3-chloroaniline
(m-chloroaniline)


1,3-dimethylbenzene
(m-xylene)


1-Chloro-2-ethylbenzene
(o-Choloroethylbenzene)

6.4.3 Physical Properties of Benzene

In general, benzene resembles other hydrocarbons in its physical properties. Benzene is a flammable, colourless, and volatile liquid with a characteristic smell. It is less dense than water, non-polar and insoluble in water, but soluble in non-polar solvents like ether and carbon tetrachloride. At one time, benzene was widely used as a solvent, but exposure to it causes leukemia, a blood cancer. That is benzene is a carcinogenic (cancer-causing) substance.

6.4.4 Chemical Properties of Benzene

Benzene and other aromatic hydrocarbons are more stable than alkenes and alkynes due to the stability of the aromatic ring. It neither decolorizes bromine water (Br_2 in CCl_4) nor is reactive towards cold potassium permanganate solution. However, benzene undergoes the following reactions:

1. **Combustion reaction:** Benzene is highly inflammable. It burns with a smoky luminous flame to form CO_2 and H_2O

$$2C_6H_6 + 15O_2 \longrightarrow 12CO_2 + 6H_2O + \text{Heat}$$

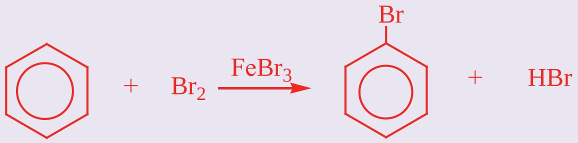
2. **Substitution reaction:** This is a characteristic reaction of benzene and other aromatic compounds. In this reaction, the hydrogen atom from the benzene ring is replaced by another atom or group.

(a). Halogenation: Benzene reacts with bromine and chlorine in the presence of iron (III) chloride or aluminium chloride catalyst to form substitution products.

Example 6.26:

$$C_6H_6 + Br_2 \xrightarrow{FeBr_3} C_6H_5Br + HBr$$

or

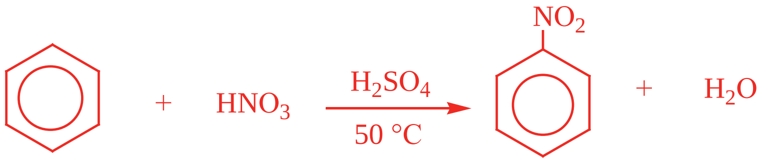


Benzene Bromobenzene

(b). Nitration: A mixture of concentrated nitric acid and concentrated sulphuric acid reacts with benzene at moderate temperatures to form nitrobenzene.

$$C_6H_6 + HNO_3 \xrightarrow[50\ °C]{H_2SO_4} C_6H_5NO_2 + H_2O$$

or

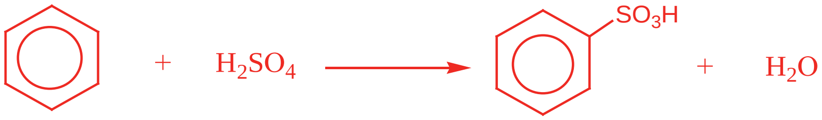


Benzene Nitrobenzene

(c). Sulphonation: Benzene reacts with concentrated sulphuric acid, H_2SO_4, at room temperature to form benzenesulphonic acid.

$$C_6H_6 + H_2SO_4 \longrightarrow C_6H_5SO_2OH + H_2O$$

or



Benzene Benzenesulphonic acid

Experiment 6.5

Chemical Reaction of Toluene

Objective: To distinguish aromatic hydrocarbons from other unsaturated hydrocarbons

Materials required: Toluene, Br_2 in CCl_4, $KMnO_4$, concentrated H_2SO_4, test tubes, test tube rack, test tube holder, measuring cylinder, Bunsen burner and dropper.

Procedure:

1. Take three test tubes and place them in the test tube rack.
2. Add 5 mL of toluene (methylbenzene) to each of the three test tubes.
3. Add one or two drops of Br_2 in CCl_4 in the first test tube and add the same amount of cold $KMnO_4$ solution in the second test tube and observe the changes.
4. Add a few drops of concentrated sulphuric acid to the third test tube and, if necessary, heat it gently, holding it with a test tube holder.

Observations and analysis:

In which test tube did a reaction occur?

Write a complete report of your observations and submit it to your teacher.

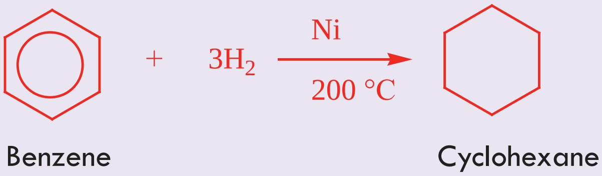
3. **Addition reaction:** Under normal conditions, aromatic compounds do not undergo addition reaction. Benzene undergoes addition reactions under special conditions.

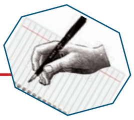
Example 6.27:

When a mixture of benzene vapour and hydrogen is passed over a finely divided nickel catalyst at 200 °C, cyclohexane is formed.

$$C_6H_6 + 3H_2 \xrightarrow[200\ ^\circ C]{Ni} C_6H_{12}$$

or





Exercise 6.18

1. What is an aromatic hydrocarbon?
2. What is the name of the simplest aromatic hydrocarbon?
3. Describe the main reactions of benzene?
4. Benzene does not change the colour of Br_2 in CCl_4 or that of $KMnO_4$ solution; why?
5. How do you prepare the following compounds from benzene?
   a. Chlorobenzene
   b. Nitrobenzene
   c. Benzenesulphonic acid.

6.5 Natural Sources of Hydrocarbons

At the end of this section, students will be able to

- list the major natural sources of hydrocarbons;
- describe natural gas;
- define crude oil;
- dxplain fractional distillation of crude oil;
- dention products of fractional distillation of crude oil;
- discuss the uses of petroleum products.

Activity 6.17

What do you think is the composition of the fuels people use for vehicles and motorbike? What about bottled gases for cooking? Does Ethiopia have an oil (crude oil) deposit? What are the natural resources Ethiopia is gifted with? Discuss whether they are renewable or nonrenewable and their environmental impact. If you were a leader of the country, would you invest more on fossil fuel extraction or in the development of renewable energy sources such as hydroelectric dams?

By far, the major sources of alkanes are the world's natural gas and petroleum deposits followed by coal. Laid down eons ago, these deposits are thought to be derived primarily from the decomposition of tiny single-celled marine organisms called foraminifera.

A. Natural Gas

Natural gas is a mixture of gases that consists chiefly of methane (more than 90%) but also contains ethane, propane, and butane. Other gases such as, CO_2, N_2, O_2 and H_2S, may also be present in natural gas. Natural gas is found in association with petroleum or alone. It is found in underground deposits several hundreds or thousands of meters below the earth's surface, where it originated from the decay and decomposition of animal and plant remains millions of years ago. The composition of natural gas varies, depending on its place of origin. Natural gas is widely used as a fuel. It is advantageous over liquid and solid fuels and also other gases, except H_2, due to its very high heat of combustion.

B. Petroleum (crude oil)

The word petroleum is derived from two Latin words, 'petra' meaning rock, and 'oleum' meaning oil. Thus, petroleum means rock oil. Petroleum (also called crude oil) is a complex mixture of fossil hydrocarbons that must be separated into fractions and then further refined before it can be used. It is generally found along with natural gas in the form of a dark-coloured viscous liquid, which is a complex mixture of hydrocarbons. Its composition varies according to its place of origin. However, crude oil mainly contains alkanes, cycloalkanes and aromatic hydrocarbons. It also contains organic compounds consisting of nitrogen, sulphur and oxygen in small amounts.

Refining of Crude Oil

Petroleum refining begins by fractional distillation of crude oil into three principal cuts according to boiling point (bp): straight-run gasoline (bp 30–200 °C), kerosene (bp 175–300 °C), and heating oil, or diesel fuel (bp 275–400 °C). Further distillation under reduced pressure then yields lubricating oils and waxes and leaves a tarry residue of asphalt. The distillation of crude oil is only the first step in gasoline production, however. ***Figure 6.8*** shows the fractional distillation set up for separating crude oil into its component fractions.

The major products (fractions) obtained from the fractional distillation of petroleum and their uses are given in ***Table 6.7***.

Table 6.7 Typical fractions obtained by fractional distillation of petroleum

Fractions	Number of carbon atoms per molecule	Boiling point range (°C) of fractions	Uses
Natural Gases	$C_1 \rightarrow C_4$	Below 20	As fuel in the form of bottled gas.
Petroleum ether	$C_5 \rightarrow C_7$	20 → 60	As solvent in dry cleaning
Gasoline (Petrol)	$C_5 \rightarrow C_{10}$	40 → 200	Petrol fuel for internal combustion engines
Kerosene	$C_{12} \rightarrow C_{18}$	175 → 275	Jet engine fuel, household fuel
Fuel oil (diesel oil)	$C_{15} \rightarrow C_{18}$	250 → 400	Furnace fuel, diesel engine fuel
Lubricating oil Petroleum jelly	$>C_{19}$	Non-volatile liquids	- Lubrication, grease
Paraffin wax Residue	$>C_{20}$	Non-volatile solids	- Candles, polishes - bitumen for roof sealing and road surfaces

Note that as the later fractions obtained at higher temperatures are more coloured, more viscous, and less inflammable.

Straight-run gasoline turns out to be a poor fuel in automobiles because of engine knock, an uncontrolled combustion that can occur in a hot engine.

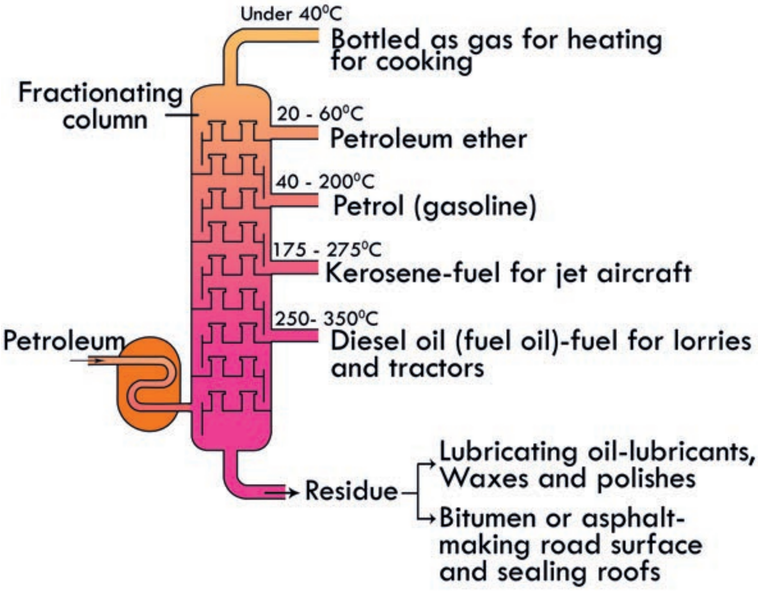
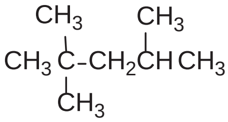


Figure 6.8 Fractional Distillation of Petroleum.

The octane number of a fuel is the measure by which its antiknock properties are judged. Straight-chain hydrocarbons are far more prone to induce engine knock than are highly branched compounds. The highly branched compound 2,2,4-trimethylpentane (called isooctane in the petroleum industry) burns very smoothly (without knocking) in internal combustion engines and is used as one of the standards by which the octane rating of gasolines is established. According to this scale, 2,2,4-trimethylpentane has an octane rating of 100. Heptane, $CH_3(CH_2)_5CH_3$, a compound that produces much knocking when it is burned in an internal combustion engine, is given an octane rating of 0. Mixtures of 2,2,4-trimethylpentane and heptane are used as standards for octane ratings between 0 and 100. A gasoline, for example, that has the same characteristics in an engine as a mixture of 87% 2,2,4-trimethylpentane and 13% heptane would be rated as 87-octane gasoline.

$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$

Heptane
(octane number = 0)



$$CH_3\overset{CH_3}{\underset{CH_3}{C}}-CH_2\overset{CH_3}{CH}CH_3$$

2,2,4-trimethylpentane
(octane number = 100)

Because straight-run gasoline burns so poorly in engines, petroleum chemists have devised numerous methods for producing higher-quality fuels. One of these methods, catalytic cracking, involves taking the high-boiling kerosene cut ($C_{11}-C_{14}$) and ‘cracking’ it into smaller branched molecules suitable for use in gasoline. What does cracking mean? Cracking is the decomposition of large hydrocarbon molecules into smaller ones by the application of heat (thermal cracking or pyrolysis)

$$\underset{\text{Hexadecane}}{C_{16}H_{34}} \xrightarrow{\text{heat}} \underset{\text{Octane}}{C_8H_{18}} + \underset{\text{Octene}}{C_8H_{16}}$$

or in the presence of catalysts (catalytic cracking). Hydrogen gas can be added during cracking to saturate the alkenes formed in the process.

Another process, called reforming, is used to convert C_6-C_8 alkanes to aromatic compounds such as benzene and toluene, which have substantially higher octane numbers than alkanes. The final product that goes into the tank has an approximate composition of 15% C_4-C_8 straight-chain alkanes, 25% to 40% C_4-C_{10} branched-chain alkanes, 10% cyclic alkanes, 10% straight chain and cyclic alkenes, and 25% arenes (aromatics).

C. Coal

At the end of this section, students will be able to
- identify the composition of coal;
- describe destructive distillation of coal.

Coal is formed naturally by the decomposition of plant matter over several million of years. It is not a pure form of carbon. It is an important source of aromatic hydrocarbons.

Heating mineral coal in the absence of air, or oxygen is called destructive distillation of coal or coking of coal. When coal is heated in the absence of air it gives volatile products and coke. The volatile products that are separated into coal gas and a liquid are called coal tar when cooled. Aromatic hydrocarbons and many other substances are isolated from coal tar by fractional distillation. Coke, which is a solid and relatively pure form of carbon, is used as a fuel in the blast furnace during the extraction of iron. It is also used to produce gaseous fuels, such as water gas (mixture of H_2 and CO) and producer gas (mixture of N_2 and CO).

Exercise 6.19

1. Discuss the importance of natural gas for a country's economic development.
2. What are the major natural sources of hydrocarbons?
3. Which hydrocarbons are the main constituents of petroleum?
4. What is destructive distillation?
5. Describe how we can obtain aromatic hydrocarbons from coal.
6. Which fraction of petroleum is in very high demand?
7. What process should be carried out during fractional distillation of crude oil to maximize the output of petrol or gasoline?
8. Describe the main uses of a) petroleum ether; b) kerosene; c) gasoline; d) diesel oil; e refinery gases.

Key Terms

- addition reaction
- alkane
- alkene
- alkyl halide
- alkyl radicals
- alkynes
- benzene
- biogas
- carboxylic acid
- chain isomerism
- cracking
- crude oil
- cycloalkanes
- cycloalkene
- decarboxylation
- dehydrohalogenation
- destructive distillation
- elimination reaction
- functional group
- geometrical isomerism
- halogenation
- hydrohalogenation
- homologous series
- hydrogenation
- inorganic compound
- isomers
- ketone
- life force (vitalism) theory
- Markovnikov's rule
- natural gas
- organic compound
- petroleum
- polymerization
- position isomerism
- saturated hydrocarbons
- substitution reaction
- unsaturated hydrocarbons
- Wurtz reaction

Unit Summary

- The 'vitalism' theory states that living organisms (plants and animals) have a special life force which helps them to synthesize organic compounds and that organic compounds cannot be synthesized artificially by man in the laboratory.
- Organic chemistry is the study of carbon compounds including natural and synthetic ones.
- Organic compounds are classified based on the functional group they contain in their molecules.
- The functional group is a part of a molecule that determines the chemical and physical properties of a compound.
- Hydrocarbons are compounds composed of carbon and hydrogen atoms only.
- Saturated hydrocarbons contain only single bonds between carbon atoms
- Unsaturated hydrocarbons are those containing one or more double or triple bonds between carbon atoms.
- A homologous series is a group of compounds in which each member of the group differs from the next member by $-CH_2-$ group.
- Alkanes are saturated hydrocarbons represented by the general formula C_nH_{2n+2}. Alkanes are named by using prefixes that indicate the number of carbon atoms and the suffix '-ane'.
- Isomerism is the existence of two or more compounds with the same molecular formula but different structures. These compounds are called isomers.
- Alkenes are unsaturated hydrocarbons containing a carbon-carbon double bond as their functional group. Their general formula is C_nH_{2n}. The suffix '-ene' in their names indicates the presence of a carbon-carbon double bond.
- Alkynes are unsaturated hydrocarbons containing a carbon-carbon triple bond as their functional group. They are represented by the general formula C_nH_{2n-2}. The suffix '-yne' indicates the carbon-carbon triple bond.
- Alkenes and alkynes undergo mainly addition reactions while alkanes undergo substitution reactions.
- Aromatic hydrocarbons are compounds containing benzene rings. Benzene is the simplest aromatic hydrocarbon.
- The natural sources of hydrocarbons are crude oil, natural gas and coal.
- Crude oil is a complex mixture of hydrocarbons, mainly alkanes, and also contains cycloalkanes and aromatics.
- Crude oil is separated into several useful fractions by fractional distillation.
- Organic chemicals are also used as fuels and solvents for dry cleaning.

Review Exercise

Part I: Multiple Choice Questions

1. The first organic compound that is artificially synthesized from inorganic starting materials is______________.
   a. benzene b. methane c. urea d. hydrocarbons

2. Which of the following hydrocarbons does not belong to the homologous series of alkanes?

   a. C_3H_8 b. C_8H_{18} c. C_6H_{10} d. $C_{10}H_{22}$

3. Which of the following alkanes is a liquid at room temperature?

   a. CH_4 b. $C_{10}H_{22}$ c. C_4H_{10} d. C_3H_8

4. The hydrocarbon used as a fuel in an oxyacetylene torch is:

   a. C_2H_2 b. C_2H_4 c. C_4H_{10} d. CH_4

5. The general formula that represents the olefin homologous series is:

   a. C_nH_{2n+1} b. C_nH_{2n+2} c. C_nH_{2n-2} d. C_nH_{2n}

6. The most common reaction that alkenes undergo is:

   a. elimination b. substitution c. addition d. decomposition

7. A compound having molecular formula C_7H_8 that does not decolorize Br_2 in CCl_4 or $KMnO_4$ solution most likely belongs to a class of hydrocarbons called ________

   a. Alkynes c. Alkenes
   b. Alkanes d. Aromatic hydrocarbons

8. Which hydrocarbon is the main constituent of natural gas and biogas?

   a. C_4H_{10} b. CH_4 c. C_2H_6 d. C_6H_{14}

9. During the fractional distillation of crude oil, the process of cracking is carried out to:

   a. separate the crude oil into different fractions
   b. vaporize the liquid components
   c. maximize the output of petrol or gasoline
   d. separate alkanes from alkenes

10. Ethyne decolorizes Br_2 in CCl_4. This is due to the formation of:

    a. 1,1-dibromoethane
    b. 1,1,2,2-dibromoethene
    c. 1,2-dibromoethane
    d. 1,1,2,2-tetrabromoethane

11. The following reaction can be classified as:

$$CH_4 + Br_2 \rightarrow CH_3Br + HBr$$

    a. an addition reaction c. a substitution reaction
    b. an elimination reaction d. a saponification reaction

12. Which one of the following substances is not used in dry cleaning?

    a. CCl_4 b. C_6H_6 c. $Cl_2C{=}CCl_2$ d. CH_3COOH

13. The first organic compound was synthesized in the laboratory by:

    a. Friedrich Wohler c. John Dalton
    b. A. Kekule d. Wurtz

14. Which reagent is used to test for unsaturation of hydrocarbons?

    a. Wurtz reagent c. Alkaline $KMnO_4$
    b. Bromine in CCl_4 d. B & C

15. Which alkene doesn't exhibit cis-trans isomerism?
    - a. 2-butene
    - b. 2-pentene
    - c. 2,3-dimethylbut-2-ene
    - d. All
16. Which one of the following compounds does not decolorize bromine water?
    - a. C_3H_4
    - b. C_2H_2
    - c. C_6H_{14}
    - d. All

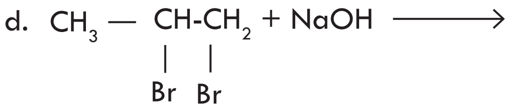
Part II: Supply Type Questions

17. Complete the following reaction equations:

a. $CH_2 = CH_2 + H_2 \xrightarrow[\text{Heat}]{\text{Ni}}$

b. $CH_3COONa + NaOH \xrightarrow{\text{CaO, heat}}$

c. $CH_3 - CH_2Br + 2Na \longrightarrow$

d. 
$CH_3 - \underset{Br}{\underset{|}{CH}}\text{-}\underset{Br}{\underset{|}{CH_2}} + NaOH \longrightarrow$

e. $CH_3CH_2CH_2OH \xrightarrow[\text{Heat}]{H_2SO_4}$

f. $CH_3CH_2CH_2Cl + KOH \xrightarrow{\text{Heat}}$

g. $CH \equiv CNa + CH_3CH_2Br \longrightarrow$

h. $CH_3\text{-}CH = CH_2 + HBr \longrightarrow$

i. $HC \equiv C\text{-}CH_2 - CH_3 + 2HCl \longrightarrow$

j. $CH_3CH_2\text{-}CH = CH_2 + H_2O \xrightarrow{H^+}$

k. $CH_3CH_2\text{-}CH = CH_2 + H_2O \xrightarrow[KMnO_4]{\text{Alkaline}}$

Part III. Workout Type Questions

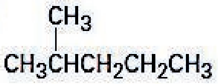
18. Draw structures for the following:
    - a. 2-Methylheptane
    - b. 4-Ethyl-2,2-dimethylhexane
    - c. 4-Ethyl-3,4-dimethyloctane
    - d. 2,4,4-Trimethylheptane
    - e. 3,3-Diethyl-2,5-dimethylnonane
    - f. 4-Isopropyl-3-methylheptane
    - g. 4-ethyl-2,2,3-trimethylhexane
    - h. 3-bromo-2-chlorohex-1-ene
19. Name the five isomers of C_6H_{14}.
20. Explain why each of the following names is incorrect:
    - a. 2,2-Dimethyl-6-ethylheptane
    - b. 4-Ethyl-5,5-dimethylhexane
    - c. 3-Ethyl-4,4-dimethylhexane
    - d. 5,5,6-Trimethyloctane

(e) 2-Isopropyl-4-methylheptane

21. Propene, $CH_3CH=CH_2$, is an alkene. Write the structure of a structural isomer of propene that is not an alkene. (Hint: It does not have a double bond.)
22. Propose structures for two structural isomers of cyclopentene that do not contain a ring.

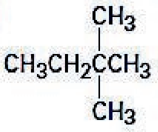
23. Give IUPAC names for the following compounds:

(a)



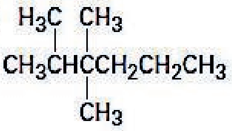
$CH_3\overset{\overset{CH_3}{|}}{C}HCH_2CH_2CH_3$

(b)



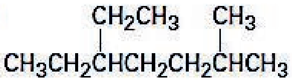
$CH_3CH_2\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{C}}CH_3$

(c)



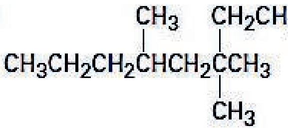
$CH_3\overset{\overset{H_3C}{|}}{C}H\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{C}}CH_2CH_2CH_3$

(d)



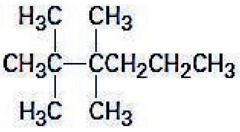
$CH_3CH_2\overset{\overset{CH_2CH_3}{|}}{C}HCH_2CH_2\overset{\overset{CH_3}{|}}{C}HCH_3$

(e)



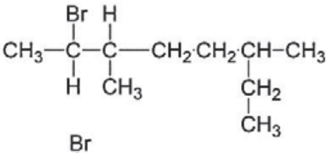
$CH_3CH_2CH_2\overset{\overset{CH_3}{|}}{C}HCH_2\overset{\overset{CH_2CH_3}{|}}{\underset{\underset{CH_3}{|}}{C}}CH_3$

(f)



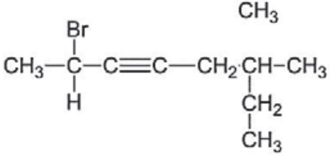
$CH_3\overset{\overset{H_3C}{|}}{\underset{\underset{H_3C}{|}}{C}}-\overset{\overset{CH_3}{|}}{\underset{\underset{CH_3}{|}}{C}}CH_2CH_2CH_3$

(g)



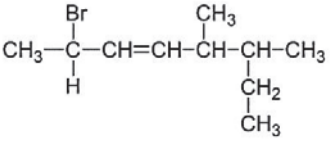
$CH_3-\overset{\overset{Br}{|}}{\underset{\underset{H}{|}}{C}}-\overset{\overset{H}{|}}{\underset{\underset{CH_3}{|}}{C}}-CH_2\cdot CH_2\cdot\underset{\underset{\underset{\underset{CH_3}{|}}{CH_2}}{|}}{CH}-CH_3$

(h)



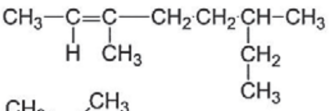
$CH_3-\overset{\overset{Br}{|}}{\underset{\underset{H}{|}}{C}}-C\equiv C-CH_2\cdot\underset{\underset{\underset{\underset{CH_3}{|}}{CH_2}}{|}}{CH}-CH_3$

(i)



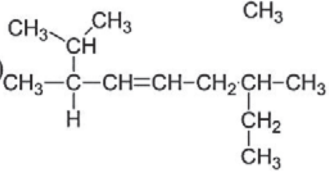
$CH_3-\overset{\overset{Br}{|}}{\underset{\underset{H}{|}}{C}}-CH=CH-\overset{\overset{CH_3}{|}}{CH}-\underset{\underset{\underset{\underset{CH_3}{|}}{CH_2}}{|}}{CH}-CH_3$

(j)



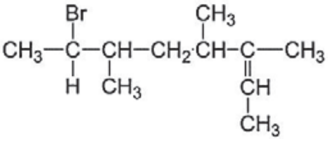
$CH_3-\underset{\underset{H}{|}}{C}=\underset{\underset{CH_3}{|}}{C}-CH_2\cdot CH_2\cdot\underset{\underset{\underset{\underset{CH_3}{|}}{CH_2}}{|}}{CH}-CH_3$

(k)



$CH_3-\overset{\overset{CH_3\diagdown\overset{\diagup CH_3}{CH}}{|}}{\underset{\underset{H}{|}}{C}}-CH=CH-CH_2\cdot\underset{\underset{\underset{\underset{CH_3}{|}}{CH_2}}{|}}{CH}-CH_3$

(l)



$CH_3-\overset{\overset{Br}{|}}{\underset{\underset{H}{|}}{C}}-\underset{\underset{CH_3}{|}}{CH}-CH_2\cdot\overset{\overset{CH_3}{|}}{CH}-\underset{\underset{\underset{\underset{CH_3}{|}}{CH}}{\|}}{C}-CH_3$