UNIT 2

Important Inorganic Compounds

Unit Outcomes

After completing this unit, you will be able to:

- *understand the classification of inorganic compounds on the basis of their composition and/or their chemistry;*
- *know the types of oxides and their chemical properties;*
- *understand the Arrhenius, Brønsted-Lowry and Lewis concepts of acids and bases;*
- *understand the classification of acids and salts;*
- *know the general properties, preparations and uses of common acids, bases and salts;*
- *understand 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;*
- *develop skills in calculating, pH, pOH, H^+ ion and OH^- ion concentration of a solution;*
- *know essential plant nutrients, fertilizers and pesticides; and*
- *demonstrate scientific inquiry skill: observing, classifying, comparing and contrasting, inferring, predicting, communicating, measuring, asking questions, interpreting data, drawing conclusion, applying concepts, relating cause and effect and problem solving.*

MAIN CONTENTS

2.1 INTRODUCTION

Competencies

By completing this section, you will be able to:

- define inorganic compounds; and
- classify inorganic compounds as oxides, acids, bases and salts.

Activity 2.1

In the laboratory, take samples of different compounds and classify them as organic and inorganic. When these compounds are burnt in air, it has been found that most organic compounds burn when heated in air, but most inorganic compounds just melt or vaporize; why? Discuss your observations in groups and share your opinion with the class.

Inorganic compounds are the compounds consisting of mineral constituents of the earth or generally found in non-living things. The term inorganic compound refers to all compounds that do not contain carbon. Although, carbon dioxide, carbon monoxide, carbonates and hydrogen carbonates are carbon-containing compounds, which are classified as inorganic compounds. Inorganic compounds are mostly found in nature as silicates, oxides, carbonates, sulphides, sulphates, chlorides and nitrates, etc.

There are different ways for the classification of inorganic compounds. They can be classified on the bases of their composition. For example; they can be classified on the basis of the:

i) metal they contain such as copper compounds, aluminium compounds, etc.

ii) non-metal they contain such as sulphur compounds, nitrogen compounds, etc.

iii) group they contain such as sulphates, nitrates, carbonates, etc.

This unit emphasizes on the four groups of inorganic chemicals namely oxides, acids, bases and salts.

Exercise 2.1

1. List some inorganic compounds with which you are familiar.
2. Which branch of chemistry is concerned with the study of all elements?

2.2 OXIDES

Competencies

After Completing this section, you will be able to:

- define oxides;
- classify oxides as acidic, basic, amphoteric, neutral and peroxides;
- define acidic oxides and give examples;
- explain the chemical properties of acidic oxides;
- define basic oxides and give examples;
- explain the chemical properties of basic oxides;
- conduct experiments to distinguish acidic oxides from basic oxides;
- compare and contrast acidic and basic oxides;
- define amphoteric oxides and give examples;
- explain the chemical properties of amphoteric oxides;
- discuss the salt forming nature of acidic, basic and amphoteric oxides;
- define neutral oxides and give examples;
- define peroxides and give examples;
- explain the chemical properties of peroxides; and
- conduct an experiment to distinguish peroxides from other oxides.

What are oxides? 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*). Binary compounds are those consisting of only two elements. Examples of oxides are calcium oxide, CaO, aluminium oxide, Al_2O_3, sulphur dioxide, SO_2, and carbon monoxide, CO.

The classification of oxides into different groups is based on their chemical behaviour; they are classified as:

- Acidic oxides
- Basic oxides
- Amphoteric oxides
- Neutral oxides, and
- Peroxides

Exercise 2.2

Decide whether the following compounds are oxides or not:

a Na_2O d P_4O_6 g Na_2CO_3
b KOH e N_2O_4 h KNO_3
c Ga_2O_3 f H_2O i H_2SO_4

Activity 2.2

1. The oxides, CO_2, N_2O_5, P_4O_{10}, SO_2 are acidic oxides.
   a Are these oxides formed by the combination of oxygen with
      i) metals ii) non-metals or iii) metalloids?
   b What general conclusion can you draw about the composition of acidic oxides?
2. The presence of CO_2 is confirmed by reacting CO_2 with lime water.
   a What is lime water? Is it a base or an acid?
   b Write a balanced chemical equation for the above reaction.
   c What type of oxides react with bases?

   Discuss in your group and present your findings to the class.

A Acidic Oxides

Acidic oxides are the oxides formed by the chemical combination of oxygen with non-metals. Thus, acidic oxides are non-metal oxides. 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.

Chemical Properties of Acidic Oxides

Acidic oxides undergo the following reactions:

1. Acidic oxides (*acid anhydrides*) dissolve in water to form acidic solution (*acid*).

   Acid anhydride + water → Acid

Examples

$CO_2 + H_2O \rightarrow H_2CO_3$ (*Carbonic acid*)

$N_2O_3 + H_2O \rightarrow 2HNO_2$ (*Nitrous acid*)

$SO_2 + H_2O \rightarrow H_2SO_3$ (*Sulphurous acid*)

2. Acidic oxides react with basic or metallic oxides to form salt.

   Acidic oxide + Basic oxide → Salt

Examples

$CO_2 + Na_2O \rightarrow Na_2CO_3$

$SO_3 + CaO \rightarrow CaSO_4$

3. Acidic oxides react with bases to form salt and water. This reaction is called neutralization reaction.

   Acidic oxide + Base → Salt + Water

Examples

$SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$

$CO_2 + 2LiOH \rightarrow Li_2CO_3 + H_2O$

Exercise 2.3

1. List five examples of acidic oxides.
2. Complete and balance the following equations.

   a $P_4O_6 + H_2O \rightarrow$

b	SO_3	+	H_2O	→	
c	CaO	+	CO_2	→	
d	$Ca(OH)_2$	+	SO_3	→	
e	$NaOH$	+	CO_2	→	

B. Basic Oxides

Oxides that are composed of metals and oxygen are basic oxides. 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.

Chemical Properties of Basic Oxides

What common reaction do basic oxides undergo? What products do they form in their reactions with water, acidic oxides and acids?

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

$BaO + SO_3 \rightarrow BaSO_4$

$CaO + CO_2 \rightarrow CaCO_3$

$Na_2O + CO_2 \rightarrow Na_2CO_3$

3. Basic oxides react with acids to form a salt and water.

   Basic oxide + Acid $\rightarrow$ salt + water

Example

$CaO + 2HCl \rightarrow CaCl_2 + H_2O$

$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

Exercise 2.4

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 $Li_2O + SO_2 \rightarrow$

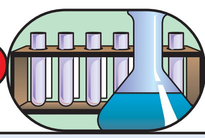
   e $BaO + H_2SO_4 \rightarrow$

   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

Experiment 2.1



Test for Acidity and Basicity of Oxides

Objective: To identify basic and acidic oxides.

Materials required: Deflagrating spoon, gas jar and gas jar lid, test tubes. Sulphur, magnesium or calcium metal, water, universal indicator, litmus paper (*blue* and *red*).

Procedure:

1. Ignite a small amount of powdered sulphur on a deflagrating spoon and insert in a gas jar of oxygen. Add 10 mL water after ignition is complete. Cover the gas jar with a lid and shake. 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 of oxygen. 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.

Observations and analysis:

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.

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. Table 2.1 shows some common indicators and the colour they develop in acidic and basic solutions.

Table 2.1 Some common indicators and their colours in acidic and basic solution.

Indicator	Colour in aqueous solution of acidic oxide	Colour in aqueous solutionof 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

In addition to their effects on indicators, acidic and basic oxides can be identified by their chemical properties. 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.

C. Amphoteric Oxides

Activity 2.3

1. How would you classify the oxides of period III elements?
2. 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?

   Discuss this in your group and present to the class.

There are oxides which exhibit both acidic and basic properties. These are known as amphoteric oxides. In their reaction with acids, they behave as bases and, in their reaction with bases they act as acids. The following reaction shows the amphoteric behaviour of aluminium oxide, Al_2O_3.

Amphoteric oxide	+	Acid	$\rightarrow$	salt	+	water
Al_2O_3 (s)	+	6HCl (aq)	$\rightarrow$	$2AlCl_3$ (aq)	+	$3H_2O$ (l)
Amphoteric oxide	+	base	$\rightarrow$	salt	+	water
Al_2O_3 (s)	+	2NaOH (aq)	$\rightarrow$	$2NaAlO_2$ (aq)	+	H_2O (l)
				sodium aluminate		

Some other examples of amphoteric oxides are ZnO, 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 substances. 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 2.5

Write chemical equations to show the amphoteric properties of ZnO and PbO when they react with:

a HCl

b NaOH

c HNO_3

d KOH

Experiment 2.2

Investigating Amphoteric Behaviour of Oxides

Objective: To observe the amphoteric behaviour of Al_2O_3.

Materials required:

Spatula, reagent bottles, beakers, glass rod, Al_2O_3, HCl, NaOH, Universal indicator and water.

Procedure:

1. Prepare solutions by:
   a mixing 20 mL concentrated HCl and 80 mL water in one reagent bottle;
   b dissolving 8 g NaOH in 100 mL water in another reagent bottle.
   c Add universal indicator to the acid and base, and observe the colour change.
2. Take two beakers and place a spatula full of Al_2O_3 in each of the beakers.
3. 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.
4. Add universal indicator in the two beakers and observe the colour change.

Observations and analysis:

1. Does Al_2O_3 react with the substances in both solutions?
2. What does the change in colour of the indicator in the mixtures indicate?
3. Write chemical equations to show what has happened?

Write a laboratory report and present to the class.

D. Neutral Oxides

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.

Activity 2.4

1. In a beauty saloon which chemical is used to decolorize hair? What is the oxidation number of oxygen in this compound?

2. If Na_2O_2 and CaO_2 are reacted with water, do you get the same compound, which is used to decolorize hair? Compare the oxidation number of oxygen in Na_2O and MgO with the oxidation number of the compound used for decolorizing hair.

   Discuss your findings in group and present to the class.

E. Peroxides

In acidic, basic, amphoteric and neutral oxides, the oxidation state of oxygen is –2, but in peroxides it is –1. In peroxides, the two oxygen atoms are linked to each other and with atoms of other elements. They contain the peroxide, "– O – O –" link. In the oxides discussed above, oxygen atoms are linked directly with atoms of other elements.

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.

$$2Na\ (s) + O_2\ (g) \rightarrow Na_2O_2\ (s)$$
$$Ca\ (s) + O_2\ (g) \rightarrow CaO_2\ (s)$$

Chemical Properties of Peroxides

What chemical properties do the peroxides exhibit? Some of the chemical properties of peroxides include:

a. Peroxides are powerful oxidizing agents; they react with different substances by losing oxygen.

Examples

$$PbS\ (s) + 4H_2O_2\ (aq) \rightarrow PbSO_4\ (s) + 4H_2O\ (l)$$
$$2KI\ (aq) + H_2SO_4\ (aq) + H_2O_2\ (aq) \rightarrow I_2\ (s) + K_2SO_4\ (aq) + 2H_2O\ (l)$$

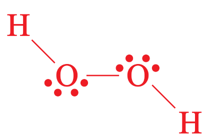
b. Peroxides react with aqueous acids to form hydrogen peroxide.

Examples

$$Na_2O_2\ (s) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + H_2O_2\ (aq)$$
$$CaO_2\ (s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (s) + H_2O_2\ (aq)$$

Hydrogen Peroxide

The structure of the hydrogen peroxide molecule is:



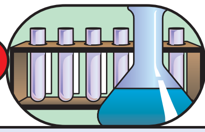
Hydrogen peroxide decomposes to release oxygen. This reaction is slow but can be speeded up by the addition of manganese (IV) oxide, MnO_2, as a catalyst.

$$2H_2O_2\,(aq) \rightarrow 2H_2O\,(l) + O_2\,(g)$$

Hydrogen peroxide is a strong oxidizing agent. Its oxidizing power is responsible for its effectiveness as an antiseptic for mouthwash and cleansing wounds. It is also used as a bleaching agent. When hydrogen peroxide is added to a coloured dye, the molecule responsible for the colour will oxidize and so the colour will disappear. 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 2.3



Distinguishing Peroxides from Other Oxides

Objective: To identify peroxides from other oxides.

Materials required: 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, 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. Prepare starch solution in the third beaker by boiling 1g starch in 100 mL water.
2. 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. 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.

Observations and analysis:

a In which beaker do you see a colour change?

b What is the cause of the colour change?

c Write a balanced chemical equation for the change?

Write a laboratory report and present to the class.

Exercise 2.6

1. Classify the following oxides as acidic, basic, amphoteric, neutral and peroxides:

a K_2O	c SO_2	e CaO
b CO	d Al_2O_3	f NO_2
g MgO	k NO	n CaO_2
h N_2O	l PbO	o Li_2O
i ZnO	m Na_2O_2	p P_4O_6
j BaO_2		

2. Complete and balance the following equations:

a $BaO + P_4O_{10} \rightarrow$	e $Al_2O_3 + HNO_3 \rightarrow$
b $SrO + SO_3 \rightarrow$	f $CaO + HCl \rightarrow$
c $CO_2 + KOH \rightarrow$	g $PbO + NaOH \rightarrow$
d $ZnO + H_3PO_4 \rightarrow$	h $MgO + H_2SO_4 \rightarrow$

3. Which acidic oxide should react with water to form each of the following acids?

a HNO_3	d H_2SO_3
b H_2CO_3	e H_3PO_4
c H_2SO_4	f H_3PO_3

4. Identify the basic anhydrides that react with water to form each of the following bases:

a $Ba(OH)_2$	e KOH
b $Sr(OH)_2$	f $CsOH$
c $Mg(OH)_2$	
d $LiOH$	

5. How do peroxides differ from other groups of oxides?
6. How can you identify whether an oxide is acidic or basic?

2.3 ACIDS

Competencies

After completing this section, you will be able to:

- define acids in terms of Arrhenius, Brønsted-Lowry and Lewis;

- give examples of acids based on Arrhenius, Brønsted-Lowry and Lewis concepts;
- classify acids as monoprotic and polyprotic based on the numbers of ionizable hydrogen atoms;
- group acids as binary and ternary based on the number of elements they contain;
- explain the general properties of acids;
- define strong and weak acids;
- distinguish between strong and weak acids;
- define concentrated and dilute acids;
- describe the conceptual difference between strong and concentrated acids;
- use the necessary precautions while working with acids;
- define pH and 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 indicator or a pH-meter;
- calculate the pH of a given acidic solution;
- calculate the hydrogen ion concentration from the given information;
- perform activities to investigate some physical properties of acids;
- perform activities to investigate some chemical properties of acids;
- explain the direct combination of elements, the reaction of acidic oxides with water, formation of volatile acids from non-volatile acids as the three methods of preparation of acids;
- conduct simple experiments to prepare acids in a laboratory; and
- describe the uses of three common laboratory acids.

Activity 2.5

1. How do you describe an acid?
2. Give some examples of acids and bases that you have encountered in your everyday life?
3. List as many sour foods as possible which you have ever tasted.
4. In which form does the hydrogen ion, H^+, exist in aqueous solutions?

Discuss with your group and present it to the class.

Acids are among the most familiar of all chemical compounds. 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.

Definitions of Acids

a 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.

Example

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

or,

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Some examples of Arrhenius acids are HCl, HNO_3, H_2SO_4, HI, HBr, and HF. For a substance to be called an Arrhenius acid, it should contain ionizable hydrogen.

Despite its early success and usefulness, the definition of an acid, first proposed by Arrhenius has some limitations. The Arrhenius definition of acids explains their behaviour only in aqueous solution. In addition, it does not explain why some substances show acidic behaviour in the gaseous state and in non-aqueous solutions.

For example, when ammonia and hydrogen chloride gases are brought together, they react to form ammonium chloride as follows:

$$NH_3\,(g) + HCl\,(g) \rightarrow NH_4Cl\,(s)$$

In this reaction, the hydrogen chloride provides hydrogen ion (or proton), H^+, which reacts with NH_3 to form NH_4^+. But, this happens in the gaseous state not in aqueous solution. Due to the limitations of Arrhenius' definition, chemists tried to define acids in a more general way.

Brønsted-Lowry Definition of Acids

A broader and more general definition of acids was provided independently in 1923, by Johannes Brønsted and Thomas M. Lowry. According to Brønsted-Lowry definition; an acid is a substance that donates protons, H^+, to some other substance.

Examples

$$HCl\,(aq) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)$$

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

$$H_3PO_4\,(aq) + 2H_2O\,(l) \rightleftharpoons HPO_4^{2-}\,(aq) + 2H_3O^+\,(aq)$$

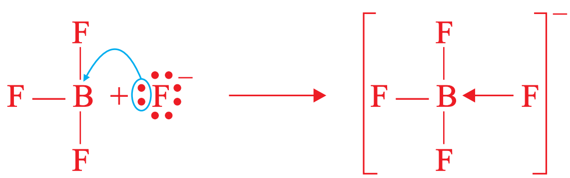
In the above reactions, HCl in the first, H_2O in the second and H_3PO_4 in the third are Brønsted-Lowry acids because they are proton donors. Here it is very important to realize that all Arrhenius acids are Brønsted-Lowry acids. However, the reverse is not true because Brønsted-Lowry acids include substances that are not acids according to Arrhenius.

Although the Brønsted-Lowry definition of an acid is more general than the definition proposed by Arrhenius, it still depends on the transfer of protons. However, the presence of acids that are not proton-donors, initiated chemists to search for another definition of acids.

b Lewis Definition of Acids

In 1923, the American chemist Gilbert Newton Lewis produced a definition of acid that extends the concept of acids even further than that of Brønsted-Lowry.

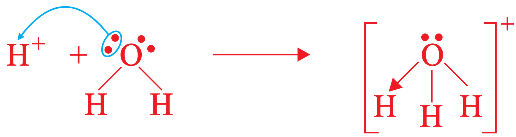
According to Lewis, an *acid is a substance that can form a coordinate covalent bond by accepting an electron pair from another* substance. In other words, an acid is an electron-pair acceptor. Consider the following reactions:



Lewis acid

In the above reaction, between BF_3 and F^-, BF_3 acts as a Lewis acid since it accepts a pair of electrons from F^-.

The Lewis definition of an acid is also valid for the Arrhenius concept of an acid. This is because as the acid releases an hydrogen ion, H^+, in aqueous solution, the released proton or H^+ accepts a pair of electrons from a water molecule to form an hydronium, H_3O^+ ion.

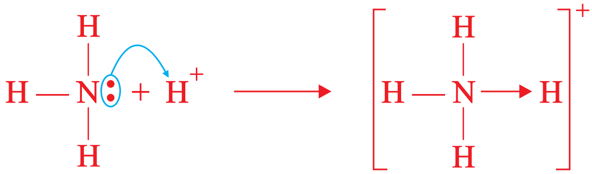


Lewis acid Hydronium ion

The Lewis definition of an acid is also valid for the Brønsted-Lowry concept of acids. Since Brønsted-Lowry defined an acid as a proton donor, the donated proton can be accepted by a molecule or species that has lone pair of electrons. So the proton, H^+, is an electron pair acceptor and hence a Lewis acid. Consider the following reaction.

$$NH_3\,(g) \quad + \quad H_2O\,(l) \quad \rightleftharpoons \quad NH_4^+\,(aq) \quad + \quad OH^-\,(aq)$$

In this reaction, H_2O is a proton donor and hence a Brønsted-Lowry acid. The donated proton from water to ammonia, accepts a pair of electrons from the nitrogen atom in NH_3 to form NH_4^+ as shown below:



Lewis acid

The Lewis concept of acids not only includes H^+ as an acid, but also ions or molecules capable of accepting an electron pair that neither release H^+ in aqueous solution nor donate a proton.

Classification of Acids

Activity 2.6

Phosphoric acid, H_3PO_4, has three hydrogen atoms and it is a triprotic acid. Acetic acid, CH_3COOH, has four hydrogen atoms and is classified as a monoprotic acid. Discuss the reason in your group and present it to the class.

There are different ways of classifying acids; they can be classified depending on the number of replaceable (*ionizable*) hydrogen atom(s) they contain per molecule as monoprotic or polyprotic acids.

Monoprotic acids are the acids containing only one ionizable (replaceable) hydrogen atom per molecule or those acids that can furnish only one hydrogen *ion per molecule in aqueous solution*. Common examples of monoprotic acids are HCl, HNO_3, HBr, HI, and CH_3COOH. The ionization of a monoprotic acid in aqueous solution is presented using HCl and HNO_3 as specific examples, as shown below:

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$

$$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$$

Polyprotic acids are those acids containing more than one ionizable (*replaceable*) hydrogen ion in aqueous solution. Example of acids in this category includes H_2SO_4, H_2CO_3, and H_3PO_4.

The ionization of polyprotic acids in aqueous solution is shown in the following chemical equations:

$$H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2-}\ (aq)$$

$$H_2CO_3\ (aq) \rightleftharpoons 2H^+\ (aq) + CO_3^{2-}\ (aq)$$

$$H_3PO_4\ (aq) \rightleftharpoons 3H^+\ (aq) + PO_4^{3-}\ (aq)$$

Polyprotic acids which contain two ionizable hydrogen atoms such as H_2SO_4, H_2S, and H_2CO_3 are also called diprotic acids; those containing three ionizable hydrogen atoms like H_3PO_4 are called triprotic acids.

Activity 2.7

Ammonia and phosphoric acid have three hydrogen atoms but ammonia does not behave as triprotic acid. Why? Discuss in your group and present it to the class.

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 elements. 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. They usually contain hydrogen, oxygen and a non-metal. Examples are H_2SO_4, H_2CO_3, $HClO_4$, and H_3PO_4.

General Properties of Acids

Activity 2.8

Add a drop of lemon juice on your tongue and try to identify its taste. Repeat your observation with orange juice. Do the two juices taste bitter or sour? Try to classify them as acid or base, depending on their taste. Share your observations with the rest of class.

Acids generally have the following properties:

1. *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 is harmless. However, most concentrated acids are corrosive and poisonous. So it is strictly forbidden to attempt to identify such acids by tasting them.

2. *Acids change the colour of indicators*

The common indicators available in high school laboratories are litmus, phenolphthalein, methyl orange or methyl red and universal indicator. The effect of acids on the colour of indicators is summarized in Table2.1, Section 2.2.

Experiment 2.4

Effect of Acids on Indicators

Objective: To detect acidity of a solution using indicators.

Materials required: Lemon juice, dilute HCl, dilute HNO_3, dilute H_2SO_4 phenolphthalein, litmus, methyl red, universal indicator, 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.

Observations and analysis:

What colours have you observed when each indicator was added to each of the four acid solutions? Use the following Table to record your observation:

Indicator	**Colour of indicator in**			
	Lemon juice	**dilute HCl**	**dilute HNO_3**	**dilute H_2SO_4**
Phenolphthalein				
Litmus				
Methyl red				
Universal indicator				

Write a laboratory report and present to the class.

3. *Acids react with active metals to form salt and hydrogen gas*

Metals like magnesium, zinc, and iron react with dilute acids to form salt and liberate hydrogen gas.

$$\text{Acid} + \text{Active metal} \rightarrow \text{Salt} + \text{Hydrogen gas}$$

Examples

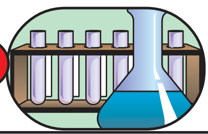
$$2HCl + Zn \rightarrow ZnCl_2 + H_2$$
$$H_2SO_4 + Mg \rightarrow MgSO_4 + H_2$$

Very active metals like sodium, potassium, and calcium react very violently with dilute acids. The reaction is very dangerous and should not be performed.

Acids reacting with metals do not necessarily produce hydrogen gas. For example, 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 2.5



Investigating the Reactions of Metals with Dilute Acids

Objective: To investigate the reaction between active metals and dilute acids.

Materials required: Dilute H_2SO_4, dilute HCl, zinc, magnesium and iron, test tubes, test tube holder, test tube rack, burner, wooden splint, 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 piece of cardboard for a few seconds. Bring a burning splint close to the mouth of the test tube and remove the cardboard.

Repeat the experiment with dilute H_2SO_4 after placing each of the three metals in three different test tubes.

Observations and analysis:

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?

Write a laboratory report on your observations and present to the class.

4. *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

Examples

$$2HCl + MgCO_3 \rightarrow MgCl_2 + H_2O + CO_2$$

$$H_2SO_4 + Ca(HCO_3)_2 \rightarrow CaSO_4 + 2H_2O + 2CO_2$$

$$2HNO_3 + Na_2CO_3 \rightarrow 2NaNO_3 + H_2O + CO_2$$

Experiment 2.6

Reactions of Acids with Carbonates and Hydrogen Carbonates

Objective: To investigate the reaction between acids and carbonates and bicarbonates.

Materials required: Sodium carbonate, calcium carbonate, calcium or magnesium or sodium bicarbonate, dilute HCl, dilute H_2SO_4, dilute HNO_3, lime water ($Ca(OH)_2$ solution) blue and red litmus paper, conical flask, delivery tube, rubber stopper, beaker.

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 2.1.

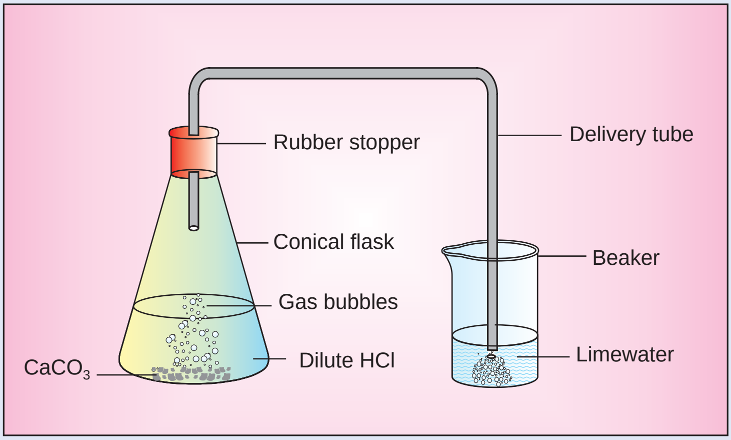


Figure 2.1 Test for Carbon dioxide.

2. Repeat the experiment using the same carbonates and hydrogencarbonate with dilute H_2SO_4 and dilute HNO_3.

Observations and analysis:

a What does the formation of bubbles in the conical flasks indicate?

b Does the colour of the damp blue litmus paper change when you hold 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 bubble 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.

5. *Concentrated acids react with sulphites liberating sulphur dioxide gas and forming a salt and water*

Sulphite + Acid → Sulphur dioxide + Salt + Water

Examples

$$Na_2SO_3 + 2HCl \rightarrow SO_2 + 2NaCl + H_2O$$

$$CaSO_3 + H_2SO_4 \rightarrow SO_2 + CaSO_4 + H_2O$$

6. *Acids neutralize basic oxides and bases or alkalis to form salt and water*

The reaction of acids with basic oxides or bases to form salt and water is called neutralization reaction.

Acid + Basic oxide → Salt + Water

Examples

$$2HCl + MgO \rightarrow MgCl_2 + H_2O$$

$$H_2SO_4 + CaO \rightarrow CaSO_4 + H_2O$$

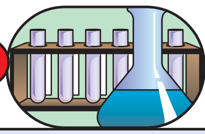
$$\textit{Acid} + \textit{Base} \rightarrow \textit{Salt} + \textit{Water}$$

Examples

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

$$HNO_3 + KOH \rightarrow KNO_3 + H_2O$$

Experiment 2.7



Neutralization Reaction

Objective: Investigate the reaction between acids and bases

Materials required: 1 M HCl, 1 M NaOH, 0.5 M $Ca(OH)_2$, red and blue litmus papers, 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.

Observations and analysis:

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?

Write a laboratory report on your observations and submit to your teacher.

7. ***Acids are electrolytes*** They conduct electricity in aqueous solutions.

Strength of Acids (Strong and Weak Acids)

Activity 2.9

1. Add a few drops of citric acid solution and nitric 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. You are allowed to taste some acids like acetic acid (in the form of vinegar), and citric acid at home, but you are never allowed to taste any kind of acids in the laboratory. What is the reason?

Discuss your findings in your group and present to the class.

Acids can be classified as strong acids and weak acids *depending on their degree of ionization in aqueous solution*, i.e., the extent to which the acids dissociate to form hydrogen or hydronium ions when they dissolve in water. A strong acid is one that ionizes almost completely in aqueous solution. Examples of strong acids are $HClO_4$, HI, HBr, HCl, HNO_3, and H_2SO_4. These acids dissociate to a very high extent as shown below:

$$HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-\,(aq)$$

$$HNO_3\,(aq) \rightarrow H^+\,(aq) + NO_3^-\,(aq)$$

$$HClO_4\,(aq) \rightarrow H^+(aq) + ClO_4^-\,(aq)$$

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 that releases few hydrogen ions in aqueous solution is a weak acid. The aqueous solution of a weak acid contains hydronium ions, anions and dissolved molecules of the acid.

Organic acids which contain the acidic carboxyl group –COOH, are generally weak acids. For example, CH_3COOH, ionizes slightly in water to give hydronium ions and acetate ions, CH_3COO^-.

$$CH_3COOH\ (aq) + H_2O(l) \rightleftharpoons CH_3COO^-\ (aq) + H_3O^+(aq)$$

Concentrated and Diluted Acids

Activity 2.10

1. In a solution, if there is 95% H_2O and the rest is HCl, what kind of solution is it?
2. In a car battery, the electrolyte used is 35% of H_2SO_4. How many percent of it is water? What do you think about the concentration of the electrolyte?

"Concentrated" and "dilute" are terms used to describe the relative amount of acid in a given acid solution.

A concentrated acid contains a relatively large amount (*percentage*) of acid and a small amount of water. For example, concentrated sulphuric acid is 98% H_2SO_4 and 2% water. Concentrated acetic acid is 99% acetic acid and 1% water. But, H_2SO_4 is a strong acid and CH_3COOH a weak acid.

A dilute acid contains a relatively small amount of acid dissolved in large amount of water. For example, a diluted solution of sulphuric acid may contain 10% H_2SO_4 and 90% H_2O.

The concentration of an acid is the measure of the number of moles of the acid in one litre of acid solution and is expressed in mol/L. This unit of concentration is called Molarity, denoted by M. Mathematically,

$$\text{Molarity (M)} = \frac{\text{Number of moles of the substance dissolved}}{\text{Volume of solution in litres}}$$

$$\text{Number of moles of the substance} = \frac{\text{Mass of the substance dissolved in grams}}{\text{Molar mass of the substance}}$$

For example, the concentration of 98% H_2SO_4 is about 18 mol/L. If one litre of this acid is added to 2 litres of water to make 3 litres of solution, the concentration will become 6 mol/L. The new solution obtained is a dilute solution of H_2SO_4. Thus, concentrated acids contains a greater number of moles of acid while dilute acids contain less number of moles of the acid per litre of the acid solution.

Example

How many moles of HCl are present in 0.8 L of a 0.5 *M* HCl solution?

Solution:

$M = \frac{n}{V} \Rightarrow n = MV = 0.5\ M \times 0.8\ \text{L} = 0.4$ mol HCl, where *M* is the molarity, *n*, the number of moles and *V*, volume of solution in litre.

An important aspect of the property of acids is that they conduct electricity in aqueous solutions, i.e., the acids are electrolytes. The extent of conduction of electricity depends on the amount of ions present in the solution. Thus, aqueous solutions of strong acids are better conductors of electricity than the same concentrations of weak acid solutions. The conduction of electricity through acid solution is also used as one means of identification between weak and strong acids.

Experiment 2.8

Can Aqueous Acid Solution Conduct Electricity?

Objectives: To investigate the conductivity of acids

Materials required: 1 M HCl, 1 M HNO_3, 1 M CH_3COOH, 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 2.2. 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.

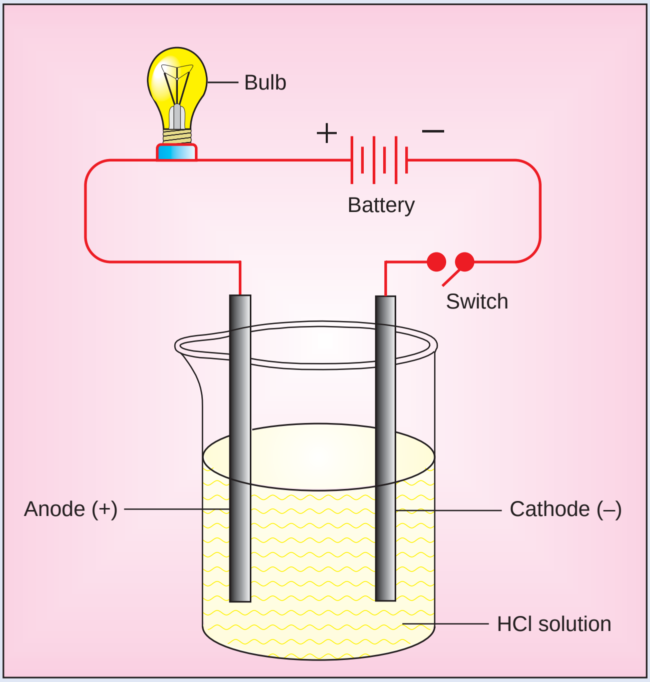


Figure 2.2 Conductivity of acid solution.

Observations and analysis:

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.

Precautions when Handling Acids

Concentrated acids are extremely corrosive and poisonous. They can destroy metals and clothes; produce a chemical burn on skin or inside the body. If taken internally they can be fatal. So acids must be handled with care.

The following precautions are helpful while working with 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 If concentrated acid is spilled on to cloth, immediately wash it with running water.

d If an acid enters your eye, wash with water repeatedly and then consult a doctor.

e If corrosive acids are swallowed, administer weak bases such as $Mg(OH)_2$ or $Al(OH)_3$.

f Use bellows to pipette acid instead of sucking using yours lips.

g To dilute concentrated acids, pour the concentrated acid in to water and not water in to the acid.

pH and pH scale

Expressing acidity as the concentration of H_3O^+ can be cumbersome because the values tend to be very small. A more convenient quantity called pH is used to indicate the concentration of H_3O^+. The letter pH stands for the French words *pouvoir hydrogène*, meaning "hydrogen power".

pH is defined as the negative logarithm of the *hydrogen ion* concentration (*hydronium ion concentration*). The mathematical expression is:

$$pH = \log \frac{1}{[H^+]} = -\log [H^+]$$

The square bracket [ ] means concentration in mol/L.

To calculate the hydrogen ion concentration of a solution from its pH we can proceed as follows.

$$[H^+] = \text{antilog}(-pH) = 10^{-pH}$$

The pH of a solution can be measured using universal indicator solution, pH indicator paper or a pH-meter. When universal indicator solution or pH indicator papers are added to an acid solution, they develop different colours depending on the pH of the solution. So, to identify the pH, we need to compare the colour developed with the standard colour chart. A pH meter is an electronic device, which reads the pH of a solution directly. (Figure 2.3).

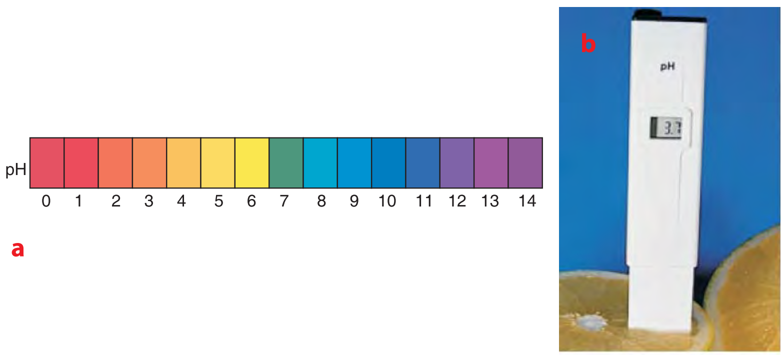
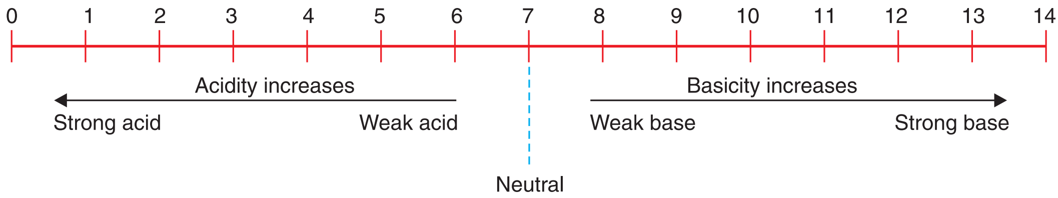


Figure 2.3 ***a) The colour of universal indicator in solutions of pH from 1 to 12, b) A pH meter with its electrical probe dipped into an orange.***

Acidity and alkalinity are usually expressed on the pH scale. The range is between 0 and 14.



A neutral solution has a pH value of 7 at 298 K. Acidic solutions have pH values less than 7. 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.

The following examples will guide us through how to calculate pH values of solutions.

Example

Calculate the pH of a 0.001 M solution of HCl?

Solution:

HCl is a strong acid and dissociates almost completely.

The dissociation of 0.001 M (0.001 mol/L) of HCl produces 0.001 mol/L of H^+ and 0.001 mol/L of Cl^- as follows:

$$HCl\,(aq) \rightarrow H^+(aq) + Cl^-\,(aq)$$

0.001 mol/L   0.001mol/L   0.001 mol/L

Thus, $[H^+] = 0.001\text{ mol/L} = 1 \times 10^{-3}\text{ M}$

$pH = -\log[H^+] = -\log 1 \times 10^{-3} = 3$

Example

The hydrogen ion concentration in a dilute solution of vinegar is 4.6×10^{-5} M. What is the pH of the solution?

Solution:

$pH = -\log[H^+] = -\log(4.6 \times 10^{-5}) = -(\log 4.6 + \log 10^{-5})$

$= -(\log 4.6 - 5) = -(0.663 - 5) = 4.337$

Example

The pH of lemon juice was determined to be 3. What is the hydrogen ion concentration in the lemon juice?

Solution:

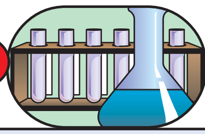
$pH = -\log[H^+]$; $\quad -pH = \log[H^+]$

$[H^+] = \text{antilog}(-pH) = 10^{-pH} = 10^{-3}M$

Exercise 2.7

1. The pH of a given solution is 8. What is the hydrogen ion concentration of the solution?
2. 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 2.9



pH of Solutions of Common Substances

Objective: To determine the pH of different substances

Materials required: Lemon juice, vinegar, tonic water, tomato juice, 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 pour 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.

Observations and analysis:

a What is your conclusion based on your observations?

b Are the substances used in this experiment acidic or neutral? Why?

c Record your observations using the following Table:

Substance	Colour developed	pH
Lemon juice		
Vinegar solution		
Tonic water		
Tomato juice		

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

Examples

SO_2 (g) + H_2O (l) → H_2SO_3 (aq)

N_2O_5 (s) + H_2O (l) → $2HNO_3$ (aq)

P_4O_{10} (s) + $6H_2O$ (l) → $4H_3PO_4$ (aq)

2. **Direct combination of some non-metals like S, Cl, Br with hydrogen:** This method is mostly used to prepare binary acids (acids consisting only two elements)

Examples

H_2 (g) + Cl_2 (g) → 2HCl (g)

H_2 (g) + Br_2 (g) → 2HBr (g)

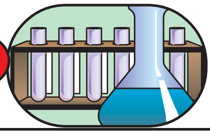
When gaseous hydrogen chloride and hydrogen bromide dissolve in water, they form hydrochloric acid and hydrobromic acid respectively.

3. **Using a non-volatile acid:** Volatile acids can be prepared by heating their salts with a non-volatile acid such as concentrated sulphuric acid, H_2SO_4. Hydrochloric acid HCl and nitric acid, HNO_3 can be prepared by this method according to the following equations.

   NaCl (s) + H_2SO_4 (l) → $NaHSO_4$ (s) + HCl (l)

   $NaNO_3$ (s) + H_2SO_4 (l) → $NaHSO_4$ (s) + HNO_3 (l)

Experiment 2.10



Preparation of Chlorous Acid

Objective: To investigate the product formed from $Ba(ClO_2)_2$ and H_2SO_4

Materials required: Two 250 mL beakers, glass rod, water, dropper, test tubes, test tube rack, blue and red litmus papers, methyl red, $Ba(ClO_2)_2$ and concentrated H_2SO_4.

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.

Observations and analysis:

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.

Reading assignment

Refer to chemistry books from the library and write the uses of sulphuric acid, nitric acid and hydrochloric acid.

Uses of some important acids

Activity 2.11

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.

Hydrochloric Acid, HCl is present naturally in the gastric juice of our body and helps in the digestion of food. Industrially, hydrochloric acid is important for pickling of iron and steel (*to remove surface impurities*) before galvanizing and tin plating. It is also used to produce aniline dyes, drugs, photographic films, plastics like polyvinyl chloride (PVC), and to recover magnesium from sea water.

Nitric Acid, HNO_3 is 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.

Sulphuric acid, H_2SO_4 is the leading industrial chemical. It is widely used in the production of sulphate and phosphate fertilizers, synthetic fibres, paints, drugs, detergents, paper and dyes. It is also used in petroleum refining, production of metals and as electrolyte in car batteries.

Exercise 2.8

1. Which one of the following substances are Arrhenius acids:

   a $HClO_4$ d BF_3

   b HI   e HF

   c HBr   f PCl_5

2. In each of the reactions listed below, identify the Brønsted-Lowry acid:

   a $CH_3COOH\ (aq) + H_2O\ (l) \rightleftharpoons CH_3COO^-\ (aq) + H_3O^+\ (aq)$

   b $HCN\ (aq) + SO_4^{2-}\ (aq) \rightleftharpoons CN^-\ (aq) + HSO_4^-\ (aq)$

   c $HF\ (aq) + NH_3\ (g) \rightleftharpoons F^-\ (aq) + NH_4^+\ (aq)$

   d $HClO_4\ (aq) + N_2H_4\ (aq) \rightleftharpoons ClO_4^-\ (aq) + N_2H_5^+\ (aq)$

3. Identify the Lewis acids in each of the following reactions:

   a $AlCl_3 + Cl^- \rightarrow AlCl_4^-$

   b $BF_3 + NH_3 \rightarrow F_3B - NH_3$

   c $SiCl_4 + 2Cl^- \rightarrow SiCl_6^{-2}$

   d $PF_5 + F^- \rightarrow PF_6^-$

4. 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 HF
   e HCl
   f HNO_2
   g HCN
   h HNO_3
   i CH_3COOH
   j H_2SO_3
   k H_2S
   l H_3PO_4

5. A reagent bottle (labelled as A) is filled with HCl solution and the other (labelled as B) is filled with water. Both liquids in the bottles are colourless. What method do you recommend to identify the acid and water?

6. What is the basis for the classification of acids as strong and weak?

7. 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

8. Calculate the hydrogen ion concentrations in solutions having the following pH values:

   a 4   b 2   c 5

9. How many moles of H_2SO_4 are present in 0.500 L of a 0.150 M H_2SO_4 solution?

2.4 BASES

Competencies

After completing this section, you will be able to:

- define bases in terms of the concepts of Arrhenius, Brønsted-Lowry and Lewis;
- give examples of bases based on Arrhenius, Brønsted-Lowry and Lewis concepts;

- explain 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 solution;
- calculate the concentration of hydroxide ions from the given information;
- conduct activities to investigate some chemical properties of bases;
- explain 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;
- conduct simple experiments to prepare bases in a laboratory;
- describe the uses of the three common laboratory bases (NaOH, $Ca(OH)_2$ and NH_3).

Activity 2.12

1. 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.
2. Discuss why farmers sometimes need to put crushed limestone on their soil.
3. While taking bath, have you ever had soap in your mouth? How does it taste?

Share your experience with your class.

Bases belong to the class of inorganic compounds that include most oxides and hydroxides of metals. 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.

Definition of Bases

There are different definitions of bases, as suggested by different chemists namely Arrhenius, Bronsted-Lowry and Lewis.

Arrhenius definition of bases: *A base is any substance that ionizes (dissociates) in aqueous solution to release hydroxide,* OH^- *ions*. Arrhenius bases are mostly ionic metal hydroxides such as, LiOH, NaOH, KOH and $Ba(OH)_2$. The dissociation of a few bases is shown by the following equations:

$$NaOH\ (aq) \rightarrow Na^+\ (aq) + OH^-\ (aq)$$

$$Ba(OH)_2\ (aq) \rightarrow Ba^{2+}\ (aq) + 2OH^-\ (aq)$$

Brønsted-Lowry definition of bases: *A base is a substance that is capable of accepting proton(s),* H^+*. Thus, a base is a proton,* H^+ *acceptor*.

Note:

Note that not all Brønsted-Lowry bases are Arrhenius bases, but all Arrhenius bases contain the hydroxide ion which is a Brønsted-Lowry base. This is because OH^- can accept a proton, H^+, to form water.

According to the Brønsted-Lowry definition, an acid-base reaction is the transfer of a proton from an acid to a base.

Example

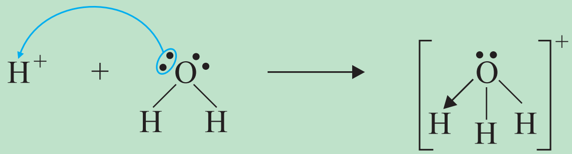
$$\underset{\text{Base}}{NH_3\ (aq)} + \underset{\text{Acid}}{HCl\ (aq)} \rightarrow NH_4^+\ (aq) + Cl^-\ (aq)$$

$$\underset{\text{Acid}}{HCO_3^-\ (aq)} + \underset{\text{Base}}{OH^-\ (aq)} \rightleftharpoons CO_3^{2-}\ (aq) + H_2O\ (l)$$

Lewis Definition of Bases

The Lewis definition is a more general definition than the Arrhenius and Brønsted-Lowry definitions. According to Lewis, *a base is a substance that is capable of donating a pair of electrons. In other words, a base is an electron-pair donor*. Brønsted-Lowry bases are also bases according to the Lewis concept. This is because for Brønsted-Lowry bases to accept a proton, H^+ they should have a lone pair. However, Lewis base include other species that can form a coordinate covalent bond by sharing their lone pair with substances other than a proton, H^+. It is important to note that Lewis acid-base reactions include many reactions that do not involve protons. The following equation shows an acid-base reaction according to Lewis concept:

Example



Lewis acid   Lewis base

Activity 2.13

Consider the following reactions:

$NH_3 + H^+ \rightarrow NH_4^+$

$BF_3 + F^- \rightarrow BF_4^-$

In the above reactions, name the Lewis acids and Lewis bases. Do NH_3 and F^- act as Lewis acid or Lewis base? Why?

General Properties of Bases

Activity 2.14

Dissolve one table-spoon of washing soda in a glass of water. Try to wash your hands with this solution. How do you feel? Does it give a soapy touch or an oily touch. Share your observations with the rest of the class.

Bases like NaOH, KOH, and $Ba(OH)_2$ are readily soluble in water, while others such as $Mg(OH)_2$ are slightly soluble. Bases that are soluble in water are also called alkalis.

Bases or alkalis show the following properties:

1. **Bases are slippery to the touch and have a bitter taste:** Bases have a bitter taste; feel slippery 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 in to contact with the skin nor tasted.
2. **Soluble bases change the colour of indicators:** Aqueous solutions of bases turn the colour of red litmus to blue, phenolphthalein to pink (*red*), methyl red to yellow and universal indicator blue (*purple*).

3. **Soluble bases 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.

Examples

$$KOH\,(aq) \rightarrow K^+\,(aq) + OH^-\,(aq)$$

$$Ca(OH)_2\,(aq) \rightleftharpoons Ca^{2+}\,(aq) + 2OH^-\,(aq)$$

4. **Bases neutralize acids or acidic oxides to form salt and water:**

Base + Acid (or acidic oxide) $\rightarrow$ Salt + Water

Examples

$$2NaOH\,(aq) + H_2SO_4\,(aq) \rightarrow Na_2SO_4\,(aq) + 2H_2O\,(l)$$

$$2KOH\,(aq) + SO_3\,(g) \rightarrow K_2SO_4\,(aq) + H_2O\,(l)$$

$$Ca(OH)_2\,(aq) + 2HCl\,(aq) \rightarrow CaCl_2\,(aq) + 2H_2O\,(l)$$

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.

Strength of Bases (Strong and Weak Bases)

Activity 2.15

Is a concentrated base a strong base and a dilute base a weak base? Discuss in your group and present to the class.

Both acids and bases are classified as strong and weak depending on their degree of ionization (*dissociation*) in aqueous solution.

Strong bases are those which ionize (*dissociate*) completely or almost completely in water solutions. Examples of strong bases include hydroxides of alkali (Group IA) metals and lower members of alkaline earth metals such as LiOH, NaOH, KOH, and $Ba(OH)_2$.

Examples

$$LiOH\,(aq) \rightarrow Li^+\,(aq) + OH^-\,(aq)$$

$$KOH\,(aq) \rightarrow K^+\,(aq) + OH^-\,(aq)$$

Weak bases are those bases which ionize (*dissociate*) only partially or slightly in aqueous solution. Examples of weak bases include NH_3, and $Ca(OH)_2$.

The dissociation of a weak base can be shown as follows using, NH_3, as an example.

$$NH_3\,(g) \quad + \quad H_2O\,(l) \quad \rightleftharpoons \quad NH_4^+\,(aq) \quad + \quad OH^-\,(aq)$$

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%.

Concentrated and Dilute Bases

Concentrated and dilute solutions are terms used to describe the amount of a substance present in a given volume of solution. The amount can be described in terms of either percentage or number of moles. 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 of a base is the measure of the number of moles of the alkali (*base*) dissolved in one litre of solution. The concentration is expressed in terms of mole per litre (*Molarity*).

$$\text{Molarity of the base} = \frac{\text{number of moles of the base}}{\text{volume of solution in litres}}$$

The greater the number of moles of the base per litre 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 litre. 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 litre.

Precautions when Handling Bases

Strong bases such as sodium hydroxide and potassium hydroxide can attack human skin, and even damage animal and plant tissues. That is why NaOH and KOH are named caustic soda and caustic potash respectively. The word “caustic” refers to a substance that can cause burning. Thus, it is very important to avoid the contact of bases with any part of our body or clothing. Not only strong bases but weak bases are also 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 table 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 in to 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

Activity 2.16

1. Since water, H_2O, has H^+ and OH^-, it is both acidic and basic. Is this true?
2. If $[OH^-]$ increases in a solution what happens to its $[H^+]$? Share your comments with your group.

pH is the measure of the hydrogen (H^+) or hydronium (H_3O^+) ion concentration in a solution. In the same manner, pOH is the measure of the hydroxide, OH^- ion concentration in a solution. It also measures the acidity or alkalinity of a solution as pH does.

pOH is defined as the negative logarithm of hydroxide, $OH^{”}$ ion concentration.

$$pOH = -\log [OH^-]$$

To calculate the hydroxide ion concentration from a given pOH value, we proceed as follows:

$$[OH^-] = \text{antilog}(-pOH) = 10^{-pOH}$$

Relationship Between pH and pOH

How can you derive the mathematical relationship between pH *and* pOH*?* To understand the relationship between pH and pOH let us start from the ionization of water. Water undergoes ionization to a small extent as follows:

$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$$

or

$$2H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) \quad + \quad OH^-\,(aq)$$

We can write a mathematical expression for the ionic product of water or dissociation constant for water, K_w as

$$K_w = [H^+]\,[OH^-] = [H_3O^+]\,[OH^-] \qquad (2.1)$$

$$K_w = 1.0 \times 10^{-14}\,(mol/L)^2 \text{ at } 25°C$$

Since water is neutral, $[H^+] = [OH^-]$ or $[H_3O^+] = [OH^-]$. So we can write equation 2.1 as

$$K_w = [H^+]\,[OH^-] = [H^+]^2 = 1.0 \times 10^{-14}(mol/L)^2$$

Solving for $[H^+]$

$$[H^+] = \sqrt{1.0\times10^{-14}\left(mol/L\right)^2} = 1.0\times10^{-7}\,mol/L$$

Thus, in any neutral aqueous solution at 25°C,

$$[H^+] = [OH^-] = 1.0 \times 10^{-7}\ mol/L$$

In an acid solution, $[H^+] > [OH^-]$ and in a basic solution, $[OH^-] > [H^+]$. However, the product $[H^+] \times [OH^-]$ remains constant, i.e., 1.0×10^{-14} at 25°C. Thus, if the concentration of either H^+ or OH^- is known in a given solution, we can calculate the concentration of the other.

$$[OH^-] = \frac{K_w}{[H^+]}; \qquad [H^+] = \frac{K_w}{[OH^-]}$$

To find the relationship between pH and pOH let us begin with equation 2.1.

$$K_w = [H^+]\,[OH^-] = 1.0 \times 10^{-14}\,(mol/L)^2$$

Take the negative logarithm of both sides

$$-\log K_w = -\log\,[H^+]\,[OH^-] = -\log 1.0 \times 10^{-14}$$

$$-\log K_w = -\log\,[H^+] + (-\log\,[OH^-]) = -\log 1.0 \times 10^{-14}$$

but, $-\log K_w = pK_w$,

$$-\log\,[OH^-] = pOH,$$

$$-\log\,[H^+] = pH$$

$$-\log 1.0 \times 10^{-14} = 14$$

So, the above equation becomes

$$pk_w = pH + pOH = 14 \text{ (at 25°C)} \qquad (2.2)$$

Thus, the sum of pH and pOH is 14 for any aqueous solution, at 25°C. Since pH, pOH, $[H^+]$ and $[OH^-]$ are interrelated through K_w, knowing any one of the values enables us to determine the others.

Table 2.2: Acidity and Basicity in relation to $[H^+]$, $[OH^-]$, pH and pOH.

Solution	$[H^+]$ (mol/L)	$[OH^-]$ (mol/L)	pH	pOH
Acidic solution	$> 10^{-7}$	$< 10^{-7}$	< 7	> 7
Basic solution	$< 10^{-7}$	$> 10^{-7}$	> 7	< 7

Example

The hydroxide ion concentration in a solution is 5.0×10^{-5}M? Calculate:

a pOH b $[H^+]$ c pH

Solution:

a To calculate pOH we proceed as follows.

$$pOH = -\log [OH^-] = -\log (5 \times 10^{-5}) = -(\log 10^{-5} - \log 5)$$

$$= 5 - \log 5 = 4.3$$

b $[H^+]$ is calculated using the relationship

$$[H^+] = \frac{K_W}{[OH^-]} = \frac{1.0 \times 10^{-14}\ (mol/L)^2}{5.0 \times 10^{-5}\ mol/L} = 2.0 \times 10^{-10}\ M$$

c $pH = -\log [H^+] = -\log 2.0 \times 10^{-10}$

$$= 10 - \log 2 = 9.7$$

Since the pOH of the solution is found in part (a), we can also calculate the pH using the relationship

$pH + pOH = 14$

$pH + 4.3 = 14$

$pH = 14 - 4.3 = 9.7$

Exercise 2.9

1. Determine the $[OH^-]$ in 3×10^{-5} M solution of HCl.
2. Determine the pH of 2.0×10^{-2} M $Sr(OH)_2$ solution assuming complete dissociation.

Experiment 2.11

Chemical Behaviour of Bases

Objectives: To investigate the thermal stability and reaction of bases with acids.

Materials required: NaOH or KOH, $Ca(OH)_2$, HNO_3, water, blue and red litmus papers, cobalt chloride, 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. Put 5 mL of the neutral solution on a watch glass and allow the water to evaporate until the next day.

Observations and analysis:

1. What type of reaction occurred between the KOH and HNO_3? Write the balanced chemical equation for the reaction.
2. What is left on the watch glass?
3. Which hydroxide melts on heating? Which base decomposes 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 to the class.

Preparation of Bases

Bases containing hydroxide ion (*hydroxide bases*) can be prepared by the following methods:

1. 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.

Examples

$2Li\ (s) + 2H_2O\ (l) \rightarrow 2LiOH\ (aq) + H_2\ (g)$

$2K\ (s) + 2H_2O\ (l) \rightarrow 2KOH\ (aq) + H_2\ (g)$

$Ca\ (s) + 2H_2O\ (l) \rightarrow Ca\ (OH)_2 + H_2\ (g)$

2. By the reaction of Group IA or Group IIA metal oxides with water, which gives the metal hydroxides.

Metal oxide + Water $\rightarrow$ Metal hydroxide

Examples

$Na_2O\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq)$

$CaO\ (s) + 2H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$

$BaO\ (s) + 2H_2O\ (l) \rightarrow Ba(OH)_2\ (aq)$

3. By double displacement reaction.

This method involves the reaction of an aqueous solution of a soluble base and a soluble salt, which gives another soluble base and an insoluble salt as products.

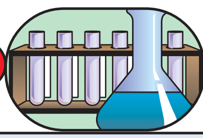
Soluble base + Soluble salt $\rightarrow$ Another Soluble base + Insoluble salt

Examples

$Ba(OH)_2\ (aq) + K_2SO_4\ (aq) \rightarrow 2KOH\ (aq) + BaSO_4\ (s)$

$Ca(OH)_2\ (aq) + Na_2CO_3\ (aq) \rightarrow 2NaOH\ (aq) + CaCO_3\ (s)$

Experiment 2.12



Preparation of Bases

Objective: To prepare bases from metal and metal oxides.

Materials required: Water, trough, a knife, tongs, filter paper, litmus paper, test tubes, test tube rack, calcium or sodium metal, CaO or MgO, Na_2SO_4

or K_2SO_4, $Ba(OH)_2$ solution, spatula, balance, three beakers, measuring cylinder, dropper, glass rod.

Procedure:

1. Pour clean water into a trough to three-fourth of its volume. Then take a piece of sodium metal or calcium metal from the bottle with tongs. If you use sodium, place it on filter paper and then blot it to remove the oil. Carefully cutoff a very small piece with a knife and drop this piece onto the surface of the water in the trough. Observe the reaction.

   After the reaction is complete, take some solution from the trough and add it to two test tubes. Test the solution in the first test tube using red litmus paper and add a few drops of methyl orange to the second.

2. Add a half a spatula measure of CaO or 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?

3. Add 6 g Na_2SO_4 or 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 Na_2SO_4 or K_2SO_4 solution, mix it with the solution in the third beaker and shake well. Observe what is happening. Continue adding Na_2SO_4 or 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 Na_2SO_4 or K_2SO_4 solution and test whether the solution in the third beaker is acidic or basic using litmus paper.

Observations and analysis:

1. Does the colour of red litmus paper change when dipped into the solutions obtained in procedure 1 and 2? Are the solutions acidic or basic?
2. What gas is given off during the reaction in procedure 1?
3. Why do you think is the white precipitate formed?
4. What is the nature of the final solution in the third beaker?
5. Write balanced chemical equations for the reactions.

Write a laboratory report on your observations and present to the class.

Uses of Some Important Bases

Sodium hydroxide, NaOH: This is known by the common name lye or caustic soda. It is one of the most important laboratory and industrial chemicals. It is used in the laboratory for absorbing carbon dioxide and other acidic gases, in a number of organic reactions and in chemical analysis. Industrially, it is used in the manufacture of soaps and detergents, pulp and paper, textiles, dyes, cosmetics, pharmaceuticals, in purifying aluminium ore and petroleum refining. It is also used to clean drains since it dissolves grease and other organic matter.

Potassium hydroxide, KOH: It is commonly known as caustic potash. It is a stronger base and is more expensive than NaOH and has limited uses. It is used as an electrolyte in some dry cells and to make soft soap.

Calcium hydroxide, $Ca(OH)_2$: It is also called slaked lime or lime water when it is dissolved in water. It is a weaker base than NaOH and KOH, and is widely used to remove soil acidity, to make mortar (*mixture of lime, sand and water*) which serves as binding material for bricks and plastering walls, to remove temporary hardness of water and in manufacturing bleaching powder. Its aqueous solution is also used to test for carbon dioxide gas.

Ammonia: A solution of gaseous ammonia, NH_3 in water is used as household cleaning agent and as a laboratory reagent. It is used to manufacture ammonium fertilizers,

Reading Assignment

The common uses of NaOH, KOH, $Ca(OH)_2$ and NH_3 have been described above. After reading reference books in your library, search for some more uses of these compounds. Present your findings to the class.

Exercise 2.10

1. Identify the Brønsted-Lowry bases and acids in each of the following chemical reactions:

   a $HCOOH(aq) + CN^-(aq) \rightleftharpoons HCOO^-(aq) + HCN(aq)$

   b $H_2S(g) + OH^-(aq) \rightleftharpoons HS^-(aq) + H_2O(l)$

2. Identify the substances acting as Lewis bases and acids in each of the following chemical reactions:

   a $AlCl_3(aq) + Cl^-(aq) \rightarrow AlCl_4^-(aq)$

   b $H^+(aq) + NH_3(g) \rightarrow NH_4^+(aq)$

3. What is the concentration of the base (in molarity) containing:
   a 80 g NaOH is dissolved in water to make 1 litre of solution?
   b 224 g KOH is dissolved in water to make 2 litres of solution?
4. Calculate the pOH, $[H^+]$ and pH of the solutions having the following hydroxide ion concentrations:
   a $[OH^-] = 4.6 \times 10^{-3}$ mol/L
   b $[OH^-] = 2.5 \times 10^{-6}$ mol/L
5. Potassium hydroxide solution is added from a burette to dilute hydrochloric acid solution in a beaker. Which of the following changes will not occur in the beaker:
   a The pH of the solution in the beaker increases.
   b The $[H^+]$ of the solution in the beaker decreases.
   c The pOH of the solution in the beaker increases.
6. How many moles of NaOH are contained in 65 mL of a 2.20 M solution of NaOH in H_2O?
7. A solution of barium hydroxide, $Ba(OH)_2$, contains 4.285 g of barium hydroxide in 100.0 mL of solution. What is the molarity of the solution?

2.5 SALTS

Competencies

After completing this section, you will be able to:

- define salts;
- give examples of salts;
- classify salts as acidic, normal and basic salts;
- explain the direct combination of elements, the reaction of acids with bases and the reaction of acids with metals as methods of salt preparation;
- conduct simple experiments to prepare a salt by neutralization;
- list some important salts and discuss their uses;
- explain the properties of salts;
- describe the chemical tests of some salts by conducting activities;
- mention the essential nutrients of plants;
- describe the functions of nitrogen, phosphorus and potassium in plant growth;

- define fertilizers;
- list some common fertilizers;
- explain the importance of fertilizers; and
- list some common inorganic compounds that are used as pesticides and herbicides.

Activity 2.17

Discuss the following in your group and present it to the class:

1. List all the salts that you know and tell where they are found.
2. Make a list of coloured salts available in your school laboratory.
3. Which salts are found in Dalol area, Ethiopia?

Salts are ionic compounds that contain positive ions (cations) derived from bases and negative ions (anions) derived from acids. Salts are also defined as ionic compounds formed when the ionizable hydrogen of acids are partly or completely replaced by metal ions or ammonium ions. For example, $Ca(HCO_3)_2$ and $CaCO_3$ are salts derived from the acid H_2CO_3 while the salts $Mg(H_2PO_4)_2$, $MgHPO_4$ and $Mg_3(PO_4)_2$ are salts derived from H_3PO_4.

- A salt is produced when an acid is neutralized by a base:

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

Examples

$$HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l)$$

$$H_2SO_4\,(aq) + Ca(OH)_2\,(aq) \rightarrow CaSO_4(s) + 2H_2O\,(l)$$

Some common examples of salts are sodium chloride (NaCl), calcium carbonate ($CaCO_3$), potassium nitrate (KNO_3), copper sulfate ($CuSO_4$) etc.

Classification of Salts

Activity 2.18

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 ion?

b What would be the nature of the salt, if all the base hydroxide ions are not replaced by the anions of the acid?

Discuss in your group and present the findings to the class.

Salts can be classified as:

- **Acidic salts:** These are salts in which not all of the hydrogen ions in an acid have been replaced by metal ions or ammonium (NH_4^+) ions.

  Example: Sodium hydrogen sulphate, $NaHSO_4$.

  When acidic salts are dissolved to make an aqueous solution, they release H^+ ions which make the solution acidic. For example, $NaHSO_4$ releases ions in aqueous solution as follows:

  $$NaHSO_4\,(aq) \rightarrow Na^+\,(aq) + H^+\,(aq) + SO_4^{2-}\,(aq)$$

- **Normal salts:** These are salts in which all of the hydrogen ions in an acid have been replaced by metal ions or ammonium ions.

  Example: Sodium sulphate, Na_2SO_4.

  Normal salts do not contain ionizable H^+ or OH^-.

Example

Na_2SO_4 release ions as follows:

$$Na_2SO_4\,(aq) \rightarrow 2Na^+\,(aq) + SO_4^{2-}\,(aq)$$

- **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.

  Example: Zinc hydroxychloride, Zn(OH)Cl.

  When basic salts are dissolved to make an aqueous solution, they release OH^- ions and this makes the solution basic.

  $$Zn(OH)Cl\,(aq) \rightarrow Zn^{2+}\,(aq) + OH^-\,(aq) + Cl^-\,(aq)$$

Preparation of Salts

Activity 2.19

List a few methods to prepare sodium chloride in the laboratory? Discuss in groups and present it to the class.

Salts can be prepared by the following methods:

1. **Direct combination of elements:** This method is used to prepare binary salts, that is, salts consisting of only two elements.

Examples

2Na (s)	+	Cl_2 (g)	→	2NaCl (s)
Fe (s)	+	S (s)	→	FeS (s)
Fe (s)	+	$3Cl_2$ (g)	→	$2FeCl_3$ (s)

2. **Reacting active metals with dilute acids:**

Active metal + Acid → Salt + Hydrogen gas

Examples

2Al (s)	+	6HCl (aq)	→	$2AlCl_3$ (aq)	+	$3H_2$ (g)
Mg (s)	+	2HCl (aq)	→	$MgCl_2$ (aq)	+	H_2 (g)
Zn (s)	+	H_2SO_4 (aq)	→	$ZnSO_4$ (aq)	+	H_2 (g)

Since, the reactions of metals of Group IA like sodium and potassium are very vigorous, it is advisable not to use the metals of Group IA for the preparation of salts by this method.

Sometimes, the reaction of active metals with dilute acids to prepare salts may not be as successful as expected. For example, calcium sulphate is only sparingly soluble in water. So in the reaction with dilute sulphuric acid, calcium tends to become coated with calcium sulphate which inhibits the reaction from proceeding.

3. **Reacting metal oxides with acids** (*by neutralizing basic oxides with dilute acids*):

Metal oxide + Dilute acid → Salt + Water

Examples

CaO (s)	+	2HCl (aq)	→	$CaCl_2$ (aq)	+	H_2O (l)
CuO (s)	+	H_2SO_4 (aq)	→	$CuSO_4$ (aq)	+	H_2O (l)
ZnO (s)	+	$2HNO_3$ (aq)	→	$Zn(NO_3)_2$ (aq)	+	H_2O (l)

4. **Reacting metal hydroxides with dilute acids:** This reaction is commonly called neutralization.

Metal hydroxides + Dilute acid → Salt + Water

Examples

$NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

$2KOH\ (aq) + H_2SO_4\ (aq) \rightarrow K_2SO_4\ (aq) + 2H_2O\ (l)$

5. **Reacting metal carbonates or hydrogen carbonates with dilute acids:** The reaction of carbonates or hydrogen carbonates with dilute acids always releases carbon dioxide gas along with the formation of a salt and water.

Metal carbonates + Dilute acid → Salt + Water + Carbon dioxide

Examples

$MgCO_3\ (s) + H_2SO_4\ (aq) \rightarrow MgSO_4\ (aq) + CO_2\ (g) + H_2O\ (l)$

$CuCO_3\ (s) + 2HCl\ (aq) \rightarrow CuCl_2\ (aq) + H_2O\ (l) + CO_2\ (g)$

$Na_2CO_3\ (s) + 2HNO_3\ (aq) \rightarrow 2NaNO_3\ (aq) + H_2O\ (l) + CO_2\ (g)$

$Ca(HCO_3)_2\ (s) + 2HCl\ (aq) \rightarrow CaCl_2\ (aq) + 2H_2O\ (l) + 2CO_2\ (g)$

6. **Double decomposition reaction** (*double displacement reaction*): In this method, two different soluble salts react to form a soluble and an insoluble salt as products that can be separated easily.

Examples

$Pb(NO_3)_2\ (aq) + NaI\ (aq) \rightarrow PbI_2\ (s) + 2NaNO_3\ (aq)$

$CaCl_2\ (aq) + Na_2CO_3\ (aq) \rightarrow CaCO_3\ (s) + 2NaCl\ (aq)$

It may be noted that each method of preparing salts cannot be used to make every salt; i.e., different methods may be used to make different salts.

Experiment 2.13

Preparing salts from Metallic Oxides and Acid

Objective: To prepare copper sulphate from copper oxide and sulphuric acid.

Materials required: Copper oxide, sulphuric acid, beaker, filter funnel, filter paper, evaporating dish.

Procedure:

1. Pour some sulphuric acid in beaker and add copper (II) oxide, warm and stir gently.
2. When the reaction is complete, filter the solution to remove excess copper oxide.
3. Transfer the filtrate to the evaporating dish and evaporate some of the water from the filtrate. Leave for the salt to crystallize out.

Observations and analysis:

1. What is the product formed? Write a balance chemical equation for the reaction.
2. What is the colour change as the reaction occurs? What does this show?

Write a laboratory report and present to the class.

Some Important Salts and their Uses

Activity 2.20

1. Have you ever taken Oral Rehydration Salt (ORS) for diarrhoea? What are the contents of the ORS?
2. Can you prepare your own ORS?

Discuss in your group and present to the class.

Sodium chloride (common salt, NaCl) is widely used in the preparation and preservation of food, as raw a material for the manufacture of sodium, chlorine, and sodium hydroxide. It is used as component of Oral Rehydration Salt (ORS) that has medical applications. It is also used to manufacture baking soda ($NaHCO_3$) and Na_2CO_3 that serve as raw materials for the production of glass.

Ammonium nitrate, (NH_4NO_3) is used as a nitrogenous fertilizer and in explosives.

Copper (II) sulphate, ($CuSO_4$) is 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. $CuSO_4$ is also useful in electroplating.

Iron (III) chloride, ($FeCl_3$) is used in the treatment of waste water, and in etching printed circuits.

Potassium nitrate, (KNO_3) is used in making gun powder (*mixture of* KNO_3, *carbon and sulphur*) and other explosives. It is also used as a fertilizer.

Calcium sulphate, ($CaSO_4 \cdot 2H_2O$, Gypsum) is used for plastering walls and supporting fractured bones.

Barium sulphate, ($BaSO_4$) is given to patients as a "barium meal" before gastro-intestinal x-ray photography; it is also used as a white pigment.

Iron (II) sulphate ($FeSO_4$) is given as iron tablets to patients who suffer from anaemia.

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. The following table summarizes some simple observations about the solubility of some groups of salts.

Table 2.2 Solubility of salts.

Soluble salts	Exceptions
• All nitrates (NO_3^-), perchlorates (ClO_4^-), chlorates (ClO_3^-) and acetates (CH_3COO^-)	• No exception
• Most halides	• Those containing Ag^+, Hg^{2+}, Pb^{2+} ($AgCl$, $AgBr$, AgI, $HgCl_2$, $HgBr_2$, HgI_2, $PbCl_2$, $PbBr_2$, PbI_2) ($PbCl_2$ is soluble in hot water)
• Most sulphates	• $BaSO_4$, $HgSO_4$, $PbSO_4$, $SrSO_4$
• Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}) and sulphides (S^{2-}) of alkali metals and ammonium ion (NH_4^+)	• All carbonates, phosphates and sulphides of other elements are insoluble
• Sulphides of alkaline-earths, Group IIA, elements	

2. Tendency to absorb water from the atmosphere or release water to the atmosphere.

Depending on their tendency to absorb water from or release water to the atmosphere, salts can be classified as hygroscopic, deliquescent and efflorescent.

Hygroscopic salts are those which absorb water from the atmosphere but remain solid.

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.

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.

Table 2.3 Examples of hygroscopic, deliquescent and efflorescent salts.

Deliquescent	Hygroscopic	Efflorescent
• Calcium chloride , $CaCl_2$	• Anhydrous Copper(II) sulphate, $CuSO_4$	• Hydrated sodium carbonate $Na_2CO_3{\cdot}10H_2O$
• Sodium Nitrate, $NaNO_3$		• Hydratedsodium sulphate, $Na_2SO_4{\cdot}10H_2O$
• Iron (III) chloride, $FeCl_3$		
• Magnesium chloride, $MgCl_2$		
• Lithium nitrate, $LiNO_3$		
• Sodium nitrate, $NaNO_3$		

3. Aqueous solutions of soluble salts are good conductors of electricity, because they release mobile positive and negative ions in solution.

Examples

$$NaNO_3\,(aq) \rightarrow Na^+\,(aq) + NO_3^-\,(aq)$$

$$CaCl_2\,(aq) \rightarrow Ca^{2+}(aq) + 2Cl^-\,(aq)$$

4. **Thermal stability of salts:** When different salts containing the same anion are heated, they may not show similar behaviour. Some salts are thermally stable while others undergo decomposition. The following examples illustrate this fact.

a Thermal decomposition of carbonates:

Carbonates are salts containing a carbonate, as an anion. Carbonates of Group IIA and most transition metals decompose on heating to give the metal oxides and carbon dioxide:

$$MgCO_3\,(s) \xrightarrow{\text{heat}} MgO\,(s) + CO_2\,(g)$$

$$CuCO_3\,(s) \xrightarrow{\text{heat}} CuO\,(s) + CO_2\,(g)$$

However, carbonates of sodium and potassium are not decomposed by heat.

b Thermal decomposition of nitrates

Nitrates are salts containing a nitrate, as an anion. Heating nitrates of Group IIA and most transition metals produces a metal oxide, nitrogen dioxide and oxygen:

$$\text{Nitrate} \xrightarrow{\text{heat}} \text{Metal oxide} + \text{Nitrogen dioxide} + \text{Oxygen}$$

Examples

$$2Mg(NO_3)_2\,(s) \xrightarrow{\text{heat}} 2MgO\,(s) + 4NO_2\,(g) + O_2\,(g)$$

$$2Pb(NO_3)_2\,(s) \xrightarrow{\text{heat}} 2PbO\,(s) + 4NO_2\,(g) + O_2\,(g)$$

Nitrates of sodium and potassium decompose on heating to give nitrites and oxygen gas;

$$2NaNO_3\,(s) \xrightarrow{\text{heat}} 2NaNO_2\,(s) + O_2\,(g)$$

$$2KNO_3\,(s) \xrightarrow{\text{heat}} 2KNO_2\,(s) + O_2\,(g)$$

The carbonate and nitrate of lithium differ from those of sodium and potassium; they decompose on heating in the following manner:

$$4LiNO_3\,(s) \xrightarrow{\text{heat}} 2Li_2O\,(s) + 4NO_2\,(s) + O_2\,(g)$$

$$Li_2CO_3\,(s) \xrightarrow{\text{heat}} Li_2O\,(s) + CO_2\,(g)$$

Chemical Tests of Some Ions in Salts

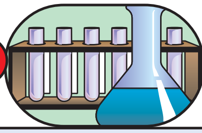
Activity 2.21

1. Do you know why fireworks give different colours? 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?

There are several simple tests which may help in the identification of metal ions and anions present in salts.

1. **Flame tests:** Certain metals give a characteristic colour 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 2.14



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.

Materials required: Platinum or nichrome wire, watch glass, Bunsen burner, distilled water, salts containing Li^+, Na^+, K^+, Ca^{2+}, Sr^{2+} and Ba^{2+}.

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.

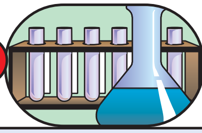
Observations and analysis:

a Write the colour of the flame produced in the following table.

Metal ion in the salt	Colour of flame produced
Lithium (Li^{+})	
Sodium (Na^{+})	
Potassium (K^{+})	
Calcium (Ca^{2+})	
Strontium (Sr^{2+})	
Barium (Ba^{2+})	

Write a laboratory report and present to the class.

Experiment 2.15



Test for cations-II

Objectives: To identify the presence of Cu^{2+}, Fe^{2+} and Fe^{3+} in salts.

Materials required: Beakers, test tubes, test tube rack, glass rod, ammonia solution, sodium hydroxide solution, salts containing each of Cu^{2+}, Fe^{2+} and Fe^{3+} ions

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.

Observations and analysis:

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:

Colour observed	Confirms the presence of

3. Write balanced chemical equations for the reactions.

Write a laboratory report in group and present to the class.

Experiment 2.16

Test for Halide Ions

Objectives: To identify the presence of Cl^-, Br^- and I^- ions.

Materials required: test tubes, test tube rack, NaCl, NaBr, NaI, $AgNO_3$, ammonia solution, reagent bottles, dilute HNO_3

Procedure:

1. Prepare 125 mL solution by dissolving 2.0 g $AgNO_3$ in the first, 1.59 g NaCl in the second, 2.50 g NaBr in the third and 3.75 g NaI in the fourth reagent bottle respectively.
2. Take three test tubes and pour about 5 mL of NaCl solution in to the first, 5 mL NaBr solution in to the second and 5 mL NaI solution in to the third. To each of the solutions in the test tube add 1 mL of dilute HNO_3 followed by addition of about 5 mL $AgNO_3$ solution. Observe if a precipitate is formed in each beaker.
3. After a precipitate has been formed, add ammonia solution to each test tube and see what happens to the precipitate.

Observations and analysis:

1. What did the formation of the white precipitate in the first test tube confirm? Name the compound formed as the white precipitate.
2. What did the formation of the yellow precipitate in the second test tube confirm? Name the compound formed as the yellow precipitate.
3. What did the formation of the deep-yellow precipitate in the third test tube confirm? Name the compound formed as the deep-yellow precipitate.

4. What happened to the contents in the three test tubes, when ammonia solution was added to each of the test tubes?
5. Write balanced chemical equations for the reactions taking place in each test tube.

Write a laboratory report and present to the class.

Experiment 2.17

Test for Sulphates

Objective: To identify for the presence of sulphate using barium salts.

Materials required: Beakers, test tubes, and test tube rack, any soluble sulphate salt (such as sodium sulphate), barium chloride or barium nitrate solution, and dilute HCl.

Procedure:

Add some sodium sulfate 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.

Observations and analysis:

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 2.18

Test for Carbonates and Hydrogen Carbonates

Objective: To distinguish betweem carbonates and hydrogen carbonates

Materials required: Na_2CO_3, $NaHCO_3$, dilute HCl, lime water, conical flasks, beakers, rubber stopper, gas delivery tube, calcium chloride solution

Procedure:

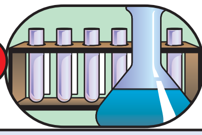
1. Take 20 mL solution of Na_2CO_3 and add it to one conical flask and 20 mL $NaHCO_3$ solution to another. Add the same amount of dilute HCl to each of the conical flasks; Fit a rubber stopper to which a delivery tube is inserted to each conical flask. Allow the gas produced to pass through lime water 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.

Observations and analysis:

1. What has happened to the lime water in each case?
2. Which solution forms a white precipitate on the addition of $CaCl_2$ solution?
3. Write balanced chemical equations for all the reactions.

Write a laboratory report and present to the class.

Experiment 2.19



Test for Nitrates (Brown Ring Test)

Objectives: To identify the presence of nitrate in a solution.

Materials Required: Test tube, test tube rack, beaker, nitrate solution, iron (II) sulphate solution, concentrated H_2SO_4

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.

Observations and analysis:

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 and present to the class.

Exercise 2.11

1. Name the acid and base pairs required to prepare each of the following salts:
   a K_2SO_4
   b $Ca(NO_3)_2$
   c $MgBr_2$
   d BaI_2
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 $FeCO_3$
   c $AgNO_3$
   d $PbSO_4$
   e Li_3PO_4
   f $AgCl$
   g $CaCl_2$
   h Na_2S
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+} ions?
   c Sulphate, SO_4^{2-} ions?
8. An unknown salt when heated with a Bunsen flame produced a bright-yellow flame. An aqueous solution of this salt when reacted with dilute HCl released a gas that turns lime water milky. Upon addition of calcium chloride solution to an aqueous solution of this salt, a white precipitate is formed. What is the unknown salt?

Plant Nutrients

Essential nutrients

Activity 2.22

1. Form a group and list the kind of fertilizers you are familiar with. Which of the listed fertilizers are inorganic salts?

2. Plants need nitrogen for their growth, and there is plenty of nitrogen in the air. If this is so, why do we apply nitrogeneous fertilizers for plants?
3. 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? Discuss in groups and present to the class.

The growth and development of plants is determined by various factors of soil, climate, and by factors inherent in the plants themselves. Minerals required by plants for their growth and development are called plant nutrients. For a soil to produce crops successfully, it must have an adequate supply of all the necessary nutrients that plants take from the soil. The required nutrient elements must be present in the soil in forms that plants can use and also there should be a rough balance among them in accordance with the amounts needed by plants. When any one of these elements is lacking or is in excess of the required proportion, the normal plant growth will not take place.

Elements required by plants for their growth and development are called essential nutrients. The elements considered as essential nutrients are carbon, hydrogen, oxygen, sulphur, nitrogen, phosphorus, potassium, calcium, magnesium, iron, manganese, boron, copper, zinc, molybdenum, cobalt and chlorine. All these seventeen elements have their own roles in the growth and development of plants. Depending on the amount required by plants, these elements are classified as macronutrients and micronutrients.

Activity 2.23

Try to identify the (a) micronutrients, (b) macronutrients, required for the growth of wheat, maize, and rice crops, by consulting the books in your library. Share your findings with the class.

Macronutrients are the elements required in relatively large amounts by plants. These are nitrogen, phosphorus, potassium, calcium, magnesium, sulphur, carbon, hydrogen and oxygen.

Micronutrients 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. The roles of nitrogen, phosphorus and potassium in the plant growth are discussed below:

Nitrogen is a very important plant nutrient. It is absorbed by plants in the form of nitrate ions (NO_3^-). In the growth and development of plants, nitrogen is utilized in the synthesis of amino acids, proteins, coenzymes and nucleic acids. The proteins formed control the metabolic processes that occur in plant cells. Nitrogen is also involved in the synthesis of chlorophyll to produce a deep green colour.

Phosphorus is an important element without which the growth of plants is impossible. It facilitates early growth and root formation, quick maturity and promotes seed or fruit production. Phosphorus is absorbed by plants mainly in the form of $H_2PO_4^-$ and HPO_3^{2-} in small amounts. Phosphorus 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 is absorbed by plants in the form of the K^+ ion. It is a component of enzymes that facilitate the process of photosynthesis and protein synthesis. It also adjusts stomata movement. It is the most important ion in controlling the turgidity of plant cells. It 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 2.12

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?

Fertilizers

During the growth and development of plants, the plants consume the essential nutrients from the soil. The concentration of these nutrients will decrease in the soil. If the plants are specially food crops, the nutrients can be lost from the soil when the plants are removed on harvest. So for the soil to be suitable for the growth of plants in the next round, the nutrients lost must be replaced.

How can we replace the nutrients lost from the soil? How can we increase the crop producing potential of soils? The nutrients lost from the soil can be replaced or the crop producing potential of soils can be increased by adding the appropriate fertilizers to the soil.

What are fertilizers? Fertilizers are materials that are added to soils to increase the growth, yield or nutrient value of crops. Fertilizers are usually classified as organic (*natural*) fertilizers and synthetic or artificial fertilizers.

Organic (natural) fertilizers are those derived from animals and plants. This includes animal dung, urine and substances obtained from the decay and decomposition of plants. These fertilizers provide readily available nutrients to plants after some period of decay and decomposition. These fertilizers can supply the elements nitrogen, phosphorus, potassium and various trace elements to the soil.

Synthetic or artificial fertilizers contain all three major plant nutrients i.e., nitrogen, phosphorous and potassium, and are known as complete fertilizers.

A mixed artificial fertilizer may contain nitrogen, phosphorus and potassium. Such a type of fertiliser is called a complete fertiliser. 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 the 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 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 2.13

1. Classify the following as organic or synthetic fertilizers:

   a NH_4NO_3 b urea c $(NH_4)_2SO_4$

2. Can the same fertilizer be applied to all types of soil? Explain.

Pesticides

Pesticides are chemicals that can be applied to agricultural crops to kill pests which affect the growth and development of plants. Parasitic fungi, insects and animals like rodents, rabbits and birds reduce crop yields or even completely destroy crops.

Pesticides are used to overcome the damage caused by fungi and pests. The majority of the pesticides used in agriculture are organic compounds. However, there are some inorganic substances that serve as pesticides. These include fumigant insecticides that exist in the vapour phase. Examples of fumigants are hydrogen cyanide (HCN), sulphur dioxide (SO_2), carbon disulphide (CS_2) and phosphine (PH_3). These pesticides are mostly used in the treatment of empty transport containers, grain stores, warehouses, harvested products prior to or during storage and to destroy pests in the soil. The other inorganic pesticides are mostly copper compounds such as $CuCl_2$, $Cu(OH)_2$, $Cu(NO_3)_2$, $CuSO_4$, Cu_2O and $CuCO_3$. Among the inorganic compounds, $CuSO_4$ and Cu_2O are used as *fungicides* where as HCN, SO_2, PH_3, $CuCl_2$, $Cu(OH)_2$, $Cu(NO_3)_2$ and $CuCO_3$ are used as insecticides.

Herbicides

Herbicides are also called weed killers. They are used to control the damage caused by weeds. Inorganic compounds used as herbicides include ammonium sulphamate ($(NH_4)SO_3–NH_2$), borax ($Na_2B_4O_7·10H_2O$), sodium chlorate ($NaClO_3$) and ferrous sulphate ($FeSO_4·7H_2O$).

Check list

Key terms of the unit

- *acidic oxide*
- *amphoteric oxide*
- *Arrhenius acid*
- *Arrhenius base*
- *acid, basic and normal salt*
- *basic oxide*
- *Bronsted - Lowery acid*
- *Bronsted - Lowery base*
- *concentrated acid*
- *dilute acid*
- *deliquescent*
- *dissociation constant for water (kw)*
- *efflorescent*
- *electrolyte*
- *flame test*
- *hygroscopic*
- *indicator*
- *Lewis acid*
- *Lewis base*
- *litmus*
- *molarity*
- *neutral oxide*
- *neutralization*
- *non-volatile acid*
- *peroxide*
- *pH*
- *plant nutrient*
- *pOH*
- *salt*
- *strong acid*
- *strong base*
- *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.*
- *A Brønsted-Lowery acid is a proton-donor.*
- *A Brønsted-Lowery base is a proton-acceptor.*
- *A Lewis acid is an electron-pair acceptor.*
- *A Lewis base is an electron-pair donor.*
- *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* $pH = -\log [H^+]$.
- pOH *is the negative logarithm of hydroxide ion concentration* $pOH = -\log [OH^-]$.
- $K_w = [H^+][OH^-] = 1 \times 10^{-14}$ *at 25°C.*
- $pK_w = pH + pOH = 14$ *at 25°C.*
- *A 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 substances absorb water from the atmosphere.*
- *Deliquescent substances absorb water from the atmosphere and dissolve in the water absorbed to form solutions.*
- *Efflorescent substances lose their water of crystallisation 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 utilised 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 EXERCISE ON UNIT 2

Part I: Write the missing words in your exercise book

Acidic	*litmus*	*salt*	*ammonia*	*Alkalis*
react	*acids*	*hydroxide*	*soluble*	

Most oxides of non-metals are __1__ oxides, acidic oxides react with __2__ to form a __3__ and water. Some acidic oxides __4__ with water to form solutions of acids. These solutions turn blue __5__ red. Alkalis are __6__ bases. Examples of alkalis are __7__ and sodium __8__. Alkalis react with __9__ to form a __10__ and water.

Part II: A Match the substances on the left with those properties stated on the right

1. Vinegar (pH ≈ 4.5) a strongly acid
2. dish-washing powder (pH ≈ 12) b weakly alkaline
3. soap (pH ≈ 7.5) c neutral
4. distilled water (pH ≈ 7) d strong alkaline
5. concentrated hydrochloric acid e weakly acidic

Part II: B Match the ions or molecules on the left with the correct test reagent on the right

6. Iron (II) ion a silver nitrate
7. Iodide ion b acidified barium nitrate
8. Sulphate ion c heat with aluminium
9. Carbon dioxide d iron (II) sulphate and sulphuric acid
10. Nitrate ion e sodium hydroxide

f limewater

Part III: Choose the correct answer from the given alternatives

11. A solution has a pH of 9. The solution is best described as:
    - a strongly alkaline
    - c weakly alkaline
    - b weakly acidic
    - d strongly acidic
12. Which one of these statements about magnesium oxide is true:
    - a it reacts with hydrochloric acid to form a salt
    - b it reacts with sodium hydroxide to form a salt
    - c it reacts with water to form an acid
    - d it reacts with basic oxide to form a base
13. Sodium hydroxide is added to solution M. A reddish-brown precipitate is formed, solution M contains:
    - a iron (II) ions
    - c copper (II) ions
    - b iron (III) ions
    - d silver (I) ions
14. Which one of these statements about fertilizers is true?

a ammonium nitrate can be used as a fertilizer

b primary fertilizers contain nitrogen, sulphur and iron

c fertilizers are added to the soil to make it more alkaline

d fertilizers are made by combining calcium with oxygen

15. Which of the following is not the characteristic of an acid:

a an acid changes the colour of an indicator

b an acid has a bitter taste

c an acid ionizes in water

d an acid produces hydronium ions in water

16. When an acid reacts with an active metal:

a hydronium ion concentration increases

b metal forms anions

c hydrogen gas is produced

d carbon dioxide gas is produced

17. Which of the following is the correct definition of a Brønsted-Lowry base?

a an electron pair donor

b an electron pair acceptor

c a proton donor

d a proton acceptor

18. The pH of a solution is 6.32. What is the pOH of the solution:

a 6.32 C 7.68

b 4.8×10^{-7} d 2.1×10^{-8}

19. A neutral aqueous solution:

a has a 7.0 M H_3O^+ concentration

b contains neither hydronium ion nor hydroxide ion

c has an equal number of hydronium ions and hydroxide ions

d has an unequal number of hydronium ions and hydroxide ions

20. Which of the following solutions would have a pH value greater than 7:

a $[OH^-] = 2.4 \times 10^{-2}$ M c 0.0001 M HCl

b $[H_3O^+] = 1.53 \times 10^{-2}$ M d $[OH^-] = 4.4 \times 10^{-9}$ M

Part V: Give Short Answers

21. Write equations for the reaction of zinc oxide with:

    a hydrochloric acid

    b aqueous sodium hydroxide

22. 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.

23. 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.

24. 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

25. 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.

26. For the following three reactions, identify the reactants that are Arrhenius bases, Brønsted-Lowry bases, and/or Lewis bases. State the type(s) of bases in each case. Explain your answer.

    a $NaOH\ (s) \rightarrow Na^+\ (aq) + OH^-\ (aq)$

    b $HF\ (aq) + H_2O\ (l) \rightarrow F^-(aq) + H_3O^+\ (aq)$

    c $H^+\ (aq) + NH_3\ (aq) \rightarrow NH_4^+\ (aq)$