STOICHIOMETRY OF CHEMICAL REACTIONS, SOLUTIONS AND SOME INORGANIC COMPOUNDS

CHEMISTRY
GRADE 10

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Introduction to the Grade 10 Chemistry

Chemistry is the study of the composition, properties, and changes that occur in matter. Anything with mass and space is considered to be matter. As a result, chemistry is the study of practically everything in our environment, including the liquids we drink, the gases we breathe, and the chemical composition of everything from the plastic on your phone's case to the ground beneath your feet.

Students must have a good exposure in chemistry in order to prepare for careers as biologists, geoscientists, doctors, nurses, soil scientists, food scientists, chemical engineers, petroleum engineers, and many other professions.

This distance material, prepared for grade 10 Chemistry covers the entire syllabus in the grade 10 curriculum. The course material is provided in two modules.

The first module deals with **stoichiometry of chemical reactions, solutions and some inorganic compounds.** The second module covers **electrochemistry and some important substances (metals, nonmetals, and hydrocarbons).** Each module is broken down into units, which are themselves divided into sections. There are **activities, self-test exercises and self-review exercises**. **Learning activities** are activities at the end of each topic. **Self-test exercises** are exercises at the end of each section. There are also check lists at the end of each section. The check lists help to confirm whether you have understood or not the list of concepts or ideas of the section. If you do not grasp the concepts properly, you should revise thoroughly until you master the basic ideas. Don't skip topics without proper understanding of the concepts. **Self-review exercises** are exercises at the end of each unit. **Answer keys** are provided for all the activities, self-test exercises and self-review exercises at the end of each unit.

As you may recognize, all these exercises are self-evaluation questions. All of the exercises should be performed by you. In other words, you take on the roles of both a teacher and a student. Checking the answers should only be done after you have properly tried and clarified. If you are not able to solve the problems, try again and again till you understand the question. You shouldn't immediately check the answers before you are confident in them. You will not learn anything if you check the answers without mastering it. You can test yourself how much you score out of a certain number. If you score below 70 %, you should revise and test yourself again.

There are experiments in some units for practical activities. It is required that you participate in the laboratory activities. Your tutor explains the steps and clarifies you. You should read all the procedures and safety precautions before you carry out or attend every laboratory activity. You are expected to do the experimental activities in the nearby high school to your residence. Ask your tutor the appropriate time and condition for carrying out the experiments.

Study Tips for Chemistry

1. **Divide the sections into smaller ones.** You'll probably become frustrated or lose confidence before you even begin if you try to master a large part all at once.
2. **Study for understanding rather than memorization.** When studying, avoid trying to learn

the material by rote. Find out patterns to remember the concepts.

3. **Develop writing skills and always carry paper and a pen with you.** Continue to write and read. Your capacity for focus and memory will significantly improve as a result. Create patterns in your work with highlighters to help you recall.
4. **Practice chemistry problems:** Practicing is the best way to improve the skills that you have gained.
5. **Create a list of every concept:** Make it simple for yourself instead of making it difficult. Make a list of the ideas that you are unclear about, and then put it in writing.
6. **Write an outline of the subject**

If the subject is confusing to you, you might begin by creating a thorough outline of it. The concept will be simple for you to grasp.

General Objectives of Grade 10 Chemistry

To develop understanding and acquire knowledge of

- Types, properties and processes of formation of solutions.
- The interconvert ion of electrical and chemical energies and their applications.
- Importance of electrochemical cells in daily life.
- Extraction, chemical properties and uses of metals and nonmetals.
- Classes of organic and inorganic compounds and some of their uses.
- Formation of chemical bonding and properties of compounds formed by different types of bonds.

To develop skills and abilities of

- Handling and using science apparatuses and laboratory substances correctly.
- Preparing solutions of specific concentration and solving quantitative problems involving solutions.
- Conducting experiments to observe and analyze the physical properties of substance and determine the type of bonding.
- Designing and conducting simple experiments appropriate to their level.
- Applying conservations of mass laws to calculate relevant quantities.

To develop the habit and attitude of

- Appreciating the roles of chemistry in energy production.
- Having an interest and curiosity towards natural resources.
- Responsible about safety of oneself, and protection of natural resources.
- Being honest and accurate in recording and validating data.

MODULE - I
Stoichiometry of Chemical Reactions, Solutions and some Inorganic Compounds

- Dear learner! Module I is divided in to three units. The first unit is about chemical reactions & reaction's stoichiometry. Under this topic, you will study the meanings of chemical reactions, rules of writing chemical equations and balancing chemical equations as well as quantitative aspects of chemical equations. The second unit deals with solutions. The focus of unit two is on properties, types, the solution process, preparation of solutions and ways of expressing their concentrations. The third unit is about important inorganic compounds. The important inorganic compounds which you will discuss under unit three are oxides, acids, bases and salts. Hence, unit three focuses mainly on the definitions, classifications, properties and uses of these compounds.

There are certain **experiments** in some units for practical observation. It is required that you participate in the laboratory activities. Your tutor explains the steps and clarifies you. You should read all the procedures and safety precautions before you cary out or attend every laboratory activity.

Module Objectives

Up on completion of this module, you will be able to

- Understand the basic ideas of chemical reactions.
- Solve stoichiometric problems, which are related to chemical equations.
- Understand solutions and solution concentration.
- Solve problems related to solution concentration.
- Understand the meanings of oxides, acids, bases and salts.
- Explain the properties and uses of oxides, acids, bases and salts.

Module Learning Strategies

Dear learner!

Since you may need to balance multiple jobs and family commitments with your learning activities, distance learning presents you its own set of challenges. In addition, learning at a distance usually has limited interaction because of geographic isolation from the instructor and other students. Therefore; you must rely on technology as a distant learner to provide information for your courses. Within a specific time frame, digest and review your reading material in small doses. Consider reading for 1 or 2 hours each day rather than 5 hours in a single day. Seek understanding, as you master the fundamentals of chemistry and gain understanding of the concepts, you'll find it much easier to memorize everything else. Take good notes, note taking forces you to write things down and will help you to determine what you do and don't understand.

The Study Time Needed

2 months, 4hrs per week

Resource Materials

Current grade 10 chemistry text

Legend

There are a number of icons, or symbols in this teaching material. The meanings of each icon is as follows.

This tells you there is an overview of the unit and what the unit is about.

? This tells you there is an in-text question to answer or think about in the text.

This tells you to take note of or to remember an important point.

This tells you there is a self-test for you to do.

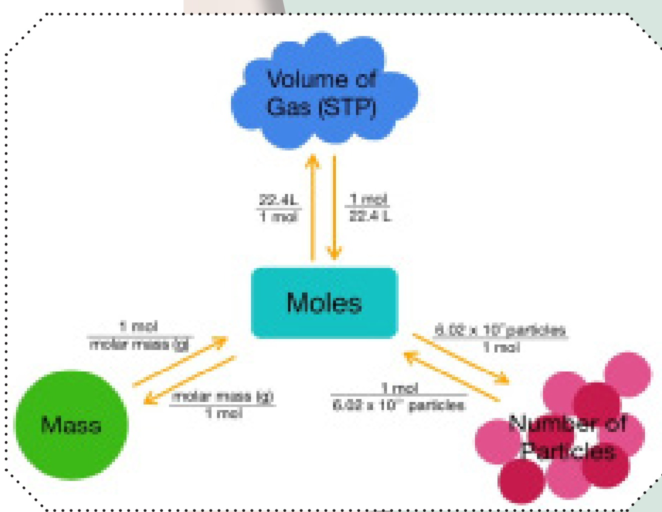
This tells you there is a checklist.

This tells you there is a written assignment.

This tells you that this is the key to the answers for the self-tests.

MODULE CONTENTS

UNIT 1



CHEMICAL REACTIONS AND STOICHIOMETRY

- Dear learner! Unit 1 is about chemical reactions & reaction's stoichiometry. It is subdivided in to four sections. In the first section you will study the meanings and techniques of writing chemical reactions and chemical equations. The second section describes the different types of reactions namely single replacement, double displacement, decomposition and direct combination reactions including redox reactions. The third section deals with the Molecular and formula mass, the mole concept and chemical formulas. The fourth section is about reaction stoichiometry. Stoichiometry deals with mass-mass, volume-volume, and mass-volume calculations, it also describes limiting & excess reactants, theoretical, actual and percentage yields.

Unit Outcomes

Upon completion of this unit, you will be able to

- define the basics of chemical reaction and describe the four major types of reactions
- develop skills in writing and balancing chemical equations
- understand oxidation - reduction reactions and analyse redox reactions by specifying the oxidizing agent, the reducing agent, the substance reduced or oxidized
- understand fundamental laws of chemical reactions and know how they are applied
- develop skills in solving problems based on chemical equations (mass - mass, volume - volume and mass - volume problems).
- develop skills in determining the limiting reactant, theoretical yield, actual yield and percentage yield.

Unit Contents

Section 1.1: Chemical reactions and Chemical Equations
Section 1.2: Types of Chemical Reactions
Section 1.3: Oxidation and Reduction Reactions
Section 1.4: Molecular and Formula Mass, the Mole Concept and Chemical Formulas
Section 1.5: Stoichiometry

The Required Study Time for the Unit - 10 Hours

Unit Learning Strategies

Dear learner, you may have family and busy schedule. Therefore, whenever you start distance learning you need to devise a clear strategy to perform the study. Though the strategy of learning varies from individual to individual. The following are more common and therefore we suggest you follow them.

1. Create a schedule and follow it.

Design a schedule on how to study the new topic. It can include a timeline for completing readings, doing exercises and assignments, as well as the specific methods and resources that will be used to learn the material. The learning schedule helps you to stay organized, motivated and confident.

2. Read roughly then deeply, take notes, memorize, understand and do the exercises.

Unit 1 involves some calculations which need through reading and understanding. First read the topic roughly, then read deeply and identify the key points. If the material is not clear find an alternative one for clarity or keep the content that is difficult for you and later on come back to it after a period of time. Take note always while you read and put descriptions and explanations in to your own words. Use internet, video and other materials for more clarity and understanding.

3. Utilize alarms and calendars to keep you abreast with deadlines

After getting your course or syllabus plans, document all important dates on the calendar. Normally, learners who utilize homework excuses in a distance learning course are not spared the consequences.

In a few scenarios, since remote learning is less social, the consequences of forgetting or using homework excuses in an assignment might be more inflexible than in a classroom surrounding.

4. Appreciate yourself

A student must reward themselves with something they cherish performing; this is a good manner to keep yourself inspired. When you reach a targeted milestone or attain what you had intended for a specific learning session, you should reward yourself with anything positive. By teaching your brain to acknowledge that learning will result in promising rewards, you will be more inspired to keep moving and cherish your learning and achievements.

5. Take part in distant discussions

As a learner, you must consider opportunities to participate with classmates in remote group activities or discussions. It would help if you remembered to be mindful or courteous of your attitude whenever remotely communicating.

Section 1.1: Chemical Reactions and Chemical Equations

- Dear learner, do you know exactly what a chemical reaction is? This section deals with about chemical reactions and chemical equations; The burning of wood in a fire place, the rusting of iron, the growth of seed into a plant, the souring of 'teji' and almost all other changes you see taking place around you are the result of chemical reactions. In this section, you will learn the meanings of chemical reactions and chemical equations, how to write chemical equations and the methods of balancing chemical equations.

Learning Outcomes

Upon completion of this section, you will be able to:

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and (→), appropriately when writing a chemical reaction
- Explain the roles of subscripts and coefficients in chemical equations.
- Explain the role of the Law of Conservation of Mass in a chemical reaction

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

1.1.1. Meanings of Chemical Reactions and Chemical Equations

Chemical reaction is the process in which reacting substances, called reactants, are converted into new substances, called products. The characteristics of the products are completely different from those of the reactants. For example, ash, water vapor, and carbon dioxide are the end products of burning wood, and they are completely different from the initial material (wood). Chemical reactions are carried out to produce new types of substances in the laboratory and industry.

For example,

In the laboratory, hydrogen gas is prepared by the reaction between zinc metal and dilute sulfuric acid.

Zinc + Dilute sulfuric acid → Zinc sulphate + Hydrogen gas

In industry, ammonia gas is manufactured from hydrogen and nitrogen by Haber's process.

Hydrogen + Nitrogen → Ammonia

? How can you tell whether a reaction is occurring or not?

The occurrence of a chemical reaction demonstrates one or more of the following characteristics:

Evolution of gas
Change in energy
Change in color
Change in temperature
Formation of precipitate
Change of odor etc

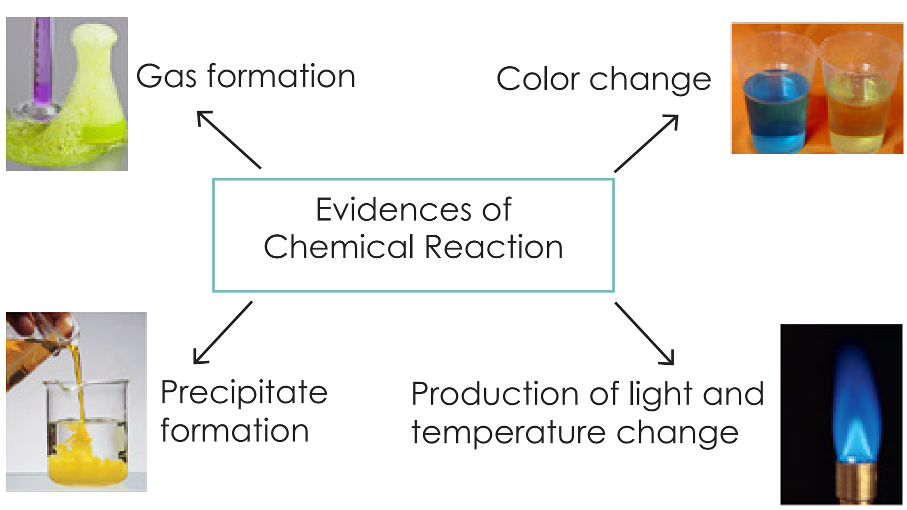


Figure 1.1 Examples of evidences for a chemical reaction

? What is a short way to show a chemical reaction?

A chemical reaction is represented by a short hand notation called chemical equation as:

Reactants → Products

A chemical equation uses chemical symbols to show what happens during a chemical reaction. A balanced chemical equation can be used to describe the relationships between the amounts of reactants and products.

1.1.2. Writing Chemical Equations

- Dear student! A chemical equation represents a chemical reaction using symbols and numbers.

For example, when hydrogen gas (H_2) burns in air (which contains oxygen, O_2) form water (H_2O). This reaction can be represented by the chemical equation:

$$H_2 + O_2 \rightarrow H_2O$$

Where the "plus" sign means "reacts with" and the arrow means "to yield." Thus, this symbolic expression can be read: "Hydrogen reacts with oxygen to yield water."

During chemical reactions, bonds are broken in the reactants to form new bonds in the products. Although they have been rearranged into different configurations, the atoms in the products are the same as those in the reactants.

A complete chemical equation requires that the reactant and product sides have the same amount of atoms. In other words, the equation should be **balanced**.

The above equation can be balanced by placing the appropriate coefficient (2 in this case) in front of H_2 and H_2O:

$$2H_2 + O_2 \rightarrow 2H_2O$$

This **balanced chemical equation** shows that "two hydrogen molecules can combine or react with one oxygen molecule to form two water molecules". In this section, you will study how to write and balance chemical equations.

Steps to Write a Chemical Equation

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

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

Hydrogen + Nitrogen → Ammonia (word equation)

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

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

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

3. Balance the equation

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

☞ The substances that take part in chemical reactions may be solids, liquids, or gases, or they may be dissolved in a solvent. Ionic compounds, in particular, frequently undergo reaction in aqueous solution that is, dissolved in water. Sometimes this information is added to an equation by placing the appropriate symbols after the formulas: Solid - (s), Liquid- (l), Gas - (g) and aqueous solution - (aq)

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

1. Word equation.

Calcium Carbonate + Sulfuric acid → Calcium Sulfate + Water + Carbon Dioxide

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

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

1. Word equation. Silver nitrate + sodium chloride → silver chloride + sodium nitrate
2. Chemical equation. $AgNO_3(aq) + NaCl\,(aq) \rightarrow AgCl(aq) + NaNO_3\,(aq)$
3. The equation is balanced.

A chemical equation has both qualitative and quantitative meanings. Qualitatively, a chemical equation indicates the types of the reactants and products in the reaction. Quantitatively, a chemical equation expresses the relative number (amount) of moles, molecules or masses of the reactants and products.

Activity 1.1

Explain the quantitative and qualitative aspect for the formation of magnesium oxide from magnesium and oxygen by the equation:

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s).$$

1.1.3. Balancing Chemical Equations

? Why we balance chemical equations?

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear nor do new atoms appear to form the products. In chemical reactions, atoms are never created nor destroyed based on the law of conservation of mass. The same atoms that were present in the reactants are present in the products – they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be balanced. That is, in a complete chemical equation, the same number of each atom must be present on the reactants and the products sides of the equation.

To balance a chemical equation means to equalize the number of atoms on both sides of the equation by putting appropriate coefficients in front of the formulas. Chemical equations can be balanced using a variety of methods. This topic will cover three methods of balancing chemical equations. Namely, **the Least Common Multiple (LCM) method**, **the inspection method**, and **the algebraic method**.

1. Inspection Method

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

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

i. Write the word equation.

ii. Write the correct symbols or formulas for the reactants and products (skeleton equation). The term "skeletal equation" refers to an equation that has the correct formulas but has not yet had the proper coefficients added.

iii. Count and tabulate the number of each type of atoms on the two sides of the skeleton equation.

iv. Make the number of each type of atoms on the left side equal to the number of corresponding atoms on the right side of the equation.

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

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

Step 2. skeleton equation: $Fe + H_2O \rightarrow Fe_3O_4 + H_2$

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

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

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

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

$$3Fe + H_2O \rightarrow Fe_3O_4 + H_2 \qquad \text{---- 2}$$

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

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + H_2 \qquad \text{--- 3}$$

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

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2 \qquad \text{--- 4}$$

Equation 4 is a balanced chemical equation. Note that number of atoms of each element is conserved (Number of atoms reactant side = Number of atoms product side).

? How?

These are shown in the table below:

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

2. The Least Common Multiple (LCM) Method

In the LCM method, the coefficients for the balanced chemical equation are obtained by taking the LCM of the total valency of reactants and products and then dividing it by total valency of reactants and products. All the necessary steps to balance a chemical equation by the LCM method are shown by the following examples.

Example 1.4

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

Step 1. Represent the reaction by a word equation.

Aluminium + Oxygen → Aluminium oxide

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

$$Al + O_2 \rightarrow Al_2O_3$$

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

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

From the equation we see that the valency of aluminium is 3.

The total valency of oxygen is 2 × 2 = 4.

The total valency of aluminium in Al_2O_3 is 3 × 2 = 6.

The total valency of oxygen in Al_2O_3 is 2 × 3 = 6

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

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

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

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

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

II. When iron reacts with water, iron (III) oxide and hydrogen are produced. Write the balanced equation.

Step 1: Iron + water → Iron (III) oxide + hydrogen

Step 2: $Fe + H_2O \rightarrow Fe_2O_3 + H_2$ (Unbalanced equation)

Step 3: $\overset{2}{Fe} + \overset{2\ 2}{H_2O} \rightarrow \overset{6\ \ 6}{Fe_2O_3} + \overset{2}{H_2}$

Step 4: $\overset{3}{Fe} + \overset{2}{H_2}\overset{2}{O} \underset{LCM}{\overset{6}{\rightarrow}} \overset{6}{Fe_2}\overset{6}{O_3} + \overset{2}{H_2}$

Step 5: $\underline{2}Fe + \underline{3}H_2O \rightarrow Fe_2O_3 + \underline{3}H_2$ (Balanced)

Checking: There are 2 iron, 6 hydrogen, and 3 oxygen atoms on each side of the equation. Thus, the equation is balanced

III. The reaction of ammonium sulphate with aluminium nitrate would form aluminium sulphate and ammonium nitrate.

Solution:

Step 1:

Ammonium sulphate + Aluminum Nitrate → Aluminum Sulphate + Ammonium Nitrate

Step 2: $(NH_4)_2SO_4 + Al(NO_3)_3 \rightarrow Al_2(SO_4)_3 + NH_4NO_3$

Step 3: $\overset{2}{(NH_4)_2}\overset{2}{SO_4} + \overset{3}{Al}\overset{3}{(NO_3)_3} \rightarrow \overset{6}{Al_2}\overset{6}{(SO_4)_3} + \overset{1}{NH_4}\overset{1}{NO_3}$

Step 4: $\overset{2}{(NH_4)_2}\overset{2}{SO_4} + \overset{3}{Al}\overset{3}{(NO_3)_3} \underset{LCM}{\overset{6}{\rightarrow}} \overset{6}{Al_2}\overset{6}{(SO_4)_3} + \overset{1}{NH_4}\overset{1}{NO_3}$

Step 5: $3(NH_4)_2SO_4 + 2Al(NO_3)_3 \rightarrow Al_2(SO_4)_3 + 6NH_4NO_3$ (balanced)

There are 12 nitrogen, 24 hydrogen, 3 sulphur, 30 oxygen and 2 aluminium atoms on both sides of the equation. Thus, the equation is correctly balanced.

Activity 1.2

1. Write the balanced chemical equation to represent the following reactions.
   A. Potassium chlorate when heated produces potassium chloride and oxygen
   B. Sodium carbonate reacts with hydrochloric acid to form water, carbon dioxide and sodium chloride
   C. Silver oxide decomposes to silver and oxygen gas

3. Balancing Equations by Algebraic Method

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

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

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

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

$$aN_2 + bH_2 \rightarrow cNH_3$$

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

N: $2a = c$

H : $2b = 3c$

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

In the above case, a is the smallest coefficient. Assuming a = 1, the values of b and c can be obtained as follows. $c = 2 \times 1 = 2$, $2b = 3c = 3 \times 2 = 6$; $b = 6/2 = 3$

Since a, b, and c have no common multiples, they can be substituted into the equation as follows.

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

Example 1.5: Balance the reaction: $Al + O_2 \rightarrow Al_2O_3$ using the algebraic method.
Assign variables for each element

$$aAl + bO_2 \rightarrow cAl_2O_3$$

The equation for Aluminum: $a = 2c$,
The equation for oxygen: $2b = 3c$
Assuming a = 1, we get:

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

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

Since fractional values of b and c are obtained, the lowest common denominator between the variables a, b, and c must be found and multiplied with each variable. Since the lowest common denominator is 4, each of the variables must be multiplied by 4.

Therefore, $a = 4 \times 1 = 4$; $b = (3/4)\times 4 = 3$; $c = (1/2)\times 4 = 2$
Substituting the values of a, b, and c in the unbalanced equation, the following balanced chemical equation is obtained.

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

Example 1.6: Balance the equation: $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$
Step 1. $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$
Step 2. $aPCl_5 + bH_2O \rightarrow cH_3PO_4 + dHCl$

P: $a = c$

Cl : $5a = d$

H: $2b = 3c + d$

O : $b = 4c$

Assume a =1 , therefore c=1, 5a = d d =5, b =4.
Substituting the values of a, b, c and d in the unbalanced equation:

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

Checklist -1.1

Please put a tick (√) if your answer is 'yes' if not go back and revise it to refer to relevant reference books. I can ...

Competencies		Check
1.	write symbols of elements & chemical formulas of reactants & Products	
2.	differentiate between chemical reaction & chemical equation	
3.	derive chemical equation from word equation	
4.	explain the qualitative aspect of chemical reaction	
5.	describe quantitative aspect of chemical reactions	
6.	balance chemical equations by inspection method	
7.	balance chemical equations by LCM method	
8.	balance chemical equations by algebraic method	

Self-Test Exercise 1.1

1. Say true or false
   i. In any chemical reaction, each type of atom is conserved.
   ii. In a balanced chemical equation, the total number of moles of reactants is equal to the total number of moles of products.
2. To balance a chemical equation means to equalize the number of molecules on both sides of the equation by putting appropriate subscripts under the formulas
3. Fill in the blanks
   a. ______ uses chemical symbols to show what happens during a chemical reaction.
   b. An equation stated in words is ______equation.
   c. ____________ has both qualitative and quantitative aspects.
4. Write the balanced chemical equation to represent the following reactions.
   a. Sodium bromide reacts with chlorine to form sodium chloride and bromine
   b. hydrochloric acid reacts with sodium carbonate to form sodium chloride, water and carbondioxide
   c. Potassium chlorate when heated produces potassium chloride and oxygen
   d. Sodium carbonate reacts with hydrochloric acid to form water, carbon dioxide and sodium chloride
   e. Silver oxide decomposes to silver and oxygen gas
   i.
5. Balance the following chemical equation, using the inspection method:
   a. $Na + H_2O \rightarrow NaOH + H_2$
   b. $KClO_3 \rightarrow KCl + O_2$
   c. $H_2O_2 \rightarrow H_2O + O_2$
   d. $Al + H_3PO_4 \rightarrow AlPO_4 + H_2$
6. Balance the following equations by any method.
   a. $PCl_5 + H_2O \rightarrow H_3PO_4 + HCl$
   b. $Mg + H_2O \rightarrow Mg(OH)_2 + H_2$

c. $Zn(NO_3)_2 \rightarrow ZnO + NO_2 + O_2$
d. $H_2SO_4 + NaOH \rightarrow Na_2SO_4 + H_2O$
e. $NH_3 + O_2 \rightarrow NO + H_2O$
f. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
g. $FeCl_3 + MgO \rightarrow Fe_2O_3 + MgCl_2$
h. $BaCl_2 + K_3PO_4 \rightarrow Ba_3(PO_4)_2 + 6KCl$
i. $P_4O_{10} + H_2O \rightarrow H_3PO_4$

Section 1.2: Types of Chemical Reactions

- Dear student! Anything that is categorized into types makes the study easier. As you may have observed certain reactions give light, while others show color change or form solid, while still others produce vapor, and so on. Here, you will learn about the classification and the different types of chemical reactions with examples.

Learning Outcomes

Upon completion of this section, you will be able to

- Classify a chemical reaction as a direct combination(synthesis), decomposition, single replacement, or double replacement.
- Predict the products of simple reactions.
- Classifications of chemical reactions.

Chemical reactions are classified into types to help us analyze them and also to help us predict what the products of the reaction will be. Different elements and compounds react in different ways to produce different kinds of new substances. Many chemical reactions can be classified as one of the four basic types of reactions. The four major types of chemical reactions are synthesis, decomposition, single replacement and double replacement. By analyzing the reactants and products of a particular reaction, you can classify them into one of these categories. Some reactions may be categorized into more than one category. A complete understanding of these reaction types will help predict the products of unknown reactions.

1.2.1. Direct Combination (Synthesis) Reactions

A synthesis reaction is one in which two or more reactants combine to make one type of product.

General equation: $A + B \rightarrow AB$

Synthesis reactions occur as a result of two or more simpler elements or molecules combining to form a more complex molecule.

Look at the example below.
Here two elements (hydrogen and oxygen) are combining to form one product (water).

Example: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

- Dear learner! You can always identify a synthesis reaction because there is **only one product of the reaction**.

You should be able to write the chemical equation for a synthesis reaction if you are given a product by picking out its elements and writing the equation. Also, if you are given elemental reactants and told that the reaction is a synthesis reaction, you should be able to predict the products.

Example 1.7:

(a) Write the chemical equation for the synthesis reaction of silver bromide, AgBr.

(b) Predict the products for the following reaction: $CO_2(g) + H_2O(l)$

Solution:

(a) $2Ag + Br_2 \rightarrow 2AgBr$

(b) $CO_2(g) + H_2O(l) \rightarrow H_2CO_3$

1.2.2. Decomposition Reactions

- When one type of reactant breaks down to form two or more products, we have a decomposition reaction. The best way to remember a decomposition reaction is that for all reactions of this type, there is only one reactant.

General Equation: $AB \rightarrow A + B$

Look at the example below for the decomposition of ammonium nitrate to dinitrogen oxide and water.

Example 1.8: $NH_4NO_3 \rightarrow N_2O + 2H_2O$

Notice the one reactant, NH_4NO_3, is on the left of the arrow and there is more than one on the right side of the arrow. This is the exact opposite of the synthesis reaction type.

When studying decomposition reactions, we can predict reactants in a similar manner as we did for synthesis reactions. Look at the formula for magnesium nitride, Mg_3N_2. What elements do you see in this formula? You see magnesium and nitrogen. Now we can write a decomposition reaction for magnesium nitride. Notice there is only one reactant.

$$Mg_3N_2 \rightarrow 3Mg + N_2$$

Example 1.9: Write the chemical equation for the decomposition of:

(a) Al_2O_3 (b) Ag_2S (c) MgO

Solution:

(a) $2\ Al_2O_3 \rightarrow 4Al + 3O_2$

(b) $Ag_2S \rightarrow 2Ag + S$

(c) $2MgO \rightarrow 2\ Mg + O_2$

1.2.3. Single Replacement Reactions

- A third type of reaction is the single replacement reaction. In single replacement reactions one element reacts with one compound to form products. The single element is said to replace an element in the compound when products form, hence the name single replacement. Metal elements will usually replace other metals in ionic compounds or hydrogen in an acid. Nonmetal elements will usually replace another nonmetal in an ionic compound.

General equation: $A + BC \rightarrow B + AC$

Consider the following examples.

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$$

Notice that the metal element, Zn, replaced the metal in the compound $Cu(NO_3)_2$. A metal element will always replace a metal in an ionic compound. Also, note that the charges of the ionic compounds must equal zero. To correctly predict the formula of the ionic product, you must know the charges of the ions you are combining.

$$Zn(s) + 2HBr(aq) \rightarrow ZnBr_2(aq) + H_2(g)$$

When a metal element is mixed with acid, the metal will replace the hydrogen in the acid and release hydrogen gas as a product. Once again, note that the charges of the ionic compounds must equal zero. To correctly predict the formula of the ionic product, you must know the charges of the ions you are combining, in this case Zn^{2+} and Br^-.

$$Cl_2(g) + 2KI(aq) \rightarrow 2KCl(aq) + I_2(s)$$

When a nonmetal element is added to an ionic compound, the element will replace the nonmetal in the compound. Also, to correctly write the formulas of the products, you must first identify the charges of the ions that will be in the ionic compound.

Example 1.10: What would be the products of the reaction between solid aluminum and iron(III) oxide?

The reactants are: $Al + Fe_2O_3 \rightarrow$

Solution: In order to predict the products we need to know that aluminum will replace iron and form aluminum oxide (the metal will replace the metal ion in the compound). Aluminum has a charge of +3 and oxygen has a charge of -2. The compound formed between aluminum and oxygen, therefore, will be Al_2O_3. Since iron is replaced in the compound by aluminum, the iron will now be the single element in the products. The unbalanced equation will be:

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

and the balanced equation will be:

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

More Examples of Single-Displacement Reactions

1. Active metals displace hydrogen from acids

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

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

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

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

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

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

   Zinc displaces copper from copper (II) sulphate solutions

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

Iron displaces copper from copper (II) sulphate solutions

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

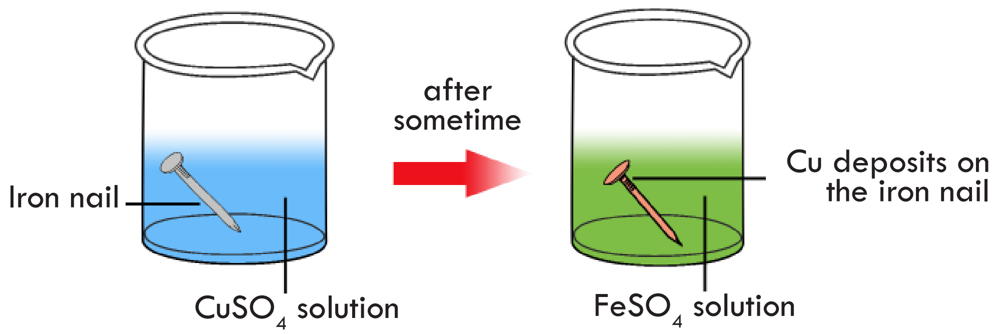


Figure 1.2 Replacement of copper by iron metal from copper sulphate solution.

Experiment 1.1

Investigation of Single Displacement Reaction

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

Apparatus: Iron rod and beaker

Chemicals: $CuSO_4$.

Procedure:

1. Clean a piece of iron rod or iron knife with emery paper to remove any rust.
2. Take saturated $CuSO_4$ solution in a beaker.
3. Dip the iron rod into the $CuSO_4$ solution as shown in ***Figure 1.3*** and wait for a few minutes. What did you observe on the iron rod?
4. Allow the reactants to stand for one day and observe any change on the iron rod.

Questions

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

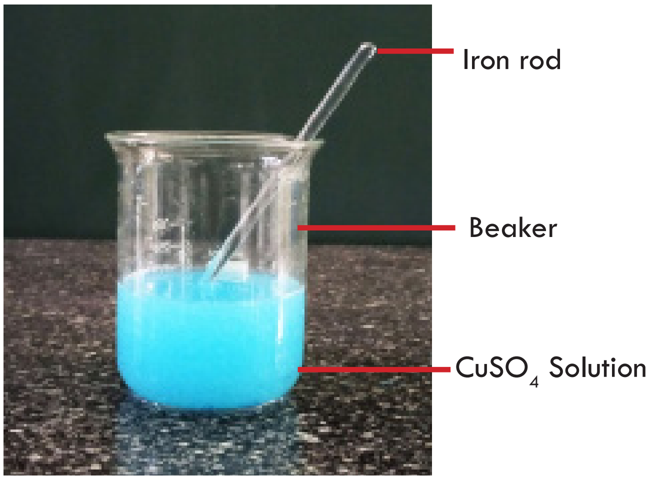


Figure 1.3 Reaction between iron and copper (II)

1.2.4. Double Replacement Reactions

For double replacement reactions two ionic compound reactants will react by having the cations exchange places, forming two new ionic compounds. The key to this type of reaction, as far as identifying it over the other types, is that it has two compounds as reactants. This type of reaction is more common than any of the others and there are many different types of double replacement reactions. Precipitation and neutralization reactions are two of the most common double replacement reactions.

General equation: AB + CD → AD + CB

For example, when solutions of silver nitrate and sodium chloride are mixed, the following reaction occurs:

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

This is an example of a precipitate reaction. Notice that two aqueous reactants form one solid, the precipitate, and another aqueous product.

An example of a neutralization reaction occurs when sodium hydroxide, a base, is mixed with sulfuric acid:

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

In order to write the products for a double displacement reaction, you must be able to determine the correct formulas for the new compounds. Remember, the total charge of all ionic compounds is zero. To correctly write the formulas of the products, you must know the charges of the ions in the reactants. Let's practice with an example.

Example 1.11: A common laboratory experiment involves the reaction between lead(II) nitrate and sodium iodide, both colorless solutions. The reactants are given below. Predict the products.

$$Pb(NO_3)_2(aq) + NaI(aq) \rightarrow$$

Solution:

We know that the cations exchange anions. We now have to look at the charges of each of the cations and anions to see what the products will be. In $Pb(NO_3)_2$, the nitrate, NO_3^- has a charge of -1. This means the lead must be +2, Pb^{2+}. In the sodium iodide, we are combining Na^+ and I^-.

Now we switch ions and write the correct subscripts so the total charge of each compound is zero. The Pb^{2+} will combine with the I- to form PbI_2. The Na^+ will combine with the NO_3^- to form $NaNO_3$.

Only after we have the correct formulas can we worry about balancing the two sides of the reaction. The final balanced reaction will be:

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

Experiment 1.2

Investigation of Double Displacement Reaction

Objective: To observe the displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$.

Apparatus: Beaker, stirrer, filter paper, filter funnel.

Chemicals: Na_2SO_4, $Ba(NO_3)_2$ and water.

Procedure:

1. Take solution of $Ba(NO_3)_2$ into a beaker and add drop-wise Na_2SO_4 solution as shown in ***Figure 1.4***. Then stir it continuously.
2. Filter the precipitate using a filter paper and funnel. Collect the filtrate or the solution in a clean beaker.

Questions

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

Figure 1.4 Double displacement reaction between Na_2SO_4 and $Ba(NO_3)_2$

Activity 1.3

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

Checklist - 1.2

Please put a tick (√) if your answer is 'yes' if not go back and revise it to refer to relevant reference books. I can ...

Competencies	No
1. describe combination reactions?	
2. describe single displacement reactions?	
3. describe decomposition reactions?	

4.	describe double replacement reactions?	
5.	predict the products of simple reactions?	
6.	classify a chemical reaction as a direct combination(synthesis), decomposition, single replacement, or double replacement?	
7.	define and describe oxidation reactions?	
8.	define and describe reduction reactions	
9.	assign oxidation number of elements and compounds?	
10.	define reducing agent?	
11.	define oxidizing agent?	
12.	Identify a reducing and oxidizing agent in a reaction?	
13.	balance reduction- oxidation reaction by redox method?	
14.	describe non redox reactions with examples?	

Self-Test Exercise 1.2

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

Section 1.3: Oxidation and Reduction Reactions

Dear student! In our day to day activity, we are familiar with the chemical processes like rusting of iron, burning of substances, respiration, digestion of food and so on. All such types of processes or reactions are known as oxidation and reduction or redox reactions. Additionally, most metallic and nonmetallic elements are obtained from their ores by the process of oxidation or reduction. In this section, you will study about the meaning of oxidation-reduction, oxidizing and reducing agents and balancing redox reactions using oxidation number method.

Learning Outcomes

Upon completion of this section, you will be able to

- Describe oxidation-reduction reactions.
- Balance oxidation-reduction reactions using redox method.
- Define redox reactions.
- Define the terms oxidation and reduction in terms of electron transfer.
- Define oxidation number (oxidation state).
- State oxidation number rules.

↳ Determine the oxidation number of an element in a given formula.

1.3.1. Oxidation-Reduction

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

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

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

Oxidation and reduction reactions can also be defined in terms of oxidation number. Oxidation is an increase in the oxidation number of an element and reduction is a decrease in the oxidation number.

For example,

$$CuO + 2Ag^{+} \rightarrow Cu^{2+} + Ag_2O$$

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

1.3.2. Oxidation Number or Oxidation State

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

Rules for Assigning Oxidation Numbers

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

Example 1.12: The oxidation number of Be = 0, Cu = 0, Br in Br_2 = 0, O in O_3 = 0, S in S_8 = 0. P in P_4 = 0, O in O_2 = 0

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

Example 1.13: $Na^{+} = +1$, $Mg^{2+} = +2$, $S^{2-} = -2$.

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

Exceptions

The oxidation number of oxygen in:

A. peroxides is –1. Example: Na_2O_2

B. superoxides is –1/2. Example: KO_2

C. oxygen diflouride is +2 Example: OF_2

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

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

Example 1.14: Assign oxidation number to all elements in

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

Solution

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

$$\overset{+1}{H}\overset{x}{N}\overset{-2}{O}_3$$

$$1 + x + (-2x3) = 0$$

$$x = +5$$

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

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

iii. MnO_4^-: Let the oxidation number of Mn be x. $MnxO_4^{2-}$. The sum of the oxidation numbers of Mn and O in MnO_4^- is –1 (Rule 6)

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

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

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

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

$$+2 + (4 \times (+1)) + (2 \times x) + (8 \times (-2)) = 0$$

$$2 + 4 + 2x - 16 = 0$$

$$2x - 10 = 0 \text{ or } x = +5$$

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

1.3.3. Oxidizing and Reducing Agents

? What are Oxidizing and Reducing Agents?

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

Reducing agent	Oxidizing agent
1. Loses one or more electrons	1. Gains one or more electrons
2. Causes reduction	2. Causes oxidation

3. Undergoes oxidation	3. Undergoes reduction
4. Becomes more positive	4. Becomes more negative

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

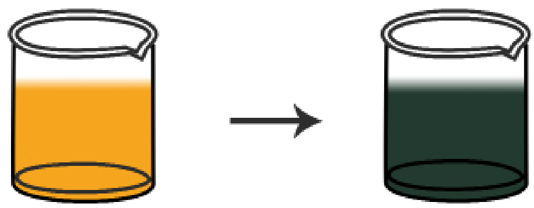
1. Permanganate ion (MnO_4^-) in acidic solution changes color from purple to colorless

MnO_4^- (purple) Mn^{2+}(colorless)

$$2MnO_4^- + 16H^+ + 10e^- \longrightarrow 2Mn^{2+} + 8H_2O$$

Figures 1.5 The reduction of Mn^{7+} to Mn^{2+} in acidic media.

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



$Cr_2O_7^{2-}$ (orange) Cr^{3+}(Green)

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \longrightarrow 2Cr^{3+} + 7H_2O$$

Figure 1.6 The reduction of Cr^{6+} to Cr^{3+} in acidic media.

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

I. A moist starch solution changes potassium iodide paper to blue-black to show that iodine is formed, $2I^- \rightarrow I_2$. That is potassium iodide is a reducing agent.

II. Hydrogen sulphide bubbled through a solution of an oxidizing agent forms a yellow precipitate, $S^{2-} \rightarrow S$. Thus, H2S is a reducing agent.

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

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

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

Example: Oxidizing agents $K\overset{7+}{Mn}O_4$, $Na\overset{7+}{Cl}O_4$, $K_2\overset{6+}{Cr_2}O_7$, Mn^{7+}, Cl^{+7}, Cr^{6+} These ions have no valence electrons and cannot lose but they can gain electrons. In the process, they get reduced and make others oxidized.

Reducing agents $Fe^{+2}S$, $C^{+2}O$, $Na_2S^{+4}O_3$, Fe^{2+}, C_2^+, S_4^+ can lose electrons.

For example $F^{2+} \rightarrow Fe^{3+} + e$

$$C_2^{+} \rightarrow C_4^{+} + 2e$$

$$S_4^{+} \rightarrow S_6^{+} + 2e$$

As they become oxidized and make others reduced.

1.3.4. Balancing Redox Reactions: Oxidation-Number-Change Method

The oxidation-number change method for balancing redox reactions is based on the principle that the total number of electrons lost by an element(reducing agent)must be equal to the total number of electrons gained by another element(oxidizing agent) in any redox reaction.

The steps of the oxidation number method are as follows:

Step 1: Correctly write the formula for the reactants and the products of the chemical reaction as in the example below:

Step 2: Assign oxidation numbers to each of the atoms in the equation and write the numbers above the atom.

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

Step 4: Use a line to connect the atoms that are undergoing a change in oxidation number. On the line, write the oxidation-number change.

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

Step 5: Use coefficients to make the total increase in oxidation number equal to the total decrease in oxidation number. In this case, the least common multiple of 2 and 3 is 6. So the oxidation-number increase should be multiplied by 3, while the oxidation-number decrease should be multiplied by 2. The coefficient is also applied to the formulas in the equation. So 3 is placed in front of the CO and in front of the CO_2. 2 is placed in front of the Fe on the right side of the equation. The Fe_2O_3 does not require a coefficient because the subscript of 2 after the Fe indicates that there are already two iron atoms.

$$\overbrace{\overset{+3}{Fe_2}\overset{-2}{O_3} + \underbrace{\overset{+2}{C}\overset{-2}{O} \rightarrow \overset{0}{Fe}}_{+2*3=6}}^{-3*2=6} + \overset{+4}{C}\overset{-2}{O_2}$$

Step 6: Check whether each element is balanced. Occasionally, a coefficient may need to be placed in front of a molecular formula that was not involved in the redox process. In the current example, the equation is now balanced.

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

Example 1.15: Balance the chemical equation:

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

1. Identify the oxidation number of each atom.

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

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

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

   N: +5 → +2; Change = -3

   As: +3 → +5; Change = +2

3. Make the total increase in oxidation number equal to the total decrease in oxidation number. We need 2 atoms of N for every 3 atoms of As. This gives us total changes of -6 and +6.

4. Place these numbers (2 & 3) as coefficients in front of the formulas containing those atoms

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

Activity 1.4

Balance the following using oxidation number change method

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

1.3.5. Non-redox Reactions

? What are non-redox reactions?

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

Example 1.16:

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

Checklist - 1.3

Please put a tick (√) if your answer is 'yes' if not go back and revise it to refer to relevant reference books. I can ...

Competencies	Check
1. define and describe oxidation reactions?	
2. define and describe reduction reactions?	
3. assign oxidation number of elements and compounds?	
4. define reducing agent?	
5. define oxidizing agent?	
6. identify a reducing and oxidizing agent in a reaction?	
7. balance reduction- oxidation reaction by redox method?	
8. describe non redox reactions with examples?	

Self-Test Exercise 1.3

Give the correct answers for the following questions.

1. Find the oxidation numbers to the underlined species for the following compounds or ions

a. $K_4[\underline{Fe}(CN)_6]$
b. $K_2\underline{Cr}_2O_7$
c. $H_2\underline{Pt}Cl_6$
d. $Na_2\underline{S}_4O_6$
e. $\underline{S}_2O_8^{-2}$
f. $H_2\underline{P}_2O_7^{2-}$
g. $H_2\underline{S}O_4$
h. $H\underline{Au}Cl_4$
i. $Fe_2(\underline{S}O_4)_3$

2. Determine whether the following processes are oxidation or reduction reactions

a. $Cu^{2+} + 2e^- \rightarrow Cu$
b. $K \rightarrow K^+ + e^-$
c. $O + 2e^- \rightarrow O^{2-}$
d. $S^{2-} \rightarrow S + 2e^-$
e. $Fe^{2+} \rightarrow Fe^{3+} + e^-$
f. $N + 3e^- \rightarrow N^{3-}$

3. Where do the most easily reduced and oxidized elements found in the periodic table of the elements?

4. For the following substances, tell whether the oxidation number increases or decreases in a redox reaction:

a. An oxidizing agent
b. A reducing agent
c. A substance undergoing oxidation
d. A substance undergoing reduction

5. In the following reactions, label the oxidizing agent and the reducing agent.

a. $ZnO + C \rightarrow Zn + CO$
b. $8Fe + S_8 \rightarrow 8FeS$
c. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
d. $PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$

6. Balance the following using oxidation number change method

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

Section 1.4: Molecular and Formula Mass, the Mole Concept and Chemical Formulas

- The mass of a molecule is called its molecular mass. The molecular mass is calculated by summing the masses of all the atoms shown in the chemical formula. Ionic compounds are not comprised of discrete molecules, but rather, ions. The smallest unit of an ionic compound that retains the identity of an ionic compound is called a formula unit. The mass of a formula unit is called its formula mass.

The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of the number of atoms or molecules in a bulk sample of matter. Because atoms and molecules are so small, the mole concept allows us to count atoms and molecules by weighing macroscopically small amounts of matter.

The chemical formula of a compound is a symbolic representation of its chemical composition. There are three main types of chemical formulas: empirical, molecular and structural. Empirical formulas show the simplest whole-number ratio of atoms in a compound, molecular formulas show the number of each type of atom in a molecule, and structural formulas show how the atoms in a molecule are bonded to each other.

Learning Outcomes

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

- Explain and determine molecular mass, formula mass.
- Define mole and explain the concept using examples.
- Define molar mass.
- Calculate the molar masses of substances.
- Determine molecular masses of molecules.
- Determine formula masses of ionic compounds.
- Explain the mole concept using examples.
- Calculate molar masses of some substances.
- Define empirical and molecular formulas.
- Determine empirical and molecular formulas.
- Calculate percentage composition of an element in a substance.

1.4.1. Molecular & Formula Mass, the Mole Concept

Everything in nature is chemically or physically connected to other substances. To determine the amount of substance in a sample, you need to know the fraction of the sample. The composition of a compound may be determined from its chemical formula and the atomic masses of the elements that make up the compound. The mass of a molecule is called its molecular mass.

Molecular mass (MM) and Formula Mass (FM)

? Why is understanding composition essential? Everything in nature is either chemically or physically combined with other substances. To find the amount of a material in a sample, you need to know what fraction of the sample it is. Some simple applications of composition are: the amount of sodium in sodium chloride for a diet, the amount of iron in iron ore for steel production, the amount of hydrogen in water for hydrogen fuel, and the amount of chlorine in freon to estimate ozone depletion.

? How the molecular and Formula masses are deterimined? The composition of a compound may be determined from its chemical formula and the atomic masses of the elements that make up the compound. If we know the atomic masses of the component atoms, we can calculate the mass of a molecule.

Example 1.17:

MM of H_2O = 2 x 1 amu (2 hydrogen atoms) + 16 amu (1 oxygen atom) = 18 amu, MM of CO_2 =12 amu (1carbon atom) +2 x 16 amu (2 oxygen atoms) = 44 amu

Note; amu = atomic mass unit

? What are molecular and Formula masses?

The molecular mass (sometimes called molecular weight) is the mass (in amu) of one molecule. The mass of a molecule is determined by adding the masses of all of its constituent atoms.

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

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

Example 1.18:
FM of NaCl = 23 amu (1Na) + 35.5 amu (Cl) = 58.5 amu.
FM of $Ca(OH_2)_2$ = 40 amu (1Ca) + 2 x [16 amu(1O) + 2 amu (2H)] = 74 amu

1.4.2. The Mole Concept

Dear student ! Do you know the meaning of mole and its concept in chemistry? The Mole Concept is the concept adopted as a convenient way to deal with the large quantities of very small entities such as atoms, molecules, or other specified particles. A mole is defined as the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12. 12 g of carbon-12 contains $6.022x10^{23}$ atoms. This number is known as Avogadro's number (NA), in honor of Amedeo Avogadro (1776-1856, who first proposed the concept). The term mole, like a dozen or a gross, refers to a particular number of things. A dozen of eggs equals 12 eggs, a gross of pencils equals 144 pencils, and a mole of any substance equals $6.02x10^{23}$ units of the substance.

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

Molar mass is the molecular mass or formula mass expressed in grams. It is the mass of 1 mole of a substance.

Molar mass of substance A = mass of one mole of A = 6.02 x10^{23} units of substance A

Example 1.20: molar mass of H_2O = 18 g = mass of 1 mole of H_2O = 6.02 x 10^{23} molecules of H_2O
Molar mass of NaCl = 58.5 g = mass of one mole of NaCl =6.02 x 10^{23} formula units of NaCl

1.4.3. Percentage Composition(%), Empirical and Molecular Formula

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

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

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

1.4.4. Percent Composition by Mass (%)

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

Mathematically,

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

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

$$\%\,H = \frac{2 \times 1.00\text{ g H}}{34.02\text{ g }H_2O_2} \times 100\% = 5.926\%$$

$$\%\,O = \frac{2\text{x}16\text{g O}}{34.02\text{ g }H_2O_2} \times 100\% = 94.06\%$$

1.4.5. Determination of Empirical and Molecular Formulas

Steps to determine empirical formula are:

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

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

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

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

Step 1: Derivation of moles of each element,

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

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

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

Step 2: Smallest molar amount is 1.41

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

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

Empirical formula is P_2O_5

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

Step 1: C= (79.9 g)/(12 g/mol) = 6.7 mol, H = 20.1g/(1g/mol) = 20.1 mol

Step 2: C= 6.7/6.7 = 1, H = 20.1/6.7 = 3

Step 3: Empirical formula = CH_3

Molecular formula = Empirical formula x n

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

Example 1.24: What is the molecular formula of the oxide of phosphorus that has the empirical formula P_2O_5 if the molecular mass of this compound is 284?

Molar mass of P_2O_5 = 142, n = 284/142 = 2

Therefore, molecular formula = $(P_2O_5)_2 = P_4O_{10}$

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

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

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

1. Calculate the moles of each element present:

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

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

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

2. After calculating the moles of each element present, identify the smallest molar amount and divide each element's molar amount by the smallest molar amount.

C : 3.22/3.22 =1

H : 9.67/3.22 =3

O : 3.22/3.22 =1

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

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

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

Checklist - 1.4

Please put a tick (√) if your answer is 'yes' if not go back and revise it to refer to relevant reference books. I can ...

Competencies	Check
1. determine molecular masses of molecules.	
2. determine formula masses of ionic compounds.	
3. explain the mole concept using examples.	
4. calculate molar masses of some substances.	
5. define empirical and molecular formulas.	
6. determine empirical and molecular formulas.	
7. calculate percentage composition of an element in a substance.	
8. calculate molar masses of some substances.	

Self-Test Exercise 1.4

1. Define and discuss each of the following terms: mole, Avogadro's number, molar mass, molecular mass, formula mass, percentage composition
2. Distinguish between (a) formula mass and molecular mass of a compound (b) empirical and molecular formula.
3. How many atoms and ions are there in one formula unit of
   a. K_2CO_3
   b. $Ba_3(PO_4)_2$
4. How many moles of each element are present in
   a. 1.00 mole of $Cu_3(PO_4)_2$
   b. 2.5 mole of Na_2CO_3
   c. 2 moles of $Al_2(SO_4)_3$
5. Calculate the empirical formula for:
   a. 39.3% Na, 60.7% Cl   b. 56.5% K, 8.7% C, 34.8% O
6. a. A compound of carbon and hydrogen contains 92.3% C and has a molar mass of 78.1g/mol. What is its molecular formula?
   b. A compound of carbon, hydrogen and chlorine has a molar mass of 99 g/mol. Analysis of a sample shows that it contains 24.3% carbon and 4.1% hydrogen. What is its molecular formula?
7. Calculate the molar mass of a compound if 0.372 mole of it has a mass of 152 g.
8. Calculate the molecular mass or formula mass (in amu) of each of the following substances:
   a. CH_4 b. NO_2 c. SO_3 d. C_6H_6 e. Na_2SO_4 f. $Ca_3(PO_4)_2$
9. Calculate the percentage composition by mass of the following compounds:
   a. C_2H_5OH b. Na_3PO_4 c. H_2SO_4 d. $CaCO_3$ e. $K_2Cr_2O_7$

Section 1.5: Stoichiometry

❁ Dear learner! A basic question raised in the chemical laboratory is, "How much product will be formed from specific amounts of starting materials (reactants)?" Or in some cases, we might ask the reverse question, "How much starting material must be used to obtain a specific amount of product?" To interpret a reaction quantitatively, we need to apply our knowledge of molar masses and the mole concept. The quantitative analysis of the reactants and products in a chemical reaction is called stoichiometry. Stoichiometry examines the quantity or ratio of reactants to products in terms of moles, mass, or volumes (for gases).

Stoichiometric calculations are built on the basis of the following two fundamental principles.

A. In a chemical equation, the composition of any substance must be expressed by a definite formula.

B. The law of conservation of mass must be followed (the total mass of reactants should be equal to the total mass of products)

Learning Outcomes

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

- Deduce mole ratios from balanced chemical equations.
- Solve mass-mass problems based on the given chemical equation.
- Define molar volume.
- State Avogadro's principle.
- Solve mass -volume problems based on the given chemical equation.
- Solve volume -volume problems based on the given chemical equation.
- Describe limiting and excess reactants.

1.5.1. Mass – Mass, Volume – Volume & Mass – Volume Relationships

Stoichiometric relationships may be mass-mass, Volume –Volume or mass - volume Relationships

A balanced chemical equation provides information about the nature, masses, number of moles, number of molecules/atoms of reactants and products. In this section, you will discover more details on stoichiometric relationships.

A. Mass – Mass Relationships

Mass-mass calculations are the most practical of all mass-based stoichiometry problems. In this type of problem, the mass of the given substance A is converted into moles of A by use of the molar mass of A. Then, the moles of the given substance A are converted into moles of the unknown B by using the mole ratio from the balanced chemical equation. Finally, the moles of the unknown B are converted to mass of B by use of its molar mass as shown below:

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

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

The balanced chemical equation for the reaction of aluminum metal with oxygen to give aluminum oxide is

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

The conversion steps are

grams of Al_2O_3 → moles of Al_2O_3 → moles of Al → grams of Al

$$\text{gram } Al_2O_3 \times \frac{1\text{mole } Al_2O_3}{\text{molar mass } Al_2O_3} \times \frac{\text{mole Al}}{\text{mole } Al_2O_3} \times \frac{\text{molar mass Al}}{\text{mole Al}} = \text{gram Al}$$

Coefficients of Al and Al_2O_3 are 4 and 2 respectively Therefore, conversion factor = mole ratio = 4 mol Al/2 mol Al_2O_3.

Given molar masses, Molar mass of Al_2O_3 = 102 g,

Molar mass of Al = 27 g

Using the conversion steps, we can find the mass of Al.

$$20.4 \text{ g } Al_2O_3 \times \frac{1\text{mole } Al_2O_3}{102 \text{ g } Al_2O_3} \times \frac{4 \text{ mole Al}}{2 \text{ mole } Al_2O_3} \times \frac{27 \text{ g Al}}{\text{mole Al}} = 10.8 \text{ gram Al}$$

Example 1.27: At high temperatures, magnesium reacts with nitrogen gas according to the equation: $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

How many grams of magnesium are needed to produce 25.0 g of Mg_3N_2?

Solution: The conversion steps are:

$$\text{gram } Mg_3N_2 \times \frac{1\text{mole } Mg_3N_2}{\text{molar mass } Mg_3N_3} \times \frac{\text{mole Mg}}{\text{mole } Mg_3N_2} \times \frac{\text{molar mass Mg}}{\text{mole Mg}} = \text{gram Mg}$$

Given: Conversion factor from the balanced equation = 3mol Mg/1mol Mg_3N_2

Molar mass Mg_3N_3 = 100 g, Mass of Mg_3N_2 = 25 g Molar mass Mg = 24 g,

Substituting the values, we can solve for the mass of Mg as follows:

$$25 \text{ g } Mg_3N_2 \times \frac{1\text{mole } Mg_3N_2}{100 \text{ g } Mg_3N_3} \times \frac{3\text{mole Mg}}{1\text{mole } Mg_3N_2} \times \frac{24 \text{ Mg}}{\text{mole Mg}} = 18 \text{ gram Mg}$$

B. The Mole - ratio Method

In this method, calculations are made in terms of moles; therefore the given mass is converted into mole. The obtained mole can be converted back to mass if required.

In general, for mole –mole problems, mole A x mole ratio = mole B

If substance A is given in grams, it should be converted to moles as follows:

Gram of A x mole of A/ Molar mass of A x mole A x mole ratio = mole B

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

$$CaO + CO_2 \rightarrow CaCO_3$$

gram CO_2 → Mole CO_2 → mole CaO → gram CaO

$$\text{gram } CO_2 \times \frac{1\text{mole } CO_2}{\text{molar mass } CO_2} \times \frac{\text{mole CaO}}{\text{mole } CO_2} \times \frac{\text{molar mass CaO}}{\text{mole CaO}} = \text{gram Mg}$$

$$22.0\text{ g }CO_2 \times \frac{1\text{mole }CO_2}{44.0\text{ g }CO_2} \times \frac{1\text{mole CaO}}{1\text{mole }CO_2} \times \frac{56\text{ g CaO}}{\text{mole CaO}} = 28\text{ gram CaO}$$

Alternative method for solving mass-mole relations.

Step 1: Write the balanced chemical equation.

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

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

Step 4: Set up the proportion.

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

Solving example 1.27 using the above steps

Step 1: $CaO + CO_2 \rightarrow CaCO_3$

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

$$\begin{array}{lcccc} & x & & 0.5\text{mol} & \\ \textbf{Step 3:} & CaO & + & CO_2 & \rightarrow CaCO_3 \\ & 1\text{mol} & & 1\text{mol} & \end{array}$$

Step 4: x/1mol = 0.5mol/1mol

x = 0.5 mol

Step 5: Convert the moles into grams of CaO

Mass of CaO produced = No. of moles CaO x molar mass CaO = 0.5mol x 56 g/mol = 28 g

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

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

$$\text{gram Fe} \times \frac{1\text{mole Fe}}{\text{molar mass Fe}} \times \frac{\text{mole CO}}{\text{mole Fe}} \times \frac{\text{molar mass CO}}{\text{mole CO}} = \text{gram CO}$$

$$28.0\text{ g Fe} \times \frac{1\text{mole Fe}}{56.0\text{ g Fe}} \times \frac{3\text{mole CO}}{2\text{mole Fe}} \times \frac{28\text{ g CO}}{\text{mole CO}} = 21\text{ gram CO}$$

28 g x 1mole Fe/ 56 g Fe x 3mole CO/2mole Fe x molar mass of CO/mole CO. Find the mass of CO using the stepwise method.

C . Volume – Volume Relationships

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

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

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

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

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

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

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

$$\begin{array}{ccccccccc} & & x & & & & 44.8\text{ L} \\ CH_4 & + & 2O_2 & \rightarrow & 2H_2O & + & CO_2 \\ & & 44.8\text{ L} & & & & 22.4\text{ L} \end{array}$$

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

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

Step 5: Solve for the unknown volume, x   x = 89.6 L of oxygen is needed.

1.5.2. Limiting & Excess Reactants, Theoretical, Actual & % Yields

Chemical reaction equations give the ideal stoichiometric relationship among reactants and products. However, the reactants for a reaction in an experiment are not necessarily a stoichiometric mixture.

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

Concept of Limiting Reactant

Consider a nonchemical example. Suppose you want to make some cheese sandwiches. Each is made from two slices of bread and a slice of cheese as shown below:

2 slices bread + 1 slice cheese → 1 cheese sandwich

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

Now consider the following chemical reaction as an example of a limiting reactant.

Example 1.31: Magnesium metal reacts with hydrochloric acid by the following reaction. If 0.30 mol Mg is added to hydrochloric acid containing 0.52 mol HCl, how many moles of H_2 are produced?

The actions listed below must be taken in order to identify the reaction's limiting reactant

1. Determine the reaction and make sure it is balanced

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

2. Find the number of moles of the product using each reactant.

A. Quantity of H_2 produced using Mg

$$\begin{array}{lllll} 0.3\text{ mol} & & & & x \\ Mg & + \ 2HCl & \rightarrow & MgCl_2 & + \ H_2 \\ 1\text{ mol} & & & & 1\text{ mol} \end{array}$$

0.3mol/1mol = x/1mol

0.3 mol/1 mol = x/1 mol   x = **0.30 mol of H_2**

B. Quantity of H_2 produced using HCl

$$\begin{array}{lllll} & 0.52\text{ mol} & & & x \\ Mg & + \ 2HCl & \rightarrow & MgCl_2 & + \ H_2 \\ & 2\text{mol} & & & 1\text{mol} \end{array}$$

0.52 mol/2 mol = x/1mol

0.52 mol/2 mol = x/1 mol   x = **0.26 mol of H_2**

3. Compare the products obtained by the two reactants to determine which reactant is limiting. The reactant which produces the smallest number of moles of product is the limiting reactant. In the above reaction, HCl is the limiting reactant because it produced less quantity of H_2

4. Calculate the number of moles of product that can be obtained from the limiting reactant.

The product yield is the quantity calculated using the limiting reactant. In the above reaction, the yield is **0.26 mol of H_2 because it is the quantity calculated using the limiting reactant(HCl).**

Example 1.32: 10.0 g of oxygen gas was bubbled into 20.0 g acetaldehyde, CH_3CHO, into a reaction vessel under pressure at 60 °C , in a process for producing acetic acid, based on the reaction below:

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

A. How many grams of acetic acid can be produced? B. How many grams of the excess reactant remain after the reaction is complete?

Solution

1. The balanced equation is given in the question.
2. Convert grams in to moles

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

$$= \frac{20\text{ g}}{40\text{ g/mol}} = \mathbf{2\ mol}$$

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

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

Now, find the number of moles of the product using each reactant.

A. Using the quantity of CH_3CHO

$$\begin{array}{lcccc} 0.454 \text{ mol} & & & & x \\ 2CH_3CHO & + & O_2 & \rightarrow & 2HC_2H_3O_2 \\ 2 \text{ mol} & & & & 2 \text{ mole} \end{array}$$

0.454 mol/2 mol = x/2 mol   x = 0.454 mol of $HC_2H_3O_2$

B. Using the quantity of O_2

$$\begin{array}{lcccc} & & 0.312 \text{ mol} & & x \\ 2CH_3CHO & + & O_2 & \rightarrow & 2HC_2H_3O_2 \\ & & 1 \text{ mol} & & 2 \text{ mol} \end{array}$$

0.312 mol/1 mol = x/2 mol   x = **0.624 mol of $HC_2H_3O_2$**

CH_3CHO is the limiting reactant. The quantity of acetic acid ($HC_2H_3O_2$) produced is 0.454 mol. Why? Because it is the quantity determined using the limiting reactant.

Convert moles of $HC_2H_3O_2$ to grams using molar mass of $HC_2H_3O_2$

0.454 mol x 60 g $HC_2H_3O_2$/ mol = 27.24 g $HC_2H_3O_2$

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

$$\begin{array}{lcccc} & & x & & 0.454 \text{ mol} \\ 2CH_3CHO & + & O_2 & \rightarrow & 2HC_2H_3O_2 \\ & & 1\text{mol} & & 2 \text{ mol} \end{array}$$

Amount of O_2 consumed, x = 0.454 mol x 1 mol/2 mol = **0.227 mol**

We started with 10 g of O_2 or 0.312 mol, so the excess O_2 = 0.312 - 0.227 = 0.085mol or 2.7 g.

No. of moles of O_2 = given mass of O_2 / molar mass of O_2 = 10 g/32 g/mol = 0.312 mol

1.5.3. Theoretical, Actual and Percentage Yields

The amount of limiting reactant present at the start of a reaction determines the theoretical yield of the reaction- that is, the amount of product that would result if all the limiting reactant reacted. The theoretical yield, then, is the maximum obtainable yield, predicted by the balanced equation.

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

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

$$\text{Percentage yield} = \frac{\text{Actual(measured) yield}}{\text{Theoretical (calculated)yield}} \times 100\%$$

Example 1.33: What is the percent yield if 24.8 g of $CaCO_3$ is heated to give 13.1 g of CaO?

Solution

The balanced chemical equation is : $CaCO_3 \rightarrow CaO + CO_2$

The actual yield of CaO is 13.1 g (given in the problem!). Determine the theoretical yield using mass-mass relationship:

$$\text{gram } CaCO_3 \times \frac{1\text{mole } CaCO_3}{\text{molar mass } CaCO_3} \times \frac{1\text{mole CaO}}{1\text{mole } CaCO_3} \times \frac{\text{Molar mass CaO}}{1\text{mole CaO}} = \text{gram CaO}$$

$$24.8 \text{ g} \times \frac{1\text{mole } CaCO_3}{100 \text{ g}} \times \frac{1\text{mole CaO}}{1\text{mole } CaCO_3} \times \frac{56 \text{ g}}{1\text{mole CaO}} = 13.9 \text{ gram CaO}$$

13.9 g is the theoretical yield of CaO

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

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

Checklist - 1.5

Please put a tick (√) if your answer is 'yes' if not go back and revise it to refer to relevant reference books. I can ...

Competencies	Check
1. deduce mole ratios from balanced chemical equations.	
2. define molar volume.	
3. state Avogadro's principle.	
4. calculate molar masses of some substances.	
5. solve mass -mass problems based on the given chemical equation.	
6. solve mass -volume problems based on the given chemical equation.	
7. solve volume-volume problems based on the given chemical equation.	
8. differentiate limiting and excess reactant.	
9. differentiate limiting reactant and reaction yield.	
10. calculate the theoretical and percentage yield of a given reaction.	
11. calculate molar masses of some substances.	
12. describe the meanings of actual, theoretical and percentage yields.	
13. explain why percentage yields cannot be 100 %.	

Self-Test Exercise 1.5

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

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

   a. 6 g of H_2O are produced for every 4 g of NH_3 reacted.
   b. 1 mole of NO is produced per mole of NH_3 reacted.
   c. 2 moles of NO are produced for every 3 moles of O_2 reacted

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

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

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

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

6. What mass of nitrogen dioxide is produced by the decomposition of 182 g of magnesium nitrate?
7. Solve the following using either mass-mole or mole-mass relations
   a. What mass of hydrogen sulphide gas is burned in a reaction that produces 4 mol SO_2 and water vapor: $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$
   b. How many moles of CO_2 will be produced when 50 g of $CaCO_3$ are reacted according to the chemical equation:

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

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

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

11. Calculate the volume of oxygen produced in the decomposition of 5 moles of $KClO_3$ at STP?
12. How many moles of water vapour are formed when 10 litres of butane gas, C_4H_{10} is burned in oxygen at STP?
13. Calculate the mass of calcium carbide that is needed to produce 100 cm^3 of acetylene according to the equation:

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

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

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

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

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

a. Which substance is the limiting reactant?
b. How many grams of the reactant remain unreacted?
c. How many grams of hydrogen would be produced?

16. What mass of Na_2SO_4 is produced if 49 g of H_2SO_4 reacts with 80 g of NaOH?
17. If 20 g of $CaCO_3$ and 25 g of HCl are mixed, what mass of CO_2 is produced?

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

18. If 3 moles of calcium react with 3 moles of oxygen, then
    a. Which substance is the limiting reactant?
    b. How many moles of calcium oxide are formed?
19. For the reaction: $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$, how many grams of hydrogen are produced if 0.8 mole of aluminum reacts with 1.0 mole of sulphuric acid?
20. When 21.6 g of aluminum reacts with oxygen, 20.8 g of aluminum oxide is produced. What is the percentage yield of aluminum oxide?
21. When 14.5 g of SO_2 reacts with 21 g of O_2, what will be the theoretical yield and percentage yield of the reaction if the actual yield is 12 g?
22. In the reaction: $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2$, when 52.7 g of octane (C_8H_{18}) burns in oxygen, the percentage yield of carbon dioxide is 82.5%. What is the actual yield in grams?

Unit Summary

Chemical equations are representations of chemical reactions using chemical symbols and formulas.

In any reaction, the total mass of reactants equals the total mass of products according to the law of conservation of mass. Thus, we should always balance chemical equations.

Chemical reactions can be classified as combination, decomposition, single displacement or double displacement reactions.

When a substance loses electron (s) or increases its oxidation number in a reaction, the process is oxidation and when a substance gains electron (s) or decreases its oxidation number, the process is reduction.

Oxidizing agents are substances reduced, and reducing agents are the substances oxidized in reduction –oxidation (redox) reaction.

Reactions in which neither oxidation nor reduction takes place are non-redox reactions. Neutralization and double displacement reactions are examples of non-redox reactions.

Molecular mass or formula mass expressed in grams is called molar mass. Molar mass is the mass of one mole of a substance.

A mole is defined as the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12. 12 g of carbon-12 contains $6.022x10^{23}$ atoms. This number is known as Avogadro's number (NA).

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

In chemical reactions, the quantitative relationship between reactants and products is called stoichiometry. There are mainly mass-mass, mass volume and volume-volume stoichiometric relations

The reactant completely consumed first in a reaction is called the limiting reactant. The reactants present in quantities greater than necessary to react with the quantity of the limiting reactants are excess reactants

In chemical reactions, the actual (experimentally determined) yield of a product is usually less than the theoretical (calculated) yield. The percentage yield is the ratio of the actual yield to the theoretical yield multiplied by 100.

Self-Assessment Exercises

I. Multiple choice questions

1. On immersing an iron nail in $CuSO_4$ solution for a few minutes, you will observe-
   a. No reaction takes place
   b. The colour of the solution fades away
   c. The surface of iron nails acquires a black coating
   d. he colour of the solution changes to green
2. What happens when dilute hydrochloric acid is added to iron filings?
   a. Hydrogen gas and iron chloride are produced.
   b. Chlorine gas and iron hydroxide are produced.
   c. No reaction takes place.
   d. Iron salt and water are produced.
3. Chlorine gas is passed in an aqueous potassium iodide solution to form potassium chloride solution and solid iodine. Identify the type of reaction:
   a. Decomposition reaction
   b. Double Displacement reaction
   c. Displacement reaction

d. None of these

4. The formula weight of the compound, $Al_2(SO_4)_3$ $18H_2O$ is:
   a. 394.4 g
   b. 666.4 g
   c. 110,900 g
   d. 466.8 g
5. Which of the following statements is(are) FALSE?
   a. The percent by mass of each element in a compound depends on the amount of the compound.
   b. The mass of each element in a compound depends on the amount of the compound.
   c. The percent by mass of each element in a compound depends on the amount of element present in the compound.
   d. None
6. What is the percent, by weight, of carbon in 154 g of $C_4H_8O_3$?
   a. 46%
   b. 31%
   c. 72
   d. 27%

II. Fill in the blank questions

7. Write the empirical formula of the following compounds:
   a. C_4O_{12} __________
   b. SiO_2 __________
   c. $N_4H_8Cl_2$ __________
8. The limiting reactant is the reactant that __________ .
9. When the actual yield is equal to the theoretical yield, the percentage yield will be _____
10. An oxidizing agent is a substance that ____________ electrons and a reducing agent is a substance ____________electrons.
11. The oxidation state of carbon in $C_6H_{12}O_6$ is ____________
12. In a chemical reaction, the total mass of the reactants is equal to the total mass ______ hence the mass remains conserved in a chemical reaction.
13. In __________ reaction, a single substance reacts to make multiple substances as products.

III. Short answer questions

14. Select (i) combination reactions (ii) decomposition reactions and displacement reactions from the following-
    a. $ZnCO_3 \rightarrow ZnO + CO_2$
    b. $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$
    c. $H_2 + Cl_2 \rightarrow 2HCl$
    d. $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$
    e. $3H_2 + N_2 \rightarrow 2NH_3$
15. Why is it always essential to balance a chemical equation?
16. A mixture of 128.5g of P and 132.8g of O_2 reacts completely to form P_4O_6 and P_4O_{10}.

Find the masses of P_4O_6 and P_4O_{10} that are formed by the reaction based on the reaction

$2P_4 + 5O_2 \rightarrow P_4O_6 + P_4O_{10}$

4. 448g of iron(III) oxide, Fe_2O_3, reacts with an excess of carbon monoxide, CO in the following reaction. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g$.Calculate the mass of iron solid, Fe, produced and use the following molar masses in your calculation: Fe_2O_3=159.7g mol^{-1}. Fe=55.85g mol^{-1}.

17. Identify the substance oxidised and substance reduced in the following reactions-
    a. $ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$, Zn reduced, C is oxidized
    b. $2Na(s) + O_2(g) \rightarrow 2Na_2O(s)$, Na is oxidized, O reduced
    c. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$. Cu reduced, H_2 oxidised.
18. Balance the following equations
    a. $Ba_3N_2 + H_2O \rightarrow Ba(OH)_2 + NH_3$;
    b. KNO□ + $H_2CO_3 \rightarrow K_2CO_3 + HNO_3$
    c. $CaCl_2 + Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + NaCl$
    d. FeS + O□ → Fe□O□ + SO□
    e. $FeCl_3 + NOH_5 \rightarrow Fe(OH)_3 +$ NH4Cl
    f. $NaOH + FeCl_3 \rightarrow NaCl + Fe(OH)_3$
    g. $Al + HCl \rightarrow AlCl_3 + H_2$
    h. $FeS + O_2 \rightarrow Fe_2O_3 + SO_2$
19. Determine the empirical formula consisting of
    a. 55.3% K, 14.6% P, and 30.1% O.
    b. 52.14%C, 13.13%H, and 34.73% O.
20. Cyclobutane has the empirical formula CH2. Its molar molar mass is 42g. What is its molecular formula?
21. A 10.00 g sample of vitamin C was analyzed and found to contain 4.092 g of C, 0.458 g of H, and 5.450 g of O. Given that the molar mass of the compound is 176g, determine the molecular formula of vitamin C.

Assignment for Submission (12 points)

Answer the following questions correctly

1. In an industrial process, hydrogen chloride, HCl, is prepared by burning hydrogen gas, H_2, in an atmosphere of chlorine, Cl_2. Write the chemical equation for the reaction. Below the equation, give the molecular, molar, and mass interpretations.
2. How many moles of calcium (Ca) are required to react with 2.5 mole of chlorine (Cl) to produce the compound $CaCl_2$ (calcium chloride)?
3. Photosynthesis is a chemical reaction that is vital to the existence of life on Earth. During photosynthesis, plants and bacteria convert carbon dioxide gas, liquid water, and light into glucose ($C_6H_{12}O_6$) and oxygen gas.
    a. Write down the equation for the photosynthesis reaction.
    b. Balance the equation.
    c. If 3 moles of carbon dioxide are used up in the photosynthesis reaction, what mass of glucose will be produced?
4. A. What is the percentage composition of chloroform, $CHCl_3$, a substance once used

as anesthetic? B. How many moles of Na_2CO_3 are in 132 g Na_2CO_3?

5. A colorless liquid used in rocket engines, whose empirical formula is NO_2, has a molecular mass of 92.0. What is its molecular formula?
6. Aluminum reacts with oxygen to aluminum oxide Al_2O_3, as shown below:

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

Calculate the number of grams of Al_2O_3 that could be formed if 12.5 g O_2 react completely with aluminum.

7. Calculate the formula mass of each of the following to three significant figures, using a table of atomic masses. iron(III) sulfate, $Fe_2(SO_4)_3$.
8. Zinc metal reacts with hydrochloric acid by the following reaction:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

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

9. Aluminum chloride, $AlCl_3$, isused as a catalyst in various industrial reactions. It is prepared from hydrogen chloride gas and aluminum metal shavings.

$$2Al(s) + 6HCl(g) \rightarrow 2AlCl_3(s) + 3H_2(g).$$

Suppose a reaction vessel contains 0.15 mol Al and 0.35 mol HCl. How many moles of $AlCl_3$ can be prepared from this mixture?

10. Sodium carbonate, has molecular formula Na_2CO_3
    a. How many grams does 0.250 mol Na_2CO_3?
    b. How many moles of Na_2CO_3 are in 132 g Na_2CO_3?
11. What is the empirical formula of a compound composed of 43.7% P and 56.3% O by weight?

Answers to Self-Review Exercises

I.

1. d. The color of the solution changes to green
2. a. Hydrogen gas and iron chloride are produced.
3. c. Displacement reaction
4. a. $ZnCO_3 \rightarrow ZnO + CO_2$ (decomposition)
   b. $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$ (displacement)
   c. $H_2 + Cl_2 \rightarrow 2HCl$ (direct combination)
   d. $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$(displacement)
   e. $3H_2 + N_2 \rightarrow 2NH_3$(direct combination)
5. The oxidation state of carbon in $C_6H_{12}O_6$ is zero
6. to the total mass of the products
7. decomposition

II.

8. a. $ZnCO_3 \rightarrow ZnO + CO_2$ (decomposition)
   b. $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$ (displacement)
   c. $H_2 + Cl_2 \rightarrow 2HCl$ (direct combination)
   d. $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$(displacement)
   e. $3H_2 + N_2 \rightarrow 2NH_3$(direct combination)

9. To obey the law of conservation of mass.

10. P_4O_{10}= 147.26 g, P_4O_6 = 114.04 g

11. 312.76 g

12. A. $ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$, Zn reduced, C is oxidized

    B. $2Na(s) + O_2(g) \rightarrow 2Na_2O(s)$, Na is oxidized, O reduced

    C. $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$. Cu reduced, H_2 oxidized.

13. Balance the equations

    A. $Ba_3N_2 + 6H_2O \rightarrow 3Ba(OH)_2 + 2NH_3$

    B. $2KNO_3 + H_2CO_3 \rightarrow K_2CO_3 + 2HNO_3$

    C. $3CaCl_2 + 2Na_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6NaCl$

    D. $4FeS + 7O_2 \rightarrow 2Fe_2O_3 + 4SO_2$

    E. $FeCl_3 + 3NOH_5 \rightarrow Fe(OH)_3 + 3NH_4Cl$

    F. $3NaOH + FeCl_3 \rightarrow 3NaCl + Fe(OH)_3$

    G. $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

    H. $4FeS + 7O_2 \rightarrow 2Fe_2O_3 + 4SO_2$

14. A. K_3PO_4 B. C_2H_6O

15. C_3H_6

16. $C_6H_8O_6$

Answer Key to Self Test Exercises

Answe Key to Activity 1.1

Qualitatively, magnesium burns in oxygen to give magnesium oxide

Quantitavely, 2 mole of Mg cobines with one mole of oxygen to give 2 moles of magnesium oxide.

Answe Key to Activity 1.2

A. $2KClO_3 \rightarrow 2KCl + 3O_2$

B. $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

C. $2Ag_2O \rightarrow 4Ag + O_2$

Answe Key to Activity 1.3

1. Aminoacids and glucose = decomposition
2. Alcohol (CH_3CH_2OH) and carbondioxide (CO_2) = decomposition
3. photosynthesis: Reactants; H_2O and CO_2 & products; $C_6H_{12}O_6$ (glucose) & O_2(oxygen) = combination
4. H_2O and CO_2 = decomposition

Answe Key to Activity 1.4

Balance the following using oxidation number change method

A. $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$

B. $2KClO_3 \rightarrow 2KCl + 3O_2$

C. $3MnO_2 + 4Al \rightarrow 3Mn + 2Al_2O_3$

D. $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$

ANSWER KEY TO SELF TEST EXERCISES

Self-Test Exercise -1.1

1. Fill in the blank space
   a. Chemical equations
   b. Word equation
   c. Chemical equation
2. Write the balanced chemical equation to represent the following reactions.
   a. $2NaBr + Cl_2 \rightarrow 2NaCl + Br_2$
   b. $2HCl + Na_2CO_3 \rightarrow 2NaCl + H_2O + CO_2$
   c. $2KClO_3 \rightarrow 2KCl + 3O_2$
   d. $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$
   e. $2Ag_2O \rightarrow 4Ag + O_2$
3. Say true or false
   a. True　　b. False　　c. True
4. Balance the following chemical equation, using the inspection method:
   a. $2Na + 2H_2O \rightarrow 2NaOH + H_2$
   b. $2KClO_3 \rightarrow 2KCl + 3O_2$
   c. $2H_2O_2 \rightarrow 2H_2O + O_2$
   d. $2Al + 2H_3PO_4 \rightarrow 2AlPO_4 + 3H_2$
5. Balance the following equations by any method.
   a. $PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl$
   b. $Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$
   c. $2Zn(NO_3)_2 \rightarrow 2ZnO + 4NO_2 + O_2$
   d. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
   e. $4NH_3 + 5O_2 \rightarrow NO + 6H_2O$
   f. $C_6H_{12}O_6 + 6O_2 \rightarrow 6\ CO_2 + 6H_2O$
   g. $2FeCl_3 + 3MgO \rightarrow Fe_2O_3 + 3MgCl_2$
   h. $3BaCl_2 + 2K_3PO_4 \rightarrow Ba_3(PO_4)_2 + 6KCl$
   i. $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

Self-Test Exercise -1.2

1. Classify the following reactions as combination, decomposition, single or double displacement reactions.

a. Single – displacement reaction　　b. combination reaction
c. Double displacement reaction　　d. Decomposition reaction
e. Double displacement reaction　　f. Decomposition reaction

2. What type of reaction does usually take place in each of the following reactions?

a. Single – displacement reaction　　b. Combination reaction
c. Double displacement reaction　　d. Decomposition reaction

Self-Test Exercise - 1.3

1. Find the oxidation numbers to the underlined species for the following compounds or ions.

a. +2　　c. +4　　e. +7　　g. +6
b. +6　　d. +2.5　　f. +5　　h. +3

i. +6

2. Determine whether the following processes are oxidation or reduction reactions

a. Reduction
c. Reduction
e. Oxidation
b. Oxidation
d. Oxidation
f. Reduction

3. Transition elements at the right of the periodic table such as silver and copper can easily be reduced (strong oxidizing agents) from their compounds. Alkali metals like potassium can easily be oxidized (strong reducing agents)

4. For the following substances, tell whether the oxidation number increases or decreases in a redox reaction:

a. decrease
c. increase
b. increase
d. decrease

5. In the following reactions, label the oxidizing agent and the reducing agent.

a. Zn = oxidizing agent, C = reducing agent

b. Fe = reducing agent, S_8 = oxidizing agent

c. Fe_2O_3 = oxidizing agent, CO = reducing agent

d. PbS = reducing agent, H_2O_2 = oxidizing agent

6. Balance the following using oxidation number change method

a. $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$
b. $2KClO_3 \rightarrow 2KCl + 3O_2$

Self-Test Exercise - 1.4

1. Refer to the text
2. Refer to the text
3. A. K_2CO_3 = 2K atoms + 1C atom + 3O atoms = 6 atoms
   = 2K+ ions + 1 CO_3^{2-} = 3 ions
   B. $Ba_3(PO_4)_2$ = 3Ba atoms + 2P atoms + 8 O atoms = 13 atoms

4. a. 1.00 mole of $Cu_3(PO_4)_2$ = 3 moles of Cu + 2moles of P + 8 moles of O

b. 2.5 mole of Na_2CO_3 = 2 (2.25 moles) Na, 2.5 moles of C, 3 ×2.5 moles of O

c. 2 moles of $Al_2(SO_4)_3$ = 2×2 moles of Al, 2×3 moles S + 2×12 moles of O

5. a. NaCl b. K_2CO_3

6. a. C_6H_6 b. $C_2H_4Cl_2$

7. 408.6g/mole

8. a. CH_4 = 16 b. NO_2 = 46 c. SO_3 =80 d. C_6H_6 =78 e. Na_2SO_4 =142 f. $Ca_3(PO_4)_2$ = 290

9. a. C = 52.2% , H=13 %, O = 34.8 % b. Na= 42.1 %, P=18.8 %, O= 39 %

c. H = 2.04%,S =32.65 %,O =65.31 % d. $CaCO_3$, Ca = 40 %, C = 12 %,O = 48 %,

e. $K_2Cr_2O_7$, K = 26.53 %. Cr = 35.37 %, O = 38.1 %

Self-Test Exercises 1.5

1. B
2. 2.25 mol
3. 6.6 moles
4. 0.03 mol
5. 20.82 g
6. 9.77 g
7. a. 136 g of CO_2 b. No of moles of CO_2 = 0.5mol

8. 16.8 liters of N_2
9. 6 liters of SO_3
10. 116 liters of NH_3
11. 168 liters of O_2
12. 2.23 moles
13. 336 liters of O_2
14. 298.7 g
15. 0.286 g
16. 3,046 milliliters
17. a. HCl b. 2.05 g c. 0.14 g
18. 71 g
19. 8.8 g of $CaCO_3$
20. a. Ca is the limiting reactant b. 3 moles of CaO
21. 2 g of H_2 is produced
22.

Yield = 63.4 %

23. $2SO_2 + O_2 \rightarrow 2SO_3$, the limiting reactant is SO_2 and 16.9 g of SO_3 is produced,

% yield = 12/16.9 x 100 %

=71 %

24.

Yield = Actual yield/ theoretical yield,

the calculated (theoretical) yield from the equation is 160 g,

therefore the actual yield = % yield x theoretical yield

= 0.855 x 160.6g

= 137 g

Experiment 1.1

Investigation of Single Displacement Reaction

1. Reddish brown color on the iron rod
2. $Cu^{2+} + Fe \rightarrow FeSO_4 + Cu$
3. Fe can displace copper from copper sulphate ($CuSO_4$) solution

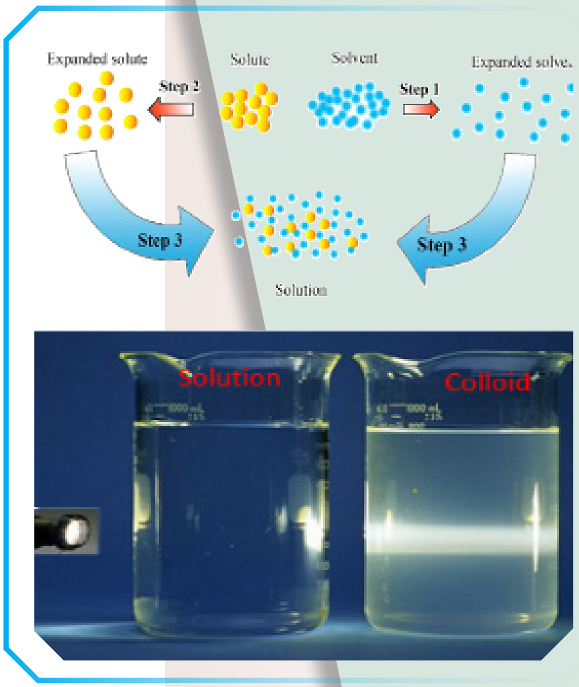
Experiment 1.2

Investigation of Double Displacement Reaction

1. The precipitate is barium sulphate and the compound exist in the form of its ions is sodium nitrate
2. The color of the precipitate is white
3. $Na_2SO_4 + Ba(NO_3)_2 \rightarrow BaSO_4 + 2NaNO_3$

References

The current Ethiopian Grade 10 Chemistry Textbook for face-to-face modality.



UNIT 2

SOLUTIONS

Introduction

❁ Dear learner, welcome to Grade 10 Chemistry Lesson on Solutions! You know that in chemistry, we study about the composition and properties of matter and the changes it undergoes. Some part of chemistry focuses on mixtures because some mixtures are even more important than pure substances they are composed of. For instance, the stainless steel that you use at your home as a spoon or cooking utensils is stronger and resistant to corrosion than the component substances (Fe, Cr, Ni, and C). Most of the medications for human and animal treatments need to be prepared as solutions-that are also mixtures. Many laboratory investigations in chemistry, biology and medicine are carried out in solutions. Thus, the matter of interest to us in this unit is a type of mixtures called solutions.

Before we proceed to our main topic, let's have a brief review about matter and its classification because a solution is a kind of matter. Recall from the previous Grades that matter is classified into pure substances and mixtures. Pure substances are further classified into elements and compounds. In chemistry, a substance is a kind of matter that consists of **one type of atom or molecule or formula units** and has definite composition and properties (melting point (m.p.), boiling point (b.p.), density, etc.). On the other hand, mixtures are combinations of two or more substances in which the component substances retain their distinct identities and thus can be separated by physical means. Mixtures have variable proportions and properties. Mixtures are classified into two-heterogeneous mixtures and homogeneous mixtures. They can further be classified into solutions, suspensions, and colloids.

Take the next minute to look at the pictures below.

(a) (b) (c) (d) (e)

Figure 2.1 (a) the IV solution being administered to a patient. (b) barely juice (c) a cup of coffee (d) a glass of milk (e) a cup of tea

? **What do you observe?**

What comes to your mind is most probably the IV fluid being administered to a patient, barely juice, coffee, milk, and tea. Don't you wonder how important these stuffs are to our lives?

? Which of these would form separate layers, which would coagulate, and which would remain unaltered upon standing? Which of these would keep a uniform composition (and hence taste) from top to bottom even upon standing for long time? Why do you think we have a different observation in each case?

If we analyze these stuffs carefully, we can identify at least two different components in each case. The IV fluid consists of water and some dissolved substances such as glucose and medications. The barely juice consists of particles of barely grains, dissolved nutrients (proteins, carbohydrates, etc), and water. The milk consists of water and globules of fat and proteins, carbohydrates, etc. The coffee and tea consists of water, caffeine, and other dissolved components. Since we have at least two different components in each case, the pictures represent the different types of mixtures. But, can you sort them apart as solutions, suspension, and colloids or homogeneous and heterogeneous mixtures? Do you know how some medical preparations such as IV fluids and ORS made? Can you prepare them? This unit will help you answer such questions and more. Our focus will be on solutions- their properties, types, the solution process, preparation of solutions and ways of expressing their concentrations.

Unit Outcomes

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

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

Unit Contents

The Required Study Time (60 Hours)

You should complete this unit in 2 months. If you schedule 2 hours a day, this will cost you a total of 60 hours.

Unit Learning Strategies

Dear learner, you can adapt to different learning strategies depending on the circumstances. We kindly advise you to read the content under each section carefully and try to associate to your daily experiences so that you will have a stabilized memory of the learned material. Be assisted with examples. Frequently check your progresses with the help of activity and self-test exercises provided under each sections. Be assisted with internet especially video lessons. You can also find a friend attending similar program arrange a meeting and hold joint discussion on topics that you find difficult at first hand. We assume that you can find schools in your locality. Feel free to contact school chemistry teachers. They will help you. Similarly students of grade 10 and above can also help you enrich your understanding about the subject matter. Hiring a tutor is another alternative that you can think of is deemed necessary. Please be damn sure that you understood the current topic very well before you proceed to the next section. Sketch your own study time table and try to respect it.

Section 2.1: Heterogeneous and Homogeneous Mixtures

- In this section you will study about the definitions of the terms mixtures, homogeneous mixtures, and heterogeneous mixtures. You will also study about the distinctions between these types of mixtures and examples of them.

At the end of this section, you should be able to
- Define the terms mixture, homogeneous and heterogeneous mixtures.
- Distinguish between homogeneous and heterogeneous mixtures.

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

A homogeneous mixture is a mixture in which the composition is uniform throughout the mixture. i.e., it has no visible boundaries the components are evenly distributed throughout the entire mixture. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that a substance has definite/constant composition regardless of the source or the way it is made. NaCl, for instance, is always formed in a ratio of 23 is to 35.5 by mass of Na and Cl atoms. Any sample of table salt (NaCl) whether a spoonful or a barrel contains that ratio. But a homogeneous mixture of NaCl and water may be formed by combining them in various proportions; 5 g in 100 g or 10 g in 100 g water; would all give a homogeneous mixture of NaCl and water called a solution. The proportion of NaCl and H_2O is different in the two samples. When a spoonful of sugar dissolves in water, the composition of the mixture, after sufficient stirring, is uniform throughout the solution. This solution is a homogeneous mixture but not a substance.

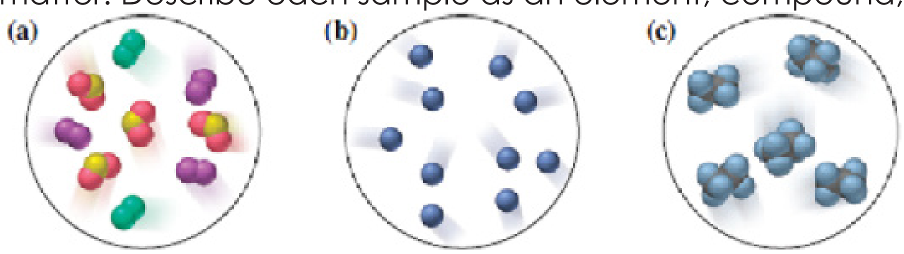
A heterogeneous mixture is a mixture that consists of physically distinct parts, each with different properties. This type of mixtures has at least two visible phases. Mixture of oil and water, wood ash and water, sand and water, etc. are good examples.

Mixtures can be converted into component substances by physical process and vice versa. Whereas a substance cannot be separated into other kinds of matter by any physical process. Thus, substances (compounds and elements) are inter-converted through chemical process only.

Now that you understand the difference between a substance and a mixture as well as a heterogeneous mixture and a homogeneous mixture, let me give you a chance to check your understanding. Please take the next minute to perform the following activity.

Activity 2.1

The scenarios below represent atomic-scale views of three samples of matter. Describe each sample as an element, compound, or mixture.



Self-Test Exercise 2.1

1. Define these terms: mixture, homogeneous mixture and heterogeneous mixture.
2. Classify each of the mixtures- sand + water, coffee, tea; seawater, air, brass; steel, natural gas, pizza, vinegar, vegetable salad, and fruit punch as a. Homogeneous if it has a uniform composition or b. Heterogeneous if it is an insoluble mixture of substances and has non-uniform composition. c. Where do you categorize blood, milk, butter and cloud: to homo or hetero?
3. A substance has "definite (or constant) composition" whereas a solution or a homogeneous mixture has a "uniform composition." Explain, using examples, the differences between the underlined phrases as applied in the definitions of substances and mixtures.

Section 2.2: Solutions, Suspensions, and Colloids

- Solutions, suspensions, colloids are similar as consisting of dispersions of one substance in another substance but have characteristics that set each one apart from the others.

At the end of this section, you should be able to

- Describe suspension and colloids.
- Define the terms solute, solvent, solution.
- Explain the different types of solutions.
- Give examples for each types of solutions.
- Present a report on how jewelry gold is made to class after a visit to nearby goldsmith; " describe a colloid.

- A solution is a homogenous mixture of solute and solvent. It is formed when a solute dissolves in a solvent. Generally, the major component of the solution is called solvent, and the minor component is called solute. If both components in a solution are 50%, the term solute can be assigned to either component. When a gaseous or solid material dissolves in a liquid, the gas or solid material is called the solute. When two liquids dissolve in each other, the major component is called the solvent and the minor component is called the solute.
- A solution has a single uniform phase and has no visible boundaries. This is because the solutes are evenly and completely dispersed throughout the solvent.
- Solutions are very stable. That means, the solutes do not settle upon standing or separate by filtration.
- Solutions are transparent to light. Solutions do not scatter light.
- The particles of a solution are very small; i.e they are of the size of individual molecules, ions, or atoms that can pass through the pores of even fine filter paper. Consequently, solutes cannot be separated from the solvent by filtration. Note that, although the components of a solution may not be separated by settling (sedimentation) or

filtration the solute can be separated from the solvent by other physical processes. For example, a solution of sodium chloride and water can be separated by distillation or evaporation.

Types of Solutions

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

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

ii. **Liquid solutions:** Solids, liquids and gases dissolve in a liquid solvent to form liquid solutions. Solutions in water, called aqueous solutions, are typical example of liquid solutions that are especially important in the chemistry laboratory and comprise a major portion of the environment and of all organisms. Other examples of liquid solutions include carbonated beverage (gas and liquid), alcoholic beverage (liquid in liquid), and sea water (solid in liquid). Note that in liquid solutions the solute can be a solid, a liquid, or a gas but the solvent is always a liquid.

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

Dear student, the following project is assigned to you to help you get practical experience about solutions. Please make a self-managed plan and gather experiences on how jewellery gold is made. Please include the learned materials into your note book and study for your exam.

Project 2.1

Do you know how jewellery gold is made? Make a self-managed plan, pay a visit to the nearby goldsmith, and write your experiences.

Suspensions

☞ If you take a glass of water and add in a handful of sand or dust, stir it for a while, and allow it to stand you will observe that the sand or dust do not dissolve in water, and, though it may look homogenous for a few moments, the sand or dust particles gradually sinks to the bottom of the glass. Therefore, a mixture of sand and water is a good example of suspension. Some medications are delivered as suspensions and must be mixed well before the doses measured to make sure the patient is receiving the correct amount of medication.

Definition: A suspension is a heterogeneous mixture that consists of a dispersion of fine solid particles in a liquid or gas, removable by filtration. The dispersed phase is a solute (solid) and the dispersing medium (liquid or gas) is called a solvent. The solute particles do not dissolve, but get suspended throughout the bulk of the solvent, left floating around freely in the medium. The particles in a suspension are far larger than those of a solution, so the components of a suspension could also separate under the influence of gravity (sedimentation) if left undisturbed. A mixture of sand and water is an example. Suspensions are considered heterogeneous because the different substances in the mixture will not remain uniformly distributed if they are not actively being mixed.

Colloids

Colloids exhibit properties intermediate between those of suspensions and solutions. The suspended particles are smaller and **do not** settle (separate) under the influence of gravity. This distinguishes a colloid from a suspension in which the suspended particles settle (separate) under the influence of gravity. Fog is an example of a colloid: it consists of very small water droplets (dispersed phase) in air (continuous phase). Other examples of colloids include milk and blood. Colloids are characterized as according to their state (solid, liquid, gas) of the dispersed phase and the continuous phase. We classify gaseous colloids as aerosols (smoke, fog, mist, clouds), liquid as emulsion (milk, blood, paints), foam (Soap suds, whipped cream), or sol, and solid colloids as gel, foam, or solid sol.
Aerosols are gaseous colloidal mixtures. They contain small particles of liquid or solid dispersed in a gas (e.g. smoke contains solid particles in gas while **fog** and **mist** contain liquid droplets in gas). Similarly, clouds are gaseous colloidal mixtures composed of air and water droplets that are small enough that they do not settle out.

Liquid sol is a colloidal suspension of solid particles in a **liquid** (e.g. Paints, cell fluids). In emulsion, both the dispersed phase and dispersing media are liquids. Examples of emulsion include milk, and a well shaken mixture of oil and water. Liquid foam is formed when many gas particles are dispersed in a liquid (e.g. Soap suds, whipped cream).

A gel is a solid colloidal mixture of particle of liquid in a solid. Similarly, a cheese is a solid emulsion of liquid in a solid. Solid foam is formed when many gas particles are dispersed in a solid. Lava and pumice are typical examples of solid foams. ***Figure 2.2*** illustrates some examples of colloidal systems. ***Table 2.1*** lists various types of colloids and some examples of each.

Gaseous colloids (aerosol) | Liquid Colloids | Colloidal emulsion | Solid colloid

Figure 2.2. Some practical examples of colloids.

Table 2.1 Types of colloids.

Dispersed Phase	Dispersing media	Name	Example
Liquid	Gas	aerosol	Fog, clouds, mist
Solid	Gas	aerosol	Smoke, automobile exhaust
Gas	Liquid	foam	Whipped cream, soapsuds, Shaving cream
Liquid	Liquid	emulsion	Mayonnaise (oil dispersed in water), Milk, face cream
Solid	Liquid	sol	Paint, Milk of magnesia, mud, AgCl (s) dispersed in H_2O
Gas	Solid	foam	Pumice, plastic, foams, rubber, sponge
Liquid	Solid	Gel	Jelly, cheese, butter, opal
Solid	Solid	Solid sol	Ruby (milky) glass, colored gemstone

Properties of Colloids

Colloids exhibit properties intermediate between those of suspensions and solutions.

A colloid (or colloidal solution) appears to be homogeneous but actually it is heterogeneous upon closer inspection.

The size of particles in a colloid is bigger than those in a true solution but smaller than those in a suspension. It is between 1 nm and 100 nm in diameter.

The particles of a colloid can pass through a filter paper. So, a colloid cannot be separated by filtration.

Although the dispersion appears uniform, even under a microscope, it is not as clear as a true solution. Thus, most colloids appear cloudy or opaque unless they are very dilute.

There are two phases in colloidal solution. They are known as the dispersed phase and the dispersion medium. The component present in smaller proportions is the dispersed phase while the one present in greater proportion is the dispersion medium.

Colloids are quite stable. The particles do not separate on keeping.

A colloid (or a colloidal solution) scatters a beam of light passing through it (because its particles are fairly large). As a result, these rays as well as colloidal particles become visible. The phenomenon in which the particles in a colloid scatter the beams of light that are directed at them and make the path of the light beam visible is called Tyndall effect (***Figure 2.3***). This effect is not observed in true solutions. You may expect suspensions to display Tyndall effect. But the Tyndall effect in suspensions is unstable and disappear as the particles settle down.

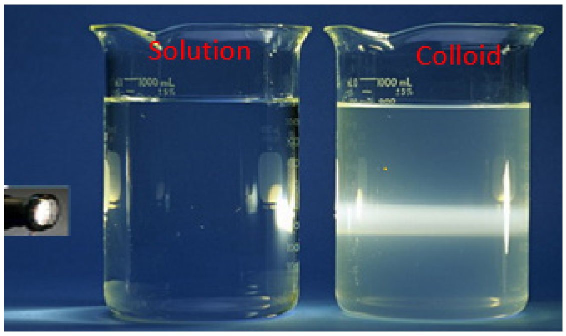


Figure 2.3 The light beam is not visible as it passes through true solution (left) but it is clearly visible when it passes through a colloid (right).

The other two industrially important properties of colloids are their coagulation and association. So, let's have a brief look at these very important properties of colloids one by one.

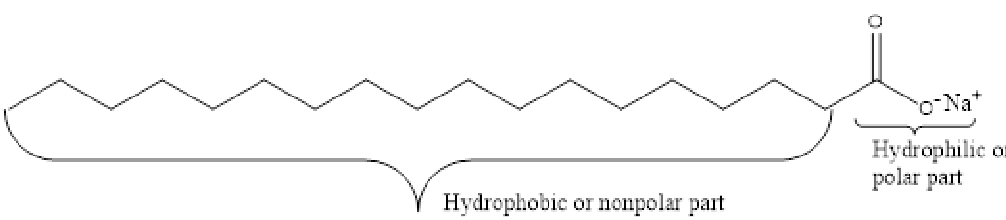
☞ **Coagulation**

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

☞ **Association of colloids**

When molecules or ions that have both a hydrophobic (nonpolar, water hating) and a hydrophilic (polar, water loving) ends are dispersed in water, they associate, or aggregate, to form colloidal-sized particles called micelles. The hydrophobic end point inward toward one another and the hydrophilic ends are on the outside of the micelle facing the water molecules. A colloid in which the dispersed phase consists of micelles is called **an association colloid**.

Ordinary soap in water provides an example of an association colloid. Soap consists of compounds such as sodium stearate, $C_{17}H_{35}COONa$. The stearate ion has a long hydrocarbon end that is hydrophobic (because it is nonpolar) and a carboxyl group (COO-) at the other end that is hydrophilic (because it is ionic).



Sodium stearate (common soap)

In water solution, the stearate ions associate into micelles (***Figure 2.4***) in which the hydrocarbon end point inward toward one another and away from the water, and ionic carboxyl groups are on the outside of the micelle facing the water.

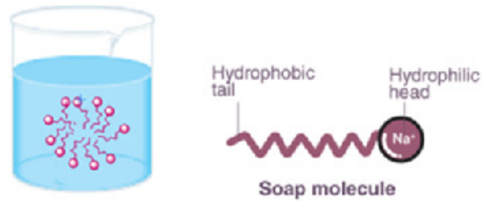


Figure 2.4. Micelle formation in water. The soap micelles represented by a polar head and zigzag hydrocarbon tail.

The cleansing action of soap occurs because oil and grease can be absorbed into the hydrophobic center of the soap micelles and washed away (*Figure 2.5*).

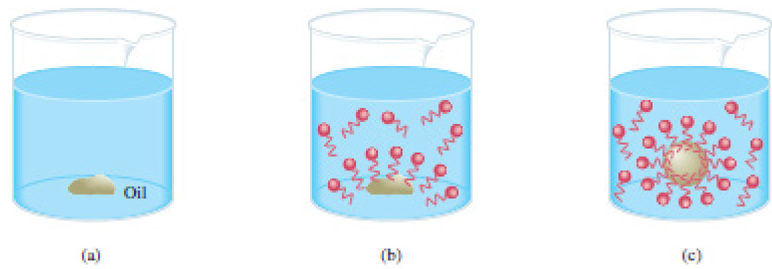


Figure 2.5 The cleansing action of soap. An oily spot (a) can be removed by soap (b) because the nonpolar tail dissolves in the oil, and (c) the entire system becomes soluble in water because the exterior portion is now ionic and washed away.

So far, we have studied about solutions, suspensions, and colloids. We have tried to characterize these tree types of mixtures by considering their appearance, particle size, and extent of dissolution, components, and effect on light, sedimentation, and separation by filtration. Now, let's summarize this in *Table 2.2*.

Table 2.2. Summary of characteristics of solutions, suspensions, and colloids

Characteristics	Solutions	Suspensions	Colloids
Appearance	Clear, transparent and homogeneous	Cloudy, heterogeneous, at least two phases visible	Cloudy, uniform, homogeneous
Particle Size	Atomic or molecular in size	larger than 10,000 Angstroms	10-1000 Angstroms
Dissolution	Completely dissolved	Not dissolved but dispersed	Not dissolved but dispersed finely
Components	Ions, atoms, or molecules dissolved in liquid, solid, or gas	fine solid particles suspended in liquid	Fine Particles suspended in liquid, solid, or gas
Effect of Light (Tyndall Effect)	Transparent (None)	Temporary (disappears when parties settle down)	Yes
Sedimentation	none	Yes	none
Filtration	No	yes	No

We have completed our discussion of the first section of this unit. I hope you have understood the difference between homogeneous and heterogeneous mixtures and the way we classify mixtures as solutions, suspensions, and colloids and you should be able to sort out most of the materials you encounter in your day-to-day life as solutions, suspensions, or colloids. To help you enrich the understanding about these stuffs, I kindly advise you to take some time to perform the following activity.

Activity 2.2

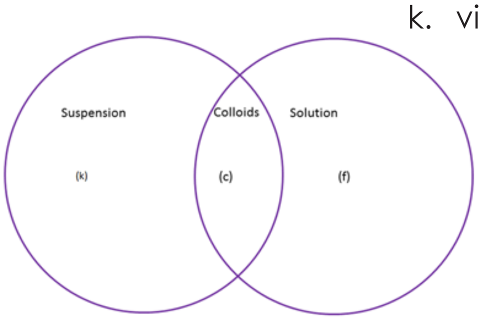
Read on "cloud seeding" technology and understand the mechanism. Then, prepare for a conversation to argue for or against on the intervention of this technology on the natural water cycle to produce artificial rain. You may seek information from Internet.

1. Make your own plan and pay a visit to nearby municipality water treatment plant. Ask if coagulation is involved and the chemicals used for this purpose.
2. Recall that oil is insoluble in water and forms a distinct layer. Perform the following activities at home and write your findings.
3. Take half a cup of oil and half a cup of water. Pour them carefully into a glass. Pour the water first and the oil later.
    1. Write your observation. Do they mix?
    2. If not, stir them thoroughly with a stirring rod or any dry stick. Then put aside to allow them stand. What do you observe?
    3. Now add a tea spoonful of soap powder and mixt thoroughly. Write your observation and explain.
    4. What is the analogy between the use of soap in this experiment and that of bile in the digestion of fat in your body?
4. Make your own hair gel using locally available materials. You may be assisted with internet. Write all the procedures you followed including the name of the materials, the source, the amount and the processing procedures and steps. You may consider a combination of gelatin or flaxeed with some kind of essential oils. For the essential oil, ethanol (''areki'') extracts from lavender, rosemary, lemongrass, thyme, or tea tree can be used.

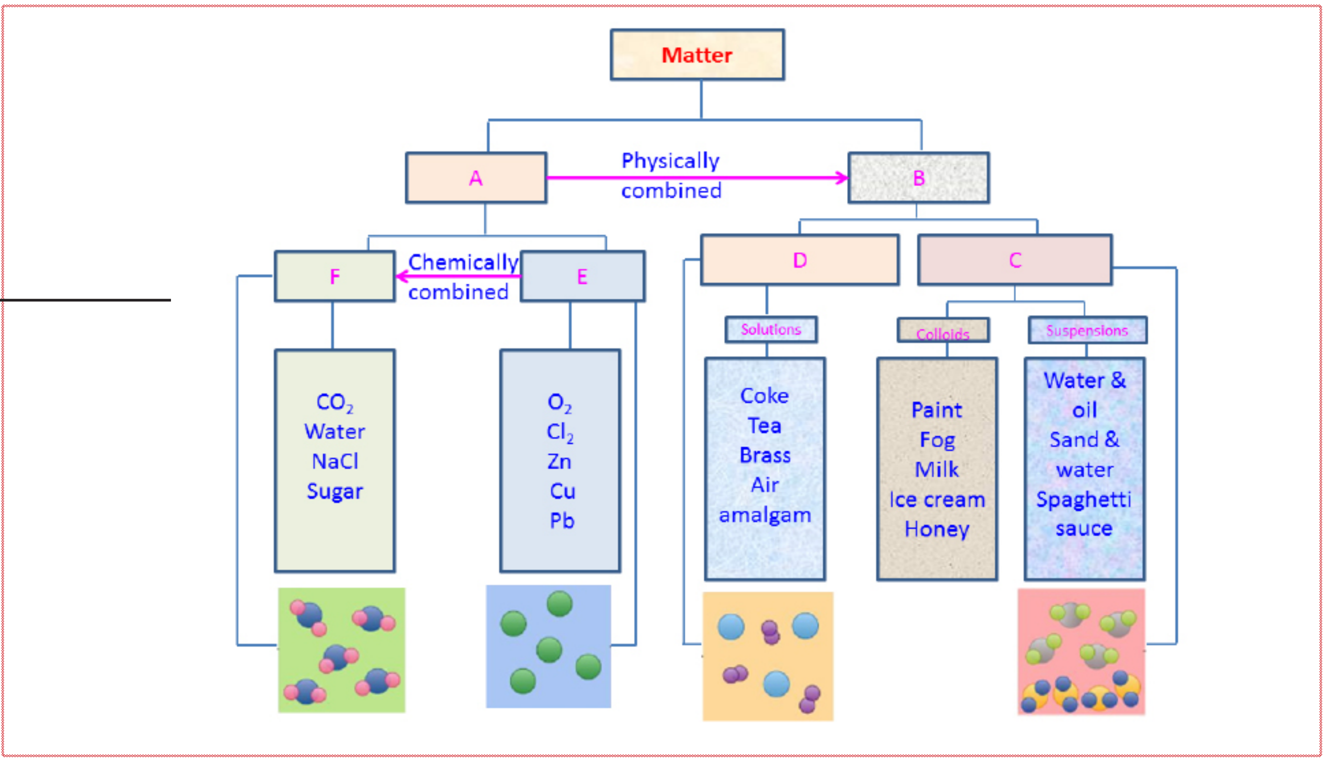
Now, spend about 30 minutes to do the following self-test exercise.

Self-Test Exercise 2.2

1. Consider a patient whose health condition is deteriorated and require a quick recovery. a)If you were a physician treating the patient, would you recommend an IV (intravenous) fluids or a capsule? b) With the help of internet, discuss the differences among isotonic, hypotonic, and hypertonic solutions. Which solution would you administer to a patient who is suffering from dehydration due to excessive diarrhea?
2. Classify each of the following as homogeneous or heterogeneous mixtures and discuss their composition.
   a. Glass
   b. sand
   c. cement
   d. water gas
   e. producer gas
3. Explain the cleansing action of soap and detergents.
4. Discuss the difference between
   a. suspension and colloid
   b. solution and colloid
   c. pure substance and mixtures
   d. compound and homogeneous mixtures
5. When a corona virus infected patient coughs or sneezes, fine droplets of respiratory fluid containing the viral particles get dispersed in the atmosphere as a colloidal particle called______. (choose correct answer)
   a. emulsion
   b. aerosol
   c. fog
   d. smoke
6. Classify each of the following as suspension, solution, or colloid using the vein diagram shown below. (c), (f), and (k) are already classified as examples.
   a. Paint
   b. Yogurt
   c. Toothpaste
   d. Honey
   e. Soda water
   f. Sea water
   g. Blood
   h. Ruby
   i. Urine
   j. Soap suds
   Orange juice (+pulp)
   k. vinegar



7. Complete the missing information labeled as A, B, C, D, E, and F in the concept map shown below.



Section 2.3: The Solution Process

- In this section you will study about the useful rule-of-thumb "like dissolves like" and its application to determine whether a given solute dissolves in a given solvent. You will also study about the importance of inter-particle forces in the three types of interactions- solute-solute, solvent-solvent, and solute-solvent involved in the solution formation process.

Although solutions can have either solid, liquid, and gaseous states, we will consider only liquid solutions. More importantly, the solutions of solids in liquids and liquids in liquids will be discussed. Remember, that solids can be ionic compounds like NaCl or molecular compounds like sucrose (table sugar), $C_{12}H_{22}O_{11}$. This section will also help you understand the dissolution of these types of solids in water to form aqueous solutions and factors affecting the rate of dissolution. You will get introduced to terms like miscible and immiscible. Then, you will study about energy changes in solution process.

Learning Outcomes

At the end of this section you should be able to

- Explain how the "like dissolves like" rule depends on inter-particle forces of interactions;
- Predict relative solubility.
- Define rate of dissolution.

- When one substance (the solute) dissolves in another (the solvent), particles of the solute disperse throughout the solvent. The solute particles occupy positions that are normally taken by solvent molecules. The ease with which a solute particle replaces a solvent molecule depends on the relative strengths of three types of interactions (*Figure 2.6*):

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

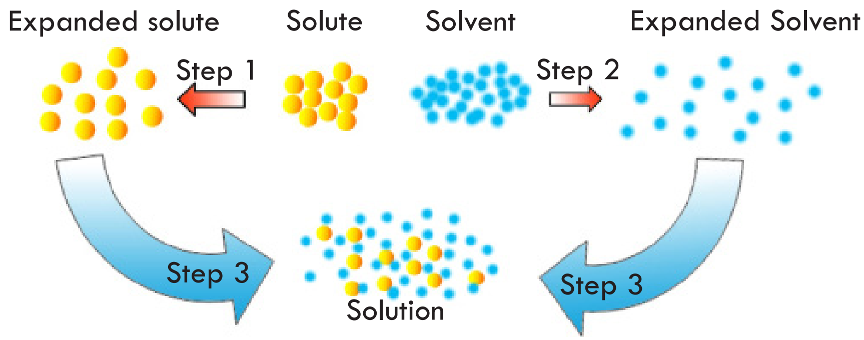


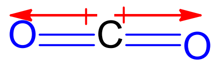
Figure 2.6 A molecular view of the solution process, portrayed as taking place in three steps.

☞ Solutes dissolve in a solvent when the inter-particle forces of interaction between solute molecules and solvent molecules are replaced by solute-solvent interactions. The useful rule-of-thumb "like dissolves like" says that substances with similar types of intermolecular forces dissolve in each other. That means, polar solutes dissolve in polar solvents and nonpolar solutes dissolve in nonpolar solvent.

As you read about covalent and ionic compounds in Grade 9, you learned that ionic compounds have the highest polarity forming full cations and anions within each formula unit as electrons are transferred from one atom to another. You also learned that a covalent bond is polar when the bonded atoms have different electronegativities and, thus, share the bonding electrons unequally.

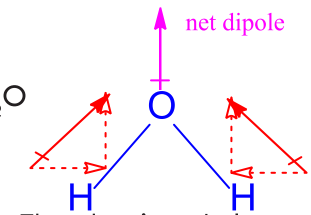
In diatomic molecules, such as HF, there is only one bond that is polar and the molecule is also polar. In larger molecules, molecular polarity is determined by the net dipole moment (imbalance of charge over the molecule); which in turn depends on both shape and bond polarity. Dipole moment (μ) is a measure of molecular polarity, given in the unit debye (D) derived from SI units of charge (coulomb, C) and length (m): 1 D = 3.34×10^{-30} C.m. Substances with zero or low electronegativity difference in between the bonding atoms and having zero net dipole moment such as H_2, O_2, N_2, CH_4, CCl_4 are nonpolar whereas H_2O, NH_3, CH_3OH, NO, CO, HCl, H_2S, PH_3, etc. having high electronegativity difference in between the bonding atoms and having net dipole moment are polar compounds. See illustrations below where the dipole moments cancel out in carbon dioxide (CO_2) but add-up to give a net dipole in water molecule (H_2O). Note that the presence of polar bonds does not always result in a polar molecule; we must also consider shape and the atoms surrounding the central atom.

(a) CO_2



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

(b) H_2O



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

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

Many ionic compounds are soluble (at least 1 g dissolves in 100 mL water) in water; however, not all ionic compounds are soluble. Chemists refer to substances as soluble, slightly soluble, or insoluble in a qualitative sense. For instance, KCl is readily soluble in water whereas AgCl is insoluble. $CaOH_2$ is slightly (sparingly) soluble. Ionic compounds that are soluble in water exist in their ionic state as hydrated cations and anions within the solution (See ***Figure 2.7***). When NaCl is added into water it splits into Na^+ and Cl^- ions. These charges get neutralized by interaction with water molecules. The cation (Na^+) gets neutralized by the partial negative charges on oxygen of water molecules and the anion (Cl^-) gets neutralized by the partial positive charges on hydrogens of water molecules as illustrated in ***Figure 2.7***.

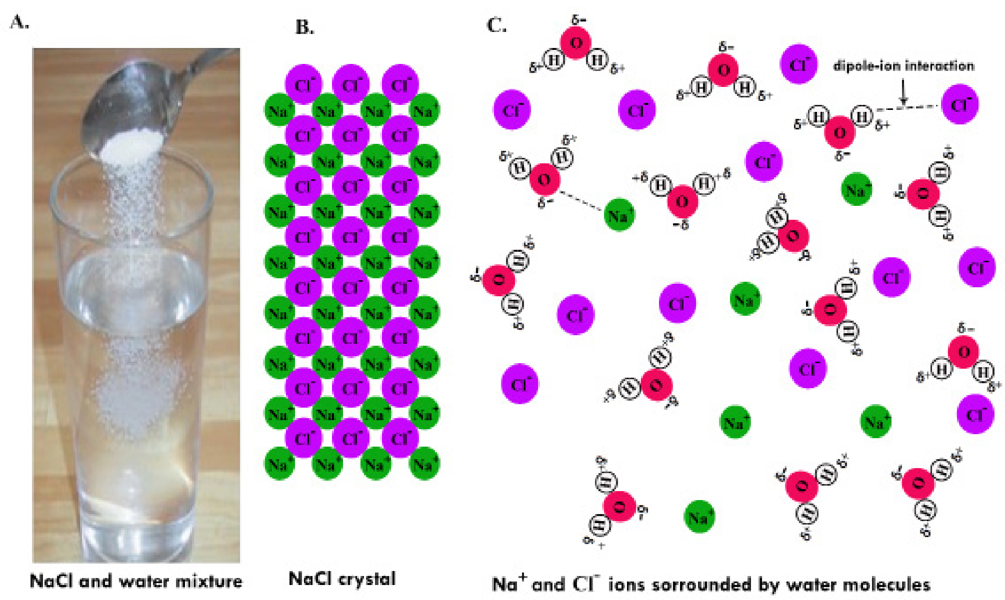


Figure 2.7 Hydration of Na^+ and Cl^- ions as NaCl dissolves in water.

2.3.1 The Solubility Rules for Ionic Solids in water

- Solubility is defined as the maximum amount of a substance that can be dissolved in a given volume (usually 100 mL) of solvent at a given temperature. Many ionic compounds are soluble (at least 1 g dissolves in 100 mL of water) in water; however, not all ionic compounds are soluble. Ionic compounds that are soluble in water exist in their ionic (as cations and anions) state within the solution. The solubility rule can be used to predict which ionic compounds will be soluble in water.

☞ The solubility rules for common ionic compounds are listed below.

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

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

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

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

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

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

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

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

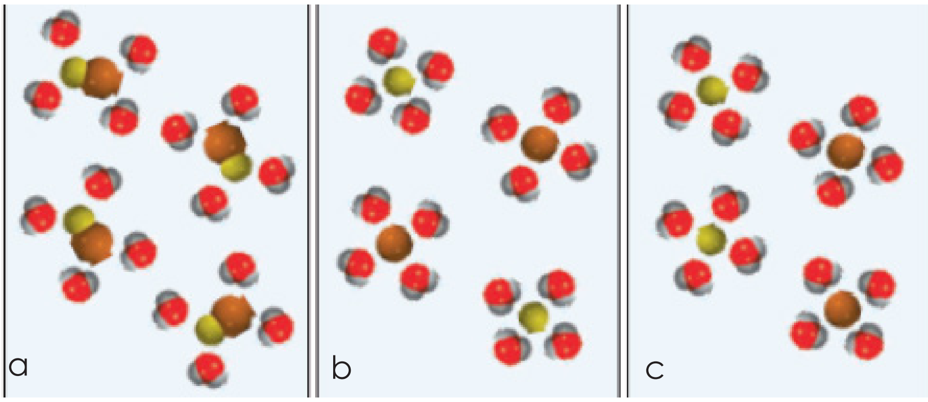
? How fast does a solute dissolve in a given solvent? What are the major factors that affect the rate of dissolution?

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

Dear student, at this stage of your course you should have understood what happens when you put NaCl (table salt) in a glass of water. You also learnt about solubility rules. To check your understanding about these topics, take 5 minutes to do the following activity. The first activity is related to the dissolution of ionic solid in water whereas the latter activity requires you to apply your understanding about solubility rules.

Activity 2.3

1. Which of the following diagrams best represents the hydration of NaCl when dissolved in water? The Cl^- ion is larger in size than the Na^+ ion.



2. Read the issue presented below and discuss how lime removes fluorides from water based on what you have learned in this section. Fluoride (F-) has a significant alleviating effect against dental caries if the concentration is approximately 1 mg/l. However, continuing consumption of higher concentrations (>1.5 mg/l) can cause dental fluorosis and in extreme cases even skeletal fluorosis. One of the methods used for the removal of excess fluoride from drinking water is precipitation by employing lime [quicklime (CaO) or slacked lime ($Ca(OH)_2$)] followed by sedimentation and/or filtration.

Dental fluorosis

Skeletal fluorosis

Self-Test Exercise 2.3

For question 1 to 3, choose the correct answer and then check your answer against the answer key provided at the end of this section.

1. Which of the following exists predominantly in the water (H_2O) molecule?
   a. van der Waal's force
   b. ion-dipole force
   c. hydrogen bond
   d. none
2. Which of the following has the strongest intermolecular dipole interactions?
   a. He
   b. Ar
   c. Ne
   d. Kr
3. The process of dissolution is always accompanied by energy change.
   a. True
   b. False

Answer questions 4 to 7 accordingly and your answer against the answer key provided at the end of this section.

4. Oxygen (O) and sulphur (S) are in the same group in the periodic table. They form compounds with hydrogen, H_2O and H_2S. However, H_2O is a liquid whereas H_2S is a gas at a room temperature. Explain.
5. Classify each of the following as soluble or insoluble in water:
   a. $BaCO_3$
   b. $AlCl_3$
   c. MgO
   d. $Al(OH)_3$
   e. KI
   f. $Al_2(SO_4)_3$
   g. $Ca_3(PO_4)_2$
   h. $PbCl_2$
   i. $AgNO_3$
   j. $(NH_4)NO_3$
   k. NH_4Cl
6. Classify each of the following pairs as miscible or immiscible:
   a. Water and grain alcohol
   b. "Tej" and water
   c. Water and oil
   d. Water and honey
   e. Water and petrol
   f. Gasoline and ''naphtha''
7. What type (s) of intermolecular (interparticle) forces exist when the following pairs interact:
   a. HBr and H_2S
   b. Cl_2 and CBr_4
   c. I_2 and NO_3^-,
   d. NH_3 and C_6H_6(Benzene)

2.3.2 Energy Changes in Solution Process

- Now that you are familiar with the solution formation as a process involving three types of interactions - solute-solute, solvent-solvent, and solute-solvent forces of interaction. In this subsection, you will study about energy changes involved in the dissolution process and the definitions of some important terms.

Learning Outcomes

At the end of this subsection you should be able to

- Define heat of solution, solvation energy and hydration energy.
- Apply the concept of heat of solution to the solution of ammonium nitrate crystal.
- Apply the concept of heat of solution to the solution of sodium hydroxide crystal.
- Explain how heat of solution is influenced by the inter particle interaction forces.

Study Notes and Important Points

- A solution is formed when the solute particles are completely dissolved in the solvent. We have noted previously that gases mix freely. But the formation of liquid and solid solutions requires overcoming the solute-solute and solvent-solvent inter-particle forces of attraction before the mixing step. To understand the enthalpy changes involved during the dissolution, therefore, we think of a hypothetical three-step process each accompanied by an enthalpy change

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

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

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

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

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

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

We combine the three individual enthalpy changes to find the enthalpy of solution (ΔH_{soln}), the total enthalpy change that occurs when solute and solvent form a solution:

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

Endothermic and Exothermic Dissolution Processes

Depending on the relative magnitude of ΔH_{solute} + $\Delta H_{solvent}$ & ΔH_{mix}, the ΔHsoln can be zero (ideal solution), negative (exothermic) or positive (endothermic). For instance, if you mix ammonium chloride in water in a flask, the flask gets cold. But when sodium hydroxide (NaOH) is mixed in water, the flask gets hot. Why? The reason is attributed to the energy changes that accompany the dissolution process. As discussed earlier, the solution process involves three steps: breaking solute-solute interaction, breaking solvent-solvent interaction, and forming solute-solvent interactions or the mixing step. When the energy released during the third step in solution process, i.e. the mixing step is larger than the energy required to break the solute-solute and solvent-solvent forces, energy is released-the process is exothermic. This energy heats up the surrounding and that is why the flask gets hot. On the contrary, if the energy released in the mixing step is less than the energy required to break the solute-solute and solvent-solvent forces, the dissolution process absorbs energy from the surrounding and the surrounding (the flask) gets cold- the dissolution process is endothermic. The energy involved in the mixing step is called **solvation energy**.

Solvation refers to the process of attraction and association of a solute molecule with solvent molecules.

☞ Solvation energy is the energy released when solute molecules are solvated by solvent molecules. Solvation in water is called hydration. Thus, hydration energy is the energy released when solute molecules or ions are solvated by water molecules.

☞ **Ideal Solutions (ΔHsoln = ΔHsolute + ΔHsolvent + ΔHmix = 0)**

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

☞ **Examples of Endothermic and Exothermic Processes**

When sodium hydroxide is dissolved in water, the solution becomes hot (the solution process is exothermic). On the other hand, when ammonium nitrate is dissolved in water, the solution becomes very cold (the solution process is endothermic). This cooling effect from the dissolving of ammonium nitrate in water is exploited in instant cold packs used in hospitals. An instant cold pack consists of a bag of NH_4NO_3 crystals inside a bag of water (*Figure 2.8*). When the inner bag is broken, NH_4NO_3 dissolves in the water. Heat is absorbed, so the bag feels cold. Hot packs, by contrast, containing either $CaCl_2$ or $MgSO_4$, produce heat when the salts dissolve in water (*Figure 2.8*).

Figure 2.8 The instant cold compress and hot compresses

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

$q = m \times Cg \times \Delta T$ Eq 2.1

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

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

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

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

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

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

☞ **Enthalpy of Hydration (Ionic Solids in Water)**

We have already discussed that Ionic solids that are soluble in water exist in their ionic state as hydrated cations and anions within the solution. For ionic compounds that are not soluble in water, the ions are so strongly attracted to one another in the crystal lattice that they cannot be broken apart by the partial charges of the water molecules. Thus, for solutions of ionic solids in liquid the lattice energy and hydration energy must be considered.

Therefore, $\Delta H_{soln} = \Delta H_{solute} + \Delta H_{solvent} + \Delta H_{mix} = \Delta H_{lat} + \Delta H_{hydr}$

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

? How do lattice energy and hydration energy affect the Solubility of Ionic Solids?

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

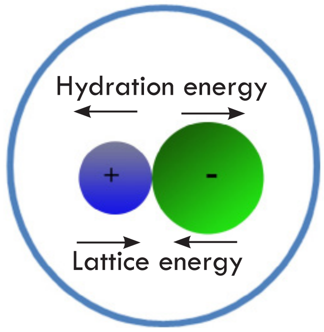


Figure 2.9 The effects of lattice energy and hydration energy on the solution process of ionic solids.

- Lattice energies depend on the charge on the ions and also the distance between the centers of the neighboring positive and negative ions. As the magnitude of the charge on the ions increases the lattice energy also increases. For this reason, you can expect substances with single charged ions to be more comparatively soluble, and those with multiple charged ions to be less soluble in water. Lattice energy is inversely proportional to the inter-particle distance between the centers of the two ions. Thus, CsCl has less lattice energy than NaCl.
- Enthalpy (heat) of solution can be determined in the laboratory **by measuring the temperature change** of the solvent when solute is added. **A calorimeter** is a device used to measure the amount of heat involved in a chemical or physical process.

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

Dear student, most of the concepts you have studied in this section are related to your experiences. To substantiate your understanding of the learned material, let's contextualize it to your experiences. Take the next time to perform the following project work in your convenience time.

Activity 2.4

1. Make a self-managed plan and arrange a visit to nearby hospital or other health center. Your role is to contact the professionals on the usage of instant cold packs and instant hot packs. Ask them to get one each, discuss the usage, working principle, and the composition. Write the dissolution equation in each case.
2. Discuss how people traditionally alleviate inflammation (tenderness), and pains caused by muscle or joint damage in relation to the medical uses of cold packs and hot packs.

Self-Test Exercise 2.4

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

Section 2.4: Solubility as an Equilibrium Process

❁ So far we have discussed that whether a solute dissolves in a given solvent is determined by inter-particle or intermolecular forces and we predict whether a given solute dissolves in a given solvent by the "Like Dissolves Like" rule. In this section you will study about the dissolution process as an equilibrium process and preparations of saturated, unsaturated, and supersaturated solutions. We will consider how some factors (variables) affect the solubility of a solid or liquid solute in a liquid or a gaseous solute in a liquid. The application of Henry's law for quantitative determination of solubility of gases at a given pressure is presented with adequate examples. For simplicity, we will consider aqueous solutions only.

Learning Outcomes

At the end of this section, you should be able to
- Define solubility.
- Describe the distinctions among saturated, unsaturated, and supersaturated solutions.
- Prepare unsaturated and saturated solutions of sodium sulphate.
- Prepare supersaturated solution of sodium thiosulphate.
- Explain the equilibrium nature of saturated solution.

Study Notes and Important Points

- Solution equilibrium is the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate.

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

To understand this, suppose that you have 1 litre of water in a beaker or any similar container to which you add a spoonful of table salt (NaCl) and stir the mixture until the NaCl crystals dissolve as illustrated in ***Figure 2.10***. You keep adding more and more salt and continue stirring. Eventually, you reach a point where no more of the salt will dissolve in the presence of undissolved solid no matter how long or how vigorously you stir it. When that maximum solubility point is reached, the rate at which NaCl crystals are dissolving to supply Na^+ and Cl^- ions into the solution and the rate at which the ions are crystalizing back and precipice as NaCl crystals becomes equal. At this point, the total amount of dissolved salt remains unchanged-equilibrium is established. The liquid phase above the precipitate is a **saturated solution**. That means, a saturated solution is at equilibrium and contains the maximum amount of a solute in a given solvent, at a specific temperature in the presence of undissolved solute. Therefore, the formation of a saturated solution involves an equilibrium process.

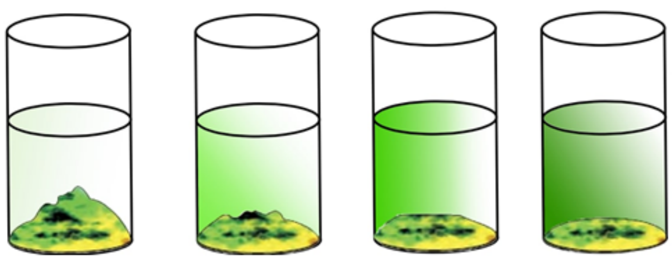
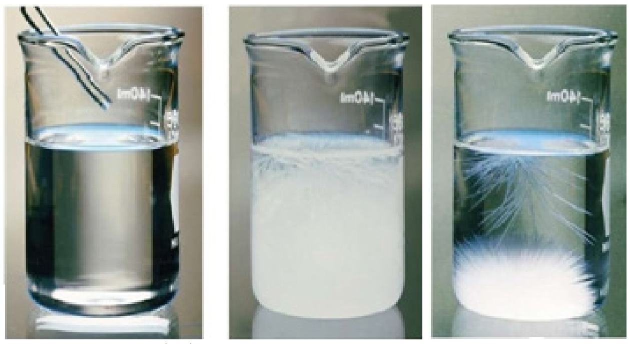


Figure 2.10 Illustrates the formation of saturated solution. (a) excess solid added to a solvent (b) The solute starts to dissolve (c) The solute continues to dissolve (d) No more solute is dissolving indicating that the process has reached a dynamic equilibrium where the rate of dissolution is equal to the rate of crystallization.

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

A solution that has not reached its maximum solubility is called **an unsaturated solution**. This means that more solute could still be added to the solvent and dissolving would still occur.

In special circumstances, a solution may be **supersaturated**. Supersaturated solutions are solutions that have dissolved solutes beyond the normal saturation point. Usually a condition such as increased temperature or pressure is required to create a supersaturated solution. For example, sodium acetate has a very high solubility at 270 K. When cooled, such a solution stays dissolved in what is called **a meta-stable state**. However, when a seeding crystal is added to the solution, the extra solute will rapidly solidify, leaving a saturated solution (***Figure 2.11***). During the crystallization process, heat is evolved, and the solution becomes warm. Common hand warmers use this chemical process to generate heat.



(a) (b) (c)

Figure 2.11 (a) a hot supersaturated solution (formed at elevated temperature) (b) the scenario when the supersaturated solution is cooled (c) cold supersaturated solution after seeding.

Experiment 2.2

Preparation of Unsaturated, Saturated and Supersaturate Solutions

Objectives: To prepare unsaturated, saturated and supersaturate solutions

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

Procedure:

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

Activity 2.5

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

Self-Test Exercise 2.5

1. Discuss how you would prepare saturated, unsaturated, and supersaturated solutions.

Section 2.5: Factors Affecting Solubility of Substances

- In this section you will study about the effect of temperature on solubility equilibrium. More importantly the effect of temperature and pressure will be discussed along with practical examples. You will also learn about Henry's law-its statement and applications to calculations involving effect of pressure on solubility of gases.

Learning Outcomes

At the end of this subsection you should be able to

- Describe the factors that affect solubility of substances.
- Investigate the effect of temperature on solubility of sodium sulphate.
- Conduct an experiment to determine solubility of table salt and sugar.

Study Notes and Important Points

Solubility of a substance in a given solvent depends on

inter-particle forces,

temperature, and

pressure (for gases only)

2.5.1 Effect of Temperature on Solubility of Solids

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

Experiment 2.1

Determination of the Solubility of NaCl

Objective: To determine the solubility of NaCl

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

Chemicals: Sodium chloride and water.

Procedure:

1. Take an evaporating dish and weigh it.
2. Take 100 mL of water in a beaker and add about 20 g of sodium chloride to it. Stir the solution vigorously with a glass rod, add more sodium chloride while stirring and continue this process until undissolved sodium chloride is left in the beaker.
3. Take 50 mL of the supernatant saturated solution (assume that 50 mL of solution is equal to 50 mL of the solvent) and transfer it to an evaporating dish.

4. Heat the solution in the evaporating dish as shown in the Figure, till all the water has evaporated and dry sodium chloride is left in the evaporating dish.
5. Cool the evaporating dish containing dry sodium chloride to room temperature and weigh it again.

Observations and analysis:

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

2.5.2. Effect of Temperature on Solubility of Gases

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

Note: Pressure has virtually no effect on the solubility of solids or liquids in liquid solvents

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

At the end of this subsection you should be able to

- State Henry's law
- Use Henry's law to calculate concentration of gaseous solute in a solution

Study Notes and Important Points

- External pressure has very little effect on the solubility of liquids and solids because they are incompressible. In contrast, the solubility of gases increases as the partial pressure of the gas above a solution increases (Henry's Law). The qualitative effect of a change in pressure on the solubility of a gas can be predicted from Le Châtelier's principle which states that when a system in equilibrium is disturbed by a change of variables such as temperature, pressure, or concentration, the system shifts in equilibrium composition in a way that tends to counteract this change of variable. Let us see how Le Châtelier's principle can predict the effect of a change in pressure on gas solubility.

Imagine a cylindrical vessel that is fitted with a movable piston and contains carbon dioxide gas over its saturated water solution (***Figure 2.12***). The equilibrium is

$$CO_2\,(g) \rightleftharpoons CO_2\,(aq)$$

This system has initial equilibrium pressure P_1. Suppose you increase the partial pressure of CO_2 gas from P1 to P_2 by pushing the piston down. How would the system counteract the extra pressure applied? This change of partial pressure, according to Le Châtelier's principle, shifts the equilibrium composition in a way that tends to counteract the pressure increase. This is possible if some gaseous CO_2 enter the solution so that the partial pressure of CO_2 gas above the liquid decreases as more CO_2 dissolves. The system comes to a new equilibrium, in which more CO_2 has dissolved.

$$CO_2\,(g) \rightleftharpoons CO_2\,(aq)$$

From the preceding equation, you see that the partial pressure of CO_2 gas above the liquid decreases if more CO_2 dissolves.

☞ **Note:** Pressure above the liquid is directly proportional to number of gaseous molecules above the liquid.

Similarly, a further increase in pressure from P_2 to P_3 will have the same effect. So you can predict that carbon dioxide is more soluble at higher pressures. Conversely, when the partial pressure of carbon dioxide gas is reduced, its solubility is decreased. A bottle of carbonated beverage fizzes when the cap is removed: as the partial pressure of carbon dioxide is reduced, gas comes out of solution.

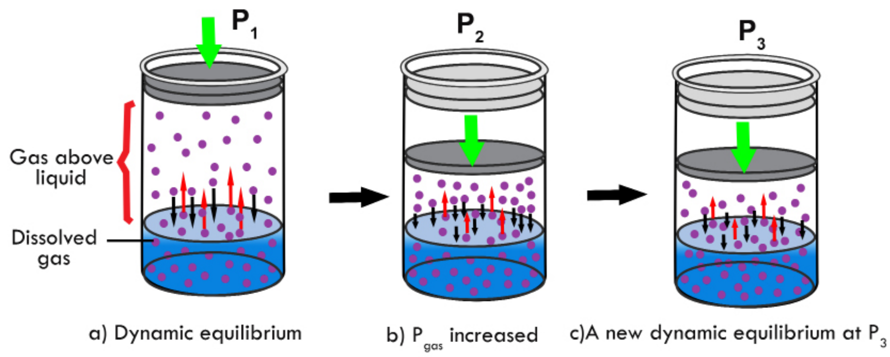


Figure 2.12 A model depicting why the solubility of a gas increases as the partial pressure increases at constant temperature.

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

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

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

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

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

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

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

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

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

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

Example 2.2

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

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

Example 2.3

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

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

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

Now, it is your turn to do some calculations and exercises.

Activity 2.6

1. The concentration of CO_2 in a solution is 0.032 M at 3.0 atm. What is the concentration of CO_2 at 5.0 atm of pressure?
2. Discuss the effect of climate change (rise in temperature of the environment) on aquatic life.
3. As nations shift from agriculture-led to industry- led economy, environmental pollution becomes a serious concern. Water bodies, for instance, can be polluted by chemicals released to the environment by industries along with wastes. Discuss how pollution of water bodies by chemicals can affect aquatic life such as fish and its productivity. From this lesson, what precautions would you make when disposing wastes of personal hygiene materials such as sanitary pads (Modess) or hair gel containers and other plastic materials?
4. State Henry's law and explain why (a) Carbonated beverages such as coke, Pepsi, fizzes (sparkles, Figure to the right) when the cap is removed (b) Multicellular organisms need hemoglobin in red blood cells to bind and carry O_2 to support the energy needs. Why this mechanism of replenishing O_2 in biological fluid is necessary? What would happen if hemoglobin is absent?

By now, you should be able to state Henry's law and do some calculations based on it. Let me give you a chance to check this. Take five minutes to perform the following self-test exercise.

Self-Test Exercise 2.6

1. Which of the following gases has the greatest Henry's law constant in water at 25 °C?
   a. CH_4
   b. Ne
   c. NH_3
   d. H_2

Section 2.6: Ways of Expressing Concentrations of Solutions

In this section you will study about the definition of concentration, and the different ways of expressing concentrations of solutions. You will also study about the conversion of concentration units.

Learning Competencies

At the end of this subsection you should be able to

- Define concentration of a solution.
- Define mass percentage, ppm and ppb of a solute in a solution.
- Calculate the mass percentage, ppm and ppb of a solute in a solution from a given information.

Study Notes and Important Points

In chemistry, concentration is defined as the relative quantity of a solute against a total quantity of solution or solvent. There are various ways of expressing concentrations of solutions. These include percent by mass or volume, mole fraction, molarity, molality, and normality. We will study about these one by one. Let's start with percent by mass or volume.

2.6.1 Percent by Mass/Volume

The percent by mass (also called the percent by weight or the weight percent) designated as %(w/w) is defined as

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

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

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

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

Given	Required	Solution
Mass of solute (glucose) = 13.5 g Mass of solvent = 0.100 kg =100 g	%(w/w)	$\%(w/w)_{solute}$ $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100\%$ But $Mass_{solution} = mass_{solute} + mass_{solvent}$ Substituting the values of masses of glucose (solute) and solution in the above equation gives: $\%(w/w)_{solute} = (13.5\ g/113.5\ g)\ x100\%$ $= \mathbf{11.9\%}$

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

$$\text{Volume percent} = \frac{\text{Volume of solute}}{\text{Volume of solution}} x100\%$$

Example 2.5

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

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

Example 2.6

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

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

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

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

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

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

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

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

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

Example 2.7

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

Given	Required	Solution
Mass of solute (Pb^{2+}) = 5.4 µg = 5.4×10^{-6} g Mass solution = 2.5 g	Concentration in ppm	$\text{ppm} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$ = $[(5.4 \times 10^{-6}\ g)/(2.5\ g)] \times10^6$ = **2.16 ppm**

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

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

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

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

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

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

Substituting in the above equation yields,

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

Therefore, for dilute solutions

1 ppm = 1 mg/L

Dear student, I hope you have understood the percent by mass and percent by volume, ppm, and ppb based calculations. Now it is your turn to do some calculations. Take 10 minutes to do the following calculations.

Self-Test Exercise 2.7

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

2.6.2 Mole Fraction

At the end of this subsection you should be able to

- Define mole fraction.
- Calculate mole fraction of a solute and a solvent in a solution.

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

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

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

Note: Usually a solution composed of solute and solvent. So $X_{solute} + X_{solvent} = 1$, Therefore, if mole fraction of the solute is 0.25, then mole fraction of solvent = 1- 0.25 = 0.75. In other words, if solute is 25% by mole, then solvent is 75%.

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

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

Example 2.8

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

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

Now that you understand how to calculate mole fractions, please spend the next minutes to do the following calculations.

Self-Test Exercise 2.8

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

2.6.3 Molarity

At the end of this subsection you should be able to

- Define molarity.
- Prepare molar solutions of different substances.
- Calculate molarity of a solution from a given information.

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

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

1 M = (1 mole of solute)/(1 L of solution)

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

Eq 2.11

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

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

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

Effect of temperature: A liquid expands when heated, so a unit volume of hot solution contains less solute than that of cold solution; thus, the molarity is different.

Effect of mixing: Because of solute-solvent interactions that are difficult to predict, volumes may not be additive: adding 500 mL of one solution to 500 mL of another may not give 1000 mL of final solution.

Example 2.9

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

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

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

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

Let me give you an opportunity to try some calculations. Take ten minutes to do the exercise below.

Self-Test Exercise 2.9

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

2.6.4 Molality

At the end of this subsection you should be able to

- Define molality.
- Prepare molal solutions of different substances.
- Calculate molality of a solution from a given information.

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

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

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

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

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

Example 2.10
Calculate the molality of a solution prepared by dissolving 32.0 g of $CaCl_2$ in 271g of water.

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

☞ **Now it is your turn to do solution preparation calculations.** Take five minutes to do the self-test exercise 2.10.

Self-Test Exercise 2.10

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

2.6.5 Normality

At the end of this subsection you should be able to
- Define the terms equivalent weight, number of equivalents and normality.
- Calculate normality of a solution from a given information.

☞ The normality of a solution is the gram equivalent weight of a solute per liter of solution.

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

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

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

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

Substituting eq 2.16 in eq 2.15 gives:

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

Substituting eq 2.17 in eq 2.14 gives:

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

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

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

Example 2.11

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

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

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

Example 2.12

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

a. Completely neutralized to give PO_4^{3-}
b. Converted to $H_2PO_4^-$

c. Converted to HPO_4^{2-}

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

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

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

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

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

Given	**Required**	**Solution**
(a) No. of equivalets = 1.5 eq V = 100 mL = 0.1 L	Normality	a) N=Eq/L= (1.5 Eq)/(0.1 L)=15 Eq/L=15 N b) N=Eq/L= (0.5 Eq)/(0.1 L)=5 Eq/L=5 N c) N=Eq/L= (1 Eq)/(0.1 L)=15 Eq/L=**10 N**

(b) No. of equivalents = 0.5eq

(c) No. of equivalents = 1eq

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

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

The easiest way to find normality is from molarity. All you need to know is the value of 'z'. N = M.z; where N is normal

Example 2.13

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

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

Now it is your turn to do some calculations involving normality. Take 15 minutes to do the self-test exercise 2.11.

Self-Test Exercise 2.11

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

2.6.6 Conversion of Concentration Units

At the end of this section you should be able to interconvert various concentration expressions.

- The interconversion of concentration units is very simple. Because they are intrinsic units. That means, any amount of liquid taken from a 5% (v/v) ethanol solution, for instance, always contains ethanol to solution in ratio of 5 to100 regardless of the volume taken.
- Therefore, 5%(v/v) read as 5 mL of ethanol per 100 mL of solution. Thus, if you want to calculate the molarity of 5%(v/v) ethanol solution, you 5 mL ethanol to number of moles of ethanol and 100 mL solution to litres of solution.
- Molarity of 5%(v/v) _etanol=(number of moles of 5 mL ethanol)/(volume of solution (100 mL=0.1 L))

$$\text{Molarity of } 5\% \left(\frac{v}{v}\right)_{etanol} = \frac{number\ of\ moles\ of\ 5\ mL\ ethanol}{volume\ of\ solution\ (100\ mL = 0.1\ L)}$$

$$= \frac{mass/_{molar\ mass}}{0.1\ L} = \frac{vd/_{molar\ mass}}{0.1\ L} = \frac{0.78\frac{g}{mL}x5mL\Big/_{46\ g/mole}}{0.1\ L}$$

where 46 g/mole is molar mass of ethanol, v = volume, d = 0.78 g/mL is density of ethanol

☞ Similarly, 5 ppm glucose means 5 g of glucose per 10^6 g of solution. And 5 ppb of glucose means 5 g of glucose per 10^9 g of solution.

Example 2.14: Converting (w/w) percent to Molarity

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

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

Example 2.15

Converting Mole Fractions to Molality

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

Given	Required	Solution
$X_{solute(glucose)}$ = 0.15 $X_{solvent(water)}$ = 0.85	Molality = m	Molality = m = $\frac{\text{number of moles of solute}}{\text{mass of solvent in kg}}$ Let's consider 0.15 moles of glucose and 0.85 moles of water which is consistent with the given mole fractions. Thus, to calculate the mass of solvent in kg we convert the mole solvent into mass solvent as follows: $Mass_{H_2O}$ = mole x molar mass = 0.85 mole x18 g/mole =15.3 g = **0.0153 kg** Substituting the values in the above equation gives: Molality = m = $\frac{0.15 \text{ mole}}{0.0153 \text{ kg}}$ =**9.8 m**

Example 2.16

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

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

Now it is your turn to do some calculations involving interconversion of concentration units. Take 15 minutes to do the self-test exercise 2.12.

Self-Test Exercise 2.12

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

Section 2.7: Preparation of Solutions

- So far you have learnt various ways of expressing concentrations of solutions. You have also studied how to convert from one concentration unit to the other. Note that in most of the examples discussed above, we considered moles of solutes. But there is no piece of equipment that can measure out the moles of a substance. For this, we need to convert the number of moles of the sample into the number of grams represented by that number. We can then easily use a balance to weigh the amount of substance needed for the solution. In this section, we will consider preparation of stock solutions using a standard solid compounds and preparation of dilute solutions using stock solutions.

Learning Outcomes

- At the end of this subsection you should be able to prepare molar, molal, or normal solutions of different substances.

Study Notes and Important Points

2.7.1 Preparation of Stock Solutions from Pure Compound

A) preparation of liquid solution using a solid solute

- The solution preparation generally involves 6 steps; part of which are shown in (***Figure 2.13***):

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

Mass required = n x molar mass

Step 2: Weigh the required mass using balance.
Step 3: Transfer the measured amount into a beaker of enough capacity. Add some solvent and mix (shake) to dissolve. The amount of solvent added at this stage should not exceed half the final volume of the solution to be prepared. Heating over a hot plate while stirring may be required if the solid is not readily soluble.
Step 4: Transfer the dissolved mixture into volumetric flask of desired capacity and swirl to mix.
Step 5: Dilute to the mark by filling the remaining portion of the flask with the solvent. The solvent is usually water as we often prepare aqueous solutions.
Step 6: Cap it and label your flask containing the prepared solution with the key information (concentration, name of solute, date of preparation, and owner (your name). This step is very important even if you need the solution immediately

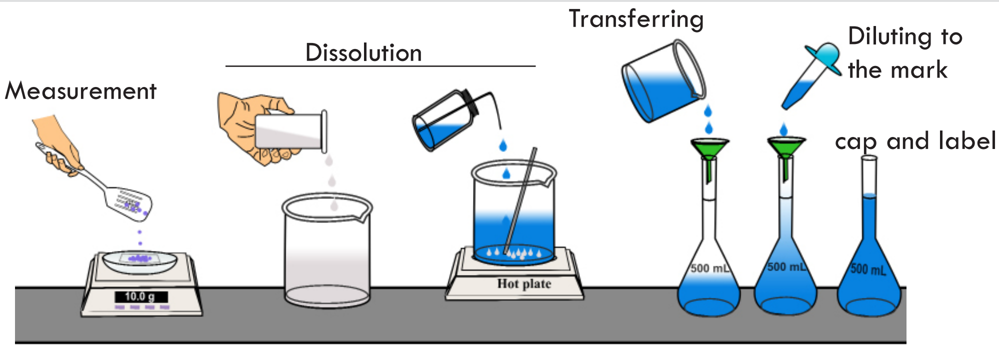


Figure 2.13 Illustration of steps involved in making a 500 mL liquid solution of given concentration by dissolving a solid solute in a liquid solvent.

Example 2.17

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

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

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

B) Preparation of a liquid solution using a liquid solute

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

V= m/d = (n x molar mass)/d; where d density of the solute.

Then you use measuring cylinder to measure the volume of liquid and the measured volume of liquid is directly transferred to a volumetric flask of desired capacity and diluted to the mark.

2.7.2 Diluting a Solution

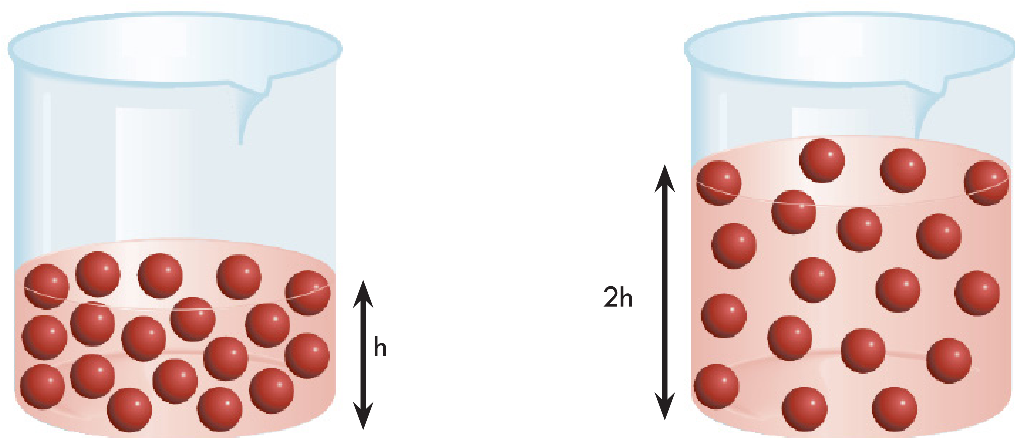
At the end of this subsection you should be able to

- Explain dilution process.
- Calculate the volume or concentration changes during dilution of solution.
- Prepare a dilute solution from concentrated solution.

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

Dilution is the procedure for preparing a less concentrated solution from a more concentrated one by simply adding more solvent. Often, a small portion (V_i) of stock solution is withdrawn from the more concentrated solution, transferred to volumetric flask having a capacity of V_f, and diluted to final volume V_f by simply adding more solvent.

To understand what changes and what do not during dilution, imagine we have an aqueous salt solution containing 18 particles in a certain volume V_1 of solution as shown in ***Figure 2.14***. Let's assume that this solution has a certain concentration M_1. That means we have 18 particles (18/(6.02x10^{23} moles)) dissolved in the volume V_1 of the solution. Let's dilute this solution by adding more water to double the volume of the solution to $V_2 = 2V_1$ as illustrated in ***Figure 2.14***. What stays the same and what is changed when more water is added? Did the number of particles (number of moles) change? What about the concentration?



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

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

Figure 2.14. The dilution of a more concentrated solution (solution I) to a less concentrated one (solution II) by doubling the volume of solution by adding more solvent does not change the total number of solute particles.

☞ The molarity of solution I is

$$M_1 = Moles_1/V_1$$

Rearranging to find moles we obtain:

$$Moles_1 = M_1V_1$$

The molarity of solution II is:

$$M_2 = Moles_2/V_2$$

Rearranging to find moles we obtain:

$$Moles_2 = M_2V_2$$

However, the number of particles (number of moles) of solute is the same in both solutions.

$$Moles_1 = Moles_2$$

☞ $\mathbf{M_1V_1 = M_2V_2}$

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

Note that as you can see in ***Figure 2.14***, adding more solvent changed the volume of the solution; doing so we also changed its concentration because the number of particles (number of moles) per unit volume (per unit litre or per unit mL) is changed. However, the total number of particles (number of moles) of solute is the same in both solutions.

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

Example 2.18

Describe how you would prepare 2.50×10^2 mL of a 2.25 M H_2SO_4 solution, starting with a 7.41 M stock solution of H_2SO_4.

Given	Required	Solution
$M_1 = 7.41$ M $V_1 = ?$ $M_2 = 2.25$ M $V_2 = 250$ mL	V_1	$M_1 V_1 = M_2 V_2$ $V_1 = \frac{M_2 V_2}{M_1}$ $V_1 = (2.25 \text{ M} \times 250 \text{ mL})/(7.41 \text{ M}) =$ **75.9 mL**

Thus, we must withdraw 75.9 mL from the 7.41 M H_2SO_4 solution, transfer it to a 250 mL volumetric flask half-filled with distilled water, and dilute it with sufficient amount of water to the mark.

Now it is your turn to do some exercises and calculations. Take 30 minutes to do the self-test exercise 2.13.

Self-Test Exercise 2.13

1. Outline the preparation of 2.00×10^2 mL of a 0.866 M KOH solution, starting with a 5.07 M stock solution.
2. What is the difference between dilute and unsaturated? Or concentrated and saturated solutions?
3. Outline the preparation of 100 mL of (a) hypotonic (0.1 M) and (b) hypertonic (0.25 M) aqueous NaCl solution and (c) 0.1 M H_2SO_4 using a stock solution of 98% (w/w) H_2SO_4. Density = 1.84 g/ml. (**Hint:** First calculate the molarity of the 98% (w/w) H_2SO_4 solution. Then use the dilution formula).

Section 2.8: Solution Stoichiometry

☸ In Unit 1 you studied about stoichiometric calculations in terms of the mole method, which treats the coefficients in a balanced equation as the number of moles of reactants and products. In this section you will learn about solving stoichiometry problems for reactions in solution. This requires the additional step of converting the volume of reactant or product in solution to amount (mole).

Learning Outcomes

- At the end of this subsection you should be able to use stoichiometrically equivalent molar ratios to calculate amounts of reactants and products in a reaction of pure and dissolved substance.

Study Notes and Important Points

- Aqueous solution chemistry is central to laboratory chemistry. Many environmental reactions and almost all biochemical reactions occur in solution. Solution stoichiometry refers to the study of the quantitative relationships in substances and their reactions in solution. In other words, it deals with the numerical relationships between two reactants or between two products or between a reactant and a product. Solving stoichiometry problems for reactions in solution requires the additional step of converting the volume of reactant or product in solution to amount (mole) using the formula:

Number of moles of solute = Molarity x Volume of solution (litre)

Generally, you can follow these 4 steps while dealing with solution stoichiometry:

1. Balance the equation. Note that the coefficients in chemical equation indicate number of moles.
2. Find the amount (mol) of one substance from the volume and molarity.
3. Relate it to the stoichiometrically equivalent amount of another substance.
4. Convert to the desired units.

Mole-Mole Calculation

Example 2.19

What is the concentration of sodium hydroxide required to react completely with equal volume of 0.104 M sulfuric acid? As you can see in the question, we are given molar concentration of solution. So, we can set n = MV.

Given

$V_{NaOH} = V_{H_2SO_4}$

$M_{H_2SO_4} = 0.104$ M

Required

$M_{NaOH} = ?$

Solution

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

Stoichiometry	1 mole	2 mole			
Equation	H_2SO_4	+ 2NaOH	$\rightarrow$	Na_2SO_4	+ $2H_2O$
Given	nH_2SO_4=0.104 mol/L x VH_2SO_4	nNaOH=MNaOH x VNaOH			

Stoichiometry:= $n_{NaOH}/n_{H_2SO_4} = 2/1$

$= M_{NaOH}V_{NaOH}/M_{H_2SO_4}V_{H_2SO_4}$

$= M_{NaOH}V_{NaOH}/0.104MV_{H_2SO_4}$

Since volume is the same, it cancels out and the equation becomes: $2/1 = M_{NaOH}/0.104$ M

$M_{NaOH} = 0.104$ M x 2

$=$ **0.208 M**

As you can see in the question, we are given molar concentration of solution. So, we can set n = MV.

Now it is your turn to do a mole-mole calculation. Take five minutes to do exercise 2.14.

Self-Test Exercise 2.14

What is the concentration of sodium hydroxide that is required to react completely with equal volume of 0.104 M hydrochloric acid?

Mole – Mass

Example 2.20

What mass of solid magnesium hydroxide can be produced if 45 mL of a 0.63 M $Mg(NO_3)_2$ solution reacts completely with excess NaOH?

Given	Required	Solution
$V_{Mg(NO_3)_2}$ = 45 mL $M_{Mg(NO_3)_2}$ = 0.63 M Note that M and V are given means number of moles (n) is known: n=MV	V_{AgNO_3} = ?	see below

Stoichiometry		1 mole		58.321 g	
Equation	2NaOH(aq) +1	$Mg(NO_3)_2$(aq)		$Mg(OH)_2$(s) +2$NaNO_3$(aq)	
Given		0.63 M, 45 mL		Mass =?	

Stoichiometry = $n_{Mg(NO_3)_2}/mass_{Mg(OH)_2}$=1 mole/58.321 g

= 0.63 mol/L x 0.045 L/$mass_{Mg(OH)_2}$

$mass_{Mg(OH)_2}$ = 0.63 mol/L x 0.045L x 58.321 g/1 mole

=1.65g

Mass – Volume

Example 2.21

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$Al(NO_3)_3\ (aq) + 3Ag(s) \rightarrow 3AgNO_3(aq) + Al(s)$$

Given	Required	Solution
w_{Ag}=3.66 g $M_{Al(NO3)3}$ = 0.0995 M	$V_{Al(NO3)3}$= ?	see below

Stoichiometry	3x107.97 g	1 mole		
Equation	3Ag(s)	+ $Al(NO_3)_3$(aq) →	3 $AgNO_3$(aq)	+ Al(s)
Experimental	3.66 g	n $_{Al(NO_3)_3}$ = $M_{Al(NO_3)_3}$ x $V_{Al(NO_3)_3}$		

Stoichiometry: =$w_{Ag}/n_{Al(NO3)3}$=3x107.97 gAg/1$mole_{Al(NO3)3}$

= 3.66g/$M_{Al(NO3)3}V_{Al(NO3)3}$=3.66g/0.0995MxV$_{Al(NO3)3}$

3x107.97g/1mole=3.66g/(0.0995MxV$_{Al(NO3)3}$); where MV = mole

$V_{Al(NO3)3}$= 3.66 g x1 mole/3x107.97 g x0.0995 mol/L=**0.114 L**

Now it is your turn to do a mass-volume calculation. Take five minutes to do exercise 2.15.

Self-Test Exercise 2.15

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$H_2C_2O_4(s) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(l)$$

Mole - Volume Calculation

Example 2.22

What volume of a 0.35 M $AgNO_3$ is required to react completely with 55mL of a 0.24 M NaCl solution?

Given	Required	Solution
V_{NaCl}=55 mL M_{NaCl}=0.24 M **Note** that M and V are given means number of moles (n) is known: n = MV	V_{AgNO_3}=?	see below

Stoichiometry	1 mole	1 mole		
Equation	NaCl(aq) + $AgNO_3$(aq)		$NaNO_3$(aq) + AgCl(s)	
Experimental	0.24 M, 55 mL	0.35 M, V=?		

Stoichiometry: = n_{NaCl}/n_{AgNO_3}=1/1

= (0.24 mol/L x0.055L)/(0.35 mol/L x V_{AgNO_3})

=1

V_{AgNO_3}= (0.24 mol/L x0.055L)/(0.35 mol/L) = 0.0377 L=**37.7 mL**

Mole - Number of Particles

Example 2.23

1. How many HCl particles are there in 0.1 mole of the HCl solution?

Strategy: 1 mole HCl contains Avogadro's number of HCl particles, then 0.1 mole contains how many?

Answer: 0.6022 x 10^{23} particles = 6.022 x 10^{22} HCl particles

Now it is your turn to do a mole-number of particles calculation. Take five minutes to do exercise 2.16.

Self-Test Exercise 2.16

What volume of 0.5 M glucose ($C_6H_{12}O_6$) contains 3.01 × 10^{22} molecules of glucose?

Section 2.9: Describing Reactions in Solution

- This is the last section to be considered in this unit. It is perhaps, the simplest of the topics we have covered so far in the unit. Here, you will study about ways of describing reactions in aqueous solutions by writing molecular equation, ionic equation, and net-ionic equation.

Learning Outcomes

At the end of this subsection you should be able to

- Explain the relationship between reacting ions, spectator ions, precipitation and solubility.
- Write net ionic equations.

Study Notes and Important Points

Molecular Equation

A molecular equation is a chemical equation in which the reactants and the products are written as if they were molecular substances, even though they may actually exist in solution as ions. The state (solid, liquid, gas) of each substance is indicated in parentheses after formula. (aq) is used to indicate that molecule is dissolved in water.

Consider the reaction between silver nitrate, $AgNO_3$, and sodium chloride, NaCl, in aqueous solution to give solid silver chloride, AgCl, and aqueous sodium nitrate. The equation for this reaction may be written as:

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

This is just a molecular equation. Note from the solubility rules (***Table 2.1***) that all common nitrates and chlorides of Group 1A are soluble. That is why (aq) is used after these formulae. The solubility rule also taught us that AgCl is insoluble in water. That is why (s) to indicate that it is precipitated out from the solution to the bottom of the reaction flask/test tube as a solid.

Ionic Equation

Although a molecular equation is useful in describing the actual reactants and products, it does not give any information about what is happening at the level of ions. Since this kind of information is very useful, you often need to write the molecular equation as an ionic equation. Consider the reaction of silver nitrate and sodium chloride. Both are soluble ionic substances and are strong electrolytes. When they dissolve in water, they go into solution as ions giving on the reactant side of the reaction:

$$Ag^+(aq) + \ NO_3^-(aq) \ + \ Na^+(aq) + Cl^-(aq)$$

On the product side of the equation, AgCl(s) is an ionic compound that doesn't dissolve in water, but the ions are held together in particular sites in the crystalline solid. We leave the formula as AgCl(s) to convey this information in the equation. On the other hand, $NaNO_3$ is a soluble ionic compound and is a strong electrolyte. Also it dissolves in aqueous solution to give freely moving ions. Therefore, we can write it as

$$Na^+(aq) + NO_3^-(aq)$$

The complete equation is

$$Ag^+(aq) + \ NO_3^-(aq) \ + \ Na^+(aq) + Cl^-(aq) \ \rightarrow AgCl(s) \ + Na^+(aq) \ + \ NO_3^-(aq),$$

The ions appearing on both sides of the equation (Na^+ and NO_3^-) are called spectator ions, as they do not take part in the reaction and they can be canceled on both sides to express the essential reaction that occurs.

$$Ag^{+}(aq) + \cancel{NO_3^{-}}(aq) + \cancel{Na^{+}}(aq) + Cl^{-}(aq) \quad AgCl(s) + \cancel{Na^{+}}(aq) + \cancel{NO_3^{-}}(aq)$$

The resulting equation is:

$$Ag^{+}(aq) + Cl^{-}(aq) \quad AgCl(s)$$

☞ This net ionic equation, is without spectator ions and the reaction that actually occurs at an ionic level is between silver ions and chloride ions which forms solid silver chloride.

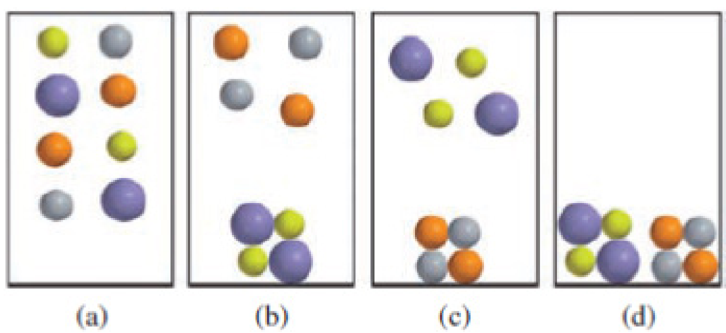
Steps for writing a net ionic equation

1. Write the balanced molecular equation
2. Write the balanced complete ionic equation
3. Cancel out the spectator ions. What remains is the net ionic equation.

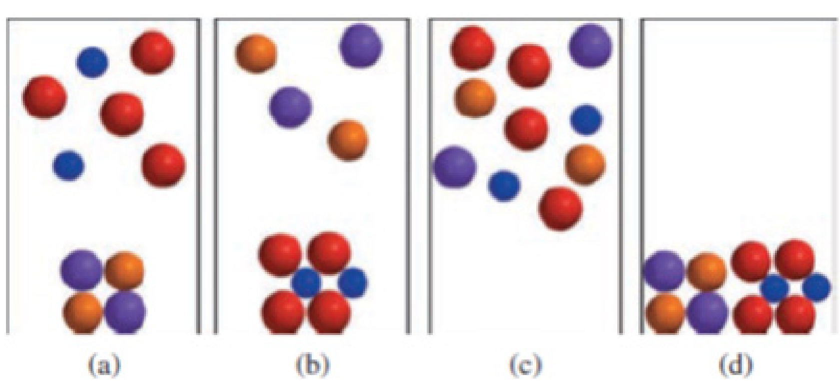
Dear student, we have completed our discussion of the last section of this unit- describing reactions in solutions. Now it is your turn to do some activities and self-test exercises to check how well you understood the concepts in this last section of the unit.

Activity 2.7

1. Two aqueous solutions of $AgNO_3$ and NaCl are mixed. Which of the following diagrams best represents the mixture? (Ag^{+} is gray; Cl- is orange; Na^{+} is green; NO_3^{-} is blue) (For simplicity, water molecules are not shown.)



2. Two aqueous solutions of KOH and $MgCl_2$ are mixed. Which of the following diagrams best represents the mixture? (K+ is purple; OH- is red; Mg^{2+} is blue; Cl^{-} is orange) (For simplicity, water molecules are not shown.)



Self-Test Exercise 2.17

1. Discuss the difference between molecular equations and ionic equations? What are spectator ions?
2. Consider the reaction between sodium carbonate and calcium hydroxide in aqueous solution:
   a. How are sodium carbonate and calcium hydroxide found in water solution, dissociated to give ions or as solid? (Refer to solubility rule). Write dissociation reaction for each of them.
   b. Write the ionic equation for the reaction between sodium carbonate and calcium hydroxide. Are the products formed soluble in water?
   c. Write the soluble product in ionic form. Are there the same ions in the reactant and product side? If yes, cancel them out!
   d. Write the net ionic equation.
3. For the following molecular equation, write ionic equation, net ionic equation, and identify the spectator ions:

   $$2AgNO_3\,(aq) + Na_2CrO_4\,(aq) \rightarrow Ag_2CrO_4\,(s) + 2NaNO_3\,(aq)$$

Unit Summary

To ensure that you understand the material in this chapter, you should review the meanings of the bold terms in the following summary and ask yourself how they relate to the topics in the chapter.

A heterogeneous mixture is a mixture that consists of physically distinct parts, each with different properties.

A homogeneous mixture is a mixture in which the composition is the same throughout. i.e., it has no visible boundaries because the components are mixed as individual atoms, ions and molecules.

A suspension is a heterogeneous mixture that consists of a dispersion of fine solid particles in a liquid or gas, removable by filtration.

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

Scattering of a beam of light by colloidal particles is called **Tyndall effect.**

A colloid can be a gel, an emulsion, a sol, foam, and aerosol.

Coagulation is the process by which the dispersed phase of a colloid is made to aggregate and thereby separate from the continuous phase.

A colloid in which the dispersed phase consists of micelles is called an **association colloid**.

A solution is a homogeneous mixture of a solute dissolved in a solvent through the action of intermolecular forces. Ion-dipole, ion–induced dipole, and dipole–induced dipole forces occur in solutions, in addition to all the intermolecular forces that also occur in pure substances.

The major component is the solvent, while the minor component is the solute. The terms miscible and immiscible, instead of soluble and insoluble, are used for liquid solutes and solvents. The statement like dissolves like is a useful guide to predicting whether a solute will dissolve in a given solvent.

Dissolving occurs by solvation, the process in which particles of a solvent surround the individual particles of a solute, separating them to make a solution. For water solutions, the word hydration is used. If the solute is molecular, it dissolves into individual molecules. If the solute is ionic it dissolves into ions.

The amount of solute in a solution is represented by the concentration of the solution. The maximum amount of solute that will dissolve in a given amount of solvent is called the **solubility of the solute**. Such solutions are **saturated**. Solutions that have less than the maximum amount are **unsaturated**. Most solutions are unsaturated, and there are various ways of stating their concentrations.

Solubility depends on two factors, predominantly **intermolecular forces** of attraction and to some extent conditions (surface area, temperature, pressure).

A rise in temperature increases solubility if the dissolving of additional solute is endothermic and the solubility of gases, almost always decreases with increasing temperature.

The solubility of most of solid solutes increase with temperature whereas those of gases invariably decrease.

Pressure has virtually no effect on the solubility of solids or liquids in liquid solvents and the solubility of gas molecules, are very markedly affected by pressure changes.

Henry's law says that the solubility of a gas is directly proportional to its partial pressure above the solution.

Dear student, we have come through long way to the end of this unit. I hope you have enjoyed the lesson and learned a lot of experiences. The activities and self-test exercises are expected to help you reach your goals. Please use them to check your understanding of the learned material.

Now, take some hours to go through the following check lists and do the review exercises. You can check your answers against the answer keys provided for you at the end of the questions.

Checklist 2.1

The following checklists are provided to you as a guide to check what you have learned in this unit. Check you understanding by putting a "✓" mark if you can remember what is presented in relation to the terms. If you don't remember, please go back to the corresponding section and make a quick review. I can ...

SN	Competencies	Check
1	define Mixtures	
2	distinguish between Homogeneous and heterogeneous mixtures	
3	explain the difference between a Solute and a Solvent	
4	give definitions of solution, suspension, and colloid	
5	describe types of solutions and give as many examples as possible of each type	
6	explain Tyndall Effect	
7	describe The solution process	
8	discuss Solute-solute, solute-solvent, and solute-solvent interactions¬	
9	describe inter-particle forces	
10	give definitions of solvation and hydration	
11	explain the difference between miscible & immiscible	
12	describe exothermic dissolution process and give examples	
13	describe endothermic dissolution process and give examples	
14	give definitions of Heat of solution	
15	explain effect of temperature on solubility of solids, liquids, and gasses in water.	
16	explain effect of pressure on solubility of gasses	
17	explain that pressure has virtually no effect on solubility of solids and liquids.	
18	state Henry's law	
19	apply Henry's law	
20	determine the solubility of gasses at a given pressure.	
21	prepare saturated, unsaturated, and supersaturated solutions	
22	perform calculatios involving, Ppm, Ppb, Mole, fraction, Molarity, Molality, and Normality	
23	prepare solutions of various concentrations and units	
24	carry out dilutions of concentrated solutions to prepare less concentrated ones.	
25	write Molecular equation, Ionic equation, Net-ionic equation	

Self-Assessment Exercises

Part I: Multiple Choice Questions

1. Soda water is an example of:
   a. liquid-liquid solution
   b. liquid-gas solution
   c. solid-liquid solution
   d. gas-gas solution
2. Which of the following type of matter can exhibit Tyndall effect?
   a. aerosol b. soda water c. brass
3. A solution that is formed when NaCl dissolves in water has the properties of-----
   a. Sodium Chloride
   b. Water
   c. both NaCl and water
   d. neither NaCl nor water
4. A simple way of determining whether a mixture is colliodial or not is through the use of ______
   a. particle size b. Tyndall effect c. non-uniform appearance d. a & b

5. A suspension is ……
   a. a homogenous mixture of a solute and a solvent
   b. a heterogeneous mixture of one phase dispersed in another phase
   c. a heterogeneous mixture that exhibits Tyndall effect
   d. a mixture that contain smaller particle size than a solution
6. Which properties of colloids is applied in the cloud seeding mechanism to produce artificial rain?
   a. Association c. Coagulation
   b. Dispersion d. Scattering
7. The cleansing action of soaps and detergents are related to which property of a colloid?
   a. Association c. Coagulation
   b. Dispersion d. Scattering
8. The continuous and dispersed phase in fog is _____ and _____ , respectively.
   a. gas, solid c. gas, liquid
   b. liquid, gas d. liquid, liquid
9. How many grams of NaOH is required to prepare 100 mL of 1 M NaOH solution?
   a. 4 b. 40 c. 0.4 d. 400
10. A solution that contain 4.9 g H_2SO_4 in 1 liter of solution has the same _____ and _____. Assume the solution has the same density as water.
   a. molarity and molality c. normality and molality
   b. molarity and normality d. none of the above
11. When the molarity of a sulfuric acid solution is doubled, its normality is _____
   a. halved b. doubled c. quadrupled d. remain the same
12. Mole fraction and percentage differ by a factor of _____
   a. 1 b. 10 c. 100 d. 0.1
13. Which of the following substances are not readily miscible within each other?
   a. C_6H_6 and CCl_4 c. C_2H_5OH and H_2O
   b. H_2O and CCl_4 d. CH_3OH and H_2O
14. Rate of dissolution is largely dependent upon:
   a. the inter-particle forces
   b. the surface area of solid solute
   c. the temperature of the system
   d. the pressure of the system
15. Which of the following statements is not correct?
   a. pressure has little effect on the solubility of liquids and solids
   b. the solubility of most solids increases with increasing temperature
   c. the solubility of gases in water increases with increasing temperature.
   d. none of the above
16. The quantitative relationship between gas solubility and pressure is given by:
   a. Raoult's law c. Henry's law
   b. Hess's law d. Dalton's law
17. A solution that has a capacity to dissolve more solute is called ___________ solution

a. dilute
b. unsaturated
c. supersaturated
d. concentrated

18. A concentrated solution is ____________
   a. same as supersaturated solution
   b. contain less solute than solvent
   c. contains relatively large amount of solute
   d. a saturated solution
19. Which of the following solutions is the most dilute?
   a. 0.1 M NaCl
   b. 0.01 M NaCl
   c. 0.001 M NaCl
   d. 10.0 mM NaCl
20. Which of the following is a saturated NaCl solution at 20°C? Solubility of NaCl is 36 g/100 mL.
   a. 0.615 M
   b. 6.15 M
   c. 36.0 M
   d. 0.0615 M
21. The ratio of the number of moles of solute divided by the total number of moles gives:
   a. the mole fraction of the solute
   b. the molarity of the solution
   c. the molality of the solution
   d. the normality of the solution
22. Phosphoric acid reacts with calcium hydroxide to form calcium hydrogen phosphate ($CaHPO_4$) and water. The equivalent weight of phosphoric acid in this reaction will be:
   a. 98 g/eq
   b. 49 g/eq
   c. 32.6 g/eq
   d. 14 g/eq
23. Which of the following is the balanced net ionic equation for the reaction between H_2SO_4 (a) and NaOH (aq)?
   a. $HSO_4^-(aq) \rightarrow H^+(aq) + SO_4^{2-}(aq)$
   b. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
   c. $SO_4^{2-}(aq) + 2Na^+(aq) \rightarrow NaSO_4^{2-}(aq)$
   d. $H_2SO_4(aq) + 2OH^-(aq) \rightarrow 2H_2O(l) + SO_4^{2-}(aq)$
24. When a given volume of concentrated solution is diluted 100 fold, which one of the following is true?
   a. the dilute solution has more number of moles of solute per unit volume than that of concentrated solution
   b. the number of moles of solute per unit volume of the dilute solution is equal to that of the concentrated solution
   c. the number of moles of solute in a given volume of a solution is the same before and after dilution
   d. all of the above

Part II: Short Answer Questions

25. Give one example of each: a gaseous solution, a liquid solution, a solid solution.
26. What are the two factors needed to explain the differences in solubilities of ionic solids in water?
27. Explain in terms of intermolecular attractions why octane is immiscible in water.
28. Give the type of colloid (aerosol, foam, emulsion, sol, or gel) that each of the following represents.

a. rain cloud b. milk of magnesia
c. soapsuds d. silt in water

29. Concrete is a mixture of _____ , _____ , _____ , and _____ .
30. Explain on the basis that "like dissolves like" why glycerol, $CH_2OHCHOHCH_2OH$, is miscible in water but benzene, C_6H_6, has very limited solubility in water.
31. Explain why ionic substances show a wide range of solubilities in water.
32. Indicate the type of solute-solvent interaction that is most important in each of the following solutions:
    a. KBr in water
    b. hexane, C_6H_{14}, in gasoline
    c. ammonia in water
33. Describe the characteritics of endothermic and exothermic dissolution processes.
34. Consider the following solutions. In each case, predict whether the solubility of the solute should be high or low. Justify your answer.
    a. KCl in H_2O
    b. HF in H_2O
    c. KCl in CCl_4
    d. H_2O in CH_3OH
    e. CCl_4 in H_2O
35. What is the usual solubility behavior of an ionic compound in water when the temperature is raised? Give an example of an exception to this behavior.
36. Give one example of each: a salt whose heat of solution is exothermic and a salt whose heat of solution is endothermic.
37. What do you expect to happen to a concentration of dissolved gas in a solution as the solution is heated?
38. Explain why a carbonated beverage must be stored in a closed container.
39. Pressure has an effect on the solubility of oxygen in water but a negligible effect on the solubility of sugar in water. Why?
40. Which of the following ions would be expected to have the greater energy of hydration, Mg^{2+} or Al^{3+}?
41. Which of the following ions would be expected to have the greater energy of hydration, F^- or Cl^-?

Part III: Work Out Problems

42. The solubility of carbon dioxide in water is 0.161 g CO_2 in 100 mL of water at 20°C and 1.00 atm. A soft drink is carbonated with carbon dioxide gas at 5.50 atm pressure. What is the solubility of carbon dioxide in water at this pressure?
43. Calculate the molarity of each of the following solutions:
    a. 10.5 g NaCl in 350.0 mL of solution
    b. 40.7 g $LiClO_4 \cdot 3H_2O$ in 125 mL of solution
44. What mass of solution containing 5.00% potassium iodide, KI, by mass contains 258 mg KI?
45. Caffeine, $C_8H_{10}N_4O_2$, is a stimulant found in tea and coffee. A sample of the substance was dissolved in 45.0 g of chloroform, $CHCl_3$, to give a 0.0946 m solution. How many grams of caffeine were in the sample? Molar mass caffeine = 194.19 g/mol
46. A 100.0-g sample of a brand of rubbing alcohol contains 65.0 g of isopropyl alcohol, C_3H_7OH, and 35.0 g of water. What is the mole fraction of isopropyl alcohol in the

solution? What is the mole fraction of water?

47. A bleaching solution contains sodium hypochlorite, NaClO, dissolved in water. The solution is 0.650 m NaClO. What is the mole fraction of sodium hypochlorite?

48. The concentrated sulfuric acid we use in the laboratory is 98.0% H_2SO_4, by mass. Calculate the molality and molarity of the acid solution. The density of the solution is 1.83 g/mL.

49. Calculate the approximate volume of water that must be added to 250 mL of 1.25 N solution to make it 0.500 N.

50. An antiseptic solution contains hydrogen peroxide, H_2O_2, in water. The solution is 0.610 m H_2O_2. What is the mole fraction of hydrogen peroxide?

51. Citric acid, $H_3C_6H_5O_7$, occurs in plants. Lemons contain 5% to 8% citric acid by mass. The acid is added to beverages and candy. An aqueous solution is 0.688 m citric acid. The density is 1.049 g/mL. What is the molar concentration?

52. A solution of vinegar is 0.763 M acetic acid, $HC_2H_3O_2$. The density of the vinegar is 1.004 g/mL. What is the molal concentration of acetic acid?

Assignment for Submission

Part I. Multiple Choice Questions

Direction: Choose The Best Answer and write the letter of Your Choice on a Separate Answer Sheet to be submitted to Your Tutor

1. A solution is characterized by
   a. Constant composition
   b. Definite proportion of solute and solvent
   c. Its ability to scatter light and exhibit stable Tyndall effect
   d. Uniform distribution of solutes in solvent
2. The components in which of the following types of mixtures can be separated by filtration
   a. Solution
   b. Suspension
   c. Colloids
   d. Homogeneous mixtures
3. The dissolution of ionic compounds in water involves which of the following interactions between solute and solvent?
   a. Dipole-dipole
   b. Ion-dipole
   c. Hydrogen bonding
   d. London force
4. Which of the following list of molecules include all nonpolar?
   a. $CHCl_3$, CCl_4, C_6H_{14}, DMF (dimethyl formamide)
   b. CH_3OH, CH_2Cl_2, CH_3CH_2OH and H_2O
   c. NH_3, H_2S, DMSO (Dimethyl Sulfoxide)
   d. CH_4, O_2, BF_3, SO_3, SF_6, PCl_5, XeF_4, and CO_2
5. All of the following compounds can dissolve in water through hydrogen bonding

except?

a. NH_3
b. H_2S
c. amino acids
d. carboxylic acids
e. alcohols

6. Some ionic compounds such as NaCl are classified as soluble while others such as $PbCl_2$ are classified as insoluble? Why? Because compounds such as $PbCl_2$
   a. Are nonpolar
   b. Have high lattice energy
   c. Have low lattice energy
   d. Have highly exothermic solvation energy
7. Solubility is defined as
   a. The minimum amount of solute that can be dissolved in a given solvent under specific condition
   b. The maximum amount of solute that can be dissolved in a given solvent under specific condition
   c. The minimum amount of solvent required to dissolve a given solute under specific condition
   d. The maximum amount of solvent required to dissolve a given solute under specific condition
8. A saturated solution is the one that
   a. Is in equilibrium and can dissolve more solutes
   b. Is in equilibrium and cannot dissolve more solutes
   c. Is metastable and cause immediate precipitation when ore solute is added
   d. Has not reached equilibrium and can dissolve more solutes
9. Which one of the following statements is not true about solubility of substances in water?
   a. Most ionic compounds are soluble in water
   b. Some polar covalent compounds are soluble in water
   c. Solubility of most solids increases with increasing temperature
   d. Solubility of gases increases with increasing pressure but decreases with increasing temperature
   e. Most nonpolar compounds are soluble in water
10. A dilute solution is the one that
   a. Contains fewer particles of solute per liter of solution
   b. Contains fewer particles of solute
   c. Contains lager number of particles per liter of solution
   d. Contains hgh number of particles of solute
11. Colloids and solutions can be distinguished by their
   a. Appearance
   b. Tyndall effect
   c. Ability to get filtered

   d. Sedimentation

12. Henry's law states that at a constant temperature, the amount of a given gas that dissolves in a liquid is
   a. directly proportional to the partial pressure of that gas in equilibrium with that liquid
   b. directly proportional to the total pressure of gas in equilibrium with that liquid
   c. inversely proportional to the partial pressure of that gas in equilibrium with that liquid
   d. inversely proportional to the total pressure of gas in equilibrium with that liquid
13. How many grams of NaCl is required to prepare 100.0 mL of 0.1 M solution?
   a. 58.5 g
   b. 5.85 g
   c. 0.585 g
   d. 585.0 g
14. Which of the following is not true about the relationship between concentration units?
   a. 1.0 M solution of H_2SO_4 is more concentrated than 1.0 N solution of H_2SO_4
   b. The molarity and molality of dilute solutions is the same
   c. 1 ppm is the same as 1.0 μg/L for dilute solutions
   d. The molarity and normality of dilute solutions are the same
15. What is the molarity of 36.5%(w/w) HCl solution (density of solution = 1.20 g/mL; molar mass HCl = 36.5 g/moL)?
   a. 12.0 M
   b. 37.0 M
   c. 18.0 M
   d. 11.0 M
16. Consider a solution formed by dissolving solute A in solvent B. If the mole fraction of A is 0.25, then the mole fraction of of B is ___________
   a. 0.75
   b. 1.75
   c. 0.5
   d. 25
17. How many moles of NaOH are needed to react with 196.0 g of H_2SO_4(molar mass 98 g/mol) according to the following equation: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
   a. 1.0 mol
   b. 4.0 mol
   c. 1.5 mol
   d. 1.2 mol
18. Four different solutions of equal volume (1L) were prepared by dissolving one mole of each of the following substances. The conduction of electricity is least in the solution containing:
   a. HCl
   b. HNO_3

c. CH_3COOH
d. KCl

19. The pH of a 0.1 M HCl solution is ______________
    a. 1
    b. 10
    c. 0.1
    d. 1.1
20. 20. Which of the following is true about the pH of 0.1 M each of HCl and CH_3COOH (acetic acid)?
    a. the pH of 0.1 M HCl is greater than 0.1 M CH_3COOH
    b. the pH of 0.1 M HCl is less than 0.1 M CH_3COOH
    c. the pH of 0.1 M HCl is equal to 0.1 M CH_3COOH
    d. Cannot be determined

Part II. Essay Type Question

21. Give three examples each of a) gaseous solutions b) liquid solutions c) solid solutions and identify the solute and the solvent
22. Discuss the mechanism of hot packs and cold packs used in health centres for alleviation of pains
23. Discuss the effect of temperature on solubility of
    a. Ionic solids in water
    b. Molecular solids in water
    c. Liquids in water
    d. Gases in water
24. Discuss how climate change (global warming) affects aquatic life and productivity?
25. Discuss the differences between
    a. Unsaturated and saturated solutions
    b. Saturated solution and supersaturated solution
    c. Dilute and concentrated solutions
    d. Molecular equation and ionic equation

Answer Key to Self Test Exercises

Answer Key to Activity 2.1

a. Mixture
b. Element
c. compound

Self-Test Exercise 2.1

1. See your module.
2.

a. Homogeneous mixtures:- coffee, tea, sea water, air, brass, steel, natural gas, vinegar,
c. Heterogeneous mixtures:- sand + water, pizza, vegetable salad, and fruit punch.
d. Blood, milk, and butter appear homogeneous but heterogeneous upon closer inspection

5. A sample of glucose, for instance, contains only glucose ($C_6H_{12}O_6$) molecules in

(6x12):(12x1):(16x6) ratio by mass. But a mixture of 10 g of glucose in 100 mL of water and a mixture of 30 g of glucose in 100 mL of water contain glucose and water molecules in a different ratio by mass. Therefore, when glucose dissolves in water you get a mixture. In earlier Grades you saw that when sodium chloride is dissolved in water, it is possible to separate the sodium chloride from the water by the physical process of distillation. However, sodium chloride is itself a substance and cannot be separated by physical processes into new materials. Similarly, pure water is a substance. No matter what its source, a substance always has the same characteristic properties. Sodium is a solid metal having a melting point of 98°C. The metal also reacts vigorously with water. No matter how sodium is prepared, it always has these properties. Similarly, whether sodium chloride is obtained by burning sodium in chlorine or from seawater, it is a white solid melting at 801°C and has a Na to Cl mass ratio of 23:35.5 regardless of the amount.

Answer Key to Activity 2.2

1. Refer to internet.
2. Municipality water treatment plant often employs sedimentation followed by coagulation. Chemicals containing highly charged metal ions such as $Al_2(SO_4)_3$ are often used for coagulation.
3. Oil is nonpolar whereas water is polar. So, they do not mix because they do not obey the "like dissolves like rule." They form separate layers upon standing. The layer get mixed when soap powder is added because the soap acts as a surfactant, emulsifier equivalent.
4. Let your friend evaluate the hair gel you have produced.

Self-Test Exercise 2.2

1. a) Normally an IV (intravenous) fluid is recommended. Note that IV fluid is a solution. You can refer to internet for the composition or ask a health professional. b) isotonic = 0.90% (w/v) of NaCl, meaning 9.0 g per liter hypotonic contains less than 0.90% (w/v) saline, and hypertonic solutions contains greater concentration of saline than 0.90% (w/v). Hypertonic solution may be administered to a patient suffering from dehydration due to excessive diarrhea to replenish the lost salts.
2. a) Glass is a homogeneous mixture of sodium silicate and calcium silicate.

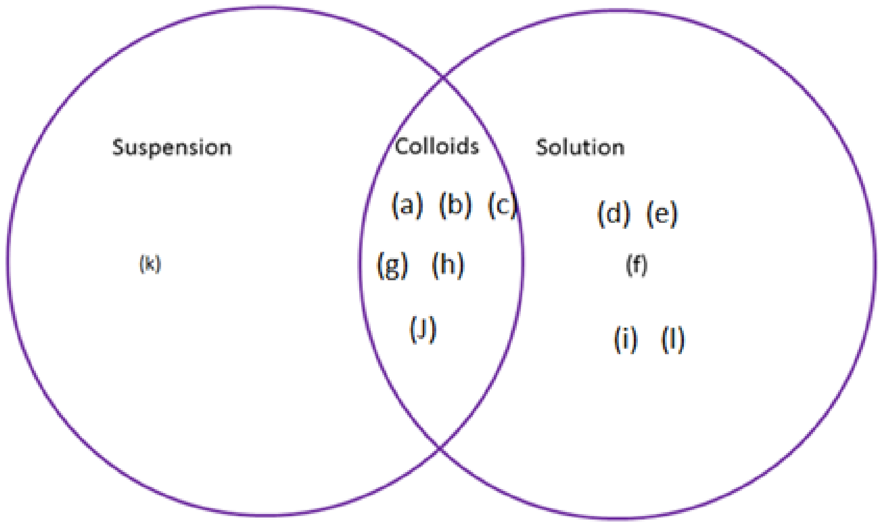
b) Sand is a heterogeneous mixture of SiO_2 and a lot of other organic and inorganic particles.

c) cement is a homogeneous mixture of calcium silicates and aluminates

d) Water gas is a homogeneous mixture of CO and H_2.

e) Producer gas is a homogeneous mixture of CO and N_2.

3. See your module
4. See your module
5. b) Aerosol
6.



7. A = Pure substances, B = Mixtures, C= Heterogeneous mixtures, D = Homogeneous mixtures, E = Elements, and F = Compounds

Answer Key to Activity 2.3

1. Answer is (b) because when NaCl dissolves in water it should split into Na^+ and Cl^- ions. These charges get neutralized by interaction with water molecules. The cation (Na^+) gets neutralized by the partial negative charges on oxygen (the red ball) of water molecules and the anion (Cl-) gets neutralized by the partial positive charges on hydrogens of water molecules.
2. Both quicklime (CaO) and slacked lime ($Ca(OH)_2$) are calcium-containing inorganic compounds. Since calcium fl uoride is insoluble in water (see solubility rule), the Ca^{2+} from the lime combines with F- in water and removes the fl uoride from the water by precipitating as CaF_2.

Self-Test Exercise 2.3

For question 1 to 3, choose the correct answer and then check your answer against the answer key provided at the end of this section.

1. c. hydrogen bond
2. d. Kr
3. B. False
4. Due to the presence of hydrogen bonding in H_2O which is stronger than the ordinary dipole-dipole force in H_2S.
5.

a. $BaCO_3$ = insoluble	g. $Ca_3(PO_4)_2$ = insoluble
b. $AlCl_3$ = soluble	h. $PbCl_2$ = insoluble
c. MgO = insoluble	i. $AgNO_3$ = soluble
d. $Al(OH)_3$ = insoluble	j. $(NH_4)NO_3$ = soluble
e. KI = soluble	k. NH_4Cl = soluble
f. $Al_2(SO_4)_3$ =soluble	

6.

a. Water and grain alcohol = miscible

b. "Tej" and water = miscible

c. Water and oil = immiscible
d. Water and honey = miscible
e. Water and petrol = immiscible
f. Gasoline and ''naphtha'' = miscible
7. a. HBr and H_2S interact via dipole-dipole forces
b. Cl_2 and CBr_4 interact via dispersion (London) forces
c. I_2 and NO_3- interact via ion-induced dipole
d. NH_3 and C_6H_6(Benzene) interact via dipole-induced dipole

Answer Key to Activity 2.4

1. Many hot packs use calcium chloride, which releases heat when it dissolves, according to the equation below.

$$CaCl_2(s) \rightarrow Ca^{2+}(aq) \ + \ 2Cl\text{-}(aq) \ + 82.8 \ kJ$$

Many cold packs use ammonium nitrate, which absorbs heat from the surroundings when it dissolves upon squeezing.

$$NH_4NO_3(s) \ + \ 25.7 \ kJ \ \rightarrow NH_4^+(aq) \ + \ NO_3\text{-}(aq)$$

2. Many people use hot and cold treatments at home using ice or hot shower to alleviate aches and pains caused by muscle or joint damage. Chemical hot packs and cold packs work because of the heats of solution of the chemicals inside them. When the bag is squeezed, an inner pouch bursts, allowing the chemical to dissolve in water. Heat is released in the hot pack and absorbed in the cold pack. Heat packs are used primarily for non-inflammatory body pain including duller and persistent pains associated with acute soreness, stiffness, cramping, and/or sensitivity. Heat allows blood vessels to expand and relaxes muscles. The soothing effect occurs because heat stimulates circulation and increases tissue elasticity. Please never apply heat to open wound, broken or infected skin or injured tissues. If the area in question is either bruised or swollen (or both), it may be better to use cold therapy.

Self-Test Exercise 2.4

1. Which one of the following pairs of ions has greater enthalpy of hydration?

i. Na^+ or Cs^+? Na^+ because of small size

ii. Mg^{2+} or Cs^+? Mg^{2+} because of multiple charges

iii. F- or Cl-? Explain. F- because of a small size.

2. ΔHsoln = -43.5 kJ mol-1, verify this!

Answer: Tf = 33.2 °C, verify this!

3. False, because the formation of ideal solutions does not involve energy change.

Answer Key to Activity 2.5

1. Dissolution of added crystals would be observed in beaker A and precipitation is expected in beakers B and C. Since, beaker C contains supersaturated solution which is unstable, the added mass would cause even more solid precipitated than the amount added.

Self-Test Exercise 2.5

1. Please refer to your module.

Answer Key to Activity 2.6

1. $C_1/P_1 = C_2/P_2$ (0.032 M)/(3.0 atm)=C_2/(5.0 atm) C_2= (0.032 M x 5.0 atm)/(3.0 atm)=0.053 M
2. The rise in temperature decreases the solubility of O_2 in water. As the level of O_2 gets depleted, fish may die short of O_2.
3. Dissolved chemicals break the O_2 and H_2O molecules interaction and force the O_2 to leave the water. Consequently, the level of dissolved O_2 decreases affecting aquatic life. Other wastes such as wastes of personal care products do also cause environmental pollution and need to be disposed appropriately. They should not be thrown irresponsibly to open places or to water bodies.
4. When the cap is removed the solution becomes exposed to atmospheric pressure which is much less than the pressure under which the coke was packed. As the partial pressure of carbon dioxide is reduced, gas comes out of solution, solubility decreases.

Self-Test Exercise 2.6

1. c. NH_3

Self-Test Exercise 2.7

1. Answer: 0.892 /(0.892 + 54.6) x 100 % = 1.61%
2. a) ppm = (0.120 g/150 g) x 106 = 800 ppm

b) ppb = (0.120 g/150 g) x 109= 800,000 ppb

3. (24.0 g/243 mL)x100% = 9.88%
4. w/v% NaCl = 13.5% = (grams NaCl/625) x 100%

Mass NaCl = (13.5 x 625)/100 = 84.4 g

5. ppm = (mass Ca^{2+}/mass pill) x106 =11,571 ppm
6. 6. 5.0% alcohol = (volume alcohol/volume solution)x100%

Volume alcohol = (5.0 x volume solution)/100 = (5.0 x 300)/100 = 15 mL

Self-Test Exercise 2.8

1. Mole isopropyl alcohol = mass of isopropyl alcohol/molar mass of isopropyl alcohol =142g/60.1g/mol = 2.36 mol

Mole water = mass of water/molar mass of water = 58.0 g/18 g/mol =3.2 mol Total mole = 2.36 + 3.2 = 5.56

Xalcohol = 2.36/5.56 = 0.42

Xwater = 3.2/5.56 = 0.58

2. Answers: A(0.25), B(0.25), C(0.50)

Self-Test Exercise 2.9

1. 5.85 g of sodium chloride (NaCl) is dissolved in water to make 250 mL of solution.

Assuming the density of the solution as 1 g/mL, calculate

a. Molarity = number of moles of solute/volume of solution in liter = (5.85g/58.5 g/mol)/0.250L =0.400 M

b. the mass percentage of the solute. Let's assume the density of solution is the same as the density of water = 1.0 g/mL Then mass of solution = density x volume = 1.0 g/mL x 250 mL = 250 g (w/w)%NaCl = (5.85/250) x100% = 2.34%

2. How would 250 ml of 0.15 M KNO_3 solution be prepared?

   Follow the following steps to prepare your solution

   Step 1. Calculate the required mass

   Let's determine number of moles first:

   Formula: Molarity = number of moles of solute/volume of solution in liter

   Molarity = 0.15 M

   Volume of solution = 0.250 L

   Number of moles of KNO_3 = 0.15 mol/Lx 0.250 L = 0.038 mol

   Then mass required is calculated as follows:

   Mass KNO_3 = mole x molar mass = 0.038 mol x 101.10 g/mol = 3.8 g.

   Step 2. Weigh out the 3.8 g KNO_3 from the stock (pure KNO_3 solid).

   Step 3. Transfer it to a beaker and dissolve by adding about 100 mL distilled water.

   Step 4. Transfer the solution to the 250.0 mL volumetric flask. We use 250 mL volumetric flask because the final volume of solution we want to prepare is 250 mL.

   Step 5. Dilute to the mark by filling the remaining portion of the flask with the solvent.

   Step 6. Label your solution

3. 36%(w/w) HCl by mass means 36 g HCl in 100 g solution.

   Now, calculate the mole of the 36 g

   Mole HCl = mass/molar mass = 36 g/36.458 g/mol = 0.99 mol

   Mass water = mass of solution – mass of HCl = 100-36 = 64 g

   Mole water = mass water/molar mass water = 64 g/18 g/mol = 3.6 mol

   XHCl = mole HCl/total mol = 0.99/(0.99 + 3.6) = 0.22

Self-Test Exercise 2.10

1. Answer: mass of glucose = 2.43 g

Self-Test Exercise 2.11

1. Normality = equivalents/volume of solution in liter

   Equivalents = Normality x Volume of solution in liter

   Equivalents = 2 eq/L x 1 L = 2

2. 0.48 = eq/0.25 Eq = 0.12 = (z x given mass)/molar mass, but z = 6 Mas = (molar mass x 0.12)/z = (342.14 x 0.12)/6 = 6.84 g

3. Molarity = mole/litre = (16.2g/399.9g/mol)/0.20 = 0.20 M

Normality = M x z = 0.20 x 6=1.20 N

4. Calculate the normality of

   a. 0.1381 M NaOH

   Z = 1

   N = M x z=0.1381 x 1=0.1381 N

   b. 0.0521 M H3PO4

   Z = 3

   N = 3 x M = 3 x 0.0521 = 0.156 N

Self-Test Exercise 2.12

1. Write the formula for molality

Determine what is given and what is missing; 30.0% by mass mean 30.0 g of H2O2 in 100 g of solution. Since we need mass of solvent for molality, we can calculate it as follows: mass (solution) = mass(solute) + mass (solvent); 100 g = 30.0 g + mass (solvent); □ mass (solvent)=70 g = 0.07kg We can also calculate the moles of solute (H2O2) as mole = given mass/molar mass; the molar mass of H2O2 is 34.02 g/mol 30.0 g/34.02 g.mol^{-1}= 0.882 mol H_2O_2 Substitute the moles of solute and mass of solvent in kg in the molality formula Ans:12.6 m H_2O_2

To calculate the mole fraction use the formula:

$$X\ (\text{solute}) = \frac{\dfrac{\text{mass (solute)}}{\text{molar mass solute}}}{\text{mole (solution)}}$$

((30.0 g H_2O_2)/(34.02 g/mol))/[((30.0 g H_2O_2)/(34.02 g/mol)))+((70 g)/(18 g/mol))] =0.184

Note: molar mass of water (H_2O) is 18 g/mol

Therefore, the mole fraction of solvent (water) = 1-0.184=0.816

(Verify this using the formula for mole fraction of the solvent)

To calculate the molarity follow these steps:

Write the formula for molarity

Identify what is given. We need number of moles of solute (H_2O_2) and volume of solution in litre.

From Q(b) we know moles of H_2O_2 = 0.882 moles,

Volume = mass/ density = 100 g/ 1.11 g.mL^{-1} = 90.1 mL = 0.0901 L,

note that we are considering 30.0 g of H_2O_2 and 100 g of solution as per the given 30.0%.

Substituting the values:

M = (0.184 moles of H_2O_2)/(0.0901 L) = 9.79 M

Self-Test Exercise 2.13

1. Given Vf = 2.00 x 10^2 mL

Cf = 0.866 M Ci = 5.07 M

Vi = ?

Answer Vi= (Cf x Vf)/Ci = (0.866 M x 200 mL)/5.07 M = 34.20 mL

Step 1. Weigh out 34.20 mL from the stock solution using graduated or measuring cylinder

Step 2. Transfer the weight volume to a 200 mL volumetric fl ask

Step 3. Dilute to the mark

Step 4. Label it. 'Done'!

2. A dilute solution has less solute present in the larger volume of solution than a concentration solution. In other words, a dilute solution contains a larger volume of solvent than a concentrated solution for a given amount of solute. A solution is said to be unsaturated if more solute could be dissolved in a given amount of the solvent at a specified temperature.

A solution is said to be saturated when no more of the solute can be dissolved in a given amount of the solvent at a specified temperature. So, the terms dilute and unsaturated are not related terms. Dilute or concentrated terms are relative to each other, but unsaturated or saturated refer to the dissolution capacity of the solution.

3. Outline the preparation of 100 mL of

(a) hypotonic (0.1 M) and

mole NaCl = VM = 0.1 L x 0.1 mol/L = 0.01;

mass NaCl required = n.Molar Mass NaCl = 0.01 mol x 58.5 g/mol = 0.585 g;

So you should dissolve 0.585g NaCl in a 100 mL volumetric flask and dilute to the mark with water.

(b) mole NaCl = VM = 0.1 L x 0.25 mol/L = 0.025; mass NaCl required = n.Molar Mass NaCl = 0.025 mol x 58.5 g/mol = 1.46 g; So they should dissolve 1.46 g NaCl in a 100 mL volumetric fl ask and dilute to the mark with water c) Outline preparation of 100 mL of 0.1 M H_2SO_4 using a stock solution of 98% H_2SO_4.

(c)

Calculate initial concentration of stock solution in molarity:

98% (w/w) means 98 g H_2SO_4 in 100 g of solution.

Volume solution = mass/density = 100 g/1.84g/ml =54.34 mL = 0.05434 L

Mole H_2SO_4 = given mass/molar mass = 98 g/ 98.079 g/mol = 0.999 mol

Molarity of the 98% H_2SO_4 = number of moles/volume solution (L)

= 0.999 mol/0.05434 L = 18.38 M

Now, you need to know what volume to withdraw from this solution

Using the dilution formula Vi = (CfVf)/Ci = (0.1 M x0.1 L)/18.38 M = 0.540 mL= 540 µL

Measure 540 µL from stock solution using micropipette, transfer it to 100 mL volumetric flask, dilute it to the mark. Label it.

Self-Test Exercise 2.14

What is the concentration of sodium hydroxide that is required to react completely with equal volume of 0.104 M hydrochloric acid?

NaOH + HCl = NaCl + H_2O

Stoichiometry	1 mole		1 mole				
	NaOH	+	HCl	→	NaCl	+	H_2O
Given	V, M		V, 0.104 M				

Stoichiometry=(1 mole NaOH)/(1 mole HCl)= (V x M NaOH)/(V x 0.104 M HCl)

Since volume is equal it cancels out

Therefore; (M NaOH)/(0.104 M HCl)=1

This give M (molarity of NaOH) = 0.104 M

Self-Test Exercise 2.15

Strategy

First calculate number of moles of glucose using the given number of molecules and Avogadro's number

We have

1 mole glucose = 6.02×10^{23} molecules

X mole = 3.01×10^{22} molecules

Simple crisscross multiplication gives the value of X (number of moles) as:

Number of moles of glucose as 0.05 moles

We can calculate volume using the formula:

n = MV

V= n/M = 0.05 mole/0.5 M = 0.1 L = 100 mL

Answer Key to Activity 2.7

1. (c)
2. (b)

Self-Test Exercise 2.16

1. See your module
2. a. Sodium carbonate is soluble and exists as $Na^+(aq)$ and $CO_3^{2-}(aq)$
Calcium hydroxide is slightly soluble and produces some $Ca^{2+}(aq)$ and $OH^-(aq)$
b. $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow 2NaCl(aq) + CaCO_3(s)$
The products NaCl is soluble and designated as NaCl(aq); $CaCO_3$ is insoluble and designated as $CaCO_3(s)$.
c. $2Na^+(aq) + CO_3^{2-}(aq) + Ca^{2+}(aq) + 2OH\text{-}(aq) \rightarrow 2Na^+(aq) + 2OH^-(aq) + CaCO_3(s)$
d. Na^+ and OH- are spectator ions.
e. Net ionic equation: $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$
3.
Balanced Molecular equation:
$2AgNO_3\ (aq) + Na_2CrO_4\ (aq) \rightarrow Ag_2CrO_4\ (s) + 2NaNO_3\ (aq)$
Ionic equation:
$2Ag^+(aq) + 2NO_3^-(aq) + 2Na^+(aq) + Cr_2O_4^{2-}(aq) \rightarrow Ag_2CrO_4(s) + 2Na_+\ (aq) + 2NO_3^-(aq)$
Net ionic equation: $2Ag+(aq) + Cr_2O_4^{2-}(aq) \rightarrow Ag_2CrO_4(s)$
The spectator ions are NO_3-, and Na^+

Answer Keys to Review Questions

Part I. Multiple Choice

1. B	6.C	11.C	16.C	21.A
2.A	7.A	12.C	17.B	22.B
3. C	8.C	13.B	18.C	23.B
4.B	9.A	14.A	19.C	24.C
5.B	10.A	15.C	20.B	

Part II: Short Answer Questions

25. Give one example for each: a gaseous solution (air), a liquid solution (beer), a solid solution (ruby). See student's textbook
26. What are the two factors needed to explain the differences in the solubilities of ionic solids in water? Answer: Lattice energy and hydration energy
27. Explain in terms of intermolecular attractions why octane is immiscible in water. Octane is non-polar while water is polar. They dislike each other.
28. Give the type of colloid (aerosol, foam, emulsion, sol, or gel) that each of the following represents.
    A. rain cloud (aerosol)
    B. milk of magnesia (sol)
    C. soapsuds (foam)
    D. silt in water (sol)
29. Sand, cement, water, and stone (rock)
30. glycerol, $CH_2OHCHOHCH_2OH$, and water are both polar. Benzene, C_6H_6, is non-polar and has very limited solubility in polar water.
31. Because of the difference in the degree of hydration and lattice energy

32. A. KBr in water, ion-dipole B. hexane, C_6H_{14}, in gasoline, dispersion force
    C. ammonia in water, dipole-dipole
33. Endothermic dissolution processes are characterized by:
    The magnitude of the energy absorbed to break up the lattice, ΔH_{lat}, is greater than the magnitude of the energy released when solute particles are surrounded by water solvent molecules, ΔH_{hyd}, so the enthalpy of solution, ΔH_{soln}, is positive, that is, the process is endothermic.
    if $|\Delta H_{lat}| > |\Delta H_{hyd}|$ then ΔH_{soln} is positive
    solute (s) + water (l) → solution(aq) ΔH_{soln} = +ve

solute (s) + water (l) + ΔH → solution(aq)

Energy is absorbed, that is, energy is a reactant.

For endothermic dissolution processes, when solute is added to water, water temperature decreases.

Exothermic dissolution processes are characterized by:

The magnitude of the energy absorbed to break up the lattice is less than magnitude of the energy released when solute particles are surrounded by water solvent molecules, so the enthalpy of solution is negative, that is, the process is exothermic.

if $|\Delta H_{lat}| < |\Delta H_{hyd}|$ then ΔH_{soln} is negative

solute (s) + water (l) → solution(aq) ΔH_{soln} = -ve

solute (s) + water (l) → solution(aq) + ΔH

Energy is released, that is, energy is a product.

For exothermic dissolution processes, when solute is added to water, water temperature increases.

34. A. KCl in H_2O, high D. H_2O in CH_3OH, high
    B. HF in H_2O, high E. NH_4Cl in C_5H_{12}, low
    C. KCl in CCl_4, low F. CCl_4 in H_2O, low
35. Increases, $LiSO_4$, $CaSO_4$, $Ca(OH)_2$.
36. NaOH (exothermic), NH_4NO_3 (endothermic).
37. Decreases
38. To control pressure
39. Oxygen is compressible whereas sugar in water is not
40. Al^{3+}, highly charged
41. F^-, small size

Part III: Work Out Problems

42. 0.886 g in 100 mL at 5.5 atm
43.
    A) 0.51 M
    B) 2 M
44. 5.160 g
45. 835 mg
46. 0.36, 0.64
47. 0.012
48. 0.23, 17 m

49. 375 mL
50. 0.01
51. 0.64 M
52. 0.796 m

Resources

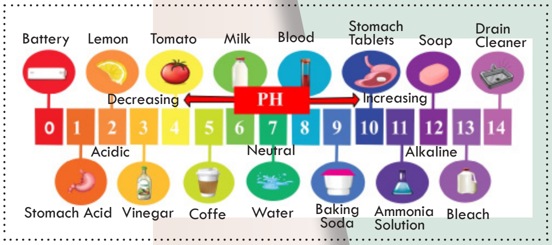
The current Ethiopian Grade 10 Chemistry Textbook for face-to-face modality.

Silberberg, Martin S. Principles of General Chemistry-3rd ed. McGraw-Hill. 2013.

Chang, R., Overby, J. General Chemistry-The Essential Concepts. 6th ed. McGraw-Hill. 2013.

Ebbing, D. D., Gammon, S. D. General Chemistry Eleventh Edition, CENGAGE Learning 2015.

UNIT 3



IMPORTANT INORGANIC COMPOUNDS

- Dear student! Unit 3 is about important inorganic compounds. The term inorganic compound refers to all compounds that do not contain carbon except simpler compounds of carbon like oxides (eg. CO_2), carbonates (eg. $CaCO_3$) and carbides (eg. SiC). Inorganic compounds are compounds consisting of mineral constituents of the earth or generally found in nonliving things. Compounds found in living things are generally organic compounds (see module 2 for the details). The majority of metal compounds are inorganic.

Inorganic compounds are mostly found in nature as silicates, oxides, carbonates, sulphides, sulphates, chlorides, nitrates, etc.

Inorganic compounds are generally classified into four groups namely oxides, acids, bases and salts.

We are familiar with the utility and importance of oxides, acids, bases and salts in our daily life. For example, we know that the carbon dioxide (CO_2) gas in CO_2 fire extinguisher is an oxide, the sour taste of vinegar is due to acetic acid (CH_3COOH), sodium hydroxide (NaOH) is a base used in the preparation of soaps, and sodium chloride (NaCl) is common salt which is useful as a food ingredient. This unit deals with the chemical nature and formation of more oxides, acids, bases and salts.

Section 3.1: deals with oxides. This section presents the classification of oxides as acidic, basic, amphoteric, neutral, and peroxides. It also describes the definitions, chemical properties and differences among the different classes of oxides.

Section 3.2: is about acids, it introduces the definition of acids. It also gives the classification of acids as monoprotic and polyprotic, binary and ternary, strong and weak, concentrated and dilute. Moreover, this unit presents the general properties of acids, safety precautions in handling acids, preparation of acids and common uses of some acids.

Section 3.3: is about bases. It introduces the definition, general properties, and safety precautions for handling of bases. It also presents the relationship between pH and pOH, preparation and uses of some bases.

Section 3.4: deals with salts. It presents the classification, preparation and properties of salts, and the uses of some important salts.

Unit Outcomes

Upon completion of this unit, you will be able to

- Classify inorganic compounds on the basis of their composition and/ or their chemistry;
- Discuss types of oxides and their chemical properties;
- Explain the Arrhenius concept of acids and bases;
- Mention the classification of acids and salts;
- Describe the general properties, preparation and uses of common acids, bases and salts;
- Distinguish the differences between strong and weak acids/ bases; and concentrated and dilute acids/ bases;

- Recognize the corrosive nature of acids and bases, and exercise the necessary precautions in handling and using them;
- Develop skills for identifying acidic, basic and neutral compounds.

Section 3.1: Oxides

? What are oxides?

Oxides are binary compounds containing oxygen and any other element (metal, non-metal or metalloid). Note that binary compounds are those consisting of only two elements.

Oxygen + Element (Metal, Non-metal or Metalloid) → Oxide

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. In this section, you will study about the types, properties, preparation and applications of oxides.

Learning Outcomes

Upon completion of this unit, you will be able to

- Define acidic oxides, basic oxides, amphoteric oxides & neutral oxides.
- Explain the chemical properties of acidic oxides, basic oxides, amphoteric oxides and neutral oxides.
- Differentiate basic oxides from acidic oxides.
- Compare and contrast acidic and basic oxides.
- Know the reactions of acid and basic anhydrides with water.
- Identify acid anyhdrides and basic anhydrides.
- Explain the salt-forming nature of acidic oxide, basic oxide and amphoteric oxide.
- Define neutral oxides and peroxides and give examples for each of them.
- Discuss the chemical properties of peroxides.
- Differentiate peroxides from other oxides.

3.1.1. Acidic and Basic Oxides

- Dear student! Most non-metals react with oxygen to produce acidic oxides. Most metals react with oxygen to produce basic oxides. In general, whether an oxide is acidic or basic depends on how electropositive its central atom is. The more electropositive the central atom the more basic the oxide. The more electronegative the central atom, the more acidic the oxide. Less electronegative elements tend to form basic oxides such as sodium oxide and magnesium oxide, whereas more electronegative elements tend to produce acidic oxides as seen with carbon dioxide and phosphorus pentoxide

A. Acidic Oxides

Acidic oxides are the oxides formed by the chemical combination of oxygen with non-metals (Groups 14 -17).

Acidic 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 . There are acidic oxides that do not react directly with water. The acids of such oxides are formed by other methods. For example; SiO_2 is not soluble in water, but it neutralizes basic oxides, thus it is acidic.

$SiO_2 + Na_2O \rightarrow Na_2SiO_3$ (sodium silicate)

The acid of SiO_2 (H_2SiO_3) is prepared by the reaction of HCl with Na_2SiO_3 as shown below:

$Na_2SiO_3\ (aq) + HCl\ (aq) \rightarrow H_2SiO_3\ (aq) + 2NaCl\ (aq)$

Chemical properties of acidic oxides

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

   Acid anhydride + Water → Acid

   $CO_2 + H_2O \rightarrow H_2CO_3$ (Carbonic 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

   $CO_2 + Na_2O \rightarrow Na_2CO_3$ (sodium carbonate)

   $SO_3 + CaO \rightarrow CaSO_4$ (calcium sulphate)

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

   Acidic oxide + Base → Salt + Water

   $SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$

   $CO_2 + 2LiOH \rightarrow Li_2CO_3 + H_2O$

Activity 3.1

1. List some examples of acidic oxides.
2. Give two examples of non-metallic oxides which are not acidic oxides.
3. Complete and balance the following equations.

a. $N_2O_5 + H_2O \rightarrow$

b. $P_4O_{10} + H_2O \rightarrow$

c. $MgO + CO_2 \rightarrow$

d. $Ca(OH)_2 + SO_3 \rightarrow$

e. $KOH + CO_2 \rightarrow$

f. $SrO + SO_2 \rightarrow$

B. Basic Oxides

Basic oxides are oxides that are composed of metals and oxygen. Most metals form oxides which exhibit basic properties and dissolve in water to give alkaline solutions. But, all metal oxides are not necessarily basic oxides; for example Al_2O_3 and ZnO are amphoteric oxides, which will be discussed in part (C). Oxides of metals that dissolve in water and react with it to form basic or alkaline solutions are called basic anhydrides. There are metallic oxides which have basic properties but are insoluble in water. These oxides react with acids to give salt and water

Example 3.1: $FeO + 2HCl \rightarrow FeCl_2 + H_2O$

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

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 3.2:

$$Li_2O + H_2O \rightarrow 2LiOH$$
$$CaO + H_2O \rightarrow Ca(OH)_2$$

2. Basic oxides react with acidic oxides to form salts.

Basic oxide + Acidic oxide → Salt

Examples 3.3:

$$BaO + SO_3 \rightarrow BaSO_4$$
$$Na_2O + CO_2 \rightarrow Na_2CO_3$$

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

Basic oxide + Acid → Salt + Water

Examples 3.4:

$$CaO + 2HCl \rightarrow CaCl_2 + H_2O$$
$$CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$$

Table 3.1: Some common indicators and the colour they develop in acidic and basic solutions

Indicator	Colour in aqueous solution of acidic oxide	Colour in aqueous solution of basic oxide
Universal Indicator	Yellow – orange (in weakly acidic) and red (in strongly acidic)	Blue (in weakly basic) and purple (in strongly basic)
Litmus	Red	Blue
Phenolphthalein	Colourless	Pink (red)
Methyl orange	Red	Yellow

Note that acidic and basic oxides can be identified either by their effects on indicators or chemical interaction they have with one another. Acidic oxides react with bases while basic oxides react with acids. But acidic oxides do not react with acids and basic oxides do not react with bases.

Experiment 3.1
Test for Acidity and Basicity of Oxides

Objective: To identify basic and acidic oxides.

Apparatus: Deflagrating spoon, gas jar and gas jar lid, test tubes.

Chemicals: Sulphur, magnesium or calcium metal, water, universal indicator, litmus paper (blue and red).

Procedure:

1. Take a small amount of powdered sulphur in a deflagrating spoon and heat it. As soon as sulphur starts burning, introduce the spoon into a gas jar. Cover the gas jar with a lid to ensure that the gas produced does not escape. Remove the spoon after some time. Add 10 mL of water into the gas jar and quickly replace the lid. Shake the gas jar well. Take two test tubes and pour 5 mL of the solution to each test tube. Add a few drops of universal indicator solution to the first test tube and blue litmus paper in the second.
2. Ignite a small amount of magnesium or calcium metal on a deflagrating spoon and insert in to a gas jar. Add 10 mL water to the ash formed and shake. Take two test tubes and pour 5 mL of the solution to each of the test tubes. Add a few drops of universal indicator in the first and red litmus paper to the second test tube.

Questions

a. What compounds are formed by the combustion of sulphur and magnesium or calcium? Write chemical equations to show the reactions.
b. What happens when water is added to the gas jars in which sulphur was burnt?
c. What colours are observed by adding drops of universal indicator and blue or red litmus to the solutions in the test tubes?
d. Why does the change in the colour of indicators occur in the various solutions? Write a laboratory report and submit to your teacher.

Note: Universal indicator and litmus paper serve as indicators. Indicators are substances used to identify whether a given solution is acidic or basic by showing colour changes.

Activity 3.2

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

3.1.2. Amphoteric, Neutral and Peroxides

Oxides are chemical compounds containing a chemical element (metal or nonmetal) bonded to one or more oxygen atoms. Concerning their properties, the key difference between neutral and amphoteric oxides is that the neutral oxides are neither acidic nor basic nature, whereas the amphoteric oxides are both acidic and basic.

Peroxides are compounds containing oxygen-oxygen (O–O) single bond or peroxide ion O_2^{2-}. Peroxides are chemically different from normal oxides. Normal oxides are oxides that contain oxide ion, O^{2-}. Peroxides fall under oxide category as they are binary compounds containing oxygen.

A. Amphoteric Oxides

An oxide which shows both acidic and basic behavior is called amphoteric oxide (see ***Figure 3.1***).

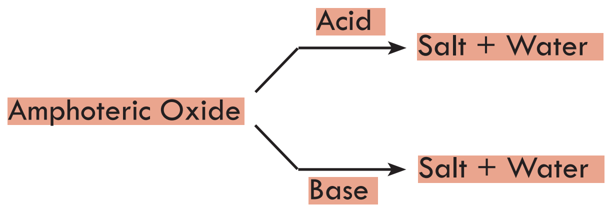
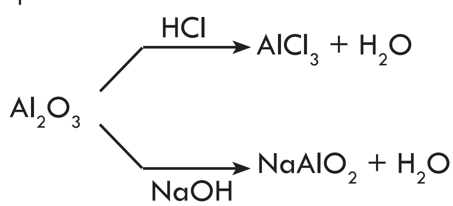


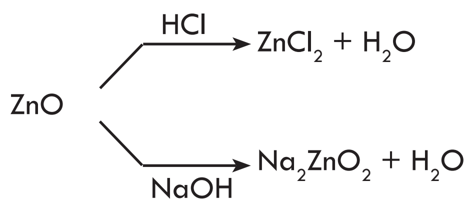
Figure 3.1 The reaction of amphoteric oxide with acids and bases.

For example, oxides of aluminium (Al_2O_3) and zinc (ZnO) are amphoteric. These oxides react with strong bases as well as strong acids.

A. Reactions showing amphoteric nature of aluminium oxide (Al_2O_3)



B. Reactions showing amphoteric nature of zinc oxide (ZnO)



Some other examples of amphoteric oxides are PbO, PbO_2, SnO, and SnO_2. It is also important to realize that hydroxides which react with both acids and bases are described as amphoteric hydroxides. For example, aluminium hydroxide, $Al(OH)_3$, reacts with both acids and bases to form salt and water. So, $Al(OH)_3$, is amphoteric in nature.

What is the common characteristic of acidic, basic and amphoteric oxides? Acidic oxides form salts when reacted with basic oxides and bases. Basic oxides also produce salts in their reactions with acidic oxides and acids. Amphoteric oxides form salts when they react with acids and bases. Thus, acidic oxides, basic oxides and amphoteric oxides are salt-forming oxides.

Experiment 3.2

Investigating Amphoteric Behaviour of Oxides

Objective: To observe the amphoteric behaviour of Al_2O_3.

Apparatus: Spatula, reagent bottles, beakers, and glass rod.

Chemicals: Al_2O_3, HCl, NaOH, Universal indicator and water.

Procedure:

1. Mix 20 mL concentrated HCl and 80 mL water in one reagent bottle;

2. Dissolve 8 g NaOH in 100 mL water in another reagent bottle.
3. Add universal indicator to the acid and base, and observe the colour change.
4. Take two beakers and place a spatula full of Al_2O_3 in each of the beakers.
5. Pour the HCl solution (which you prepared) into one of the beakers and NaOH solution into the other. Stir the mixture with a glass rod.
6. Add universal indicator in the solutions of beakers and observe the colour change.

Questions

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

Write a laboratory report and present to the class.

Activity 3.3

1. Write the chemical equations to show the amphoteric properties of, Al_2O_3 and PbO when they react with:
   a. NaOH
   b. HNO_3
   c. KOH
   d. H_2SO_4

B. 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. Neutral oxides are not salt- forming oxides.

C. Peroxides

Compounds which contain oxygen with oxidation number -1 are called peroxides. In peroxides, the two oxygen atoms are linked to each other and with atoms of other elements. Thus, peroxides contain the peroxide, "– O – O –" link. Some examples of peroxides are hydrogen peroxide, H_2O_2, sodium peroxide, Na_2O_2, calcium peroxide, CaO_2, barium peroxide, BaO_2, and strontium peroxide, SrO_2.

Most peroxides of metals are formed by burning the metals in a sufficient amount of oxygen.

Example 3.5:

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

There are many organic peroxides (ROOR') which are important activators in a process called polymerization, which creates plastic polymers.

Chemical Properties of Peroxides

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

Example 3.6:

$$PbS\,(s) + 4H_2O_2\,(aq) \rightarrow PbSO_4\,(s) + 4H_2O\,(l)$$

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

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

Example 3.7:

$$Na_2O_2\ (s) + 2HCl\ (aq) \rightarrow 2NaCl\ (aq) + H_2O_2\ (aq)$$
$$CaO2\ (s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (s) + H_2O_2\ (aq)$$

Hydrogen Peroxide

Hydrogen peroxide, H_2O_2, is a colorless liquid whose solutions are usually used as bleach and an antiseptic. Hydrogen peroxide decomposes to release oxygen. This reaction is slow but can be speed up by the addition of catalysts like MnO_2 or Pt.

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

Hydrogen peroxide is a strong oxidizing agent. Its oxidizing power is responsible for its effectiveness as an antiseptic for mouthwash and cleansing wounds. When hydrogen peroxide is added to a colored dye, the molecule responsible for the color will oxidize and so the color will disappear, showing its bleaching action. For example, if hydrogen peroxide is added to a black dye (paint) that contains lead sulphide, PbS, the black colour turns white. This is due to the oxidation of PbS to $PbSO_4$.

The equation for this process is:

$$PbS\ (s) + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O\ (l)$$

Activity 3.4

a. Oxides contain O^{2-} ions or the group –O– or =O. What is the oxidation state of oxygen in oxides?

b. What is the oxidation state of oxygen in Al_2O_3, N_2O_5, P_4O_{10}, KO_2 and SO_2?

Checklist - 3.1

The following checklists are provided to you as a guide to check what you have learned in this unit. Check you understanding by putting a "✓" mark if you can remember what is presented in relation to the terms. If you don't remember, please go back to the corresponding section and make a quick review. I can ...

SN	Competencies	Check
1	define and give examples of amphoteric & neutral oxides.	
2	explain the properties of amphoteric oxides and neutral oxides.	
3	differenciate normal oxides and peroxides.	
4	explain the chemical properties of peroxides.	
5	define acidic and basic oxides.	
6	explain the properties of acidic oxides and basic oxides.	
7	differenciate acid anyhdrides and basic anhydrides.	
8	explain the reactions of acid and basic anhydrides with water?	

Self-Test Exercise 3.1

1. Classify the following oxides as neutral, amphoteric and peroxides
   a. NO b. ZnO c. CO d. Na_2O_2 e. Al_2O_3
2. Complete and write the balanced equation for each of the following
   a. $Al_2O_3 + HNO_3 \rightarrow$
   b. $NaOH + Al_2O_3 \rightarrow$
   c. $ZnO + H_2SO_4 \rightarrow$
3. How peroxides differ from oxides? Write the structures of H_2O_2, Na_2O_2, CaO_2
4. How can you identify whether an oxide is amphoteric or neutral?
5. Write an acidic oxide that should react with water to form each of the following acids?
   a. H_2SO_3 b. H_2SO_4 c. HNO_3 d. H_2CO_3 e. H_3PO_3 f. H_3PO_4
6. Identify the basic anhydrides that react with water to form each of the following bases:
   a. $Ca(OH)_2$ b. KOH c. $Mg(OH)_2$ d. NaOH e. $Sr(OH)_2$ f. $Ba(OH)_2$
7. Complete and write the balanced equation for each of the following
   a. $MgO + H_2SO_4 \rightarrow$
   b. $CaO + H_3PO_4 \rightarrow$
   c. $K_2O + H_2SO_4 \rightarrow$
   d. $NaOH + CO_2 \rightarrow$
8. How can you identify whether an oxide is acidic or basic?

Section 3.2: Acids

- Acids play key roles in our bodies, in our homes, and in the laboratories. We seem to like the sour taste of acids; we add them to salad dressings, and spices. Many foods, including citrus fruits and some vegetables, contain acids.

Acetic acid in vinegar, citric acid in lemons and other citrus fruits, are among the acids that we encounter every day. Hydrochloric acid is the acid in gastric juice; it is essential to digestion. Phosphoric acid gives flavour to many carbonated beverages. In Latin the word acidus is used for sour. In the study of chemistry the term acid has been used to name a group of compounds which have a sour taste.

Learning Outcomes

Up on completion of this section, you will be able to

- Define acids in terms of the concept of Arrhenius.
- Give examples of acids based on Arrhenius.
- Categorize acids as monoprotic and poly protic based on the number of ionizable(replaceable) hydrogen atom.
- Group acids as binary and ternary based on the number of elements they contain
- Define strong and weak acids.
- Differentiate between strong and weak acids.

- Define concentrated and dilute acids.
- Describe the conceptual difference between strong and concentrated acids.
- Explain the conceptual difference between weak and dilute acids.
- Apply the necessary precautions while working with acids;
- Define pH.
- Describe the pH scale.
- Identify a given pH-labelled solution as acidic, basic or neutral.
- Perform activities to determine the pH of some common substances
- Using universal indicators or pH meter.
- Calculate the pH of a given acidic solution.
- Calculate the concentration of hydrogen ion from the given information.
- Discuss the direct combination of elements, the reaction of acidic oxides with water and formation of volatile acids from non -volatile acids as the three methods of preparation of acids.
- Describe the uses of the three common laboratory acids.
- Explain the general properties of acids.

3.2.1. Definition and Properties of Acids

Acids are a special group of compounds with a set of common properties. This helps to distinguish them from other compounds. Acids are solutions that are both corrosive and reactive, especially with metals. How can you define acids? What are the common properties of acids?

Arrhenius Definition of Acids

Take a look at all of the following chemical equations. What do you notice about them? What is common for each of the equations below?

Hydrochloric acid: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
Nitric acid: $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
Perchloric acid: $HClO_4(aq) \rightarrow H^+(aq) + ClO_4^-(aq)$

One of the distinguishable features about acids is the fact that acids produce H+ ions in solution. If you notice in all of the above chemical equations, all of the compounds dissociated to produce H^+ ions. This is the one main, distinguishable characteristic of acids for the Arrhenius definition of acids. Arrhenius defined an acid as a substance that releases hydrogen ion or proton, H^+, or hydronium ion, H_3O^+, in aqueous solution.

☞ H^+ in aqueous (water) solution will be hydrated and exist as H_3O^+

$$H^+ + H_2O = H_3O^+$$

In general, the ionization of acids takes place as follows.

$$HA(aq) \rightarrow H^+(aq) + A^-(aq) \text{ or } HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$$

Some more examples of Arrhenius acids are H_2SO_4, H_3PO_4, HBr, HI and HF.
All of these compounds are acids according to the Arrhenius definition because all produce H^+ ions in water.

$H_2SO_4 \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$ $H_3PO_4 \leftrightarrow 3H^+ (aq) + PO_4^{3-}(aq)$

$HBr \rightarrow H^+ (aq) + Br\text{-}(aq)$ $HI \rightarrow H^+ (aq) + I^-(aq)$

$HF \leftrightarrow H^+ (aq) + F\text{-}(aq)$

Note the double arrow (↔). The double arrow indicates the dissociation is reversible.

Activity 3.5

Show the dissociation of the following Arrhenius Acids.

$HClO_4$, H_2S, CH_3COOH, HNO_3

General Properties of Acids

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 in lemon juice and orange juice is harmless. However, concentrated acids are corrosive and poisonous. So the test 'taste' is never allowed to identify acids especially in the laboratories.
2. Acids change the color of certain acid-base indicators. The common indicators available in high school laboratories are litmus, phenolphthalein, methyl orange or methyl red and universal indicator. Note that natural dyes like red extract of cabbage leaves, petals of some colored flowers and turmeric also act as indicators to test the presence of acid or base.

Experiment 3.3

Effect of Acids on Indicators

Objective: To detect the acidity of a solution using indicators.

Chemicals: Lemon juice, dilute HCl, dilute HNO_3, dilute H_2SO_4 phenolphthalein, litmus, methyl red, universal indicator

Apparatus: Test tubes, test tube rack, test tube holder, and reagent bottles.

Procedure:

Take four clean test tubes and place some lemon juice in the first, dilute HCl in the second, dilute HNO_3 in the third and dilute H_2SO_4 in the fourth. Dip a strip of blue litmus paper into each of the four test tubes and observe. Follow the same procedure and repeat the experiment until each acid has been tested by each indicator. Record your observation.

Questions

1. What colours have you observed when each indicator was added to each of the four acid solutions? Record the colour change of each indicator in each solution (lemon juice, dilute HCl, HNO_3 and H_2SO_4) using the following table.
2. Why do all the acids show similar color changes in the same indicator?
3. What can you say about the acid-base behavior of indicators? Are indicators acidic or basic ?

Table 3.2 Colour of indicators in lemon juice, dilute HCl, HNO_3 and H_2SO_4

Indicator	Colour of indicator in			
	Lemon juice	Dilute HNO_3	Dilute HCl	Dilute H_2SO_4

4. Acids react with active metals to yield hydrogen gas. Acids react explosively with metals like sodium, potassium and calcium. The reaction is very dangerous and should not be performed. However, dilute acids (HCl, H_2SO_4) react moderately with reactive metals like: Mg, Zn, Fe and Al to form their respective salts with the evolution of hydrogen gas.

Example 3.8:

$$H_2SO_4 + Zn \rightarrow ZnSO_4 + H_2$$
$$6HCl + 2Al \rightarrow 2AlCl_3 + 3H_2$$

Experiment 3.4

Investigating the Reactions of Metals with Dilute Acids

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

Chemicals: Dilute H_2SO_4, dilute HCl, zinc, magnesium and iron

Apparatus: Test tubes, test tube holder, test tube rack, burner, match, cork and spatula.

Procedure:

Take three test tubes and place a spatula-full of powdered zinc in the first, powdered magnesium in the second and iron filings in the third. Pour dilute HCl into each of the test tubes until the metals are completely covered by the acid. To test the type of gas evolved, cover one of the test tubes with a cork which has a glass delivery (jet tube as shown in ***Figure 3.2***). Bring a burning match stick close to the glass tube mouth, the gas burns with a pale blue flame with a characteristic sound. Repeat the experiment with dilute H_2SO_4 after placing each of the three metals in three different test tubes.

Questions:

A. What does the formation of bubbles indicate?
B. What sound do you hear when the burning splint is close to the mouth of the test tube?
C. Which gas is evolved during the reaction?
D. Which metal's reaction with dilute HCl or H_2SO_4 is the most violent?
E. Write a laboratory report on your observations and present to the class.

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:

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$
$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$$

5. Acids react with carbonates and hydrogen carbonates to form salt, water and carbon dioxide gas

Acid + Hydrogen carbonate → Salt + Water + Carbon dioxide

Acid + Carbonate → Salt + Water + Carbon dioxide

Example 3.9:

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

$HCl + NaHCO_3 \rightarrow NaCl + H_2O + CO_2$

Reactions of Acids with Carbonates and Hydrogen Carbonates

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

Chemicals: Dil. H_2SO_4, dil. HCl, Zinc, Magnesium, iron

Apparatus: Test tubes , test tube holder, test tube rack, burner , match, cork and spatula

Procedure:

1. Take three conical flasks and add powdered Na_2CO_3 in the first, powdered $CaCO_3$ to the second and powdered sodium bicarbonate to the third. Pour dilute HCl into each of the three conical flasks until the acid covers the carbonates and bicarbonate. Hold damp blue litmus paper close to the mouth of the conical flasks. Repeat this with damp red litmus paper and record your observations. Bubble the gas through limewater as shown in ***Figure 3.3***.
2. Repeat the experiment using the same carbonates and hydrogencarbonate with dilute H_2SO_4 and dilute HNO_3.

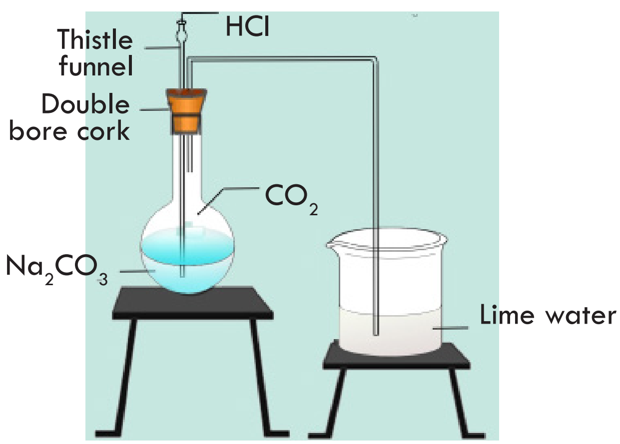


Figure 3.3 Test for Carbon dioxide.

Questions

A. What does the formation of bubbles in the conical flasks indicate?
B. Did the colour of the damp blue litmus paper change when you held it close to the mouth of the conical flasks? What about the colour of damp red litmus?
C. What happened to the colour of lime water when you bubbled the gas through it? If there was any change, what did it prove? Write a balanced chemical equation for the change?

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

6. Concentrated acids react with sulphites to form salts with the liberation of sulphur dioxide gas.

   Sulphite + Acid → Sulphur dioxide + Salt + Water

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

   $NaHSO_3 + HCl \rightarrow SO_2 + NaCl + H_2O$

7. Acids react with bases (oxides and hydroxides of metal and ammonium hydroxide) to form salts and water. The reaction of acids with basic oxides or bases to form salt and water is called neutralization reaction

Acid + Base (basic oxides or hydroxides of metal) → Salt + Water

Example 3.10: Reaction of acids with basic oxides

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

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

Example 3.11: Reaction of acids with hydroxides of metals

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

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

8. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great number of ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.

Example 3.12: A solution of hydrochloric acid (HCl) conducts electricity as it contains replaceable H atom which is furnished in solution as H+ ion and carries the current.

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

HCl is a strong electrolyte because it gives large amount of ions in aqueous solution. Strong electrolytes give bright light. CH_3COOH gives few ions in aqueous solution; hence it is a weak electrolyte. Weak electrolytes give dim light.

3.2.2. Classification and Preparation of Acids

Acids can be categorized in a number of ways. For example, acids can be classified as binary and ternary, monoprotic and polyprotic, strong and weak, organic and inorganic acids. The criterion of classification depends up on specific situation. For instance, if we need to discuss their strength we may classify acids as weak and strong, if we need to deal with the number of ionisable H^+ ions, we may classify as monoprotic, diprotic and polyprotic, etc.

The other important aspect of acids is their preparation.

Dear student! How can you prepare the required acids in the laboratories or industries? You need to know the methods of acid preparation in order to identify the appropriate one for the available materials and conditions. In this section, you will study classification of acids in different ways and the various techniques of acid preparation.

A. Classification of Acids Based on the Number of Ionizable Hydrogen

Acids are classified as monoprotic, diprotic and polyprotic depending upon the number of replaceable hydrogen ions (H^+) or hydronium ions (H_3O^+ ions) per molecule of the acid. An acid that contains only one ionisable hydrogen atom per molecule is called monoprotic.

Example 3.13: Hydrochloric acid (HCl) and nitric acid (HNO_3) are monoprotic acids because each give only one H^+ ion upon dissociation in aqueous solution.

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

Polyprotic acids are those acids containing more than one ionizable (replaceable) hydrogen ion in aqueous solution. Polyprotic acids which contain two ionizable hydrogen atoms such as H_2SO_4, H_2S, and H_2CO_3 are called diprotic acids; those containing three ionizable hydrogen atoms like H_3PO_4 are called triprotic acids. The dissociation of polyprotic acids in aqueous solution is shown in the following chemical equations:

$$H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2-}\ (aq)$$
$$H_2CO_3\ (aq) \rightleftharpoons 2H^+\ (aq) + CO_3^{2-}\ (aq)$$
$$H_3PO_4\ (aq) \rightleftharpoons 3H^+\ (aq) + PO_4^{3-}\ (aq)$$

B. Classification of Acids Based on the Number of Elements they contain

Acids can also be classified depending on the number of their constituent elements as binary and ternary acids. Binary acids are those acids composed of only two different elements. The two elements are hydrogen and a non-metal. Binary acids are generally represented as HX, where X is the nonmetal.

Examples of binary acids are HCl, HBr, HF, HI, and H_2S.

Ternary acids also called oxoacids are acids composed of three different elements. The three elements usually are hydrogen, oxygen and a non-metal. Ternary acids are generally represented as HOY, where Y is the nonmetal. Examples are H_2SO_4, H_2CO_3, $HClO_4$, and H_3PO_4.

Activity 3.6

1. Classify the following as monoprotic , diprotic or polyprotic
   $HClO_4$, H_2S, H_3PO_4, CH_3COOH, Cl_3COOH, H_2CO_3
2. Classify the following as binary or ternary (oxoacids)
   $HBrO_4$, HBr, HClO, HCl, H_2Se, HNO_3

Preparation of Acids

Acids can be prepared by:

1. The reaction of oxides of non-metals (acidic oxides) and water:

Acidic oxide + Water → Acid

Example 3.14: $N_2O_5\ (s) + H_2O\ (l) \rightarrow 2HNO_3\ (aq)$

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

2. Direct combination of some non-metals like S and Cl with hydrogen:

This method is mostly used to prepare binary acids (acids consisting only two elements

Nonmetal + Hydrogen → Binary acid

Example 3.15: $H_2\ (g) + Cl_2\ (g) \rightarrow 2HCl\ (g)$
$H_2(g) + S\ (s)\ ) \rightarrow H_2S\ (g)$

When gaseous hydrogen chloride and hydrogen sulfide dissolve in water, they form hydrochloric acid, and hydrosulfuric acid respectively.

3. Using a non-volatile acid: Volatile acids can be prepared by heating their salts with a non-volatile acid .For example, concentrated sulphuric acid (H_2SO_4) is a nonvolatile acid and it is used for the preparation of volatile acids Hydrochloric acid (HCl) and nitric acid (HNO_3) according to the following equations:

$$\underset{}{NaCl\ (s)} + \underset{\text{Nonvolatile acid}}{H_2SO_4(l)} \rightarrow NaHSO_4\ (s) + \underset{\text{Volatile acid}}{HCl\ (l)}$$

$$NaNO_3(s) + \underset{\text{Nonvolatile acid}}{H_2SO_4(l)} \rightarrow NaHSO_4(s) + \underset{\text{Volatile acid}}{HNO_3(l)}$$

Strength of Acids (Strong and Weak Acids)

Based on their degree of dissociation, acids are divided into strong acids and weak acids. If an acid completely dissociates in an aqueous solution, it is called a strong acid. Strong acids produce large amount of H^+ ions in the aqueous solution as they dissociate completely. Hydrochloric acid (HCl), nitric acid (HNO_3), and sulfuric acid (H_2SO_4) are some examples of acids that dissociate completely in aqueous solution. Their dissociation is indicated by a single arrow (→) as shown below:

$$HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$$
$$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$$
$$H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2-}\ (aq)$$

A dilute aqueous solution of strong acids contains predominantly the ions derived from the acids instead of the acid molecules. For example, HCl and HNO_3 are almost completely ionized in water.

An acid is called weak if it dissociates only partially in aqueous solution. Weak acids release few hydrogen ions in aqueous solution. The aqueous solution of a weak acid contains hydronium ions, anions and dissolved molecules of the acid. Examples of weak acids are acetic acid (CH_3COOH), oxalic acid ($H_2C_2O_4$), carbonic acid (H_2CO_3). These acids dissociate only partially as indicated by the double arrow (⇌) shown below:

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$$
$$H_2C_2O_4 \rightleftharpoons 2H^+ + C_2O_4^{2-}$$
$$H_2CO_3 \rightleftharpoons 2H^+ + CO_3^{2-}$$

Organic acids which contain the acidic carboxyl group – COOH, are generally weak acids.

Activity 3.7

1. a. Give examples of organic acids and inorganic acids. b. Classify your examples as strong and weak acids
2. Why do HCl, HNO_3, etc. show acidic characteristics in aqueous solutions while solutions of compounds like alcohol and glucose do not show acidic property?

Concentrated and Dilute Acids

The terms "concentrated" and "dilute" are used to describe the relative amount of acid present in a given solution. A concentrated acid has a relatively large amount of solute dissolved in the solvent. A dilute acid has a relatively smaller amount of solute dissolved in the solvent. Both strong acid and a weak acid may be concentrated or diluted depending on the number of moles of acid present.

For example, a dilute solution of H_2SO_4 contains 2 % H_2SO_4 and 98% water where as a concentrated solution of H_2SO_4 contains 98 % H_2SO_4 and 2 % water. A concentrated solution of CH_3COOH contains 96 % CH_3COOH and 4 % water whereas a dilute solution of CH_3COOH contains 4 % CH_3COOH and 96 % water (H_2O).

Note: H_2SO_4 is strong acid and CH_3COOH is weak acid.

pH and pH scale

The molar concentrations of H_3O^+ in aqueous solutions are generally very small and expressing acidity as the concentration of H_3O^+ becomes cumbersome. Instead, a more convenient quantity called pH is used as a measure of the acid strength of a solution. pH scale is used to rank solutions in terms of acidity or basicity (alkalinity). Concentration of hydrogen ion [H^+] in pure water is the basis for the pH scale. Water is a weak electrolyte because it ionizes very slightly into ions in a process called autoionization or self-ionization;

$$H_2O \rightleftharpoons H^+ + OH^-.$$

The ionic product constant of water 'Kw' is given as

$$Kw = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25\ °C$$

It is known that the amount of [H^+] and [OH^-] in pure water are equal

$$[H^+] = [OH^-] \text{ or } [H^+]^2 = 1.0 \times 10^{-14}$$
$$[H^+] = 1.0 \times 10^{-7} \text{ at } 25\ °C$$

As it is difficult to deal with such small figures having negative exponents, it is convenient to convert these figures into a positive figure using a numerical system. It is taking the common (base-10) logarithm of the figure and multiplying it with -1. 'p' before a symbol means' negative logarithm of the symbol. Therefore, pH is the negative logarithm of molar concentration of the hydrogen ions.

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

With reference to this equation, a scale develops according to the molar concentration of H+ ions that is called pH scale. It ranges from 0 to 14. According to this scale, pH of water is calculated as:

$$pH = -\log[H^+] = -\log(1x10^{-7}) = 7,$$

similarly $pOH = -\log[OH^-] = -\log(1x10^{-7}) = 7$

pH value normally varies from 0 to 14. Therefore:

$$pH + pOH = 14$$

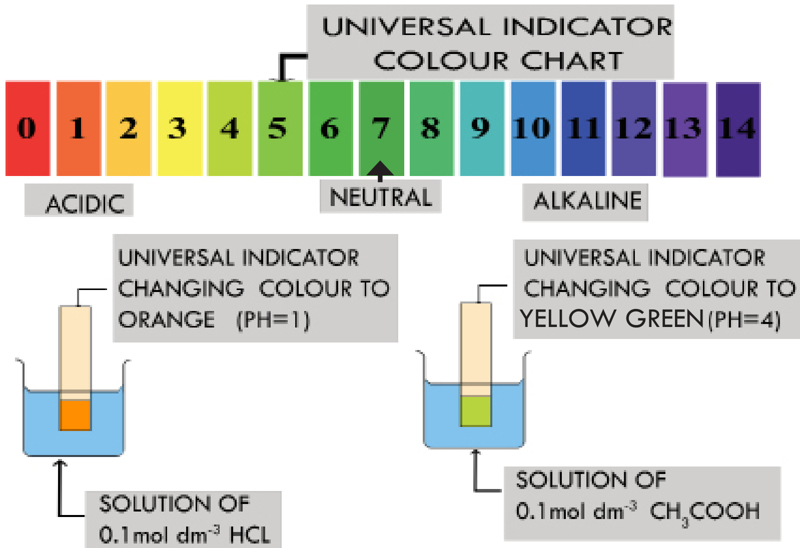


Figure 3.2. Measurement of pH using a universal indicator

Hence, the sum of the pH and pOH of the solution is always 14 at 25 °C. The pH of a solution can be measured using either a universal indicator solution or pH indicator paper or a pH-meter (see ***Figure 3.3***). When universal indicator solution or pH indicator papers are added to an acid solution, they develop different colors depending on the pH of the solution. In order to identify the pH, we need to compare the color developed with the standard color chart. A solution of a compound with pH 7 or pOH 7 is considered a neutral solution. Solutions of pH less than 7 are acidic and those more than 7 are basic. pH values of some common items are given in ***Figure 3.3*** as examples.

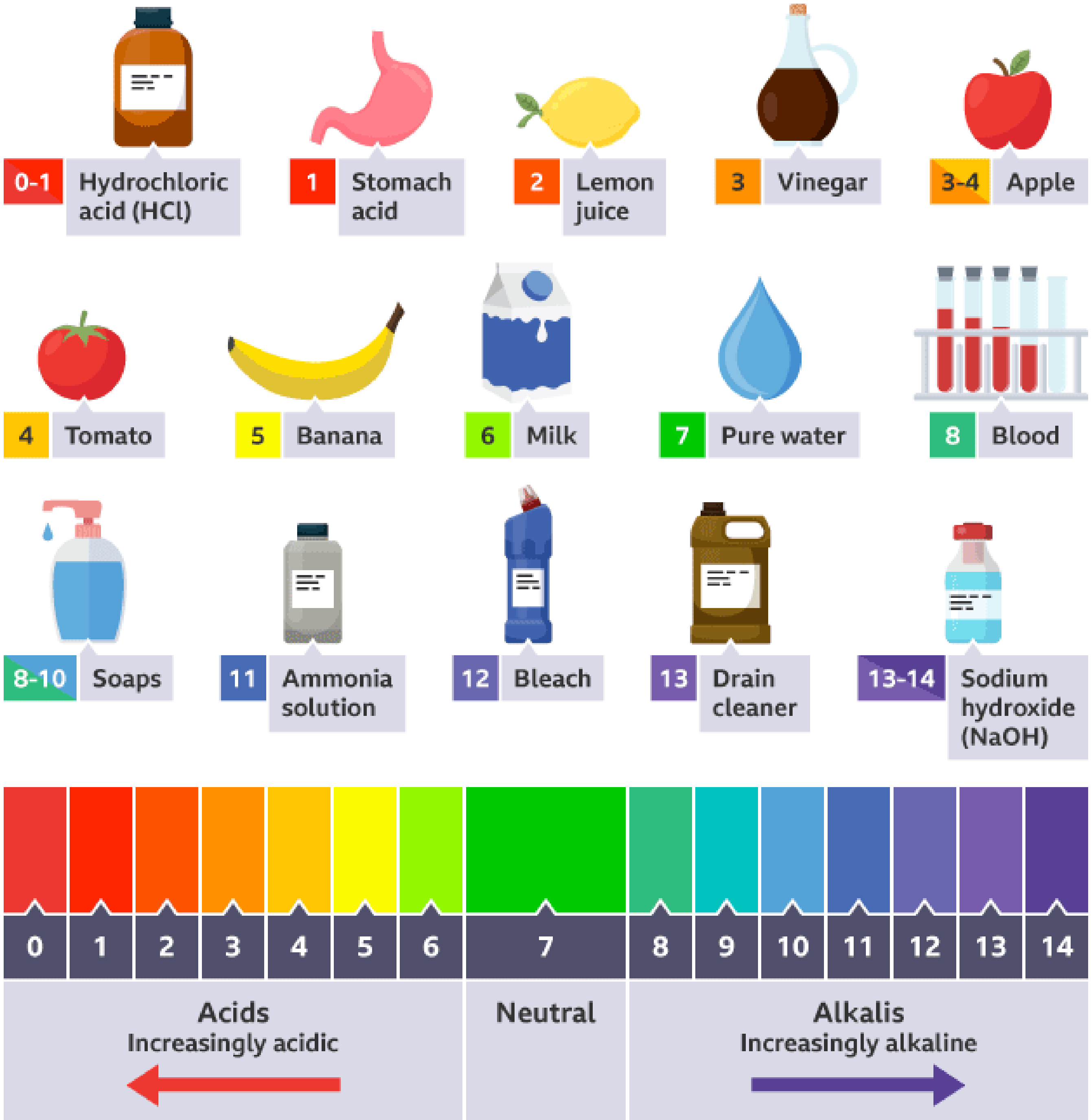


Figure 3.3. pH values of some common items.

As the pH value decreases the acidity of the solution increases. For example, if two solutions A and B have pH values of 4 and 6, respectively, then we can conclude that solution A is more acidic than solution B. The hydrogen ion concentration in solution A is one hundred times greater than that in B.

Example 3.16: Calculate the pH value of a solution when hydrogen ion concentration is 1.0×10^{-5} mo L^{-1}

Solution: pH is related to hydrogen ion concentration by an equation

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

It is given that $[H^+] = 1.0 \times 10^{-5}$ mol L^{-1} = 10^{-5} mol L^{-1}

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

$= -(-5 \log (10) = 5 \times 1 = 5$

Example 3.17: Calculate the hydrogen ion concentration of an aqueous solution when its pH is 9.

Solution: The concentration of hydrogen ion and the pH value of an aqueous solution are related by pH = – log $[H^+]$

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

$[H^+] = \text{antilog } (-pH) = 10^{-9} = 1.0 \times 10^{-9} \text{ mol L}^{-1}$

The Relationship Between pH and Concentration

Since the pH scale is logarithmic, a solution of pH 1 has 10 times higher concentration of [H+] than that of a solution of pH 2; 100 times than that of a solution of pH 3 and so on. Hence, low pH value means strong acid while high pH value means a strong base and vice versa.

Example 3.18: A solution of hydrochloric acid is 0.01M. What is its pH value?

Solution: Hydrochloric acid is a strong acid so it ionizes completely.

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

0.01 0.01 0.01

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

By putting the values of H^+ ions in the above equation:

$pH = -\log[H^+] = -\log(1.0 \times 10^{-2}) = 2$

Example 3.19: The hydrogen ion concentration in a dilute solution of citric acid is 1×10^{-6} M. What is the pH of the solution?

Solution: $pH = -\log[H^+] = -\log (1 \times 10^{-6}) = 6$

Example 3.20: Calculate the pH value of 0.05 M H_2SO_4. Assume that H_2SO_4 is completely ionized.

Solution: H_2SO_4 is a strong acid and ionizes as

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

0.05 2 × 0.05

Molarity of $H^+ = [H^+] = 0.1 \text{ mol L}^{-1}$

$pH = -\log(H^+) = -\log(0.1) = 1$

Experiment 3.6

pH of Solutions of Common Substances

Objective: To determine the pH of different substances

Chemicals: Lemon juice, vinegar, tonic water, tomato juice

Materials: Beakers, and universal indicator solution or pH indicator paper

Procedure:

Take four beakers and place lemon juice in the first, vinegar solution in the second, tonic water in the third and filtered tomato juice solution in the fourth. Then add a few drops of universal indicator solution or dip a piece of pH indicator paper into each of the solutions. Compare the colour developed with standard colour chart to decide the pH of each solution.

Questions:

a. What is your conclusion based on your observations?

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

c. Record your observations using the following Table:

Table 3.3 Colour and corresponding pH.

Substance	Color developed	pH
Lemon juice		
Vinegar solution		
Tomato Juice		
Tonic water		

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

Activity 3.8

1. Find the pH of 0.01M sulfuric acid?
2. HCl and H_2SO_4 are strong acids. Why do the equimolar solutions of the two acids have different pH?
3. Two solutions A and B have pH values of 2 and 6 respectively. How many times greater is the hydrogen ion concentration in solution A than that of solution B

Uses of Some Important Acids

Table 3.2 Uses of the common acids H_2SO_4, HCl and HNO_3

Acid	**Use**
Sulphuric acid, H_2SO_4	H_2SO_4 is the leading industrial chemical. It is used: * In production of sulphate and phosphate fertilizers, synthetic, fibers, paints, drugs, detergents, paper and dyes * In petroleum refining * In production of metals * As electrolyte in car batteries
Hydrochloric Acid, HCl	* It is present naturally in the gastric juice of our body and helps in the digestion of food * Industrially, HCl is used for pickling of iron and steel (to remove surface impurities) before galvanizing and tin plating * Used to produce aniline dyes, drugs, photographic films, plastics like polyvinyl chloride (PVC) * Used to recover magnesium from sea water
Nitric Acid, HNO_3	Used industrially in the manufacture of * Explosives such as trinitrotoluene (TNT) and trinitroglycerine * Fertilizers such as KNO_3 and NH_4NO_3 * Rubber, chemicals, plastics, dyes and drugs.

Precautions while working with Acids

Concentrated acids are corrosive and may cause visible destruction or irreversible alterations to living tissue by chemical action at the site of contact. Hence, the following safety precautions should be taken when handling acids:

a. Wear goggles, gloves and a laboratory coat
b. If a concentrated acid is spilled or splashed on your body, first wash the affected part with running water and then with 10% Na_2CO_3 solution
c. To dilute concentrated acids, pour the concentrated acid in to water and not water in to the acid.
d. If corrosive acids are swallowed, administer weak bases such as $Mg(OH)_2$ or $Al(OH)_3$.
e. Use bellows to pipette acid instead of sucking using yours lips
f. If an acid enters your eye, wash with water repeatedly and then consult a doctor
g. If concentrated acid is spilled on to cloth, immediately wash it with running water

Checklist - 3.2

The following checklists are provided to you as a guide to check what you have learned in this unit. Check you understanding by putting a "✓" mark if you can remember what is presented in relation to the terms. If you don't remember, please go back to the corresponding section and make a quick review. I can ...

SN	Competencies	Check
1	classify acids as binary and ternary, monoprotic,diprotic or poly protic?	
2	list and explain the properties of acids.	
3	write the dissociation of acid in water.	
4	the equation for the reaction of acids with metals and bases.	
5	define & give examples of acids based on the Arrhenius concept.	
6	list and explain the properties of acids.	
7	write the dissociation of acid in water.	
8	write the equation for the reaction of acids with metals and bases.	

Self-Test Exercise 3.2

Answer the following questions appropriately

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

   a. $HBrO_4$ b. BCl_3 c. H_2S d. HF e. HBr f. PCl_5

2. Complete and write the balanced equation for each of the following

   a. $Al_2O_3 + HNO_3 \rightarrow$

   b. $NaOH + Al_2O_3 \rightarrow$

   c. $ZnO + H_2SO_4 \rightarrow$

3. Classify the following acids as monoprotic, diprotic or triprotic, binary or ternary, strong or weak acids. (give three answers for each type):

   a. $HClO_4$ b. H_2SO_4 c. H_2CO_3 d. H_2SO_3 e. HF f. H_2S g. HCl

   h. H_3PO_4 i. HNO_2 j. HCN k. HNO_3 l. CH_3COOH

4. 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 colorless.

   What method do you recommend to identify the acid and water?

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

6. 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

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

8. Identify the following solutions as acidic or basic, estimate $[H_3O^+]$ values for each, and rank them in order of increasing acidity:

   a. Saliva, pH =6.5

   b. Orange juice, pH = 3.7

   c. Pancreatic juice, pH = 7.9

   d. Wine, pH = 3.5

Section 3.3: BASES

- Dear learner! What do you know about bases? Bases are either oxides or hydroxides of metals, which are therefore called basic oxides and hydroxides respectively. They are of great importance in chemical industries and in our daily lives, directly or indirectly. For example, sodium hydroxide, NaOH is used in the production of soap, paper, textile etc.

Potassium hydroxide, KOH is used to produce soft soap, fertilizers etc. Calcium hydroxide, $Ca(OH)_2$, is used to manufacture mortar and bleaching powder, to remove soil acidity etc. In this unit you will study some of the details of bases.

Learning Outcomes

After completing this section, you should be able to

- Define bases in terms of the concepts of Arrhenius.
- Give examples of bases based on Arrhenius.
- Discuss the general properties of bases.
- Define strong and weak bases.
- Distinguish between strong and weak alkalis (soluble bases).
- Define concentrated and dilute alkalis.
- Distinguish between concentrated and dilute alkalis (soluble bases).
- Use the necessary precautions while working with bases.
- Define pOH.
- Show the mathematical relationship between pH and pOH.
- Calculate the pOH of a given basic solution.
- Calculate the concentration of hydroxide ion from the given information.
- Carry out activities to investigate some chemical properties of bases.
- Discuss the reaction of active metals with water, the reaction of basic oxides with water and double displacement reactions as the three methods of preparation of bases.
- Carry out simple experiments to prepare bases in laboratory.
- Explain the uses of the two common laboratory bases.

3.3.1: Definition and Properties of Bases

? Dear learner! Do you know that the toothpaste and soap you use to clean your teeth and body are bases?

If you look around the kitchen at your home, you can also find many things which are basic in nature. Bases feel soapy or slippery on the skin and they can turn certain dyes blue. An example of a base is sodium hydroxide. Basicity is measured on a scale called the pH scale. Let's study more about bases.

Arrhenius Definition of Bases

According to the Arrhenius theory, a base is a substance that produces a hydroxide ion (OH-) when it is dissolved in water. Soluble bases are called alkalis.

For example, sodium hydroxide and potassium hydroxide are bases because they produce OH- ion when dissolved in water.

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

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

Note that the ions produced by the base can come from either of two sources. Metal hydroxides, such as NaOH, KOH, and $Ba(OH)_2$, are ionic compounds that already contain ions and merely release those ions when they dissolve in water. Ammonia, however, is not ionic and contains no ions in its structure. Nonetheless, ammonia is a base because it undergoes a reaction with water when it dissolves, producing NH_4^+ and OH^- ions:

$$NH_3\ (aq) + H_2O\ (l) \rightleftharpoons NH_4OH(aq) \rightarrow NH_4^{+}\ (aq) + OH^{-}(aq)$$

The reaction of ammonia with water is a reversible process whose equilibrium strongly favors unreacted ammonia.

Precautions in Handling Bases

Strong bases like NaOH and KOH are known as caustic alkalis. The word "caustic" refers to a substance that can cause burning. Thus, it is very important to avoid contact of these bases with any part of our body or clothing. Not only strong bases but also weak bases are corrosive. For example, concentrated ammonia solution can cause blindness if splashed into the eye.

The following safety precautions are useful in handling bases in school

a. Wear eye goggles, gloves and a laboratory coat.
b. If bases are spilled on your working tables wipe the spillages immediately.
c. Whenever bases are splashed on your cloth, wash the affected part with running water.
d. If a base enters your eyes, wash with water repeatedly as first aid treatment and seek medical advice.
e. If a base is swallowed by accident, drink 1-2% acetic acid or lemon juice immediately.
f. Whenever bases come into contact with your skin, wash the affected part with plenty of water and then wash the affected part with a very dilute solution (about 1%) of a weak acid such as acetic acid.

Activity 3.9

Answer the following

1. The chemical formula of lime water is a) CaO b) $Ca(OH)_2$ c) $CaCO_3$ d) $CaCl_2$
2. The chemical formula of caustic potash is
   a) NaOH b) $Ca(OH)_2$ c) NH_4OH d) KOH
3. Write the name given to bases that are highly soluble in water. Give an example.
4. Identify each substance as an Arrhenius acid, an Arrhenius base or neither.
   a. NaOH b. C_2H_5OH c. H_3PO_4 d. $C_6H_{12}O_6$ e. HNO_2 f. $Ba(OH)_2$
5. Write the balanced chemical equation for the neutralization reaction between
   a. KOH and $H_2C_2O_4$.
   b. between $Sr(OH)_2$ and H_3PO_4.
   c. between HCl and $Fe(OH)_3$.
   d. H_2SO_4 and $Cr(OH)_3$.

General Properties of Bases

What is the taste of carbonated beverage, coffee and tonic water? How can it be differ from the taste of lemon juice or vinegar? why a dilute solution of sodium bicarbonate can be used to treat bee stings, whereas a dilute solution of acetic acid can be used to treat wasp stings ? What are the common properties of bases?

The general properties of bases are discussed below:

1. Alkalis have a soapy feeling and a bitter taste: Bases have a bitter taste; feel slippery like soap to the skin in dilute aqueous solutions. Strong bases such as NaOH and KOH are very corrosive and poisonous. So they should be neither brought into contact with the skin nor tasted.
2. Alkalis change the color of indicators: Aqueous solutions of bases turn the color of red litmus to blue, phenolphthalein to pink (red), methyl red to yellow and universal indicator blue (purple). This property is because of the OH^- ion given by the base.
3. Alkalis release hydroxide ion in aqueous solution: The characteristic properties of bases in aqueous solutions are due to the presence of the hydroxide ion, OH^- which they release on dissolution.

Example 3.21: $NaOH\ (aq) \rightarrow Na\ (aq)^+ + OH^-\ (aq)$

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

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

Example 3.22: Base + Acid (or Acidic Oxide) → Salt + Water

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

$2NaOH\ (aq) + CO_2\ (aq) \rightarrow Na_2CO_3\ (aq) + H_2O(l)$

5. Aqueous solutions of bases conduct electricity: Soluble bases are electrolytes. Solutions of strong bases are good conductors while solutions of weak bases are poor conductors.

pOH (note that 'p' is small, O and H are capitals!)

pOH is a measure of the concentration of hydroxide ions in an acidic or a basic solution. It is defined as the negative logarithm of hydroxide, OH– ion concentration.

$$pOH = -\log10[OH^-]$$

Relationship Between pH and pOH

The pH and pOH are related mathematically as pH + pOH = 14

This equation can be derived using the ionic product of water, Kw, as follows:

The ionic product of water, Kw is the product of the concentration of $[H^+]$ or $[H_3O^+]$ and $[OH^-]$.

Electrical conductivity measurements of pure water such as distilled water show that concentrations of H_3O^+ and OH- are 1.0×10^{-7} M at 250 C. Therefore, the ionic product of pure water is $Kw = [H^+]$ or $[H_3O_+] \times [OH^-]$

$$= (1.0 \times 10^{-7}) \times (1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

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

Taking the negative logarithm of both sides of the equation $-\log10\{[H^+][OH^-]\} = -\log1.0 \times 10^{-14}$

We get $\{-\log10\ [H^+]\} + \{-\log10\ [OH^-]\} = 14$

Hence, pH + pOH = 14

Example 3.23: Find out the pH and pOH of a 0.0001M solution of NaOH?

Solution: Sodium hydroxide solution is a strong base. It ionizes completely such that one mole of NaOH gives one mole of OH ions.

$$NaOH(aq) \rightarrow Na^{+}(aq) + OH^{-}$$
$$0.0001M \quad 0.0001M \quad 0.0001M$$

Therefore, 0.0001M solution of NaOH produces 0.0001M OH ions

$$[OH^{-}] = 0.0001M = 1 \times 10^{-4}$$
$$pOH = -\log[OH^{-}] = -\log 1 \times 10^{-4} = 4$$
$$pH = 14 - 4 = 10$$

3.3.2. Classification and Preparation of Bases

Bases are usually classified on the basis of strength and concentration as strong and weak, dilute and concentrated bases.

How can you prepare the required bases in the laboratories or industries? You need to know the methods of base preparation in order to identify the appropriate one for the available materials and conditions. In this section, you will study classification of bases and the various techniques of base preparation

Strength of Bases (Strong and Weak Bases)

A strong base is almost completely ionized in aqueous solution to give its constituent cations and anions. That is the extent of ionization of a strong base is very high. Thus, the amount of hydroxide ions (OH^{-}) in the solution of a strong base is large.

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

Hydroxides of alkali (Group IA) metals such as LiOH, NaOH, KOH, and lower members of alkaline earth metals like $Ba(OH)_2$ are examples of strong bases. Strong bases are strong electrolytes as they produce large amount of ions and conduct high current.

A weak base is only partially ionized into its cations and anions in aqueous solution. That is, the extent of ionization of a weak base is small. Thus, the amount of hydroxide ions (OH^{-}) in the solution of a strong base is small.

$$NH_4OH(aq) \rightleftharpoons NH_4^{+}(aq) + OH^{-}(aq)$$

The double arrow shows that the dissociation of a weak base does not proceed to completion. This means an aqueous solution of a weak base contains only a small amount of ions derived from the dissociation of the base and a large amount of the non-ionized base. For example, a solution of ammonia containing 0.1 mole NH_3 per litre of solution ionizes only to the extent of 1.3%. Weak bases are weak electrolytes as they allow low current. Examples of weak bases include $Mg(OH)_2$ and $Ca(OH)_2$.

Concentrated and Dilute Bases

The concentration of an alkali is a measure of the number of moles of the alkali dissolved in 1 litre of solution and is therefore expressed in mol per liter. Concentrated bases contain relatively large amounts of a base in a given volume of solution while dilute base solutions contain only a small amount of base. The concentration is expressed in terms of mole per litre (Molarity).

The greater the number of moles of the base per liter of the solution, the more concentrated is the solution. For example, a solution containing ten moles of NaOH per litre is more concentrated than the solution containing two moles of NaOH per liter. The latter solution is more dilute than the former. Thus, both a strong base and a weak base may be concentrated or dilute depending on the number of moles of the base present per liter.

Preparation of Bases

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

A. By the reaction of highly reactive metals from Group IA or Group IIA (below magnesium) with water. This reaction produces the metal hydroxide with the liberation of hydrogen gas

$$2Li\ (s) + 2H_2O\ (l) \rightarrow 2LiOH\ (aq) + H_2\ (g)$$
$$2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$$
$$Ca\ (s) + 2H_2O\ (l) \rightarrow Ca(OH)_2\ (aq) + H_2\ (g)$$

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

$$\text{Metal oxide} + \text{Water} \rightarrow \text{Metal hydroxide}$$
$$Na_2O\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq)$$
$$BaO\ (s) + 2H_2O\ (l) \rightarrow Ba(OH)_2\ (aq)$$

C. By double displacement reaction

The reaction of an aqueous solution of a soluble base and a soluble salt, gives another soluble base and an insoluble salt as products

$$Ba(OH)_2\ (aq) + K_2SO_4\ (aq) \rightarrow 2KOH\ (aq) + BaSO_4\ (s)$$
$$Ca(OH)_2\ (aq) + Na_2CO_3\ (aq) \rightarrow 2NaOH\ (aq) + CaCO_3\ (s)$$

Common uses of NaOH and $Ca(OH)_2$

Sodium hydroxide, NaOH (lye or caustic soda) is used in the manufacture of soaps and detergents, degreasers and as the main ingredient in oven and drain cleaners.

Calcium hydroxide, $Ca(OH)_2$ (slaked lime) is used in the manufacture of cement and lime water. It is usually added to neutralize acidic soil. $Ca(OH)_2$ is used also for lime water test for carbon dioxide.

Experiment 3.7

Chemical Behavior of Bases

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

Chemicals: NaOH or KOH, $Ca(OH)_2$, HNO_3, water, blue and red litmus papers, cobalt chloride

Apparatus: Test-tube, test-tube rack, test-tube holder, Bunsen burner, three beakers, dropper, measuring cylinder, glass rod and watch glass.

Procedure:

1. Place 4.5 g NaOH or KOH in one test-tube and the same amount of $Ca(OH)_2$ in

another test-tube. Heat the test-tube containing NaOH or KOH gently using a Bunsen burner, by holding the test tube with a test tube holder. Hold the cobalt chloride paper partly inserted in the test tube. See whether the cobalt chloride paper shows a colour change or not. Repeat the same experiment with the second test tube that contains $Ca(OH)_2$.

2. Dissolve 3.6 g KOH in distilled water to prepare 100 mL solution in one beaker. Dilute 2 mL concentrated HNO_3 to make a 50 mL solution in another beaker. Add 10 mL KOH solution and 9.5 mL HNO_3 to the third beaker, stir thoroughly and test with blue and red litmus papers. Continue adding HNO_3 using a dropper, one drop at a time stirring after each addition and checking with red and blue litmus until the blue remains blue and the red remains red.
3. Put 5 mL of the neutral solution on a watch glass and allow the water to evaporate until the next day.

Questions:

1. What type of reaction occurred between the KOH and HNO_3? Write the balanced chemical equation for the reaction.
2. What was left on the watch glass?
3. Which hydroxide was melted on heating? Which base was decomposed on heating to give metal oxide and water? How do you know this? Write the balanced chemical equation for the reaction. Write a laboratory report and present it to the class.

Checklist - 3.3

The following checklists are provided to you as a guide to check what you have learned in this unit. Check you understanding by putting a "✓" mark if you can remember what is presented in relation to the terms. If you don't remember, please go back to the corresponding section and make a quick review. I can ...

SN	Competencies	Check
1	define bases based on the Arreheniuos concept?	
2	list and explain the properties of bases?	
3	write the dissociation of base in water?	
4	write the equation for the reaction of bases with acids & acidic oxides?	
5	identify basic materials making use of their properties?	
6	classify bases based on their strength & concentration?	
7	list and explain the classes of bases?	
8	list the techniques of preparing bases?	
9	explain some of the uses of common bases	

Self-Test Exercise 3.2

1. Write a chemical equation for each of the following.
   a. Potassium hydroxide is added to water.
   b. Calcium hydroxide is added to water.
2. Which type of medicines is used for treating indigestion?
   a. antibiotic b. analgesic c. antacid d. antiseptic
3. Which of the following pairs is an incorrect match between the base and its common name?
   a. sodium hydroxide (NaOH)- caustic soda
   b. potassium hydroxide (KOH) – potash
   c. Magnessium hydroxide($Mg(OH)_2$)- milk of magnesia
   d. Calcium hydroxide ($Ca(OH)_2$)- quick lime
4. Bases that completely ionize in water and their solutions are good electrical conductors are known as -_______
   a. concentrated bases b. strong bases c. weak bases d. dilute base
5. Bases that ionize only partially in water and their solutions are poor electrical conductors are known as -_______
   a. concentrated bases b. strong bases c. weak bases d. dilute base
6. What are alkalis? Give an example.
7. List and describe the techniques utilized for the preparation of bases.
8. Which solution has a higher [H3O+] concentration, one with pH =5, or with pH =9? Which solution has the higher OH– concentration?
9. Give the pH and pOH for each solution with the following concentrations:
   a. $[H_3O^+] = 1.0 \times 10^{-5}$
   b. $[OH^-] = 1.0 \times 10^{-9}$ M
10. Give the hydroxide ion concentrations of solutions with the following values of pH. Which of the solutions is most acidic? Which is most basic?
    a. pH 13.0 b. pH 3.0 c. pH 8.0

Section 3.4: Salts

Dear Student! We are familiar with salt, particularly, sodium chloride, or table salt. We regularly use it to season and preserve our food. However, not everyone is familiar with other types of salts and their uses, for example in manufacturing products like fertilizers, dyes, and polyester fabrics.

Learning Outcomes

Up on completion of this section, you will be able to

- Define salts.
- Give examples of salts.
- Classify salts as acidic, basic and normal salts.
- Discuss the direct combination of elements, the reaction of acids with bases,

neutralization and the reaction between acids and metals as the methods of salt preparation.

- Carry out simple experiment to prepare a salt by neutralization.
- List some important salts.
- Explain the uses of some important salts.
- Discuss the properties of salts.
- Explain the chemical tests of some salts by conducting activities.

3.4.1: Definition, Classification & Preparation of Salts

? What are salts?

Salts are ionic compounds generally formed by the neutralization of an acid with a base. Salts may be classified as acid, basic and normal salts. There are different methods utilized for the preparation of salts. In this topic, you will study more about the definition, properties and preparation of salts.

Objectives

Up on completion of this subsection, you will be able to

- Define salts.
- Give examples of salts.
- Classify salts as acidic, basic and normal salts.
- Discuss the direct combination of elements, the reaction of acids with bases, neutralization and the reaction between acids and metals as the methods of salt preparation.

Definition and Classification of Salts

Salts are ionic compounds generally formed by the neutralization of an acid with a base. A salt gets its name from the names of the metal ion (derived from the base) and the acid radical derived from the acid. For example, NaCl is formed from Na^+ derived from NaOH and Cl- is derived from HCl. Hence, NaCl (sodium cloride) is a salt of NaOH and HCl. Similarly, KNO_3 (potasium nitrate) is a salt of KOH and HNO_3. During neutralization reaction, the cation of the base combines with the anion of the acid and a salt is formed (while H^+ and OH^- combine to form H_2O) as shown below.

Salt formation

$$\underset{\text{Base}}{NaOH} + \underset{\text{Acid}}{HCl} \longrightarrow \underset{\text{Salt}}{NaCl} + \underset{\text{Water}}{H_2O}$$

Salt formation

$$\underset{\text{Base}}{KOH} + \underset{\text{Acid}}{HNO_3} \longrightarrow \underset{\text{Salt}}{KNO_3} + \underset{\text{Water}}{H_2O}$$

Salts are also defined as ionic compounds formed when the ionizable hydrogen of acids are partly or completely replaced by metal ions or ammonium ion.

Table 3.3. Examples of salts with their common names, chemical names & chemical formula

Common name	Chemical name	Chemical formula
Common Salt	Sodium chloride	$NaCl$
Washing Soda	Sodium carbonate	Na_2CO_3
Baking Soda	Sodium bicarbonate	$NaHCO_3$
Bleaching Powder	Calcium hypochlorite	$Ca(OCl)_2$

Activity 3.10

Write the base and acid pairs used to give each of the following salts

A. Na_2CO_3 B. $CaSO_4$ C. $KOCl$ D. $Ca(NO_3)_2$

Classification of Salts

Salts can be classified into three main groups namely: Normal salts, Basic salts and acid salts.

A. Normal Salts

A salt formed by the complete replacement of ionizable H^+ ions of an acid by a positive metal ion or NH_4^+ ions is called normal or neutral salt. These salts are neutral to litmus

$$HCl\ (aq) + KOH\ (aq) \rightarrow KCl\ (aq) + H_2O\ (l)$$
$$H_2SO_4(aq) + ZnO\ (aq) \rightarrow ZnSO_4(aq) + H_2O\ (l)$$
$$HNO_3(aq) + NH_4OH\ (aq) \rightarrow NH_4NO_3(aq) + H_2O\ (l)$$

B. Acidic salts

These salts are formed by the partial replacement of replaceable H^+ ions of an acid by a positive metal ion.

$$H_2SO_4\ (aq) + KOH(aq) \rightarrow KHSO_4(aq) + H_2O(l)$$
$$H_3PO_4(aq) + NaOH(aq) \rightarrow NaH_2PO_4(aq) + H_2O(l)$$

These salts $KHSO_4$ and NaH_2PO_4 turn blue litmus red. For example, when $KHSO_4$ dissolves in aqueous solution, it releases the ions K^+, H^+, SO_4^{2-}. It releases hydrogen ions hence it is acidic.

Acidic salts react with bases to form normal salts, K_2SO_4 and Na_3PO_4

$$KHSO_4(aq) + KOH\ (aq) \rightarrow K_2SO_4\ (aq) + H_2O\ (l)$$
$$NaH_2PO_4(aq) + 2NaOH\ (aq) \rightarrow Na_3PO_4(aq) + H_2O\ (l)$$

C. Basic Salts

These are salts in which not all of the hydroxide ions in a base have been replaced by the anions of the acid. Basic salts are formed by the incomplete neutralization of a polyhydroxy base by an acid.

$$Al(OH)_3\ (aq) + HCl\ (aq) \rightarrow Al(OH)_2Cl\ (aq) + 2H_2O(l)$$
$$Zn(OH)_2\ (aq) + HNO_3\ (aq) \rightarrow Zn(OH)NO_3(aq) + H_2O(l)$$

The above basic salts further react with acids to form normal salts.

$Al(OH)_2Cl\ (aq) + 2HCl\ (aq) \rightarrow AlCl_3\ (aq) + H_2O(l)$

$Zn(OH)NO_3(aq) + HNO_3(aq) \rightarrow Zn(NO_3)_2\ (aq) + H_2O(l)$

Activity 3.11

Identify each of the following as acid salt, basic or normal salt.

$NaCl$, $CaCO_3$, $Mg(H_2PO_4)_2$, $Zn(OH)Cl$, K_3PO_4, $NaHCO_3$

General Methods for the Preparation of Salts

How salts could be prepared? Usually soluble salts are prepared by methods that involve crystallization, while insoluble salts are prepared by methods that involve precipitation

1. Preparation of Soluble Salts

A. The reaction of an acid and a metal (Direct Displacement method)

This is direct displacement method in which hydrogen ion of acid is replaced by a reactive metal. Such as calcium, magnesium, zinc and iron,

e.g.

Active metal + Acid $\rightarrow$ Salt + Hydrogen

$Mg(s) + 2HCl\ (aq) \rightarrow MgCl_2(aq) + H_2(g)$

$Zn\ (s) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2\ (g)$

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

B. The reaction of an acid and a base (Neutralization method)

It is a neutralization reaction in which acid and base react to produce a salt and water

Acid + Base $\rightarrow$ Salt + Water

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

C. By the reaction of an acid and metallic oxide

Mostly the insoluble metallic oxides react with dilute acids to form salt and water

Metallic oxide + Dilute Acid $\rightarrow$ Salt + Water

$CuO(s) + H_2SO_4\ (aq) \rightarrow CuSO_4(aq) + H_2O(aq)$

$ZnO\ (s) + 2HNO_3\ (aq) \rightarrow Zn(NO_3)_2\ (aq) + H_2O\ (l)$

D. The reaction of an acid and a carbonate or bicarbonate:

Dilute acids react with metallic carbonates to produce salts, water and carbon dioxide gas.

Metal carbonates or Bicarbonates + Dilute acid $\rightarrow$ Salt + Water + Carbon dioxide

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

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

Preparation of Insoluble Salts by Double Decomposition (Precipitation)

In this method, usually solutions of soluble salts are mixed. During the reaction, exchange of ionic radicals (i.e., metallic radicals exchange with acidic radicals) takes place to produce two new salts. One of the salts is insoluble and the other is soluble. The insoluble salt precipitates (solidify in solution).

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$
$$Na_2CO_3(aq) + CuSO_4(aq) \rightarrow CuCO_3\ (s) + Na_2SO_4\ (aq)$$

Properties and Uses of Salts

? Dear student! What characteristics do salts have? Where are they utilized, and how? What sets them apart from other chemical substances?

You will learn about the typical characteristics of salts, some common applications, and the chemical tests used to determine the ions of salts under this topic.

Properties of Salts

Salts can be classified depending on the anion (negative ion) they possess, because the anion is partly responsible for the solubility of the salt.

1. Solubility of Salts

When salts are added to water or water is added to salt, different types of actions take place. Some salts are soluble in water; some are insoluble while some salts are only slightly soluble.

2. Tendency to absorb water from the atmosphere or release water to the atmosphere. Salts can be classified as hygroscopic, deliquescent and efflorescent depending on their tendency to absorb water from or release water to the atmosphere.
Hygroscopic salts are those which absorb water from the atmosphere but remain solid.
Example: Anhydrous copper (II) sulphate, $CuSO_4$

Deliquescent salts absorb water from the atmosphere to form a solution. The process of absorbing water from the atmosphere by a solid to form a solution is called deliquescence. Example: Calcium chloride($CaCl_2$), Sodium nitrate ($NaNO_3$), Iron(III) chloride ($FeCl_3$)

Efflorescent salts lose their water of crystallization to the atmosphere. The loss of water of crystallization by solid crystals to the atmosphere is known as efflorescence.

It is very important to note that all deliquescent substances are hygroscopic, but all hygroscopic substances are not necessarily deliquescent.

Example 3.24: Hydrated sodium carbonate ($Na_2CO_3 \bullet 10H_2O$), Hydrated sodium sulphate ($Na_2SO_4 \bullet 10H_2O$)

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

$$NaNO_3\ (aq) \rightarrow Na^+\ (aq) + NO_3^-\ (aq)$$
$$CaCl_2\ (aq) \rightarrow Ca^{2+}(aq) + 2Cl^-\ (aq)$$

4. Thermal stability of salts: Thermal stability of a salt is the property of a salt to resist irreversible change in its chemical or physical structure, often by resisting decomposition at a high relative temperature. When different salts containing the same anion are heated, they may not show similar behavior. Some salts are thermally stable while others

undergo decomposition. The following examples illustrate this fact.

A. Thermal Decomposition of Carbonates

All the carbonates of Group IIA undergo thermal decomposition to the metal oxide and carbon dioxide gas. The term "thermal decomposition" describes splitting up a compound by heating it. Carbonates are salts containing a carbonate as an anion.

If "M" represents any one of the elements, the following describes this decomposition:

$$MCO_3(s) \rightarrow MO(s) + CO_2$$

The carbonates of Group IIA are: $MgCO_3$, $CaCO_3$, $SrCO_3$, $BaCO_3$

The carbonates become more thermally stable down the group. In Group IA, lithium carbonate behaves in the same way, producing lithium oxide and carbon dioxide:

$$Li_2CO_3(s) \rightarrow Li_2O(s) + CO_2$$

The rest of the Group IA carbonates do not decompose at laboratory temperatures, although at higher temperatures this becomes possible. The decomposition temperatures again increase down the group.

Most transition metals decompose on heating to give the metal oxides and carbon dioxide:

Example 3.25:

$$CuCO_3\ (s) \rightarrow CuO\ (s) + CO_2(g)$$

B. Thermal Decomposition of Nitrates

Nitrates are salts containing a nitrate, NO_3^-, as an anion. The Group IIA and most transition metal nitrates undergo thermal decomposition to the metal oxide, nitrogen dioxide and oxygen gas.

Nitrate → Metal oxide + Nitrogen dioxide + Oxygen

Example 3.26: $2Mg(NO_3)_2(s) \rightarrow 2MgO(s) + 4NO_2\ (g) + O_2(g)$

$$2Pb(NO_3)_2\ (s) \rightarrow 2PbO\ (s) + 4NO_2\ (g) + O_2\ (g)$$

Nitrates of sodium and potassium decompose on heating to give nitrites (instead of the oxides) and oxygen gas;

$$2NaNO_3\ (s) \rightarrow 2NaNO_2\ (s) + O_2(g)$$
$$2KNO_3\ (s) \rightarrow 2KNO_2\ (s) + O_2\ (g)$$

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

$$4LiNO_3\ (s) \rightarrow 2Li_2O\ (s) + 4NO_2\ (s) + O_2\ (g)$$
$$Li_2CO_3\ (s) \rightarrow Li_2O\ (s) + CO_2\ (g)$$

Note that both the nitrates and carbonates of lithium are thermally unstable similar to Group IIA. This reveals that lithium behaves more like Group IIA rather than Group IA.

Table 3.4: Some important salts and their uses.

Salts	Uses
Sodium chloride (common salt, NaCl)	1. Preparation and preservation of food 2. Raw material for the manufacture of sodium, chlorine, and sodium hydroxide 3. Component of Oral Rehydration Salt (ORS) 4. Manufacture of baking soda ($NaHCO_3$) and Na_2CO_3 5. In making a freezing mixture which is used by ice cream vendors
Ammonium nitrate, (NH_4NO_3	Nitrogenous fertilizer and in explosives
Copper (II) sulphate, ($CuSO_4$)	1. Used to make Bordeaux mixture (mixture of $CuSO_4$ and $Ca(OH)_2$ and other fungicides. Bordeaux mixture is used to prevent fungal attack of leaves and vines 2. Useful in electroplating and dyeing
Iron (III) chloride, ($FeCl_3$)	1. Used in the treatment of waste water 2. Used for etching printed circuits
Potassium nitrate, (KNO_3)	1. Used in making gun powder ((mixture of KNO_3, carbon and sulphur) and other explosives 2. Used as fertilizer in agriculture
Calcium sulphate, ($CaSO_4.2H_2O$, Gypsum)	Used for plastering walls and supporting fractured bones.
Barium sulphate, ($BaSO_4$)	1. Given to patients as a "barium meal" before gastrointestinal x-ray photography 2. Used as a white pigment
Iron (II) sulphate ($FeSO_4$)	Given as iron tablets to patients who suffer from anemia

Chemical Tests of Some Ions in Salt

Flame tests are used to identify the presence of a relatively small number of metal ions in a compound. Certain metals give a characteristic color to a Bunsen flame when their solid salts or moist salts are heated directly in the flame. A flame test is commonly used to identify the presence of lithium, sodium, potassium, calcium, strontium and barium ions in salts.

Experiment 3.8

Test for Sulphates

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

Chemicals: Any soluble sulphate salt (such as sodium sulphate), barium chloride or barium nitrate solution, and dilute HCl.

Apparatus: Beakers, test tubes, and test tube rack,

Procedure: Add some sodium sulphate solution to a test tube and acidify the solution by adding a few drops of dilute HCl. Then add $BaCl_2$ or $Ba(NO_3)_2$ solution and note if a white precipitate is formed.

Questions:

1. Name the white precipitate formed?
2. Why is it necessary to add a few drops of dilute HCl?
3. Write a balanced chemical equation for the reaction.

Write a laboratory report and present to the class

Experiment 3.9

Test for Carbonates and Hydrogen Carbonates

Objective: To distinguish between carbonates and hydrogen carbonates

Chemicals: Na_2CO_3, $NaHCO_3$, dilute HCl, lime water

Apparatus: Conical flasks, Beaker

Procedure:

1. Take 20 mL solution of Na_2CO_3 and add it to one conical flask and 20 mL $NaHCO_3$ solution to another. Add the same amount of dilute HCl to each of the conical flasks; Fit a rubber stopper to which a delivery tube is inserted to each conical flask. Allow the gas produced to pass through lime water and observe the changes.
2. Again take Na_2CO_3 solution in one conical flask and $NaHCO_3$ solution in the other. Add the same amount of $CaCl_2$ solution to each of the conical flasks.

Questions:

1 What happened to the lime water in each case?

2. Which solution formed a white precipitate upon the addition of $CaCl_2$ solution?

3. Write balanced chemical equations for all the reactions.

Write a laboratory report and present it to the class.

Checklist - 3.4

The following checklists are provided to you as a guide to check what you have learned in this unit. Check you understanding by putting a "✓" mark if you can remember what is presented in relation to the terms. If you don't remember, please go back to the corresponding section and make a quick review. I can ...

SN	Competencies	No
1	define salt.	
2	give examples of salts with their common names.	
3	classify salts as acidic, basic and normal salts.	
4	discuss the various methods used for the of preparation of salts.	
5	discuss the properties of salts.	
6	explain the uses of some important salts.	
7	classify salts as acidic, basic and normal salts.	
8	explain the chemical tests of some salts.	

Self-Test Exercise 3.2

Answer the following questions appropriately

1. Identify the base and the acid used for the formation of each of the following salts
   a. Zinc sulphate ($ZnSO_4$)
   b. Calcium phosphate $Ca_3(PO_4)_2$
   c. Silver acetate (CH_3COOAg)
   d. Sodium carbonate (Na_2CO_3)
2. You are provided separate solutions of NaCl, HCl, NaOH and Phenolphtha¬lein (Marked 1). All the solutions are colorless. How would you recognize the solutions of acid, base and salt?
3. Suggest at least three methods for the preparation of salts.
4. Classify the following salts as soluble or insoluble in water:

a. NH_4Cl	e. AgNO	i. $CaSO_4$
b. K_3PO_4	f. $CaCl_2$	j. $BaCO_3$
c. $FeCO_3$	g. $PbSO_4$	
d. AgCl	h. Na_2S	

5. Why do aqueous solutions of soluble salts conduct electricity?
6. Which carbonates do not decompose on heating?
7. Nitrates mostly decompose by heat to give metal oxide, nitrogen dioxide and oxygen. Which nitrates do not give these products on heating?
8. What reagents do you use to identify the presence of the following ions in salts?
   a. Halide ions
   b. Fe^{2+} and Fe^{3+} ion
   c. Sulphate, SO_4^{2-} ions?

Unit Summary

Inorganic compounds are those compounds orginating 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. Acidic oxides produce acidic solution up on dissolution in water.

Basic oxides are oxides of metals. These metal oxides which dissolve in water are also called basic anhydrides.

Amphoteric oxides are those oxides which show the properties of both acids and bases.

Neutral oxides are those oxides which do not show basic or acidic properties.

Peroxides are oxides containing a peroxide (-O-O-) link and the oxidation state of oxygen is -1.

Arrhenius acids are substances that release hydrogen ions or protons in aqueous solution.

Arrhenius bases are substances that release hydroxide (OH^-) ions in aqueous solution.

Strong acids and strong bases ionize almost completely in aqueous solution.

Weak acids and weak bases ionize only slightly in aqueous solution.

pH is the negative logarithm of hydrogen ion concentration pH = $-\log [H^+]$.

pOH is the negative logarithm of hydroxide ion concentration pOH = $-\log[OH^-]$.

KW = $[H^+][OH^-] = 1 \times 10^{-14}$ at 25°C. pKw = pH^+ pOH = 14 at 25°C

Salt is an ionic compound containing a cation derived from a base and anion derived from an acid.

Salts may be classified as acidic, normal and basic salts. An acid salt is formed when ionizable hydrogen atoms of an acid are replaced partly by a metal ion or ammonium ion.

A normal salt is formed when all ionizable hydrogen atoms of an acid are completely replaced by a metal or ammonium ion.

A basic salt is a salt containing ionizable hydroxide ion.

Hygroscopic salt absorb water from the atmosphere.

Deliquescent salt absorb water from the atmosphere and dissolve in the ater absorbed to form solutions.

Self-Assessment Exercises

Part I. Multiple choice questions

1. Generally, metallic oxides are basic and non-metallic oxides are acidic in nature. Solution of which of the following oxides in water will change the colour of blue litmus to red?
   a. Magnesium oxide
   b. Phosphrous oxide
   c. Barium oxide
   d. Calcium oxide
2. A non-metallic oxide dissolves in water to give:
   a. alkali
   b. base
   c. Acid
   d. None
3. According to the Arrhenius concept, an acid is a substance that ________.
   a. is capable of donating one or more H^+
   b. causes an increase in the concentration of H^+ in aqueous solutions
   c. can accept a pair of electrons to form a coordinate covalent bond
   d. reacts with the solvent to form the cation formed by autoionization of that solven
4. Which one of the following is a diprotic acid?
   a. Hydrochloric acid
   b. Phosphoric acid
   c. Sulfuric acid
   d. Acetic acid
5. Which of the following is not the characteristic of an acid:
   a. An acid changes the color of an indicator
   b. An acid has a bitter taste
   c. An acid ionizes in water

d. None

6. Which of the following is true according the the Arrhenius theory of acids?
   a. A strong acid increases the OH^- ion concentration in an aqueous solution.
   b. A strong acid increases the H^+ ion concentration in an aqueous solution.
   c. A strong acid decreases the OH^- ion concentration in an aqueous solution.
   d. A strong acid decreases the H^+ ion concentration in an aqueous solution.
7. When an acid reacts with an active metal:
   a. Hydronium ion concentration increases
   b. Hydrogen gas is produced
   c. Metal forms anions
   d. Carbon dioxide gas is produced
8. The pH of a solution is 8. What is the pOH of the solution?
   a. 8
   b. 8×10^{-7}
   c. 6
   d. 2.1×10^{-6}
9. Which of the following solutions would have a pH value greater than 7?
   a. $OH^-] = 2.4 \times 10^{-2}$ M
   b. 0.0001 M HCl
   c. $[H_3O^+] = 1.53 \times 10^{-2}$ M
10. What is the chemical name for baking soda?
    a. Potassium Carbonate
    b. Potassium Hydrogen Carbonate
    c. Sodium Hydrogen Carbonate
    d. none

Part II. Fill in the blanks by using the correct words /terms given in the brackets

11. The pH value of an acid is ____ than 7. (greater/less)
12. A solution is said to have ____ character when its pH value is greater than seven (basic/acidic)
13. The pH value of a neutral solution is ____ at 298K. (Zero/7)
14. The pH of 10^{-6} is ___ than the pH of 10^{-4} HCl. (greater/less)
15. The pH of water decreases when ____ is dissolved in it. (NH_3/SO_2)
16. The hydrogen ion concentration of a solution with pH = 3 is ___ than the solution with pH = 8 (greater/less)
17. An acid produces ____ ions in aqueous solution. (Hydrogen/hydroxide)
18. The concentration of H+ ions given by HCl is ____ than that given by equal concentration of CH_3COOH (greater/less).
19. The binary compounds of metallic or non-metallic elements with oxygen are called as __oxdes_____

Part III. Short answer questions.

20. The hydrogen ion concentration in an acidic solution is 10^{-5} ? What is the pH value of the solution? pH= 5
21. An element combines with oxygen to form an oxide. This oxide dissolves in water. This aqueous solution changes blue litmus to red. Write:
    a. the nature of the aqueous solution (alkaline or acidic)

b. the nature of the element (metal or nonmetal)

22. What do you understand by the statement 'acetic acid' CH_3COOH, is a monoprotic acid?
23. Write the balanced equations for the reaction of zinc oxide with:
    a. hydrochloric acid
    b. aqueous sodium hydroxide
24. How many times will the hydrogen ion concentration change when the pH value of an aqueous solution is changed from 6 to 5?
25. Write the balanced chemical equations for the preparation of the following salts:
    a. A soluble sulfate by the action of an acid on a metal
    b. A soluble sulfate by the action of an acid on an insoluble metal oxide
    c. An insoluble sulfate by the action of an acid on another salt
26. 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.
27.
    a. What kind of salt is prepared by precipitation?
    b. Name a salt prepared by a direct combination. Write an equation for the reaction that takes place in preparing the salt you have named.
    c. Name the procedure used to prepare a sodium salt such as sodium sulphate.

Assignment for Submission (19 Points)

Answer the following questions correctly

1. Write the components of backing soda. .
2. Write the chemical formula of plaster of Paris
3. A sample of lemon juice has a hydronium-ion concentration equal to 2.5×10^{-2} M. What is the pH of this sample?
4. A sample of grape juice has a pH of 4.0. What is the hydroxide-ion concentration of this solution?
5. a) Define an acid-base indicator. Mention one synthetic acid-base indicator. b) If someone in the family is suffering from a problem of acidity after overeating, which of the following substances would you suggest as a remedy? Lemon juice,vinegar or baking soda solution. Mention the property on the basis of which you will choose the remedy.
6. Long-term CO_2 exposure to lime water results in the production of soluble calcium hydrogen carbonate, which causes the white precipitate to dissolve and the solution to clear.
7. Write one point of difference between each of the following:
    i. A hydrated salt and an anhydrous salt.
    ii. Washing soda and soda ash.
    iii. Baking soda and baking powder.
8. Define acid salt, basic salt and normal salt with examples
9. Explain the action of dilute hydrochloric acid on the following with chemical equation:
    i. Magnesium ribbon
    ii. Sodium hydroxide

iii. Crushed egg shells

10. Name the natural source of each of the following acid

i. Citric acid.

ii. Oxalic acid.

iii. Lactic acid.

iv. Tartaric acid

Answers to Self-Review Exercises

Part I.

1. b 2.c 3.b 4. c 5.b 6.b 7. b 8. c 9. a 10. C

Part II.

11. Less 12. basic 13. pH =7 14. greater 15. SO_2 16.Greater 17. hydrogen 18. Greater 19. oxides

Part III.

20. pH = 5 21. Acidic 22. Has only one ionizable Hydrogen

23. A. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

B. $Zn + 2NaOH \rightarrow Na_2ZnO_2 + H_2$

24. 10 times

25. a. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

b. $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$

c. $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

26. a. $CaO + 2HCl \rightarrow CaCl_2 + H_2O$

b. To neutralize the acid in the soil

27. a. An Insoluble salt is prepared by precipitation.

b. $2Fe + 3Cl_2 \rightarrow 2FeCl_3$, Iron (III) chloride or Ferric chloride

c. By neutralizing sodium carbonate or sodium hydroxide with dilute sulphuric acid

$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + CO_2$

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

Answer Key to Self Test Exercises

Answer Key to Activity 3.1

2. SO_2, CO_2, P_2O_5, N_2O_3

3. CO, NO,

4 a. $N_2O_5 + H_2O \rightarrow 2HNO_3$

b. $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$

c. $MgO + CO_2 \rightarrow MgCO_3$

d. $Ca(OH)_2 + SO_3 \rightarrow CaSO_4 + H_2O$

e. $2KOH + CO_2 \rightarrow K_2CO_3 + H_2O$

f. $SrO + SO_2 \rightarrow SrSO_3$

Answer Key to Activity 3.2

1. a. $K_2O + H_2O \rightarrow 2KOH$

b. $MgO + H_2O \rightarrow Mg(OH)_2$

c. $Na_2O + CO_2 \rightarrow Na_2CO_3$

d. $Li_2O + SO_2 \rightarrow Li_2SO_3$

e. $BaO + H_2SO_4 \rightarrow BaSO_4 + H_2O$

f. $CuO + HCl \rightarrow CuCl_2 + H_2O$

2. a. MgO = basic b. BaO = basic c. P_4O_{10} = acidic d. N_2O_5 = acidic

e. Cu_2O = basic f. Fe_2O_3 = basic g. K_2O = basic h. SO_2 = acidic

Answer Key to Activity 3.3

a. $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$

$PbO + 2NaOH \rightarrow Na_2PbO_2 + H_2O$

b. $Al_2O_3 + HNO_3 \rightarrow Al(NO_3)_3 + H_2O$

$PbO + HNO_3 \rightarrow Pb(NO_3)_3 + H_2O$

c. $Al_2O_3 + KOH \rightarrow 2KAlO_2 + H_2O$

$PbO + KOH \rightarrow Na_2PbO_2 + H_2O$

d. $Al_2O_3 + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2O$

$PbO + H_2SO_4 \rightarrow PbSO_4 + H_2O$

Answer Key to Activity 3.4

a. Oxidation state of oxygen is -2

b. Al_2O_3 = -2, N_2O_5 = -2, P_4O_{10}= -2, KO_2 = -1/2 , SO_2 = -2

Answer Key to Activity 3.5

$HClO_4 + H_2O \rightleftharpoons ClO_4^- + H_3O^+$

$CH_3COOH + H_2O \rightleftharpoons CH_3CO^- + H_3O^+$

$H_2S + 2H_2O \rightleftharpoons S_2- + 2H_3O +$

$HNO_3 + H_2O \rightleftharpoons NO_3^- + H_3O^+$

Answer Key to Activity 3.6

1.

Monoprotic	Diprotic	Triprotic
$HClO_4$	H_2CO_3	H_3PO_4
CH_3COOH	H_2S	
Cl_3COOH		

2.

Binary Acids	Ternary(oxoacids)
HBr	$HBrO_4$
HCl	HClO
H_2Se	HNO_3

Answer Key to Activity 3.7

1. a. Organic acids= CH_3COOH, CH_3CH_2 COOH

 Inorganic acids = HNO_3, H_2SO_4

 b. Weak acids : Organic acids= CH_3COOH, CH_3CH_2 COOH

 Strong acids: Inorganic acids = HNO_3, H_2SO_4

2. Acids are substances that contain ionizable H^+ ions

 Eg. HNO_3, H_2SO_4 are acids because they give H^+ ions whereas alcohol and sugar are not acids as they do not give H+ when dissolve in water

Answer Key to Activity 3.8

1. $H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$, $[H^+]$ = 2 x 0.01= 0.02, pH = - log 2x 10^{-2} = 0.3+ 2= 2.3

2. HCl and H2SO4 dissociate 100% in water. HCl is monoprotic (one mole of HCl gives one H+) while H2SO4 diprotic (two mole of HCl gives two moles H^+)

3. pH = 2 $\rightleftharpoons$ $[H^+]A = 1x10^{-2}$, pH =6, $[H^+]B = 1x10^{-6}$

 $[H^+]A/[H^+]B = 1x10^{-2}/1x10^{-6} = 10^4$ $\rightleftharpoons$ the concentration of A is 10,000 times morethan the concentration of B.

Answer Key to Activity 3.9

1. The chemical formula of lime water is $Ca(OH)_2$

2. The chemical formula of caustic potash is KOH

3. The name given to bases that are highly soluble in water is alkalis.

 Example: KOH, NaOH

4 a. NaOH = Arrhenius base b. C_2H_5OH = neither

 c. H_3PO_4 = Arrhenius acid d. $C_6H_{12}O_6$ = neither

 e. HNO_2 = Arrhenius acid f. $Ba(OH)_2$ = Arrhenius base

5. a. $KOH + H_2C_2O_4 \rightleftharpoons KC_2O_4 + H_2O$

b. $3Sr(OH)_2 + 2H_3PO_4 \rightleftharpoons Sr_3(PO_4)_2 + 6H_2O$

c. $3HCl + Fe(OH)_3 \rightleftharpoons FeCl_3 + 3H_2O$

d. $3H_2SO_4 + 2Cr(OH)_3 \rightleftharpoons Cr2(SO_4)_3 + 6H_2O$

Answer Key to Activity 3.10

A. $Na_2CO_3 \rightleftharpoons$ Base : NaOH & Acid: H_2CO_3

B. $CaSO_4 \rightleftharpoons$ Base: $Ca(OH_2)_2$ & Acid: H_2SO_4

C. KOCl $\rightleftharpoons$ Base: KOH & Acd : HClO

D. $Ca(NO_3)_2 \rightleftharpoons$ Base: $Ca(OH_2)_2$ & Acid : HNO_3

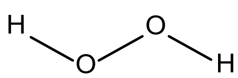
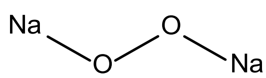
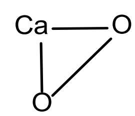
Answer Key to Activity 3.11

NaCl= Normal salt	Zn(OH)Cl= Base salt
$CaCO_3$ = Normal salt	K_3PO_4 = Normal salt
$Mg(H_2PO_4)_2$ = Acid salt	$NaHCO_3$ = Acid salt

Self-Test Exercise -3.1

1. a. NO = neutral oxide b. ZnO= Amphoteric oxide c. CO = neutral oxide
   d. Na_2O_2 = peroxide e. Al_2O_3 = Amphoteric oxide
2. a. $Al_2O_3 + 6HNO_3 \rightarrow 2Al(NO_3)_3 + 3H_2O$
   b. $2NaOH + Al_2O_3 \rightarrow 2NaAlO_2 + H_2O$
   c. $ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$
3. Peroxides contain peroxide link –O-O- and oxidation state of oxygen is -1

H_2O_2 Na_2O_2 CaO_2

H_2O_2  Na_2O_2  CaO_2 

4. Using universal indicator or pH meter
5. a. SO_2 b. SO_3 c. N_2O_5 d. CO_2 e. P_2O_3 f. P_2O_5
6. a. CaO b. K_2O c. MgO d. Na_2O e. SrO f. BaO
7. a. $MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$
   b. $CaO + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O$
   c. $K_2O + H_2SO_4 \rightarrow K_2SO_4 + H_2O$
   d. $NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$

Self-Test Exercise -3.2

1. a. $HBrO_4$ = Arrhenius acids b. BCl_3 = Not Arrhenius acids c. H_2S =Arrhenius acids d. HF = Arrhenius acids e. HBr = Arrhenius acids f. PCl_5 = Not Arrhenius acids
2. a. $Al_2O_3 + HNO_3 \rightarrow 2Al(NO_3)_3 + 3H_2O$
   b. $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$
   c. $ZnO + H_2SO_4 \rightarrow ZnSO_4 + H_2O$
3. **a.** $HClO_4$ = Monoprotic , ternary, strong acids **b.** H_2SO_4 = Diprotic, ternary, strong acid **c.** H_2CO_3 = Diprotic, ternary, weak acid **d.** H_2SO_3 = Diprotic, ternary, weak acid **e.** HF = Monoprotic, binary, weak acid **f.** H_2S = Diprotic, binary, weak acid **g.** HCl = Monoprotic, binary **h.** H_3PO_4 = Triprotic , ternary, weak acid **i.** HNO_2 = monoprotic, ternary, weak acid **j.** HCN = monoprotic , ternary, weak acid **k.** HNO_3 = Monoprotic, ternary, strong acids l. CH_3COOH = Monoprotic, ternary, weak acid

4. Blue litmus paper becomes red in acids. Universal indicator or pH meter gives lower pH for the acid and pH 7 or nearly pH 7 for water.

5. Strong acids dissociate completely or nearly completely in aqueous solutions and produce large amount of H+ whereas weak acids dissociate slightly and produce small amount of H+ in aqueous solutions.

6. What is the pH of a solution having the following hydrogen ion concentrations?

a. $[H^+] = 5 \times 10^{-3}$ M, pH = $-\log[H^+] = -\log 5\times10^{-3} = -\log5 + 3 = -0.7 + 3 = 2.3$

b. $[H^+] = 3 \times 10^{-3}$ M , pH = $-\log[H^+] = -\log^3 \times 10^{-3}$

c. $[H^+] = 2.0 \times 10^{-6}$ M, pH = $-\log[H^+] = -\log2.0\times10^{-6} = -0.3 + 6 = 5.7$

7. No. of moles = M xV(L) = 0.15mol/L x 0.5L=0.075 mol.

8. Identify the following solutions as acidic or basic, estimate $[H_3O^+]$ values for each, and rank them in order of increasing acidity:

a. Saliva, pH =6.5 ⇌ Acidic

b. Orange juice, pH = 3.7⇌ acidic

c. Pancreatic juice, pH = 7.9 ⇌ basic

d. Wine, pH = 3.5 ⇌ acidic

Self-Test Exercise – 3.3

1 a. $KOH + H_2O \rightarrow K^+(aq) + OH^-(aq)$

b. $Ca(OH)_2 + H_2O \rightarrow Ca2+(aq) + 2OH^-(aq)$.

2. (c) antacid

3. d. Calcium hydroxide $(Ca(OH)^2)^-$ quick lime

4. Strong electrolytes

5. c. weak bases

6. Alkalis are soluble bases. Example NaOH, KOH are alkalis .

7. See the text

8. pH 5 > pH9 in $[H^+]$.

pOH = 14-pH □pOH = 14-5 =9, pOH = 14-9= 5

pH5 →pOH = 9, [OH-= 1×10^{-9} M, pH 9 pOH = 5, $[OH^-] = 1\times10^{-10}$ M

9. a. $[H_3O^+] = 1.0 \times 10^{-5}$ ⇌ pH =5. pOH = 9

b. $[OH^-] = 1.0 \times 10^{-9}$ ⇌ pOH = 9, pH =5,

10. a. pH 13.0 ⇌ pOH = 1, $[OH^-] = 1\times10^{-1}$

b. pH 3.0 ⇌ pOH = 11, $[OH^-] = 1\times10^{-11}$

c. pH 8.0 ⇌ pOH = 6, $[OH^-] = 1\times10^{-6}$

Self-Test Exercise – 3.4

1. a. Zinc sulphate $(ZnSO_4)$ ⇌ ZnO & H_2SO_4

b. Calcium phosphate $Ca_3(PO_4)_2$ ⇌ $Ca(OH)_2$ & H_3PO_4

c. Silver acetate (CH_3COOAg) ⇌ Ag(OH) + CH_3COOH

d. Sodium carbonate (Na_2CO_3) ⇌ NaOH + H_2CO_3

2. Penolphthalein + NaOH ⇌ pink solution

pink solution + HCl ⇌ pink colour disappears

Phenolphthalein + NaCl ⇌ No colour change

3. see the text

4. a. NH_4Cl b. K_3PO_4 c. $FeCO_3$ d. AgCl e. AgNO f. $CaCl_2$ g. $PbSO_4$
h. Na_2S i. $CaSO_4$ j. $BaCO_3$

5. Salts are ionic compounds in which the +ve ion is derived from base & the -ve ion is derived from an acid. Soluble salts dissolve to produce their ions and conduct electricity in their aqueous solutions.

6. The carbonates of alkali metals except Li_2CO_3 are stable to heat

7. All alkali nitrates except lithium nitrate gives metal nitrite and oxygen on heating.

8. a. Halide ions $\rightleftharpoons$ $AgNO_3$
b. Fe^{2+} and Fe^{3+} ion $\rightleftharpoons$ solutions of potassium hexacyanoferrate (II), solution of potassium hexacyanoferrate (III) and potassium thiocyanate.
c. Sulphate, SO_4^{2-} ions $\rightleftharpoons$ $BaCl_2$ and HCl

Experiment 3.1

Test for Acidity and Basicity of Oxides

1. $S + O_2 \rightarrow SO_2$
$2Mg + O_2 \rightarrow 2MgO$
$2Ca + O_2 \rightarrow 2CaO$
2. When water is added to these oxides, SO_2 forms H_2SO_3, and MgO forms $Mg(OH)_2$
3. Universal indicator turns yellow-orange in a solution of SO_2 and blue-purple in a solution of $Mg(OH)_2$, and blue litmus turns red in a solution of SO_2 and red litmus blue in a solution of $Mg(OH)_2$ and CaO form $Ca(OH)_2$
4. The color change occurred because a solution of SO_2 is acidic and that of MgO is basic

Experiment 3.2

Investigating Amphoteric Behaviour of Oxides

1. Al_2O_3 reacts with both HCl and NaOH
2. The presence of chemical reaction between Al_2O_3 and HCl as well as Al_2O_3 andNaOH
3. $Al_2O_3\ (s) + 6HCl \rightarrow 2ALCl_3 + 3H_2O(l)$
$Al_2O_3(s) + 2NaOH \rightarrow 2NaAlO_2(aq) + H_2O(l$

Experiment 3.3

Effect of Acids on Indicators

Indicator	Color of indicator in			
	Lemon juice	Dilute HNO_3	Dilute HCl	Dilute H_2SO_4
phenolphthalein	colorless	colorless	colorless	colorless
litmus	Red	Red	Red	Red
Methyl red	Red	Red	Red	Red
Universal indicator	Orange-red	Red	Red	Red

Experiment 3.4

Investigating the Reactions of Metals with Dilute Acids

A. Release of gas
B. "popping" sound is heard. This proves that the gas is hydrogen.
C. Hydrogen gas

D. The reaction of powdered magnesium with HCl and H_2SO_4 is the most violent

Experiment 3.5

Reactions of Acids with Carbonates and Hydrogen Carbonates

A. The formation of bubbles indicates the release of gas.

B. The change in the color of damp blue litmus to red when it is held close to the mouth of the test tube proves the gas to be acidic

C. When the gas is passed through lime water, the clear solution turns milky, and this proves the gas to be carbon dioxide.

Experiment 3.6

pH of Solutions of Common Substances

Write from the result of the experiment

Experiment 3.7

Chemical Behavior of Bases

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

B. KNO_3

C. $Ca(OH)_2 \rightarrow CaO + H_2O$

Experiment 3.8

Test for Sulphates

1. $BaSO_4$
2. To test whether the white precipitate is BaSO4
3. Ba^{2+} + SO4^{2-} $\rightarrow BaSO_4$ (white ppt)

Experiment 3.9

Test for Carbonates and Hydrogen Carbonates

1. When dilute HCl is added to the solution of Na_2CO_3 and $NaHCO_3$, there is an evolution of carbon dioxide gas, which turns lime water milky. This is due to the formation of $CaCO_3$
2. Upon addition of calcium chloride solution, solutions of carbonates form a white precipitate of $CaCO_3$.

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

3. Refer to text

Resources

The current Ethiopian Grade 10 Chemistry Textbook for face-to-face modality.
Silberberg, Martin S. Principles of General Chemistry-3rd ed. McGraw-Hill. 2013.
Chang, R., Overby, J. General Chemistry-The Essential Concepts. 6th ed. McGraw-Hill. 2013.
Ebbing, D. D., Gammon, S. D. General Chemistry Eleventh Edition, CENGAGE Learning 2015.