MODULAR SYSTEM

SOLUTIONS

Şerafettin KÜÇÜKOĞLU

Mustafa YILMAZ

Ayhan NAZLI

Zambak PUBLISHING

www.zambak.com

Proof Reader

Andy HILTON

Digital Assembly

Zambak Typesetting & Design

Page Design

Memduh SEYMAN

Cemil ŞEN

Publisher

Zambak Basım Yayın Eğitim ve Turizm İşletmeleri Sanayi Ticaret A.Ş.

Printed in

İstanbul - TURKEY

ISBN

975-266-213-7

DISTRIBUTION

ZAMBAK YAYINLARI
Bulgurlu Mah. Libadiye Cad.
Haminne Çeşme Sk. No. 20 34696
Üsküdar / İSTANBUL

Tel : +90-216 522 09 00 (pbx)
Fax : +90-216 443 98 39
www.zambak.com

PREFACE

Chemistry is an interesting and fundamental branch of science because it gives us the chance to explain the secrets of nature. What is water? How does tap water differ from distilled water? Why is salt sprinkled on the roads in winter? Why are soft drinks kept in a freezer? How is the concentration of calcium ions measured in milk? These kinds of questions and their answers are all part of the world of chemistry. However, one does not need to be a chemist or scientist to understand the simplicity within the complexity around us. This book helps everyone to understand the nature of solutions.

The aim was to write a modern, up-to-date book where students and teachers can get concise information about the solutions. Throughout the books, colorful tables, important reactions, interesting extras, reading passages, funny cartoons and puzzles are used to explain ideas. Sometimes reactions are given in detailed form, but, in general, excessive detail has been omitted.

The book is designed to introduce a basic knowledge of solutions and measurement of their concentration. We will study principally the nature of solutions, their components with their physical and chemical properties. This book will also show you how crucial to life are solutions and how important is measuring concentration.

We would like to thank Necdet ÇELİK, Davut PİRAZ, Muhammet AYDIN, Murat DURKAYA, Ali Serhat ÖZ, Osman YILDIRIM, Nuh ÖZDİN, Uğur Hulusi PATLI and Yener EKŞİ for their thoughtful criticisms and helpful suggestions for the manuscript which have been of such great value in developing the book.

The authors would like also to thank Orhan Keskin, Ramazan Şahin and Ali Çavdar for their support and encouragement throughout.

We are particularly grateful to our spouses and children for their patience.

The Authors

CONTENTS

Chapter 1

THE PROPERTIES OF SOLUTIONS

INTRODUCTION 6

1. THE NATURE OF SOLUTIONS 7

1.1. THE DISSOLUTION PROCESS 11

1.2. HEAT CHANGE IN THE DISSOLUTION PROCESS 13

Reading : Heat Packs and Cold Packs 13

1.3. THE CONDUCTIVITY OF SOLUTIONS 14

Reading : Accumulators 15

1.4. CHANGE IN FREEZING AND BOILING POINTS 16

1.5. DILUTE AND CONCENTRATED SOLUTIONS 17

Reading : Water: The Basis of Life 19

2. SOLUBILITY 20

2.1. FACTORS AFFECTING SOLUBILITY 22

Temperature 22

Pressure 23

Reading : The Bends 23

3. MIXING AQUEOUS SOLUTIONS 27

Net Ionic Equation 27

The Application of Precipitation 29

4. HYDROLYSIS 31

SUPPLEMENTARY QUESTIONS 33

MULTIPLE CHOICE QUESTIONS 35

PUZZLES 37

Chapter 2

MEASURING CONCENTRATION

INTRODUCTION 40

1. PERCENT CONCENTRATION 41

1.1. MASS PERCENT 41

The Preparation of Dilute Solutions 42

1.2. VOLUME PERCENT 45

2. MOLARITY 47

3. THE PREPARATION OF A SOLUTION WITH A DESIRED CONCENTRATION 50

SUPPLEMENTARY QUESTIONS 54

MULTIPLE CHOICE QUESTIONS 56

PUZZLES 58

APPENDIX A 60

APPENDIX B 61

GLOSSARY 62

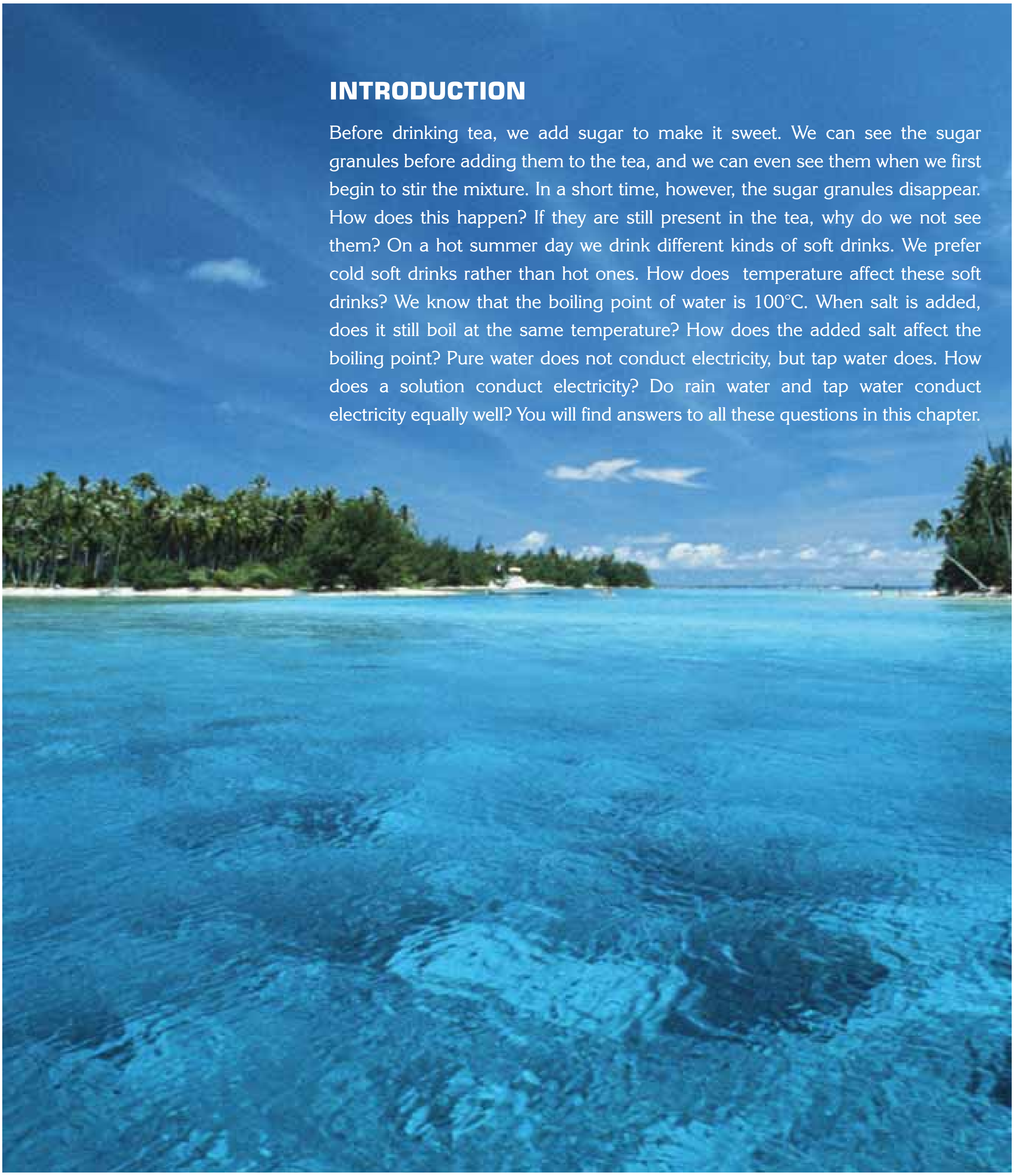
ANSWERS 64

INDEX 67

REFERENCES 70

THE PROPERTIES OF SOLUTIONS

SOLUTIONS



INTRODUCTION

Before drinking tea, we add sugar to make it sweet. We can see the sugar granules before adding them to the tea, and we can even see them when we first begin to stir the mixture. In a short time, however, the sugar granules disappear. How does this happen? If they are still present in the tea, why do we not see them? On a hot summer day we drink different kinds of soft drinks. We prefer cold soft drinks rather than hot ones. How does temperature affect these soft drinks? We know that the boiling point of water is 100°C. When salt is added, does it still boil at the same temperature? How does the added salt affect the boiling point? Pure water does not conduct electricity, but tap water does. How does a solution conduct electricity? Do rain water and tap water conduct electricity equally well? You will find answers to all these questions in this chapter.

1. THE NATURE OF SOLUTIONS

A mixture is a combination of two or more substances in which each keeps its properties. Mixtures can be classified into two main groups. One type of mixture has visible boundaries between the substances that it comprises. This type is called a **heterogeneous mixture**. Its composition differs from place to place. Sand in water, oil in water and milk and soup are some heterogeneous mixtures.

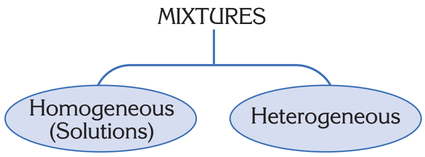
(a) Sand in water

(b) Oil in water

(c) Soup

(d) Milk

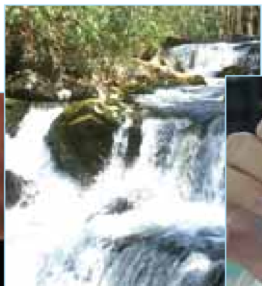
Examples of heterogeneous mixtures.



In another type of mixture, the boundaries between the substances cannot be seen. This is called a **homogeneous mixture**. Its composition is uniform throughout.

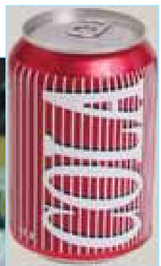
Homogeneous mixtures are also called **solutions**. Solutions are crucial to life and to many processes. We often encounter them in daily life. The air we breathe, our soft drinks, the amalgam used in dental fillings, the alloys used in the production of cars, the fog that causes traffic jams, coins, seas, lakes, and even our own body fluids are solutions.

(a) Drinking water



(b) River water

(c) Copper(II) sulfate solution

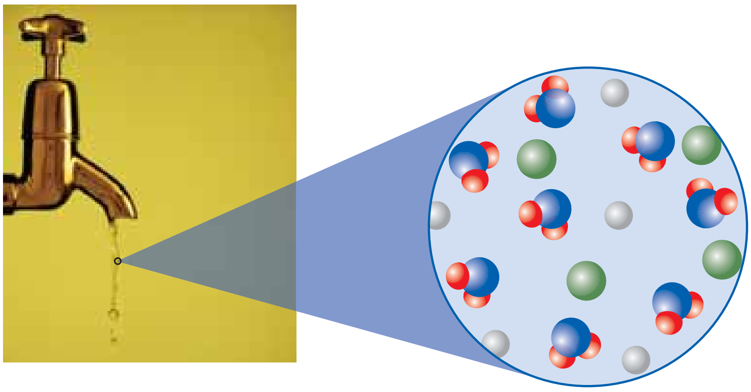


(d) Cola

Examples of homogeneous mixtures.

Solutions are impure. Even tap water is not pure water, but rather it is a solution of various substances in water. Tap water may contain chloride (Cl^-) as a disinfectant, as well as some mineral ions like sodium (Na^+), calcium (Ca^{2+}), bicarbonate (HCO_3^-), and fluoride (F^-).

	Distilled water	Drinking water
Pure	✓	✗
Homogeneous	✓	✓



Solutions are homogeneous but not pure.

The components of a solution are the solvent and the solute.

Solution = Solvent + Solute

Why is sea water salty?

Some mineral salts have been carried to the oceans and seas by rivers for 200-300 million years. These salts are soluble in water and produce ions such as, Cl^-, Na^+, SO_4^{2-}, Mg^{2+}, Ca^{2+}. These dissolved ions make sea water salty.

The component present in the largest amount or that determines the state of mixture is called the **solvent**. The other component or components said to be dissolved in the solvent are called **solutes**. For example, when we dissolve a teaspoon of sugar in a glass of water, the sugar is the solute and the water is the solvent. In tap water, the solvent is water and the ions (Cl^-, Na^+, Ca^{2+}, HCO_3^-, and F^-) are solutes.

Solutions are usually classified according to their physical state, as **solid**, **liquid**, or **gaseous**. The physical state of a solution is determined by the solvent. Many alloys are solid solutions of one metal dissolved in another. For example, brass, which is used to make musical instruments and many other objects, is a solution of copper and zinc. Air is a gaseous solution containing nitrogen, oxygen, and other gases. Carbon dioxide (a gas), alcohol (a liquid), and salt (a solid) each dissolve in water (a liquid) to form liquid solutions. Water is the most common solvent in the laboratory and in many fields. Water solutions are known as **aqueous solutions**. Because they are so important, in this section we will concentrate on the properties of aqueous solutions. Some solutions and their compositions are illustrated in Table 1.

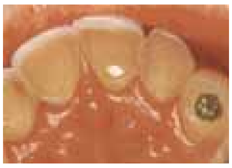
	Solute	Solvent	Example	
LIQUID SOLUTIONS	Liquid	Liquid	Alcohol in water (Antifreeze)	
	Solid	Liquid	Salt in water (Salty water)	
	Gas	Liquid	CO_2 in water (Carbonated beverages)	
GAS SOLUTIONS	Gas	Gas	Oxygen in nitrogen (Air)	
	Liquid	Gas	Moisture in the air (Fog)	
SOLID SOLUTIONS	Liquid	Solid	Mercury in silver (Dental filling)	
	Solid	Solid	Nickel in copper (Coins)	

Table 1 *Some examples of solutions with different states*

A Type of Solution: Alloys

Metals mix together by being heated under high pressure and producing alloys. Some types of alloys and their uses are listed below.

Common Alloys		
Alloy	Composition of alloy	Uses
Brass	Copper, zinc	Musical instruments
Bronze	Copper, tin, zinc	Statue building
Stainless steel	Iron, carbon, nickel, chromium	Tableware, kitchenware
Carbon steel	Iron, carbon	Auto bodies, machinery
Solder	Lead, tin	In metal plumbing
Dental amalgam	Mercury, silver, tin, copper, zinc	Dental fillings

Example 1

Decide whether the following substances are pure or impure.

a. Sea water **b.** Solder **c.** Tablespoon **d.** Copper wire

Solution

a. Sea water is a substance composed of water and some salts: therefore it is not pure.

b. Alloys are the mixtures of at least two metals. Solder is an alloy composed of lead and tin: therefore it is impure.

c. A table spoon is an example of an alloy. It is prepared by mixing certain kinds of metals, like iron, chromium, silver, and nickel. Therefore, a tablespoon is also not pure.

d. Copper wire is made up of the element copper. Elements are pure substances.

Example 2

Find the solute and solvent components of the given solutions.

a. Air **b.** Carbonated beverages **c.** Bronze **d.** Salt water

Solution

a. Air is a solution mainly composed of nitrogen and oxygen. Nitrogen (78%) is the solvent and oxygen (21%) is the solute.

b. Water (H_2O) is the major component of carbonated beverages. So, it is the solvent and the gas, carbon dioxide (CO_2), is the solute in carbonated beverages.

c. Bronze is an alloy of copper (Cu), zinc (Zn), and tin (Sn). Copper is the major component, so, it is the solvent. Zinc and tin are the minor components, so, they are the solutes.

d. The components of salt water are water and salt. The solvent is water and the solute is salt.

In carbonated beverages, the one of the solute components is gas and the solvent is water.

Exercise 1

What are the solvent and solute components in the following examples?

a. Steel **b.** Vinegar **c.** Soft drinks **d.** Tap water

1.1. THE DISSOLUTION PROCESS

What happens when we add salt to water for cooking or when we put a teaspoon of sugar in a glass of tea? As you know salt and sugar disappear in water. This is known as dissolution. **Dissolution** is the mixing of a solute in a solvent. The dissolving of solid sodium chloride in water is shown in Figure 1. When sodium chloride dissolves in water the positive parts of the water molecules (i.e., the hydrogen) attract Cl^- ions and negative parts of the water molecules (i.e., the oxygen) attract Na^+ ions. The resulting solution contains Na^+ and Cl^- ions. This can be represented as

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

where (aq) indicates that the ions are surrounded by water molecules. All of the ionic substances that dissolve in water are separated into ions. Any process in which ions are produced is called **ionization**. Acids, bases, and salts dissolve in water in this way.

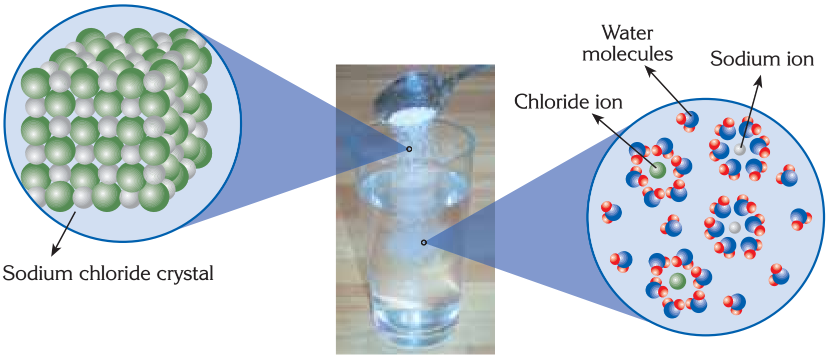


Figure 1 *Dissolution of table salt in water.*

Water also dissolves nonionic substances like sugar and alcohol because there is a chemical similarity between these molecules and water molecules (i.e., all have OH groups). Sugar and alcohol dissolve in water as molecules. When sugar dissolves in water, the sugar molecules are attracted by water molecules, as shown in Figure 2.

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$

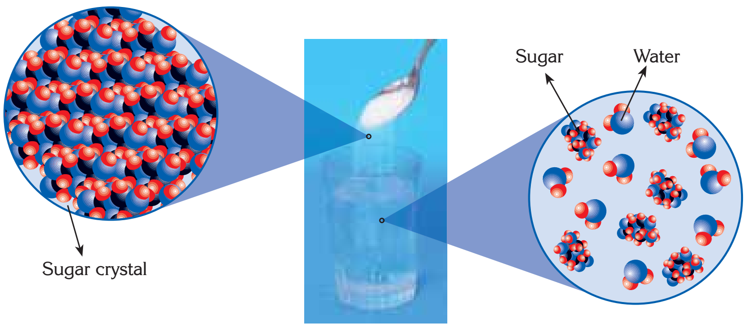


Figure 2 *Dissolution of sugar in water.*

Water and ethyl alcohol mix with each other in all proportions. Such liquids are described as **miscible**.

EXTRA

Physical or Chemical

Some dissolution processes may occur within a reaction. For example,

$$CO_2(g) + H_2O(l) \longrightarrow H_2CO_3(aq)$$

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

These kinds of dissolution processes are chemical.

Additionally, some occur without a chemical reaction. They are physical.

$$C_6H_{12}O_6(s) \xrightarrow{H_2O} C_6H_{12}O_6(aq)$$

$$C_2H_5OH(l) \xrightarrow{H_2O} C_2H_5OH(aq)$$

Therefore we can say that dissolution may be chemical or physical.

*The process of surrounding solute ions by water molecules is called **hydration**. The general term for the surrounding of a solute particle by solvent molecules is **solvation**.*

EXTRA

Dry Cleaning

Dry cleaning is a process in which clothes are cleaned without water. The cleaning fluid used is either perchlorethylene ($Cl_2C{=}CCl_2$) **'perc'** or the volatile synthetic solvent carbon tetrachloride (CCl_4). These solvents are used because their molecular structures are similar to those of the oils found in dirt and grime.

There are, however, some substances, such as oil and petroleum (Figure 3), which do not dissolve in water. Two liquids that do not mix with each other are described as **immiscible**. Just because a substance does not dissolve in one solvent, however, does not mean it will not dissolve in another. Oil, for example, dissolves in carbon tetrachloride (Figure 4).

Figure 3 *Oil and water are immiscible.*

Figure 4 *Oil and carbon tetrachloride are miscible.*

These are all examples of the rule, '**like dissolves like**'. In other words, when solute and solvent have molecules that are 'like' each other, they tend to form a solution. The dissolution of iodine is also a good example of this rule. It dissolves in carbon tetrachloride rather than water, as shown in Figure 5.

a b

Figure 5 *Iodine, a solid nonmetal, is added to the beakers in which there are water and carbon tetrachloride (picture a); it does not dissolve in water, but it does dissolve in carbon tetrachloride (picture b) because of the similarity in their structures.*

1.2. HEAT CHANGE IN THE DISSOLUTION PROCESS

When a solute is dissolved in a solvent, heat change generally occurs. A dissolution process may be exothermic or endothermic. **Exothermic processes** emit energy as heat. **Endothermic** processes absorb energy as heat. Temperature rises in an exothermic process, but falls in an endothermic one. When lithium chloride (LiCl) dissolves in water, the solution gets warmer and the temperature goes up. We can say that the dissolution of lithium chloride is exothermic. (Figure 6).

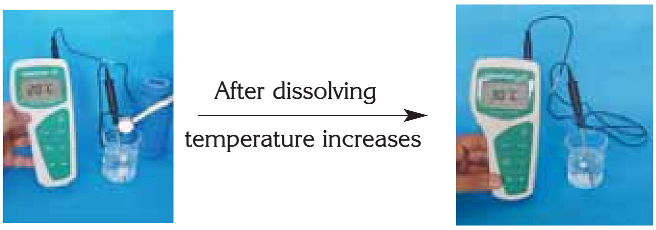


Figure 6 *When lithium chloride dissolves in water, the solution gets warmer. This process is* ***exothermic.***

On the other hand, when ammonium nitrate (NH_4NO_3) dissolves in water, the solution gets colder and the temperature goes down. Therefore, the dissolution process of ammonium nitrate is endothermic (Figure 7).

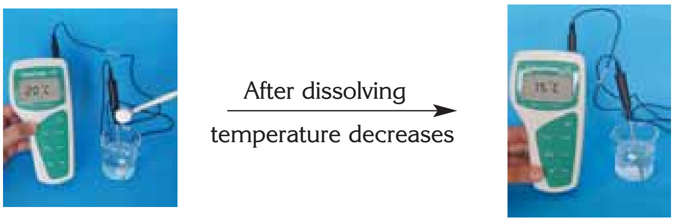


Figure 7 *The dissolving of ammonium nitrate is* ***endothermic.***

READING — HEAT PACKS and COLD PACKS

Heat packs *are used for first aid purposes. There is a supersaturated solution of sodium acetate ($NaCH_3COO$) in a packet. When the supersaturated solution is disturbed, the liquid quickly turns to a solid. This phase change causes the pack to heat up to approximately 54°C.*

It is possible to reuse a heat solution pack by immersing the solid in boiling water until it liquifies.

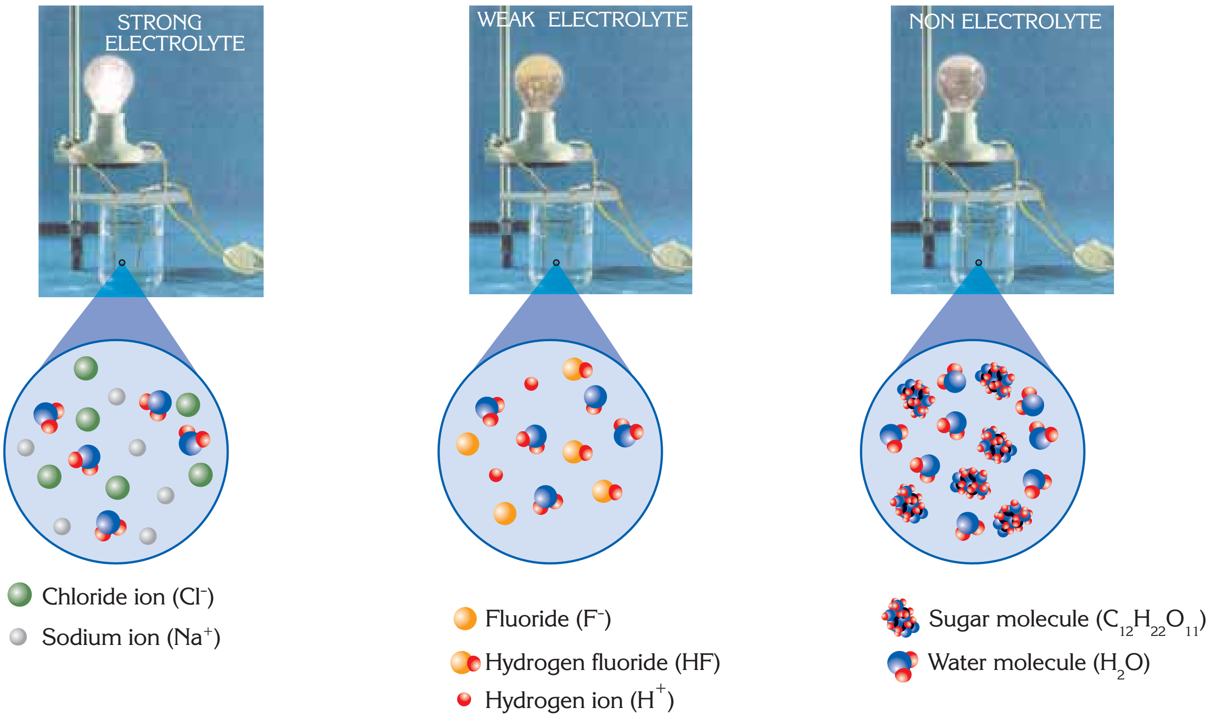
Cold packs *are made to treat pains and injures. For example, they help relieve shoulder pain and reduce swelling.*

When the bag is squeezed, the inner pouch breaks and ammonium nitrate, NH_4NO_3, and water mix. Since the dissolution of ammonium nitrate is endothermic, the water temperature decreases to about 5°C.

1.3. THE CONDUCTIVITY OF SOLUTIONS

The water solutions of some substances conduct electricity, while the solutions of others do not. The conductivity of a solution depends on its solute. The more ions a solution contains, the greater its conductivity. Solutions that conduct electricity are called **electrolytes.** Solutions which are good conductors of electricity are known as **strong electrolytes.** Sodium chloride, hydrochloric acid, and potassium hydroxide solutions are examples of strong electrolytes. If solutions are poor conductors of electricity, they are called **weak electrolytes.** Vinegar, tap water, and lemon juice are examples of weak electrolytes. Solutions of substances such as sugar and alcohol solutions which do not conduct electricity are called **nonelectrolytes**.

The conductivity of sodium chloride, hydrofluoric acid, and sugar solutions is illustrated below.



*Sodium chloride solution ionizes 100%. That is why it is a **strong electrolyte**.*

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

*Hydrofluoric acid solution is a **weak electrolyte** because it is only partially ionized.*

$$HF(g) \underset{}{\overset{H_2O}{\rightleftharpoons}} H^+(aq) + F^-(aq)$$

*In a sugar solution, no ions are produced because sugar dissolves as molecules. Thus, it is a **non-electrolyte** solution.*

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$

Tap water conducts electricity because it contains certain kinds of ions; that is, it is not pure.

READING — ACCUMULATORS

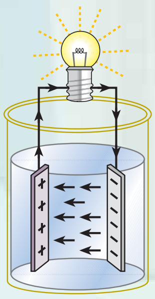
A car battery is a device used to power lighting, accessories, and other electrical systems. When the engine is shut off, it is also used to power the motor in the engine starter.

Car batteries generally have six electrochemical cells. In each cell there is a positive plate (PbO_2) and a negative plate (Pb). Both are immersed in an electrolyte acid solution of dilute sulfuric acid (H_2SO_4).

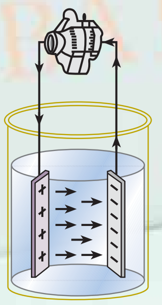
Each cell supplies 2.1 volts of electric energy, thus one battery supplies a total of 12.6 volts of energy.

A car battery stores electricity in the form of chemical energy. Since it loses its chemical energy, it must be recharged by the alternator.

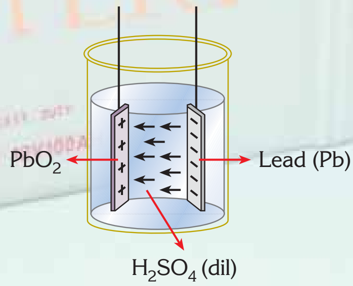
By reversing electrical current flow through the battery, the chemical process is reversed thus recharging the battery.



Discharging



Recharging



A cell example in accumulator

THE PIONEERS

Svante Arrhenius
(1859–1927)

Arrhenius is a Swedish chemist. In 1903 he won the Nobel Prize for his work with solutions. He explained why some solutions conduct electricity. In his theory, Arrhenius concluded that the 'molecule' breaks apart into a positive fragment and negative fragment, called ions. He also explained weak and strong electrolytes according to the ratio of the ions in solutions.

Salt is scattered on the road in winter. This is called de-icing.

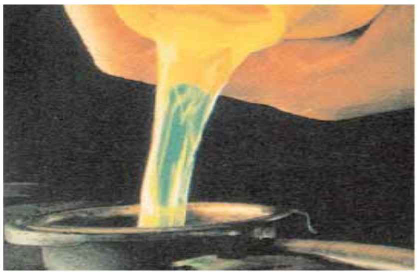
1.4. CHANGE IN FREEZING AND BOILING POINTS

People put the compound ethylene glycol (antifreeze) in their car radiators to prevent freezing in the winter and boiling in the summer. In cold climates, road crews scatter salt on roads to melt ice. Sea water, with its large salt content, freezes at a lower temperature than fresh water. These may seem to be a group of unrelated facts, but they each depend on the amount of solute particles dissolved in the solvent. Solutions freeze at lower temperatures and boil at higher temperatures than pure liquids. Water solutions freeze below 0°C and boil above 100°C. For example, solutions of one mole of ethylene glycol in one kilogram of water and of 0.5 mole of sodium chloride in one kilogram of water each begin to freeze at -1.86°C and to boil at 100.51°C.

A fresh water lake contains less ions than sea water and so freezes more easily.

At the same temperature, while a lake freezes salt water (sea water) remains unfrozen.

In sum, the lowering of the freezing temperature or the raising of the boiling temperature of a solvent is directly proportional to the amount of solute particles in the solution. The more solute in a solvent, the higher the boiling point and the lower the freezing point.

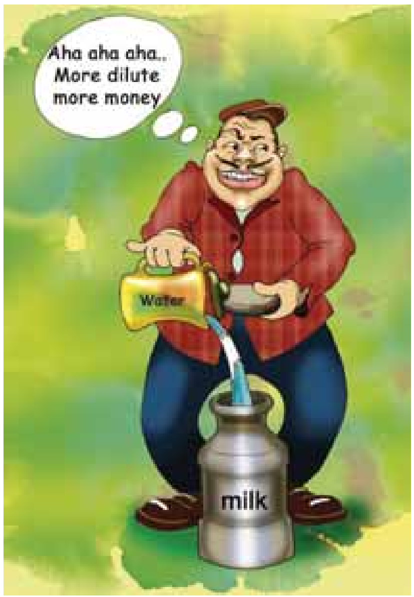


Antifreeze is used in cars both summer and winter. It prevents water from freezing in the winter and from boiling in the summer.

In winter when water pipes are frozen, ethylene glycol (antifreeze) can be used to correct the problem.

1.5. DILUTE AND CONCENTRATED SOLUTIONS

A solution can be prepared in different concentrations according to the amount of solute dissolved in it. For example, coffee can be strong or weak. Strong coffee contains more coffee dissolved in a given amount of water than weak coffee. The terms "concentrated" and "dilute" are used to describe solutions. Solutions that contain a relatively large amount of solute are called **concentrated** (strong coffee is concentrated). Those containing a relatively small amount of solute are called **dilute** (weak coffee is dilute).



Some substances are purchased in concentrated form. One obtains more dilute solutions by simply adding water (or another solvent) to these concentrated solutions. This process is called **dilution**. For example, fruit juices are sometimes packaged as concentrates, which are concentrated solutions. When you want to drink a glass of fruit juice, you should add water to make it more dilute as shown below:

a. *A concentrated fruit juice solution*

b. *Dilution process by adding water*

c. *A dilute fruit juice solution*

Example 3

Write the equations for dissolution of the following substances in water:

a. Sodium acetate ($NaCH_3COO$) **b.** Alcohol (C_2H_5OH) **c.** $CaCl_2$

Solution

a. Ionic substances dissolve in water by producing ions. Sodium and acetate ions are produced.

$NaCH_3COO(s) \longrightarrow Na^+(aq) + CH_3COO^-(aq)$

b. Molecular substances dissolve as molecules.

$C_2H_5OH(l) \longrightarrow C_2H_5OH(aq)$

c. Calcium chloride is an ionic substance. It produces calcium and chloride ions.

$CaCl_2(s) \longrightarrow Ca^{2+}(aq) + 2Cl^-(aq)$

Exercise 2

Write the equations for dissolution of the following substances in water:

a. K_2CO_3 **b.** $Al(NO_3)_3$ **c.** Glucose ($C_6H_{12}O_6$) **d.** H_2SO_4

Example 4

Classify the following solutions as either strong electrolytes, weak electrolytes or nonelectrolytes:

a. Vinegar **b.** Sugar - water solution

c. Sulfuric acid solution **d.** Lemon juice

Solution

a. Vinegar contains a weak acid (i.e., acetic acid), so it is a weak electrolyte.

b. Sugar dissolves in water as molecules. These kinds of solutions do not conduct electricity. Thus, a sugar solution is a nonelectrolyte.

c. Sulfuric acid solution is a strong acid. It ionizes nearly 100%. Therefore, it conducts electricity well. It is a strong electrolyte.

d. Lemon juice contains a weak acid (i.e., citric acid), so it is a weak electrolyte.

Exercise 3

Classify the following solutions as either strong electrolytes, weak electrolytes, or nonelectrolytes:

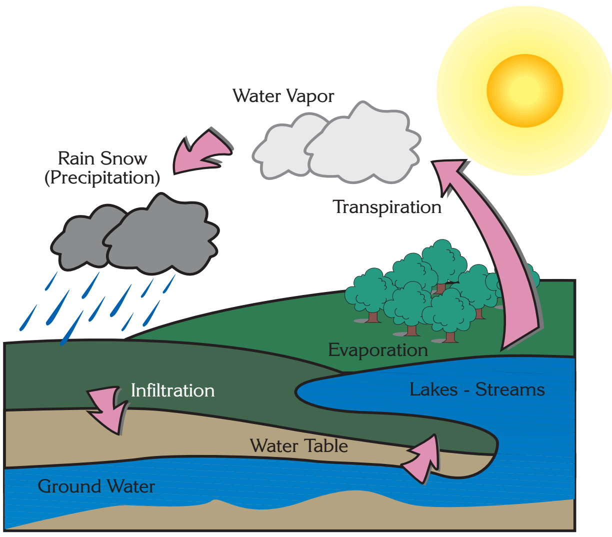
a. Carbonated beverages **b.** A hydroiodic acid solution **c.** A glucose solution

READING — WATER : THE BASIS of LIFE

All living organisms need water. Since water constitutes the majority of the body of a living organism, all reactions occurring in the cells of living organisms take place in one or another kind of water solution. Pure water is an odorless, tasteless, and colorless liquid. The taste of tap water stems from the gases and salts dissolved in it. Distilled water has no dissolved gases or salts and is not preferred for drinking. Water covers 75% of the earth's surface. It is the most abundant substance in the world. Even though the oceans and seas contain a huge amount water, this water cannot be directly used to support non-marine life. Although today, drinkable water can be produced from the seas, the most common sources of drinkable water are lakes and rivers. The most suitable source of natural water is rainwater. The water taken from natural water deposits can also be purified by removing bacteria, mud and clay. It then can be pumped through a city's water network. In everyday life, water is used for household purposes such as drinking and cleaning. Furthermore, a human body needs 3.5 liters of water to drink per day. In industry, water has many applications such as for cleaning, cooling, the dissolving of substances, and the production of steam.

Systems	% of water by mass
Seas, lakes and rivers	More than 95%
The human body	About 70%
Trees	Up to 60%
Rocks (marble)	Up to 20%
Air (in summer time)	About 1%

Properties of Water	
Freezing point	0°C (at 1 atm)
Boiling point	100°C (at 1 atm)
Density	1 g/cm^3 (at 4°C)



The Water Cycle

An electrified object attracts a stream of water.

2. SOLUBILITY

When you add a spoonful of sodium chloride (table salt) to a glass of water, it dissolves rapidly. As you continue to add more salt, however, there comes a point when it no longer dissolves. Instead, it collects at the bottom of the glass, even after stirring. When a solution contains as much solute as will dissolve, we say it is **saturated**. When the solution has some sodium chloride in it but can still dissolve more, it is said to be an **unsaturated** solution. Such solutions continue to dissolve more solute up to the point of saturation. For instance, 100 g of water can dissolve up to 13.7 g of potassium dichromate ($K_2Cr_2O_7$). No more potassium dichromate can be dissolved in 100 g of water at 20°C. Therefore, 100 g of water and 13.7 g of potassium dichromate constitute a saturated solution. If the amount of potassium dichromate is less than 13.7 g in 100 g of water, it is an unsaturated solution (Figure 8).

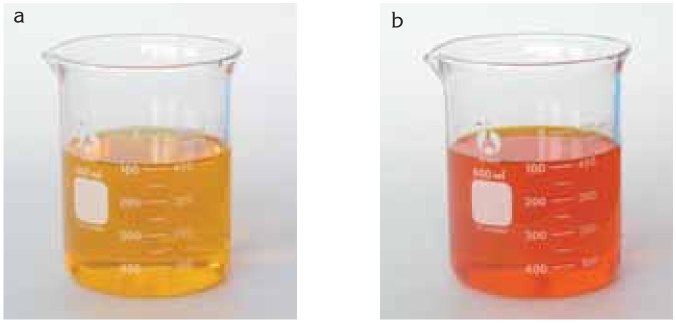


Figure 8 *a. Unsaturated potassium dichromate solution*
b. Saturated potassium dichromate solution

The **solubility** of a compound is the maximum amount solute dissolves in a given amount of solvent to form a saturated solution at a particular temperature. Solubility data are generally reported in units of grams of solute per 100 g of water. Each solid has a different solubility in water. For example, the solubility of sodium chloride is 36 g/100 g of water and of sugar is 204 g/100 g of water at 20°C.

The maximum amount of some substances that can be dissolved in 100 g of water at 100°C,
Top row (from left to right)
83 g $(NH_4)_2Cr_2O_7$
103 g $K_2Cr_2O_7$
203 g $CuSO_4$
Bottom row (from left to right)
39 g NaCl
79 g K_2CrO_4
111 g NaOH

The solubility of table salt (NaCl) is 36 g/100 g of water at 20°C.

The solubility of sugar is 204 g/100 g of water at 20°C.

The solubilities of ionic salts have wide range. For instance, in contrast to silver perchlorate ($AgClO_4$), which has a solubility of 55.7 g per 100 g of water, only 0.00018 g of silver chloride (AgCl) can dissolve in 100 g of water. If the maximum amount of solute dissolved in 100 g of water is less than 0.1 g, this solute is said to be **insoluble**. The solute that has a solubility range from 0.1 g to 1 g is called **slightly soluble**. If the amount of solute is more than 1 g, then it is **soluble**. Silver perchlorate is a soluble compound but silver chloride is an insoluble compound in water. It is not necessary to memorize the solubilities of substances. Solubility tables (e.g.,Table 2) help you determine whether a substance is soluble, insoluble, or slightly soluble in water.

Silver perchlorate ($AgClO_4$) is a soluble salt. It dissolves very well in water.

Silver chloride (AgCl) is an insoluble salt. It dissolves in water only in trace amounts.

SUBSTANCES	SOLUBLE	SLIGHTLY SOLUBLE	INSOLUBLE
NO_3^- (Nitrate), CH_3COO^- (Acetate), ClO_3^- (Chlorate)	All	Only $KClO_3$	Only $Be(CH_3COO)_2$
SO_4^{2-} (Sulfate)	Most	Ca^{2+}, Ag^+	Ba^{2+}, Sr^{2+}, Pb^{2+}
S^{2-}(Sulfide), CO_3^{2-} (Carbonate), SiO_3^{2-} (Silicate), PO_4^{3-} (Phosphate)	Only K^+, Na^+, NH_4^+		most
Cl^- (Chloride), Br^- (Bromide)	Most	Pb^{2+}	Ag^+, Hg^+
I^- (Iodide)	Most		Ag^+, Hg^+, Pb^{2+}
CrO_4^{2-} (Chromate)	Most	Hg^+	Ag^+, Ba^{2+}, Pb^{2+}
Hydroxide (OH^-)	Group IA , NH_4^+, Ba^{2+}	Ca^{2+}, Sr^{2+}	most
Hydrogen (H^+)	All		H_2SiO_3

Table 2 *Solubility Table*

2.1. FACTORS AFFECTING SOLUBILITY

The solubilities of substances do not depend only on the amount of solute or solvent. Temperature and pressure are also factors that affect solubility. As for stirring, even if it seems to increase solubility, it actually has no effect. Stirring only speeds up dissolution.

Temperature

Solubility varies with temperature. The solubilities of solids usually increase as the temperature rises. For example, more sugar dissolves in hot coffee than in cold coffee. Table 3 shows the effect of temperature on the solubility of sugar.

Temperature (°C)	0	10	20	25	30	40	45	50
Solubility (g of sugar /100 g H_2O)	179	190	204	211	219	238	248	260

Table 3 *The solubility of sugar at different temperatures*

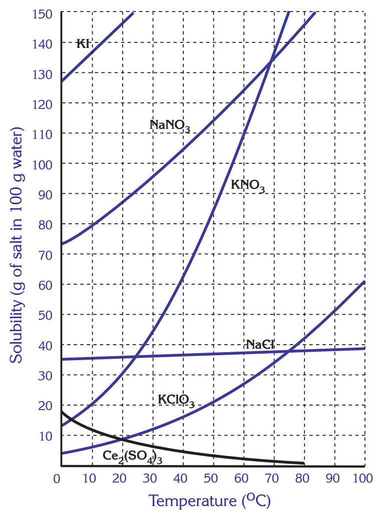


Figure 9 *The solubilities of most solids increase as temperature increases.*

Although most solids are more soluble at higher temperatures than at lower temperatures, there are some exceptions. For instance, more cerium(III) sulfate, $Ce_2(SO_4)_3$, dissolves in cold water than in hot water (Figure 9).

Gases, unlike solids, are more soluble in liquids at lower temperatures (Figure 10). This fact explains why a cold glass of cola goes "flat" upon warming, why fish seek deep and shaded places during the hottest days of summer, and why dissolved air is released as water is heated even at temperatures below the boiling point. The solubility of carbon dioxide in cola is shown below. If two bottles of cola, one cold and the other warm, are opened, the molecules of the carbon dioxide leave the warm bottle more rapidly.

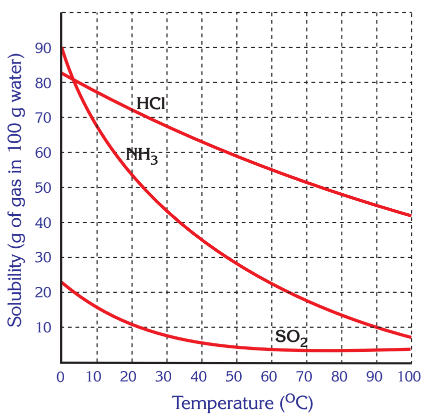


Figure 10 *The solubilities of most gases decrease with increasing temperature.*

A bottle of cola is placed on ice and another is placed in hot water. One bottle gets colder on ice and the other gets warmer in hot water.

When the bottles are opened, almost all the CO_2 in the cold cola remains in the solution, but the CO_2 gas rapidly escapes from the warm cola.

Pressure

As pressure increases, the solubilities of solids and liquids do not change much, but the solubilities of gases increase. A higher pressure above a liquid means more of the gas dissolves (Figure 11). A pressurized container for shaving cream works on this principle. Pressing a valve reduces the pressure on the dissolved gas, causing it to rush from solution carrying liquid with it as a foam. Why cola 'fizzes' in a glass and what causes the 'bends' in divers can also be explained by the effect of pressure on the solubilities of gases in liquids. (See the reading : The Bends, below).

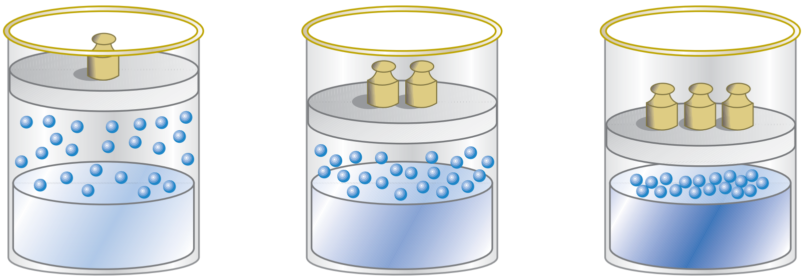


Figure 11 *As the partial pressure of the gas increases, the solubility of the gas increases.*

EXTRA

Preparation of Cola

Carbonated soft drinks like cola are generally bottled under 4 atm pressure by dissolving carbon dioxide (CO_2) gas in a soft drink solution.

When the bottle is opened, the pressure decreases to 1 atm. As a result, the carbon dioxide gas in the cola bubbles rapidly out of the solution, causing effervescence or fizzing.

READING — THE BENDS

The "bends" first recorded in 1841, is also known as decompression sickness. It is a very serious and potentially lethal condition. The bends occurs when there is a rapid and great change in blood pressure. Deep sea divers are especially vulnerable to this painful and sometimes fatal condition. There is a higher pressure environment under vast amounts of water such as in a sea or ocean. When a diver is at a significant depth, high pressure causes nitrogen to be absorbed by the fatty tissues in his or her body. If the diver ascends too quickly and the pressure drops rapidly, this liquid nitrogen in the body rapidly turns into bubbles. This event is just like uncorking a bottle of soda. These bubbles affect the nervous system by restricting the flow of blood and cause the "bends." Severe pain develops in the muscles and joints of the arms and legs. More severe symptoms include vertigo, nausea, vomiting, choking and sometimes death. To avoid the bends, the diver must rise slowly and make intermittent stops on the way up in order not to be exposed to the effects of rapidly increasing pressure.

Supersaturated Solutions

Can it be possible to dissolve more solute in an already saturated solution? You may consider it impossible. Cooling a saturated solution usually causes a solid to crystallize out of the solution. Sometimes, however, this does not happen. The excess solute stays in the solution upon cooling. This type of solution is known as supersaturated. For example, 161 g of sodium acetate ($NaCH_3COO$) can be dissolved in 100 g of water at 90°C. When this saturated solution is carefully cooled to 20°C, at which temperature the solubility of sodium acetate is 123.5 g/100 g of H_2O, all the solute may remain in the solution. The cooled solution contains more solute than it normally would. This is referred to as a **supersaturated solution**. Such a solution is unstable, and the excess dissolved solute (161 - 123.5 = 37.5 g) may crystallize by the addition of a seed crystal of the solute to the supersaturated solution (Figure 12) or of supersaturated solution onto the seed crystal of the solute (Figure 13).

Figure 13 *Crystallization of sodium acetate supersaturated solution*

Figure 12 *When a small seed crystal of sodium acetate is added to this supersaturated solution, the excess salt quickly crystallizes.*

Honey is natural, but jam and syrup are artificial supersaturated solutions in which sugar is the main solute. For example, syrup, a highly concentrated solution of sucrose (table sugar) in water, is made by heating the solution to almost the boiling point. The crystallization of the excess solute, sugar, is a common problem for these substances.

Is Your Honey Natural?

Honey is an example of a supersaturated solution. It is composed of mainly grape sugar (glucose, 35%), fruit sugar (fructose, 41%), and water (17%). Given the ratio of sugar to water, honey can, therefore, be considered as a highly concentrated solution of sugars. Sugars and water together constitute nearly 93% of honey. The remaining 7% is comprised of proteins, acids, and some elements like iron, sodium, sulfur, magnesium, and phosphorus.

The difference between honey and ordinary sugar is that honey is absorbed by the blood without digestion while ordinary sugar is absorbed by the blood after digestion. Honey is a healthy product which provides energy very soon after consumption.

Nevertheless, the production of artificial honey has increased recently. So, we must be careful when buying honey. Whether honey is natural or artificial can be identified easily. When both are tested in a flame, artificial honey forms a black substance, carbon, but natural honey does not.

Example 5

The solubility of potassium iodide is 136 g/100 g water at 10°C. How many grams of water is needed to dissolve 204 g of potassium iodide at the same temperature?

Solution

This question can be calculated by using a simple proportion.

If 136 g of KI dissolve	in 100 g of water
204 g of KI dissolve	x

$x = 100 \cdot 204/136 = 150$ g

150 g of water is needed to dissolve 204 g of potassium iodide at 10°C.

Exercise 4

How many grams of potassium iodide can be dissolved in 300 g of water at 10°C ?

(Clue : the solubility of KI at 10°C is given in Example 5 above).

Answer: 408 g

Example 6

How do the following processes affect the solubility of carbon dioxide gas in water?

a. Increasing pressure **b.** Decreasing temperature

Solution

a. As pressure increases, the solubility of carbon dioxide gas also increases. So, larger amounts of carbon dioxide gas can be dissolved at higher pressures.

b. Temperature and the solubility of gases are inversely proportional. Therefore, decreasing the temperature increases the solubility of carbon dioxide gas in water.

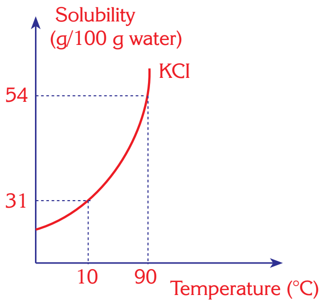
Exercise 5

How do the following processes affect the solubility of potassium iodide in water?

a. Decreasing pressure

b. Increasing temperature

c. Stirring the mixture



Example 7

A 262 g saturated solution of potassium chloride is prepared at 10°C. If the temperature of the solution were increased to 90°C, how many grams of potassium chloride would be needed to make the solution saturated?

Solution

According to the graph, 100 g of water can dissolve 31 g of KCl at 10°C, and a 131 g of saturated solution can be prepared. Then,

if 131 g of saturated solution contains	31 g of KCl
262 g of saturated solution contains	x

x = 62 g of KCl

The amount of water in this solution is

262 – 62 = 200 g

When the solution is heated up to 90°C, the amount of KCl needed to make a saturated solution with 200 g of water can be calculated as follows:

if 100 g of water can dissolve	54 g of KCl at 90°C
200 g of water can dissolve	x

$$x = \frac{200 \cdot 54}{100} = 108 \text{ g of KCl}$$

Then, the amount of KCl needed is

108 – 62 = 46 g

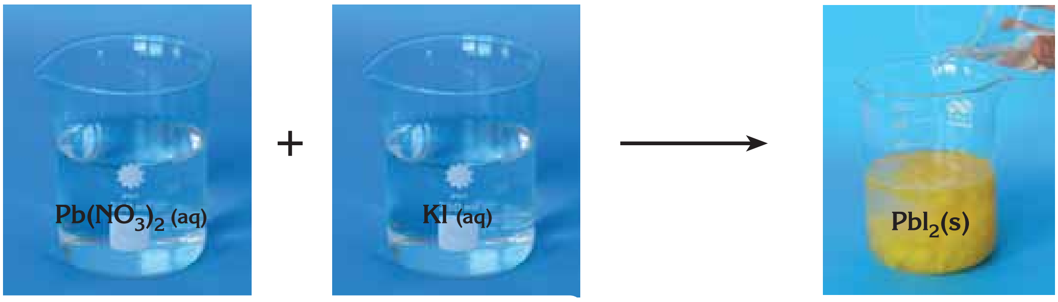
Exercise 6

A 462 g saturated solution of potassium chloride is prepared at 90°C. If the solution were cooled to 10°C, how many grams of potassium chloride could crystallize out of the solution? (Use the graph given in Example 7).

Answer: 69 g

3. MIXING AQUEOUS SOLUTIONS

When the aqueous solutions of two water-soluble compounds are mixed, there may be a reaction between the ions of these solutions. If one of the products is insoluble, crystals of this product fall from the resulting solution. This solid product is called a **precipitate**. Let us consider the reaction between the solutions of lead(II) nitrate, $Pb(NO_3)_2$, and potassium iodide, KI. According to the solubility table in Appendix A, both are soluble in water. This means that the solution of lead(II) nitrate contains Pb^{2+} and NO_3^- ions, and the potassium iodide solution contains K^+ and I^- ions. The possible products of this reaction are PbI_2 and KNO_3. According to solubility rules, potassium nitrate is soluble in water, but lead(II) iodide is not. As soon as the two reactants mix, insoluble lead(II) iodide crystals settle at the bottom of the container as a yellow precipitate.



$$Pb(NO_3)_2(aq) \quad + \quad 2KI(aq) \longrightarrow 2KNO_3(aq) \quad + \quad PbI_2(s)$$

$$Pb^{2+}(aq) + 2NO_3^-(aq) \quad + \quad 2K^+(aq) + 2I^-(aq) \longrightarrow 2K^+(aq) + 2NO_3^-(aq) \quad + \quad PbI_2(s)$$

Net Ionic Equation

A net ionic equation is an equation that includes only the actual participants in a reaction. To write a net ionic equation, first we separate all soluble salts into ions.

*The ions that are present on both sides of a reaction are called **spectator ions**.*

$$Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow 2K^+(aq) + 2NO_3^-(aq) + PbI_2(s)$$

Then, we eliminate the same ions (spectator ions) present on both sides of the equation:

$$Pb^{2+}(aq) + \cancel{2NO_3^-}(aq) + \cancel{2K^+}(aq) + 2I^-(aq) \rightarrow \cancel{2K^+}(aq) + \cancel{2NO_3^-}(aq) + PbI_2(s)$$

This yields the net ionic equation:

$Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$

Such a net ionic equation tells us that Pb^{2+} and I^{-} ions cannot both be found as ions in the same solution.

Suppose we mix solutions of potassium chloride and ammonium nitrate, and we are asked whether precipitation occurs.

$KCl + NH_4NO_3 \longrightarrow ?$

Let us rewrite the expression in its ionic form.

$K^{+} + Cl^{-} + NH_4^{+} + NO_3^{-} \longrightarrow ?$

The solubility rules show that all the possible ion combinations (for example, KNO_3 and NH_4Cl) yield water-soluble compounds. Thus, no reaction occurs and all four ions remain in the solution (Figure 14). This process is indicated as below.

$KCl + NH_4NO_3 \longrightarrow$ no reaction

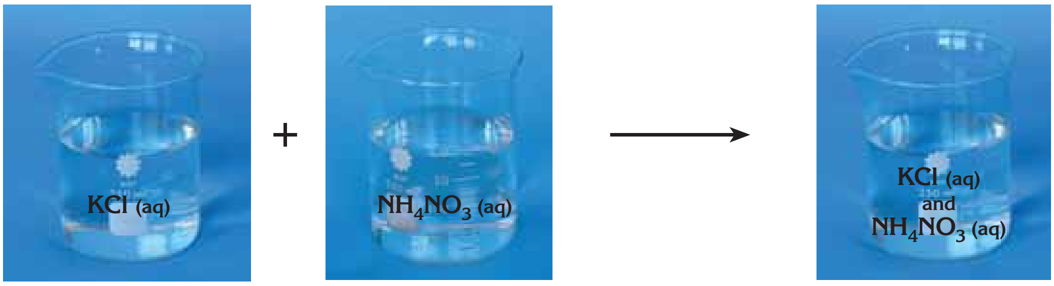


Figure 14 *Potassium chloride and ammonium nitrate solutions do not react.*

EXTRA

Colored Solutions

Salts are compounds composed of a metal ion bonded to a nonmetal ion. Their solutions may have different colors. For example, the salt solutions containing copper ions (Cu^{2+}) are usually blue, and those containing nickel ions (Ni^{2+}) are **pale green**. If a solution contains iron ions (Fe^{2+} or Fe^{3+}), it may be green or orange, white cobalt solutions (Co^{2+}) are pink.

The Application of Precipitation

One laboratory use of precipitation reactions is to determine the presence of certain ions in a solution. An illustrative example is given below. The four colorless solutions containing chloride (Cl^-), iodide (I^-), sulfide (S^{2-}), and nitrate (NO_3^-) ions can be identified by using only a silver nitrate solution. A more detailed ion analysis is given in Appendix B.

In industry, precipitation reactions are used in the manufacture of many chemicals. For example, the first step in the extraction of magnesium from sea water is to precipitate Mg^{2+} as $Mg(OH)_2(s)$.

AgCl(s)

AgI(s)

$Ag_2S(s)$

Anion	Result of Adding $AgNO_3$ Solution
Chloride, Cl^-	a white precipitate, AgCl
Iodide, I^-	a yellow precipitate, AgI
Sulfide, S^{2-}	a black precipitate, Ag_2S
Nitrate, NO_3^-	no precipitate

EXTRA

Metathesis Reactions

Metathesis reactions, sometimes called double displacement reactions, have the general form of

$$AB + CD \longrightarrow AD + CB$$

They occur not only when a precipitate is formed, but also when an insoluble gas or a weak electrolyte is formed. An acid-base neutralization reaction between sodium hydroxide and hydrochloric acid is an example.

The overall reaction : $NaOH + HCl \longrightarrow NaCl + H_2O$

Ionic equation : $Na^+ + OH^- + H^+ + Cl^- \longrightarrow Na^+ + Cl^- + H_2O$

Net ionic equation : $H^+ + OH^- \longrightarrow H_2O$

The reaction between sodium carbonate and hydrochloric acid is also illustrated below.

The overall reaction : $Na_2CO_3 + 2HCl \longrightarrow 2NaCl + CO_2 + H_2O$

Ionic equation : $2Na^+ + CO_3^{2-} + 2H^+ + 2Cl^- \longrightarrow 2Na^+ + 2Cl^- + CO_2 + H_2O$

Net ionic equation : $CO_3^{2-} + 2H^+ \longrightarrow CO_2 + H_2O$

Example 8

a. Write a balanced equation for the reaction that occurs when potassium chromate solution is mixed with silver nitrate solution.

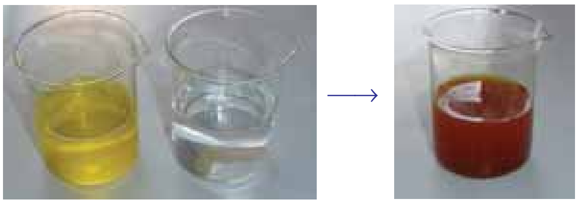
b. Write the net ionic equation for the same reaction.

Solution

a. When these solutions are mixed, there will be four ions: K^+, CrO_4^{2-}, Ag^+, and NO_3^- ions.

The solubility table shows us that Ag^+ and CrO_4^{2-} ions form an insoluble salt, Ag_2CrO_4, but that $NaNO_3$ is soluble. The equation, therefore, is

$$2AgNO_3(aq) + K_2CrO_4(aq) \longrightarrow Ag_2CrO_4(s) + 2KNO_3(aq)$$



When yellow potassium chromate and colorless silver nitrate solutions are mixed, a red precipitate, silver chromate, results.

b. The net ionic equation can be written from the balanced reaction. First, all soluble salts should be separated into ions as follows:

$$2Ag^+ + 2NO_3^- + 2K^+ + CrO_4^{2-} \longrightarrow Ag_2CrO_4(s) + 2K^+ + 2NO_3^-$$

Then, the ions present on both sides of the equation are canceled, which gives the net ionic equation.

$$2Ag^+(aq) + CrO_4^{2-}(aq) \longrightarrow Ag_2CrO_4(s)$$

Exercise 7

a. Write a balanced equation for the reaction that occurs when sodium hydroxide solution is mixed with copper(II) chloride solution.

b. Write the net ionic equation for the same reaction.

4. HYDROLYSIS

When a salt dissolves in water, the cation and anion separate from each other. Some of these ions may react with water and change the neutrality of water. The interaction of an ion with water is called **hydrolysis**. Not all ions undergo a hydrolysis reaction.

* The cations of strong bases (Li^+, Na^+, K^+, Ba^{2+}) and anions of strong acids (Cl^-, Br^-, I^-, NO_3^-) do not react with water (neutral solution).

* The cations derived from weak bases (NH_4^+, Cu^{2+}, Fe^{2+}) react with water to produce hydronium ion, H_3O^+, (acidic solution).

* The anions derived from weak acids (F^-, CH_3COO^-, CO_3^{2-}, CN^-) react with water to produce hydroxide ion, OH^-, (basic solution).

Now, let us identify NaCl, $NaCH_3COO$, and NH_4Cl solutions as acidic, basic, or neutral.

When sodium chloride dissolves in water, it produces sodium and chloride ions.

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

Na^+ is a cation of a strong base, NaOH, and Cl^- is an anion of a strong acid, HCl. Thus, neither ion hydrolyzes. Hence, the sodium chloride solution is neutral.

Sodium acetate solution contains sodium and acetate ions. Na^+ is derived from a strong base, NaOH, and CH_3COO^- is derived from a weak acid, CH_3COOH. Thus, only acetate ions (CH_3COO^-) react with water to produce an hydroxide ion.

$$Na^+ + H_2O \longrightarrow \text{no reaction}$$

$$CH_3COO^-(aq) + H_2O(l) \longleftrightarrow CH_3COOH(aq) + OH^-(aq)$$

The resulting hydroxide ion means that the sodium acetate solution is basic. It shows basic properties. For example, the solution of sodium acetate turns litmus paper blue (Figure 15).

Figure 15 *The solution of sodium acetate ($NaCH_3COO$) is basic.*

An ammonium chloride solution contains ammonium and chloride ions. Since Cl^- is an anion of a strong acid, HCl, it does not undergo a hydrolysis reaction. NH_4^+ is a cation of a weak base, NH_4OH. Therefore, it hydrolyzes to produce a hydronium ion.

$$Cl^- + H_2O \longrightarrow \text{no reaction}$$

$$NH_4^+(aq) + H_2O(l) \longleftrightarrow NH_3(aq) + H_3O^+(aq)$$

The resulting hydronium ion (H_3O^+) shows that the ammonium chloride solution is acidic. It shows acidic properties. It turns litmus paper red (Figure 16).

Figure 16 *Ammonium chloride (NH_4Cl) solution is acidic.*

In summary,

* The salts of strong bases and strong acids (for example, NaCl) do not hydrolyze. Their solutions are neutral.
* The salts of weak bases and strong acids (for example, NH_4Cl) hydrolyze. Their solutions are acidic.
* The salts of strong bases and weak acids (for example, $NaCH_3COO$) hydrolyze. Their solutions are basic.

Example 9

Predict whether the solutions of the following salts are acidic, basic or neutral.

a. $CuCl_2$

b. $NaNO_3$

c. LiCN

Solution

a. $CuCl_2$ is a salt produced from the reaction of a weak base, $Cu(OH)_2$, and a strong acid, HCl. Only the ion derived from the weak base, Cu^{2+}, hydrolyzes. Thus, the solution of $CuCl_2$ is acidic.

b. The solution of $NaNO_3$ contains sodium and nitrate ions. Both ions are derived from a strong base or a strong acid. Therefore, salt does not hydrolyze. It is a neutral salt.

c. LiCN is produced from the reaction of a strong base, LiOH, and a weak acid HCN. It hydrolyzes and produces a basic solution.

Exercise 8 Predict whether the solutions of the following salts are acidic, basic, or neutral.

a. KF

b. NH_4I

c. $Cu(NO_3)_2$

SUPPLEMENTARY QUESTIONS

Part A

1. Complete the statements below as a way of summarizing this chapter. The missing words can be found in the word list below. There may be more words in the list than necessary.

Word list

temperature	electrolyte	ethylene glycol
pressure	stirring	raises
saturated	unsaturated	lowers
heat	cold	exothermic
endothermic		

a. A solution that conducts electricity is called a(n) ....

b. .... is added to water and the resulting mixture used as an antifreeze in machines.

c. The solubility of gases is directly proportional to ....

d. .... a solution increases the rate of the dissolution process, but does not affect the solubility.

e. Salt is added on roads in winter because it .... the melting point.

f. A solution in which more solute can be dissolved is called a(n) .... solution.

g. A .... pack is used to treat shoulder pain and injures.

h. The process in which energy is given out is described as ....

2. Decide whether the following statements are true (T) or false (F).

a. The solute in a solution is always solid. ....

b. As a solution is diluted, its density does not change. ....

c. If there are no ions in a solution, it is called a weak electrolyte. ....

d. Solutions are homogeneous mixtures. ....

e. An accumulator is a device in which a dilute solution of sulfuric acid is used as an electrolyte. ....

Part B

1. Decide whether the following mixtures are solutions or not.

a. Salt + water **b.** Sugar + water
c. Oil + water **d.** HCl + water
e. Carbon dioxide + water **f.** Bronze (Cu + Zn + Sn)
g. Water **h.** Air

2. For each of the following solutions, identify the solvent and the solute.

a. Steel
b. 10 mL of alcohol mixed with 5 mL of water
c. 150 g of sugar dissolved in 100 g of water

3 Write the equations for the dissolution processes of the following substances in water.

a. $NaCl$ **b.** C_2H_5OH **c.** $Ca(NO_3)_2$ **d.** HCl

4. Hydrogen chloride gas dissolves in water very well, but iodine crystals do not dissolve in water. How can you explain this?

5. When there is ice on the roads in winter, salt is generally added to the ice. Explain why this is done.

6. Ethylene glycol is mixed with water and is used in the radiators of machines. Explain the importance of using ethylene glycol.

7. Classify each of the following substances as either strong electrolytes, weak electrolytes, or nonelectrolytes.

a. Acetic acid solution **b.** Sugar solution
c. Ammonia solution **d.** Sodium chloride solution
e. Alcohol solution **f.** Hydrochloric acid solution
g. Sodium hydroxide solution

8. In three different beakers, there are solutions of table salt, table sugar, and acetic acid. How can you identify them without tasting?

9. The solubility of table salt, NaCl is 37 g/100 g of water at 20°C. How many grams of table salt can be dissolved in 250 g of water at the same temperature?

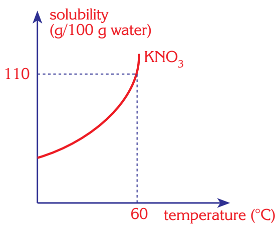
10. In two different test tubes there are two solids, sodium chloride and silver chloride. How can you identify them? (Clue : use solubility table)

11. The solubility of potassium nitrate is 64 g/100 g of water at 40°C. A solution is prepared by dissolving 50 g of potassium nitrate in 150 g of water at 40°C.

a. Is the prepared solution saturated?

b. If not, how many grams of potassium nitrate must be added to make the solution saturated?

12. Using the graph, answer the questions below.



a. How many grams of KNO_3 are dissolved in a 525 g saturated solution at 60°C?

b. 220 g of KNO_3 is dissolved in 250 g of water at 60 °C. How many grams of water should be evaporated from the solution to make the solution saturated?

13. Explain how temperature and pressure affect the solubility of carbondioxide gas in soft drinks.

14. Does stirring affect the solubilities of solids in water?

15. Explain

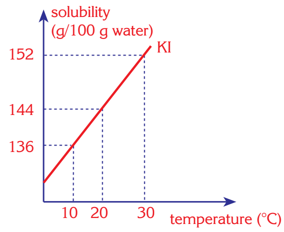
a. What causes the 'bends' in divers.

b. Why it is difficult for fish to survive during the hottest days of summer.

c. Why sugar is more soluble in hot coffee than in cold one.

16. How many grams of silver nitrate crystallize if 73.5 g of saturated silver nitrate solution at 80°C is cooled to 20°C? (The solubility of silver nitrate is 635 g/100 g H_2O at 80°C and 228 g/100 g H_2O at 20°C).

17. A 126 g of saturated solution of potassium iodide is prepared at 30°C. If it were cooled to 10°C, how many grams of potassium iodide would crystallize?



18. Answer the question below for each of these pairs of solutions.

I. NaCl and $AgNO_3$ solutions

II. NaI and $AgNO_3$ solutions

III. $MgCl_2$ and NaOH solutions

IV. $BaCl_2$ and H_2SO_4 solutions

V. NaCl and KI solutions

a. Write the equations for the reactions that take place between the given pairs. If you think that the pairs do not give a reaction, give your reasoning.

b. Write the net ionic equations for the reactions you wrote in question **a**.

19. In three different test tubes, there are the solutions of potassium chloride, potassium iodide, and potassium sulfide solutions. How can you identify them by using only one reagent?

20. Predict whether the solutions of the following salts are acidic, basic, or neutral.

a. NH_4Br **b.** KI **c.** Na_3PO_4

MULTIPLE CHOICE QUESTIONS

1. Which one of the following substances is not a solution?

A) Air B) Sugar + water mixture C) Bronze D) Water vapor E) Wine

2. Which one of the following statements is wrong for solutions?

A) They are impure substances.
B) They are homogeneous mixtures.
C) They have definite melting and boiling points.
D They have two components, a solute and a solvent.
E) In salt water, salt is the solute.

3. Which of the below dissolve(s) as molecules in water?

I. Sugar
II. Table salt
III. Alcohol

A) I only B) III only C) I and II D) II and III E) I and III

4. Which one of the following solutions is not a strong electrolyte?

A) Hydrochloric acid B) Sodium hydroxide
C) Potassium nitrate D) Ethyl alcohol
E) Sulfuric acid

5. Which one of the alternatives below is the correct comparison of the boiling points of these substances?

I. Pure water
II. Dilute salt solution
III. Concentrated salt solution

A) I = II = III B) I > II > III C) III > II > I D) I > III > II E) III > I > II

6. The solubility of table sugar is 204 g/100 g of water at 20°C. How many grams of sugar can be dissolved in 25 g of water at the same temperature?

A) 51 B) 100 C) 102 D) 204 E) 304

7. Which one of the following is not true for solubility?

A) As pressure increases, the solubilities of gases increase.
B) Solubility depends on temperature.
C) Solubility is a characteristic property of substances.
D) Solubility increases by stirring a solution.
E) Solubility is the amount of solute dissolved in 100 g of water.

8. Which one(s) of the following procedures should be done to saturate an unsaturated salt solution?

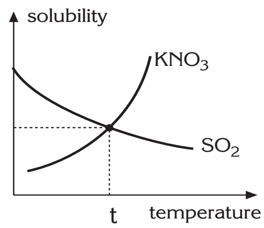
I. Adding salt
II. Adding water
III. Stirring the solution

A) I only B) II only C) I and III D) II and III E) I, II, and III

9. Which one of the following does not change during the dilution of a concentrated solution?

A) The density B) The boiling point
C) The amount of solvent D) The amount of solute
E) The amount of solution

10. The graph of solubility versus temperature for KNO_3 and SO_2 is shown in the graph. Which one of the following statements is wrong?



A) As temperature increases, the solubility of KNO_3 increases.
B) As temperature increases, the solubility of SO_2 decreases.
C) KNO_3 and SO_2 have the same solubility at the temperature (t).
D) If a saturated KNO_3 solution is heated, an amount of KNO_3 may fall out.
E) Temperature can affect the solubility of both solids and gases.

11. Which one of the following solutions does not conduct electricity?

A) Sea water B) Mineral water C) Distilled water D) Tap water E) King water (Aqua Regia)

12. Which one of the following salts gives a hydrolysis reaction?

A) $NaNO_3$ B) NH_4Cl C) KI D)LiCl E) $BaSO_4$

13. In which of the following actions does a dissolution process not occur?

A) The addition of table salt into water
B) The dropping of ethyl alcohol into water
C) The addition of olive oil into water
D) The removal of nail polish by acetone
E) The mixing of paint with petroleum ether

14. Which one of the following is the most soluble in water?

A) Ethyl alcohol B) Table salt C) Table sugar D) Carbon dioxide E) Oil

15. Solubilities of gases increase as pressure increases but decrease as temperature increases. The solubility of a gas at certain pressures and temperatures is given below.

Pressure	Temperature	Solubility
P	3T	S_1
2P	2T	S_2
3P	T	S_3

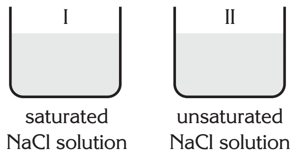
Which one of the following comparisons given below shows the correct relation among S_1, S_2, and S_3?

A) $S_1 > S_2 > S_3$ B) $S_3 > S_1 > S_2$
C) $S_1 = S_2 = S_3$ D) $S_2 > S_1 > S_3$
E) $S_3 > S_2 > S_1$

16. In which of the following does a reaction not occur? (Use the solubility table)

A) $MgCl_2(aq) + NaOH(aq) \longrightarrow$
B) $BaCl_2(aq) + H_2SO_4(aq) \longrightarrow$
C) $KI(aq) + AgNO_3(aq) \longrightarrow$
D) $NaCl(aq) + AgNO_3(aq) \longrightarrow$
E) $LiNO_3(aq) + NH_4Cl(aq) \longrightarrow$

17. Two beakers contain sodium chloride solutions at the same temperature, one saturated and the other unsaturated.



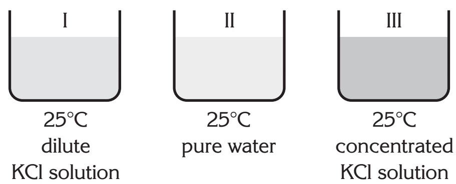
Which comparison is true for these solutions?

A) density : II > I
B) freezing point : II > I
C) electrical conductivity : II > I
D) boiling point : II > I
E) none of these

18. Which of these solutions is/are electrolytes ?
I. $(NH_4)_3PO_4(aq)$ II. $C_6H_{12}O_6(aq)$ III. NaCl(aq)

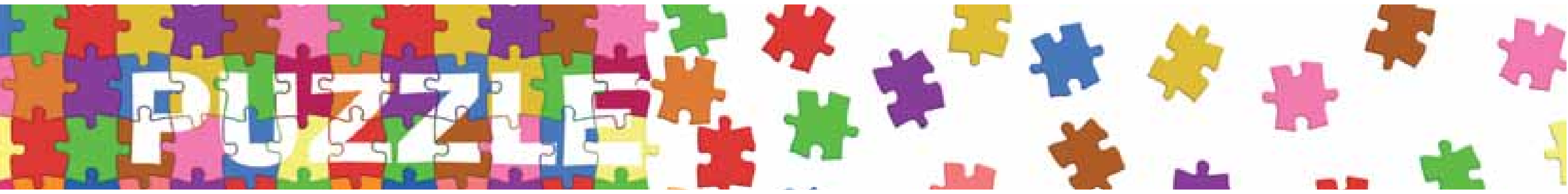
A) II only B) III only C) I and II D) I and III E) I, II, and III

19. Which of the comparison(s) below is/are true for these substances?

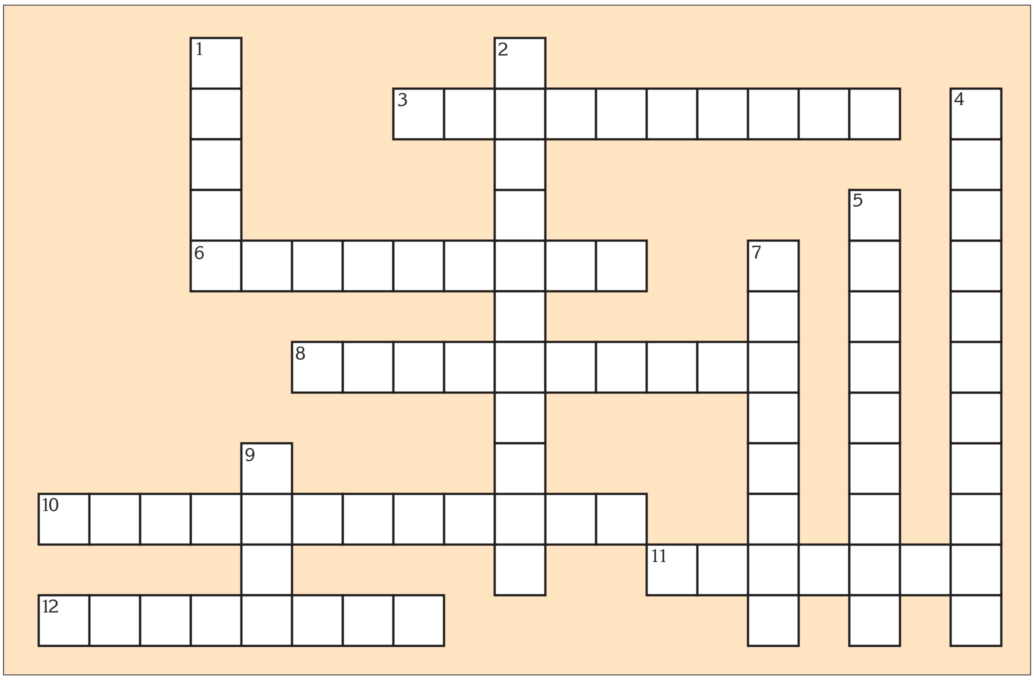


I. The boiling points : III > I > II
II. The freezing points : II > I > III
III. The electrical conductivity : III > I > II

A) I only B) I and II C) I and III D) II and III E) I, II, and III



CRISS CROSS PUZZLE



ACROSS

3 Maximum amount of solute that can be dissolved in 100 g of water at a certain temperature.

6 Ions present on the both sides of an equation.

8 Reaction of an ion with water.

10 Solutions that contain a relatively high concentration of solute.

11 Substance present in the largest amount or that determines the state of solution.

12 Homogeneous mixture composed of two or more pure substances.

DOWN

1 Problem faced by divers due to the change in pressure.

2 Solution that conducts electricity.

4 Insoluble substance that falls from a solution when two aqueous solutions are mixed.

5 Solution that contains as much solute as will dissolve.

7 Liquids that mix in all proportions.

9 Kind of pack used to give off heat.

NOTES

MEASURING CONCENTRATION

SOLUTIONS

INTRODUCTION

We have discussed the terms saturated, unsaturated, dilute, and concentrated solutions in the previous chapter. Although these terms serve a useful purpose, we often need to know the exact quantity of solute present in a given amount of solution. Why knowing the exact concentration of a solution is so important? Why should we be interested in solution concentrations? The answer is that concentration plays an important role in many areas such as chemistry, biology, and medicine. By measuring concentration, you can understand whether air pollution has reached dangerous levels, whether water is safe to drink, and whether your body has enough vitamins and minerals. The concentrations of some important solutions and their uses are given below (Table 1). When the concentrations of these solutions change, unanticipated results may arise.

Solution	Concentration (%)	Usage
Hydrogen peroxide (H_2O_2)	3	Antiseptic
Acetic acid (CH_3COOH)	5	Vinegar
Glucose ($C_6H_{12}O_6$)	5	Intravenous feeding
Ethyl alcohol (C_2H_5OH)	10	Wine
Isopropyl alcohol (C_3H_7OH)	65	Rubbing alcohol
Sodium hypochlorite (NaClO)	5	Household bleach
Carbonic acid (H_2CO_3)	1	Soft drinks

***Table 1** Concentrations of some important solutions used in daily life*

Each household cleaners contains certain amount of solutions.

Soft drinks are acidic solutions.

Sensitiveness in concentration is important not only in man-made products but also in the systems of living organisms. For instance, the concentration of oxygen in the atmosphere has a very critical value, 21% by volume. A decrease in the concentration of oxygen would make it difficult for us to survive. When the volume percentage of oxygen exceeds 21%, fires may start more easily. For example, the number of oxygen molecules per liter does not remain the same at high elevations; thus people need oxygen tanks to climb Mount Everest.


People need oxygen tanks to climb Mount Everest.

Scientists use different concentration units that describe the composition of a solution. Percent concentration, molarity, molality, normality, and parts per million (ppm) are some of the concentration units.

This chapter will examine two of these in detail: percent concentration and molarity.

1. PERCENT CONCENTRATION

Percent concentration is the simplest concentration unit. The amount of solute is compared to the amount of solution in order to measure concentration. This concentration unit is generally used for concentrated solutions of acids and bases. The percentage of solute can be expressed by mass or volume.

1.1. MASS PERCENT

One common way of describing a solution's composition is **mass percent**, otherwise known as weight percent or percent by weight. Mass percent expresses the mass of solute present in a given mass of solution. To calculate mass percent, the mass of solute is divided by the mass of the solution and multiplied by 100.

$$\text{Mass percent} = \frac{\text{mass of solute}}{\text{mass of solution}} \cdot 100$$

For example, let us prepare a solution by adding 10 g of table salt, NaCl, to 90 g of water. In this solution the mass of solute (table salt) is 10 g and the mass of solution (table salt and water) is 100 g (10 + 90 = 100 g).

The mass percent is,

$$\frac{10}{100} \cdot 100 = 10\%$$

Example 1

Glucose ($C_6H_{12}O_6$) is the main substance in a solution used for intravenous feeding. How many grams of glucose must be dissolved in water in order to prepare 500 g of 5% solution by mass?

A 5% glucose solution is used for intravenous feeding.

Solution

The mass percent and the mass of the required solution are known. The mass of solute can be found by the following equation:

$$\text{Mass percent} = \frac{m_{solute}}{m_{solution}} \cdot 100$$

$$5 = \frac{m_{glucose}}{500} \cdot 100 \quad \Rightarrow \quad m_{glucose} = 25 \text{ g}$$

This means 25 g of glucose must be dissolved in 475 g of water in order to prepare 500 g of 5% solution by mass.

Exercise 1

A solution is prepared by mixing 450 g of water and 50 g of sugar. What is the percent concentration of sugar by mass in the solution?

Answer : 10%

Exercise 2

Solder, an alloy consisting of lead and tin, is used to join metals. 70% by mass of solder is lead. Find the amounts of lead and tin in 400 g of solder.

Answer : m_{Pb} = 280 g and m_{Sn} = 120 g

The Preparation of Dilute Solutions

A solution with a high concentration is called a **stock** solution. Chemists have stock solutions in the laboratory. Adding solvent, mostly water, is one of the ways of obtaining a less concentrated solution. This process is called **dilution**. Chemists frequently use the dilution process because many common acids and bases are purchased by chemical companies as highly concentrated stock solutions (Table 2).

Great care must be taken when preparing acid solutions. According to laboratory safety rules, we must always add acid to water, never water to concentrated acid. The following example illustrates the dilution process of a concentrated acid solution.

Reagent	Formula	Concentration by mass (%)
Hydrochloric acid	HCl	36
Nitric acid	HNO_3	71
Sulfuric acid	H_2SO_4	96
Acetic acid	$HC_2H_3O_2$	99.5
Ammonia	NH_3	28
Sodium hydroxide	$NaOH$	50

Table 2 *Concentrations of some stock solutions*

Dilution of hydrochloric acid

Example 2

How many grams of water must be added to 100 g of 36% hydrochloric acid solution by mass in order to make a 10% solution?

Solution

First, let us find the mass of solute, hydrochloric acid, in the initial solution.

$$36 = \frac{m_{HCl}}{100} \cdot 100 \quad \Rightarrow \quad m_{HCl} = 36 \text{ g}$$

When water is added, the amount of solute does not change. The mass of solution becomes (100 + m) where m is the mass of the added water. Then, the mass percent formula is applied again.

$$10 = \frac{36}{100 + m} \cdot 100 \quad \Rightarrow \quad m = 260 \text{ g}$$

This means 260 g of water must be added.

Exercise 3

Household bleach is a 5% sodium hypochlorite ($NaClO$) solution by mass. How many grams of water must be added to 200 g of 5% sodium hypochlorite solution to obtain a 2% solution?

Answer : 300 g

Exercise 4

How many grams of salt should be added to 100 g of 20% salt solution in order to make it a 50% solution?

Answer : 60 g

Example 3

A 25 g sample of $CuSO_4 \cdot 5H_2O$ is dissolved in 55 g of water. What is the mass percentage of the resulting solution?

Solution

When copper(II) sulfate pentahydrate is dissolved in water, the result is copper (II) sulfate solution. The mass of this solution is 80 g (25 + 55 = 80).

The mass of solute can be calculated by proportion. The molar masses of $CuSO_4 \cdot 5H_2O$ and $CuSO_4$ are 250 g/mol and 160 g/mol, respectively.

250 g $CuSO_4 \cdot 5H_2O$ produces 160 g $CuSO_4$

25 g $CuSO_4 \cdot 5H_2O$ produces x

x = 16 g of $CuSO_4$

The mass of solute, $CuSO_4$, is found to be 16 g.

$$\text{Mass percent} = \frac{m_{solute}}{m_{solution}} \cdot 100$$

$$= \frac{16}{80} \cdot 100 = 20\%$$

The percent concentration of the solution is 20% $CuSO_4$ by mass.

$CuSO_4$ kills bacteria and microbes. It is used in purifying drinking water and preventing insects and fungi from attacking wood.

Exercise 5

How many grams of $FeSO_4 \cdot 7H_2O$ must be dissolved in water to obtain 304 g of 5% iron(II) sulfate ($FeSO_4$) solution?

Answer : 27.8 g

1.2. VOLUME PERCENT

Volume percent is similar to mass percent. It is generally preferred when working with solutions where all the components are liquids, such as antifreeze solution (ethylene glycol in water). It can be calculated by a formula similar to that of mass percent.

$$\text{Volume percent} = \frac{V_{\text{solute}}}{V_{\text{solution}}} \cdot 100$$

For example, to prepare a 50% antifreeze solution by volume, two liters of ethylene glycol should be mixed with enough water to make the final volume equal to four liters.

When two liquids are mixed, their volumes are not strictly additive. For example, when 30 mL of alcohol and 70 mL of water are mixed, the final volume is very close but not equal to 100 mL. However, we will ignore this difference here.

Example 4

Acetone (C_3H_6O) is the main substance found in nail polish removers. A solution is prepared by mixing 20 mL of acetone and 30 mL of isopropyl alcohol (C_3H_8O). Calculate

a. The volume percent of acetone in the solution,

b. The mass percent of acetone in the solution .
($\rho_{\text{acetone}} = 0.79$ g/mL, $\rho_{\text{isopropyl alcohol}} = 0.78$ g/mL)

Acetone (C_3H_6O), a solvent found in nail polish removers, will quickly dissolve a styrofoam cup.

Solution

a. Volume of the solution becomes 20 + 30 = 50 mL.

$$\text{Volume percent of acetone} = \frac{V_{\text{acetone}}}{V_{\text{solution}}} \cdot 100$$

$$= \frac{20}{50} \cdot 100 = 40\%$$

The volume percentage of acetone is 40%.

b. To calculate mass percent, the mass of each substance should be found by using the density formula

$$\rho = \frac{m}{V} \Rightarrow m = \rho \cdot V$$

Mass of acetone: $m_{\text{acetone}} = 20 \text{ mL} \cdot 0.79 \text{ g/mL} = 15.8 \text{ g}$

Mass of alcohol: $m_{\text{alcohol}} = 30 \text{ mL} \cdot 0.78 \text{ g/mL} = 23.4 \text{ g}$

$$\text{Mass percent of acetone} = \frac{m_{\text{acetone}}}{m_{\text{acetone}} + m_{\text{isopropyl alcohol}}} \cdot 100$$

$$= \frac{15.8}{15.8 + 23.4} \cdot 100 = \frac{1580}{39.2} = 40.31\%$$

The solution is 40.31% acetone by mass.

If 20 mL of acetone and 30 mL of alcohol are mixed, the result will be a 50 mL mixture. The mass of the mixture, however, will not be 50 g.

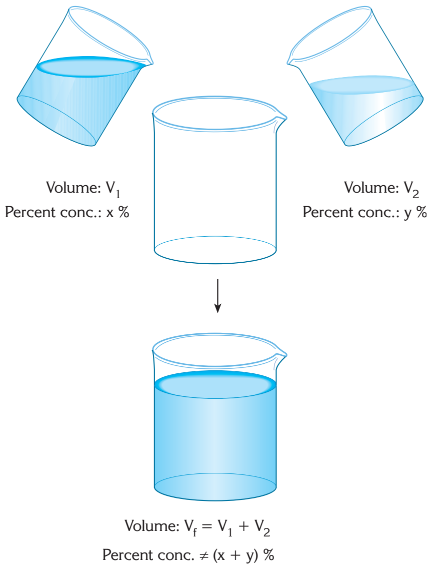
Exercise 6

What is the percent concentration of the solution, by volume, which is obtained by mixing 150 mL of alcohol and 450 g of water? The density of water is 1 g/mL)

Answer : 25%

Example 5

In a chemistry laboratory the concentration of a first nitric acid solution is 20% and that of a second nitric acid solution is 60% by volume. How many mL of each should be mixed to obtain 400 mL of a 30% nitric acid solution?



When two solutions (same) are mixed in a beaker, their volumes are additive, but percent concentrations are not.

Solution

Initially, we must determine the volume of nitric acid for the final solution.

$$30 = \frac{V_{acid}}{400} \cdot 100 \Rightarrow V_{acid} = 120 \text{ mL}$$

120 mL of HNO_3 is needed.

If a 400 mL of solution is obtained using x mL of the first solution, then $(400-x)$ mL of the second solution is needed.

$$V_1 \quad + \quad V_2 \quad = \quad V_f$$

$$\left(x \cdot \frac{20}{100}\right) + (400 - x) \cdot \frac{60}{100} = 120$$

$$0.2x + 240 - 0.6x = 120$$

$$0.4x = 120 \quad \Rightarrow \quad x = 300 \text{ mL}$$

This means 300 mL of 20% solution should be mixed with 100 mL of 60% solution.

Exercise 7

Equal volumes of two acetic acid solutions whose concentrations are 5% and 15% by volume are mixed. What is the percent concentration of the new solution by volume?

Answer : 10%

2. MOLARITY

Molarity is the most common concentration unit. It is generally used in calculations dealing with volumetric stoichiometry. Molarity can be defined as the mole number of solute dissolved per liter of solution. The abbreviation for molarity is M.

$$\text{Molarity} = \frac{\text{mole number of solute}}{\text{volume of solution in liters}} = \frac{\text{mol}}{\text{L}}$$

or

$$\boxed{M = \frac{n}{V}}$$

*Note that molarity refers to the mole number of solute per liter of **solution**, not per liter of solvent.*

The unit of molarity is mol/L or M (read as molar).

A sodium hydroxide solution that contains 1 mole of NaOH per liter of solution has a concentration of 1 molar, which is often abbreviated 1M. A more concentrated, 3M sodium hydroxide solution contains 3 moles of NaOH per liter of solution.

Keep in mind that molarity changes slightly with temperature because the volume of a solution changes with temperature.

Example 6

A water sample taken from "Lake Issyk-kul" contains 3.8 g of sodium chloride (NaCl) in one liter of solution. Find the molarity of sodium chloride in the sample.

Solution

First, we must calculate the mole number of NaCl, solute.

The molar mass (M) of sodium chloride is

$M = 23 + 35.5 = 58.5$ g/mol

Thus, the mole number of sodium chloride becomes

$$n = \frac{m}{M} = \frac{3.8\text{ g}}{58.5\text{ g/mol}} = 0.065\text{ mol}$$

Now, the molarity of solution can be calculated.

$$M = \frac{n_{solute}}{V_{solution}} = \frac{0.065}{1} = 0.065\text{ M}$$

The molar concentration of sodium chloride is 0.065 M.

The second largest crater lake of the world, Lake Issyk-kul, is located in Kyrgyzstan. It contains nearly 0.38% sodium chloride by mass.

Exercise 8

A 5.6 g sample of potassium hydroxide is dissolved in enough water to obtain 100 mL of solution. What is the molar concentration of the resulting solution?

Answer : 1 M

Potassium permanganate ($KMnO_4$) is used as an oxidizing agent, a disinfectant, and a preservative for fresh flowers and fruits.

Example 7

How many grams of potassium permanganate ($KMnO_4$) should be dissolved in water to prepare 100 mL of 0.25 M solution?

Solution

First, the mole number of $KMnO_4$ should be calculated by using the molarity equation. The volume of the solution must be measured in liters.

100 mL = 0.1 L

$$M = \frac{n_{solute}}{V_{solution}} \Rightarrow n = M \cdot V, \text{ and}$$

$$n_{KMnO_4} = 0.25 \cdot 0.1 = 0.025 \text{ mol.}$$

Then, the mass of $KMnO_4$ is calculated.

The molar mass of $KMnO_4 = 39 + 55 + 4 \cdot 16 = 158$ g/mol.

$$n = \frac{m}{M} \Rightarrow m = n \cdot M = 0.025 \cdot 158 = 3.95 \text{ g}$$

3.95 g of potassium permanganate should be dissolved.

Exercise 9

How many grams of sodium hydroxide are needed to prepare 500 mL of 0.2 M solution?

Answer : 4 grams

Example 8

Suppose that you are given a concentrated solution of HCl which is known to be 36.5% HCl by mass. If the density of the solution equals 1.2 g/mL, what is the molarity of the solution?

Solution

Let the volume of the solution be one liter (1000 mL). Then,

From the equation ($m = \rho \cdot v$)

$$m_{solution} = 1000 \text{ mL} \cdot 1.2 \text{ g/mL} = 1200 \text{ g}$$

$$m_{HCl} = 1200 \cdot \frac{36.5}{100} = 438 \text{ g}$$

The molar mass of HCl = 1 + 35.5 = 36.5 g/mol

The mole number can be calculated as

$$n_{HCl} = \frac{m}{M} = \frac{438\ g}{36.5\ g/mol} = 12\ mol$$

From the molarity formula,

$$M = \frac{n}{V} = \frac{12\ mol}{1\ L} = 12\ mol/L \text{ or } M$$

The molarity of the solution is 12 M.

Exercise 10

What is the molarity of the solution that is 63% nitric acid (HNO_3) by mass? (The density of the solution is 1.4 g/mL)

Answer : 14 M

Example 9

What is the molarity of a solution obtained by mixing 200 mL of 0.3 M ammonium nitrate solution with 400 mL of 0.6 M ammonium nitrate solution?

Solution

First, let us find the volume of the resulting solution. This volume is the sum of the volumes of both solutions.

$V_f = V_1 + V_2$ so $V_f = 200 + 400 = 600$ mL.

We know, then that the mole number of solute in the final solution is also the sum of mole numbers of solute in the first and second solutions.

$$n_f = n_1 + n_2$$

From this formula, it can be derived that

$$M_f \cdot V_f = M_1 \cdot V_1 + M_2 \cdot V_2$$

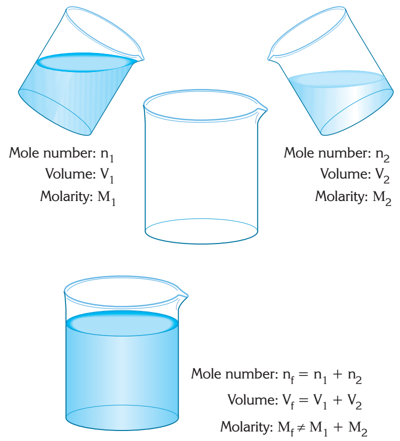
$M_f \cdot 0.6 = (0.3 \cdot 0.2) + (0.6 \cdot 0.4)$

$M_f = 0.5\ M\ NH_4NO_3$

The concentration of ammonium nitrate in the final solution is 0.5 mol/L.

Note: When more than two solutions are mixed, the following formula holds:

$$\boxed{M_f \cdot V_f = (M_1 \cdot V_1) + (M_2 \cdot V_2) + (M_3 \cdot V_3) + ... + (M_n \cdot V_n)}$$



When two solutions are mixed in a beaker, volumes and mole numbers are additive, but molarity is not.

Exercise 11

How many mililiters of 3 M and 1.5 M sulfuric acid solutions should be mixed in order to prepare 600 mL of 2 M sulfuric acid solution?

Answer : 200 mL of the 3 M solution and 400 mL of the 1.5 M solution.

3. THE PREPARATION OF A SOLUTION WITH A DESIRED CONCENTRATION

A **standard solution** is a solution whose concentration is accurately known. Chemists often need standard solutions for chemical reactions. They are prepared by using a calibrated volumetric flask.

Here, the preparation of 500 mL of 0.02 M potassium dichromate solution is explained step by step.

– Initially, the mole number of solute is calculated by using the molarity formula.

$$M = \frac{n}{V} \quad \Rightarrow \quad n = 0.02 \text{ mol/L} \cdot 0.5 \text{ L} = 0.01 \text{ mol}$$

The molar mass of $K_2Cr_2O_7$ = (2 · 39) + (2 · 52) + (7 · 16) = 294 g/mol.

Then, the mass of solute is calculated using the mole formula.

$$n = \frac{m}{M} \quad \Rightarrow \quad m = 0.01 \text{ mol} \cdot 294 \text{ g/mol} = 2.94 \text{ g}$$

– The calculated amount of solute (2.94 g) is weighed accurately (Figure 1a).

– The solute is added into a volumetric flask (Figure 1b).

– A small amount of water is added to dissolve the solute, and the flask is shaken until all the solute is dissolved (Figure 1c).

– More water is added to bring the level of solution to exactly the volume marked on the neck of volumetric flask (Figure 1d).

(a) *Weighing the amount of $K_2Cr_2O_7$*

(b) *Adding $K_2Cr_2O_7$ into a flask*

(c) *Dissolving $K_2Cr_2O_7$ with a smell amount of water*

(d) *Adding water to the volume marked on the neck of volumetric flask*

Figure 1 *Preparation of solution with desired concentration.*

Example 10

Limestone (calcium carbonate) can be removed from kettles by adding a dilute acid such as the acetic acid found in vinegar.

$CaCO_3(s) + 2CH_3COOH(aq) \rightarrow (CH_3COO)_2Ca(aq) + CO_2(g) + H_2O(l)$

What volume of 2 M acetic acid solution should be used to remove 10 g of calcium carbonate from a kettle?

Using acetic acid instead of hydrochloric acid would be better to remove limestone (calcium carbonate) from a kettle.

Solution

First, calculate the molar mass of $CaCO_3$.

$M_{CaCO_3} = 40 + 12 + (3 \cdot 16) = 100$ g/mol

Then, calculate the mole number of $CaCO_3$.

$$n = \frac{m}{M} = \frac{10\text{ g}}{100\text{ g/mol}} = 0.1\text{ mol.}$$

By using stoichiometry from the equation, 0.1 mol of $CaCO_3$ reacts with 0.2 mol of CH_3COOH.

$$\text{Finally, } M = \frac{n_{solute}}{V_{solution}}$$

$$2\text{ mol/L} = \frac{0.2\text{ mol}}{V_{solution}} \Rightarrow V_{solution} = 0.1\text{ L or } 100\text{ mL}$$

A 100 mL of acetic acid solution must be used.

EXTRA

Other Concentration Units

Molality and normality are other concentration units used in chemistry.

Molality is defined as the number of moles of solute dissolved in one kilogram of solvent. It is denoted as *m*. It is especially useful in calculating the freezing point depression and boiling point elevation.

Normality is the number of equivalence per liter of solution. It is denoted as N. Normality and molarity are related as follows:

$\boxed{N = M \cdot a}$ (a is called the mole factor)

The value of "a" is the number of H^+ ions for acids, the number of OH^- ions for bases, and the total number of positive charges for salts.

Compound	Mole Factor (a)
HCl	1
$Ba(OH)_2$	2
$Al_2(SO_4)_3$	6

Exercise 12

How many mL of 0.4 M sulfuric acid solution should be used to neutralize 8 g of sodium hydroxide?

Answer : 250 mL

A water sample that contains a high concentration of Ca^{2+} and Mg^{2+} ions is called "hard water".

Example 11

When the concentration of calcium ions reaches 120 mg in 1 L of a water sample, that water is considered to be "very hard" as shown in Table 3. If the molar concentration of calcium ions in a 1 L sample of water is 0.004 M, decide whether this sample is "very hard" or not.

The hardness of water	Concentration (in mg/L)
soft	0 - 20
moderately soft	20 - 40
slightly soft	40 - 60
moderately hard	60 - 80
hard	80 - 120
very hard	> 120

Table 3 *The hardness of water in terms of calcium content*

Solution

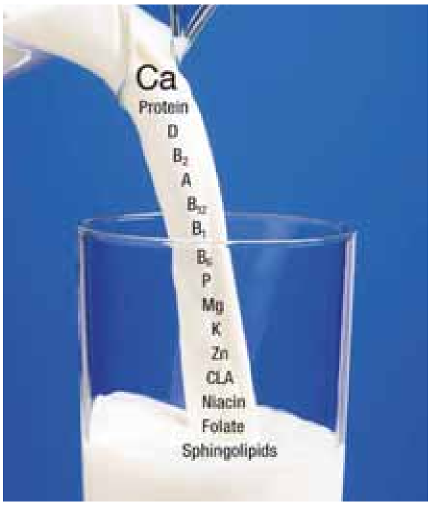
The volume and molar concentration of the water sample are known. The mole number of calcium ions can be calculated easily by using the molarity formula.

$$M = \frac{n_{\text{calcium ions}}}{V_{\text{solution}}} \Rightarrow n_{\text{calcium ions}} = M \cdot V_{\text{solution}} = 0.004 \text{ mol/L} \cdot 1 \text{ L} = 0.004 \text{ mol}$$

Thus, the mass of calcium ions is

$$m = n \cdot M = 0.004 \text{ mol} \cdot 40 \text{ g/mol} = 0.16 \text{ g} = 160 \text{ mg}$$

Since 160 mg > 120 mg, the sample is "very hard."



Milk contains a large amount of calcium, which is useful for the development of bones and teeth.

Example 12

A 500 mL sample of milk nearly contains 5 g of calcium ions, Ca^{2+}. What is the molar concentration of calcium ions in the milk?

Solution

First, the mole number of solute, Ca^{2+} ions should be calculated.

The molar mass of calcium is 40 g/mol.

Thus, the mole number of calcium ions is

$$n = \frac{m}{M} = \frac{5 \text{ g}}{40 \text{ g/mol}} = 0.125 \text{ mol}$$

Now, the molarity of solution can be calculated.

$$M = \frac{n_{\text{solute}}}{V_{\text{solution}}} = \frac{0.125 \text{ mol}}{0.5 \text{ L}} = 0.25 \text{ M}$$

The molar concentration of calcium ions, $[Ca^{2+}] = 0.25$ M.

Exercise 13

A 2 L solution is prepared by dissolving 16.4 g of $Ca(NO_3)_2$ in water. What is the molar concentration of nitrate ions?

Answer : 0.1 M

EXTRA

Parts Per Million (ppm)

When solutions are very dilute, like those of impurities in water, their concentrations are often expressed in parts per million (ppm). This means 1 g in one million grams. The following table shows the chemical composition of the ocean water in ppm.

Constituent	Concentration (ppm)
Chloride	18980
Sodium	10560
Sulfate	2560
Magnesium	1272
Calcium	400
Potassium	380
Bicarbonate	142
Bromide	65
Strontium	13
Boron	4.6
Fluoride	1.4

The chemical composition of ocean water

The concentrations of pollutants in air and water are also measured as ppm. If these concentrations exceed a fixed value, they may be considered hazardous. For instance, when the concentration of mercury in drinking water reaches 0.5 ppm, this level is regarded as harmful to health. That level refers to 0.5 g in 1000 kg of water. Lead is also dangerous at a certain level. 0.250 ppm lead can cause delayed cognitive development in children. Carbon monoxide is a significant factor in air pollution. 750 ppm of carbon monoxide by volume in the air we inhale is considered to be lethal. All these examples show that the relationship of the human body to the ecological system is very finely balanced. Even a small increase in concentrations of pollutants may result in dire consequences. Therefore, scientists have been working on the problem of pollution in order to help maintain the relative purity of our vital natural resources.

SUPPLEMENTARY QUESTIONS

Part A

1. Complete the statements below as a way of summarizing this chapter. The missing words are given in the word list below (not all the words are used).

Word List

molarity	glucose	concentrated
liters	acid	increases
mililiters	dilute	decreases
solute	solvent	

a. In the molarity equation, the unit of volume must be in ....

b. .... is the mole number of solute per liter of solution.

c. A(n) .... solution is used for intravenous feeding.

d. After the evaporation of water, a 10% salt solution becomes more ....

e. As the mass percentage of a solution increases, its density ....

f. When a concentrated stock solution is diluted to prepare a less concentrated reagent, the amount of .... is the same both before and after the dilution.

2. Fill in the blanks in the most appropriate way.

a. .... is used in the production of vinegar.

b. The concentration of a solution can be decreased by adding water. This process is called ....

c. .... flasks are used to prepare a desired solution.

d. .... is the volume of the solute divided by the volume of solution and multiplied by 100.

e. The ions, ... make water "hard."

f. A solution that is 10% by mass sodium chloride contains 10 g of sodium chloride per .... g of solution.

g. A 2 M K_2SO_4 solution contains .... moles of potassium ion and .... moles of sulfate ion per liter.

Part B

1. What is the mass percent concentration of the solution prepared by adding 20 g of sodium chloride into 380 g of water?

2. How many grams of sugar are there in 200 g of a solution which is 15% sugar by mass?

3. Brass, the alloy of copper and zinc, is used to make musical instruments. 40% of brass by mass is zinc. Find the amounts of zinc and copper in 250 g of brass.

4. A 22.4 liter sample of hydrochloric acid is dissolved in 263.5 g of water at STP. What is the mass percent of hydrochloric acid in the resulting solution?

5. The solubility of potassium nitrate is 40 g/100 g of water at 25°C. What is the mass percent of saturated potassium nitrate solution at the same temperature?

6. What mass of water must be added to 200 g of 30% sodium hydroxide solution in order to make it 6%?

7. How many grams of sodium chloride must be dissolved in 100 g of 15.5% sodium chloride solution in order to obtain a 17.5% solution?

8. What will be the mass percent of ammonium chloride if 200 g of 5% and 600 g of 30% ammonium chloride solutions by mass are mixed?

9. How many grams of water must be added to 25 g of $CuSO_4 . 5H_2O$ in order to obtain a 4% solution by mass?

10. A steel sample is made by dissolving 2.5 g of carbon and 0.75 g of nickel per 50 g of molten iron. What is the mass percentage of each component in a steel sample?

11. How many grams of $CoCl_2 \cdot 6H_2O$ and water are required to obtain 180 g of cobalt(II) chloride solution with a mass percentage of 5%?

12. The density of 500 mL of 40% nitric acid solution by mass is 1.25 g/mL. What is the mole number of nitric acid in the solution?

13. The minimum concentration of oxygen necessary for a fish to live in an aquarium is 4 mg/L. What is the minimum concentration of oxygen as mass percent in an aquarium?

(Assume that the density of the aquarium water is 1 g/mL)

14. How many milliliters of water and vinegar should be used to prepare a 900 mL solution that is 30% vinegar by volume?

15. How many grams of alcohol should be added to 400 mL solution which is 15% alcohol by volume to make the percent concentration of the solution 20% by volume?

(The density of alcohol is 0.78 g/mL)

16. How many grams of potassium hydroxide should be used to prepare 500 mL of 0.4 M solution?

17. Suppose that you have a 4 g sample of sodium hydroxide as a solid.

a. How can you prepare a 2 M sodium hydroxide solution?

b. By using the solution obtained in **a**, prepare an 0.5 M sodium hydroxide solution.

18. What will be the new molarity of 200 mL of 6 M sodium chloride solution, if 600 mL of water is added to the solution at the same temperature?

19. How many grams of sodium hydroxide should be used to neutralize 200 mL of 0.3 M hydrochloric acid solution?

20. An alcoholic iodine solution, tincture of iodine, is prepared by dissolving 10.30 g of iodine crystals in enough alcohol to make 450 mL of solution. Calculate the molarity of iodine in the solution.

21. What is the molar concentration of the solution obtained by mixing 300 mL of 6 M sulfuric acid with 200 mL of 2 M sulfuric acid solutions?

22. Formaldehyde (CH_2O) is dissolved in water; the obtained solution is called formaline. Formaline is used to fix tissue samples. What is the molarity of the solution prepared by dissolving 3 g of formaldehyde in water to obtain a 400 mL of formaline solution?

23. Calcium carbonate ($CaCO_3$) is found in some commercial antacids (materials used to decrease the amount of acid in the stomach). A 0.5 g sample of antacid is titrated with hydrochloric acid:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

If 30 mL of 0.2 M hydrochloric acid is needed for a complete reaction, what is the mass percentage of calcium carbonate in the antacid sample?

24. An acetic acid solution called "vinegar" contains nearly 5% acetic acid by mass. (The rest of solution is only water.) What is the molarity of vinegar if its density is 1 g/mL?

25. A two liter solution of 0.1 M potassium dichromate is needed to analyze the alcohol content of a certain wine. How much solid potassium dichromate must be weighed out to make this solution?

26. Find the molar concentrations of the ions in the following solutions.

a. 0.2 M sodium chloride

b. 100 mL of 0.5 mol of calcium chloride

c. 1 liter of an aqueous solution containing 8 g of sodium hydroxide

27. The density of c% nitric acid solution by mass is ρ in g/mL. Prove that the molarity of the nitric acid solution is

$$\frac{c \cdot \rho \cdot 10}{M}$$

MULTIPLE CHOICE QUESTIONS

1. A sugar solution is obtained by dissolving 100 g of sugar in 400 g of water. What is the mass percent of sugar in the resulting solution?

A) 10% B) 20% C) 25% D) 40% E) 50%

2. A 0.4 mol sample of solid sodium hydroxide is dissolved in water to prepare a 20% sodium hydroxide solution by mass. How many grams of water should be used to obtain this solution?

A) 16 B) 32 C) 64 D) 80 E) 96

3. Which of these may be done to increase the mass percentage of a salt solution?

I. Addition of water
II. Evaporation of water
III. Addition of salt

A) I only B) II only C) III only D) I and III E) II and III

4. A 30 g sample of potassium chloride is dissolved in 150 g of 10% potassium chloride solution by mass.

What is the mass percent of the new solution?

A) 10 B) 15 C) 25 D) 30 E) 40

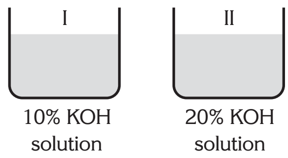
5. How many grams of water should be added to 50 g of 40% sodium nitrate solution in order to obtain a 10% solution by mass?

A) 20 B) 75 C) 100 D) 150 E) 250

6. A 375 g sample of $CuSO_4 \cdot 5H_2O$ is dissolved in 125 g of water. What is the mass percentage of the prepared solution?

A) 24 B) 32 C) 48 D) 75 E) 90

7. Which one of the following is the same for two solutions of potassium hydroxide with the same volume but different concentrations?



A) Electrical conductivity
B) Boiling point
C) Molarity of the solution
D) Mass of solute
E) None of them

8. An ethyl alcohol solution used in the laboratory is nearly 96% by volume. Which one of the following statements about this solution is false?

A) Alcohol is found in the solution as molecules.
B) It is a nonelectrolyte solution.
C) A 100 g sample of this solution contains 96 g of alcohol.
D) After the addition of water, the solution becomes more dilute.
E) A 50 L sample of this solution contains 48 L of alcohol.

9. Which one of the following is the correct definition of molarity?

A) Molarity is the mole number of solute in one liter solution.
B) Molarity is the mole number of solvent in one liter solution.
C) Molarity is the mole number of solute in one liter solvent.
D) Molarity is the mole number of solvent in one liter solvent.
E) Molarity is the mass of solute in one liter solution.

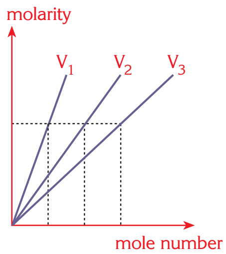
10. How many grams of silver nitrate are needed to prepare 500 mL of 0.4 M solution?

A) 17 B) 34 C) 42.5 D) 51 E) 108

11. If a 500 mL solution is prepared with 17 g of sodium nitrate, what will be the molarity of the solution?

A) 0.1 B) 0.2 C) 0.3 D) 0.4 E) 0.5

12. Three different sodium hydroxide solutions are prepared. Let their volumes be V_1, V_2, and V_3. The relationship between their mole versus molarity numbers are shown in the graph.



Which of these expresses the correct relation among V_1, V_2, and V_3 ?

A) $V_3 > V_2 > V_1$
B) $V_1 > V_2 > V_3$
C) $V_1 = V_2 = V_3$
D) $V_1 > V_2 = V_3$
E) $V_2 = V_1 > V_3$

13. When 500 mL distilled water is added to 1 L of 1 M sodium hydroxide solution, which of the following change(s)?

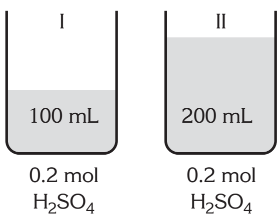
I. The molarity of the solution
II. The mass of sodium hydroxide dissolved in the solution
III. The concentration of sodium ion (Na^+) in the solution

A) I only
B) II only
C) I and II
D) I and III
E) II and III

14. A 1 liter sample of seawater contains 29.25 g of sodium chloride. What is the molar concentration of sodium chloride in seawater sample?

A) 0.1 M
B) 0.2 M
C) 0.5 M
D) 1 M
E) 10 M

15. Two different solutions, 100 mL and 200 mL, are prepared by dissolving 0.2 mol of sulfuric acid in water.



Which of the following is the same for both solutions?

A) Density
B) Molarity
C) Mass percent
D) The mass of solute
E) Electrical conductivity

16. If a concentrated nitric acid is diluted, which of the following changes?

I. Molarity II. Density III. Mass percent

A) I only
B) I and II
C) I and III
D) II and III
E) I, II, and III

17. What is the mole number of 100 mL of nitric acid solution that is 63% by mass? (The density of the solution is 1.4 g/cm^3)

A) 0.7
B) 1
C) 1.4
D) 2.8
E) 6.3

18. 100 mL of 2 M hydrochloric acid and 100 mL of 1 M sodium hydroxide solutions are mixed.

Which ion is found in trace amounts in the resulting solution?

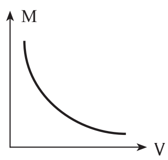
A) H^+
B) OH^-
B) Na^+
D) Cl^-
E) H^+ and OH^-

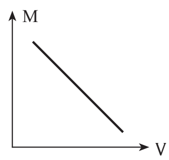
19. How many mL of 2 M hydrochloric acid should be used to dissolve 10 g of calcium carbonate completely?

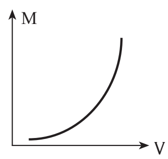
$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

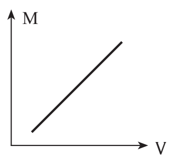
A) 0.1
B) 50
C) 100
D) 500
E) 1000

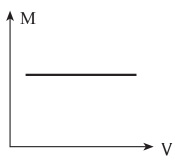
20. Water is slowly added to 10 mL of 4 M sodium hydroxide solution at the same temperature. Which one of the following shows the correct graph of the molarity of the solution versus the volume of the solution?

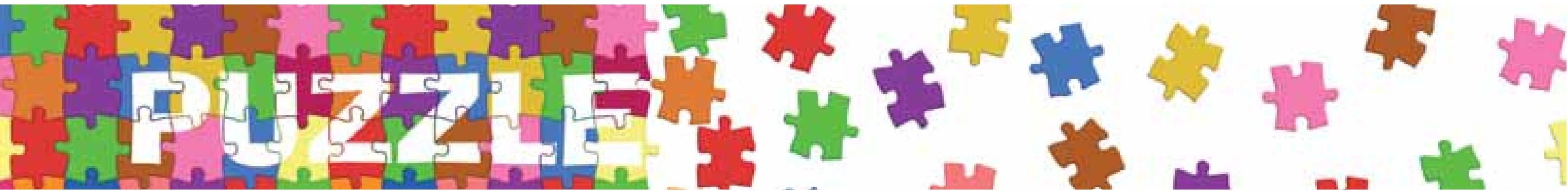
A) 

B) 

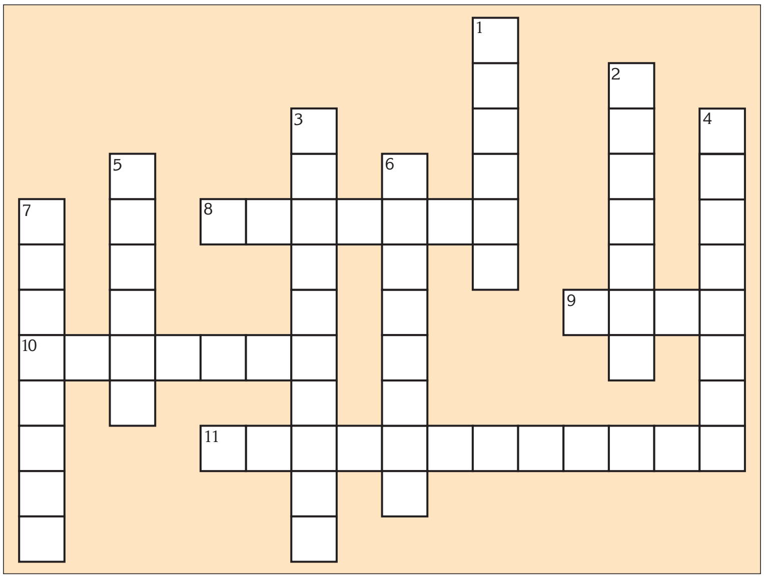
C) 

D) 

E) 



CRISS CROSS PUZZLE



DOWN

1 Percent concentration is studied as mass percent and .... percent.

2 Common name of a 5% acetic acid solution.

3 The .... flask is used in preparing a solution.

4 Solution whose concentration is known accurately.

5 Element composing 21% of atmosphere.

6 Process of getting a less concentrated solution from stock solutions.

7 Mole number of solute per liter of solution.

ACROSS

8 Element found in milk which helps the development of teeth and bones.

9 Water containing a high concentration of magnesium and calcium ions is called ....

10 Solvent used as a nail polish removers.

11 Acid found in soft drinks.

WORD BLOCK

Answer the clues and fill in the horizontal lines of the grid. A word will be formed diagonally across the shaded blocks, from top left to bottom right. A clue is also given for this word.

1. The major component of the solution.

2. The main substance found in nail polish removers.

3. To make a solution acidic by adding any acid.

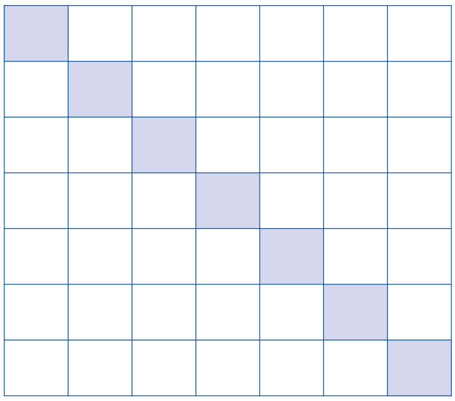
4. The common name of 5% acetic acid solution.

5. The solute in wine (wine is a 10% ethanol solution).

6. Electrolyte solutions are the solutions that .... electricity.

7. Its solution is used for intravenous feeding.

DIAGONAL: The systemized study of knowledge.



APPENDICES
GLOSSARY
ANSWERS
INDEX

SOLUTIONS

APPENDICES

Appendix A

Solubility Table

Cations \ Anions	Nitrate (NO_3^-)	Acetate (CH_3COO^-)	Chloride (Cl^-)	Hydroxide (OH^-)	Iodide (I^-)	Sulfate (SO_4^{2-})	Carbonate (CO_3^{2-})	Bromide (Br^-)	Sulfide (S^{2-})	Sulfite (SO_3^{2-})	Phosphate (PO_4^{3-})	Silicate (SiO_3^{2-})
Ammonium (NH_4^+)	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Compound does not exist or is unstable in water
Potassium (K^+)	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble
Sodium (Na^+)	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble
Barium (Ba^{2+})	Soluble	Soluble	Soluble	Soluble	Soluble	Insoluble	Insoluble	Soluble	Compound does not exist or is unstable in water	Slightly soluble	Insoluble	Insoluble
Calcium (Ca^{2+})	Soluble	Soluble	Soluble	Slightly soluble	Soluble	Slightly soluble	Insoluble	Soluble	Compound does not exist or is unstable in water	Slightly soluble	Insoluble	Insoluble
Magnesium (Mg^{2+})	Soluble	Soluble	Soluble	Insoluble	Soluble	Soluble	Insoluble	Soluble	Compound does not exist or is unstable in water	Slightly soluble	Insoluble	Insoluble
Aluminum (Al^{3+})	Soluble	Soluble	Soluble	Insoluble	Soluble	Soluble	Compound does not exist or is unstable in water	Soluble	Compound does not exist or is unstable in water	Compound does not exist or is unstable in water	Insoluble	Insoluble
Zinc (Zn^{2+})	Soluble	Soluble	Soluble	Insoluble	Soluble	Soluble	Insoluble	Soluble	Insoluble	Slightly soluble	Insoluble	Insoluble
Iron (Fe^{2+})	Soluble	Soluble	Soluble	Insoluble	Soluble	Soluble	Insoluble	Soluble	Insoluble	Slightly soluble	Insoluble	Insoluble
Iron (Fe^{3+})	Soluble	Soluble	Soluble	Insoluble	Compound does not exist or is unstable in water	Soluble	Compound does not exist or is unstable in water	Soluble	Insoluble	Compound does not exist or is unstable in water	Insoluble	Insoluble
Copper (Cu^{2+})	Soluble	Soluble	Soluble	Insoluble	Compound does not exist or is unstable in water	Soluble	Compound does not exist or is unstable in water	Compound does not exist or is unstable in water	Insoluble	Compound does not exist or is unstable in water	Insoluble	Insoluble
Silver (Ag^+)	Soluble	Soluble	Insoluble	Compound does not exist or is unstable in water	Insoluble	Slightly soluble	Insoluble	Insoluble	Insoluble	Slightly soluble	Insoluble	Compound does not exist or is unstable in water
Mercury (Hg^{2+})	Soluble	Soluble	Soluble	Compound does not exist or is unstable in water	Insoluble	Soluble	Compound does not exist or is unstable in water	Slightly soluble	Insoluble	Compound does not exist or is unstable in water	Insoluble	Compound does not exist or is unstable in water
Hydrogen (H^+)	Soluble	Soluble	Soluble	Compound does not exist or is unstable in water	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Soluble	Insoluble
Chromium (Cr^{3+})	Soluble	Soluble	Soluble	Insoluble	Soluble	Soluble	Compound does not exist or is unstable in water	Soluble	Compound does not exist or is unstable in water	Compound does not exist or is unstable in water	Insoluble	Insoluble
Lead (Pb^{2+})	Soluble	Soluble	Slightly soluble	Insoluble	Insoluble	Insoluble	Insoluble	Slightly soluble	Insoluble	Insoluble	Insoluble	Insoluble

Soluble

Slightly soluble

Insoluble

Compound does not exist or is unstable in water.

Appendix B

The Analysis of Some Ions

Tested Ion	Reagent	Result
Cl^-	Ag^+ Pb^{2+}	White precipitate
Br^-	Ag^+	Grayish-yellow precipitate
I^-	Ag^+	Yellow precipitate
SO_4^{2-}	Ba^{2+}	White precipitate
PO_4^{3-}	Ag^+	Yellow precipitate
S^{2-}	Cu^{2+}, Pb^{2+}, Ag^+ Cd^{2+}	Black precipitate Yellow precipitate
CO_3^{2-}	H^+	Formation of carbon dioxide that turns limewater milky
OH^-	Indicator	Color change
Ag^+	Cl^-, SO_4^{2-}	White precipitate
Cu^{2+}	OH^- S^{2-}	Blue precipitate Black precipitate
Mg^{2+}	OH^-	White precipitate
Fe^{2+}	OH^-	Green precipitate
Fe^{3+}	OH^-	Reddish-brown precipitate
Zn^{2+}	OH^- S^{2-}	White precipitate (soluble in excess OH^-) White precipitate
Pb^{2+}	S^{2-}	Black precipitate
NH_4^+	OH^-	Formation of ammonia gas with an irritating odor
Ba^{2+}	SO_4^{2-}, CO_3^{2-}	White precipitate
Al^{3+}	OH^-	White precipitate (soluble in excess OH^-)

GLOSSARY

Alloy : The homogeneous mixture of two or more metals.

Antifreeze : The substance added to a liquid (generally water) in order to lower its freezing point or to raise its boiling point.

Bends : A serious and lethal condition that occurs when there is a great change in pressure.

Cold packs : Packs used to treat shoulder pain or injuries by cooling.

Concentrated solution : A solution that has relatively high concentration of solute in a solution.

Dilution : The process of preparation of a less concentrated solution by adding water into stock concentrated solutions.

Dilute solution : A solution in which there is only a small amount of solute.

Dissolution : The process of mixing a solute in a solvent.

Electrolyte : A substance which undergoes dissociation into ions in solution, and thus acts as a conductor of electricity.

Endothermic : A process in which heat is absorbed by a system.

Exothermic : A process in which heat is emitted by a system.

Hard water : Water that contains a high concentration of magnesium and calcium ions.

Heat packs : Packs used for first aid purposes by giving off heat.

Heterogeneous mixture : A nonuniform mixture in which the composition varies throughout.

Homogeneous mixture : A mixture that is same throughout in its composition.

Hydration : A solvation process in which the solvent is water.

Hydrolysis : The reaction of a cation or an anion with water.

Insoluble : The term used for substances that cannot be dissolved in a liquid.

Ionization : A dissolution process in which ions are produced.

Mass percent : (also called ***weight percent***). To calculate mass percent, the mass of solute is divided by the mass of the solution and multiplied by 100.

Metathesis reactions : A reaction in which cations and anions exchange their partners.

Miscibility : The ability of a liquid to mix with another liquid.

Molality : The number of moles of solute dissolved in one kilogram of solvent.

Molarity : The concentration of a solution expressed as the number of moles of the solute per one liter of solution.

Net ionic equation : A chemical equation that does not show spectator ions but includes only the ions involved in the reaction.

Nonelectrolyte : Solutions that do not conduct electricity since they contain no ions.

Normality : The concentration of a solution expressed as the number of equivalence per one liter of solution.

Percent concentration : The proportion of solute in a solution by mass or volume.

Precipitate : The suspension of solid particles in a liquid after mixing aqueous solutions.

Precipitation : The formation of an insoluble substance in an aqueous solution.

Rubbing alcohol : A type of alcohol used for cleaning wounds or skin in hospital.

Saturated solution : A solution obtained by dissolving the maximum amount of solute in a solvent at a specific temperature.

Slightly soluble : A solute with a solubility range from 0.1 g to 1 g in 100 g of water.

Solubility : The maximum amount of solute that dissolves in 100 g of water at a certain temperature.

Solubility curve : A graphical representation of the variation of the solubility with temperature.

Solubility table : The table showing whether a substance is soluble, insoluble, or slightly soluble in water.

Soluble : The term used for substances that can be dissolved in a liquid.

Solute : The minor component of a solution.

Solution: A mixture of two or more pure substances that is uniform throughout.

Solvation : The process of surrounding solute ions by solvent molecules during the dissolution of ionic solids.

Solvent : The component of a solution that is present in large amount and that determines the physical state of the solution.

Spectator ion : An ion that is present on both sides of the reaction in an aqueous solution.

Standard solution : A solution whose concentration is known accurately.

Stock solution : A solution that is very concentrated.

Strong electrolyte : An electrolyte that contains many ions in the solution.

Supersaturated solution : A metastable solution in which the concentration of solute is higher than that of a saturated solution.

Unsaturated solution : The solution of which the concentration of solute is less than a saturated.

Volume percent : The volume of solute divided by the total volume of solution multiplied by 100.

Volumetric flask : A container used to prepare standard solutions.

Weak electrolyte : A substance that contains few ions in the solution.

ANSWERS

SUPPLEMENTARY QUESTIONS

THE PROPERTIES OF SOLUTIONS

PART A

1. a. electrolyte b. Ethylene glycol
   c. pressure d. Stirring
   e. lowers f. unsaturated
   g. cold h. exothermic

2. a. false b. false
   c. false d. true
   e. true

PART B

1. a. solution b. solution
   c. not solution d. solution
   e. solution f. solution
   g. not solution h. solution

2. a. solute : carbon
   solvent : iron
   b. solute : water
   solvent : alcohol
   c. solute : sugar
   solvent : water

3. a. $NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$
   c. $Ca(NO_3)_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2NO_3^-(aq)$

7. a. weak electrolyte b. nonelectrolyte
   c. weak electrolyte d. strong electrolyte
   e. nonelectrolyte f. strong electrolyte
   g. strong electrolyte

9. 92.5 g

11. a. No. It's unsaturated b. 46 g

12. a. 275 g b. 50 g

16. 40.7 g

17. 8 g

18. a. 1. $NaCl + AgNO_3 \rightarrow AgCl\downarrow + NaNO_3$
   3. $MgCl_2 + 2NaOH \rightarrow Mg(OH)_2\downarrow + 2NaCl$
   b. 1. $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
   3. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$

20. a. acidic b. neutral
   c. basic

MEASURING CONCENTRATION

PART A

1. a. liters b. molarity
   c. glucose d. concentrated
   e. increases f. solute

2. a. acetic acid b. dilution
   c. volumetric d. volume percent
   e. Ca^{2+} and Mg^{2+} f. 100
   g. 4 , 2

PART B

1. 5%
2. 30 g
3. $m_{Zn} = 100$ g , $m_{Cu} = 150$ g
4. 12.17%
5. 28.57%
6. 800 g
7. 2.42 g
8. 23.75%
9. 375 g
10. (%)C = 4.69%, (%)Ni = 1.41%, (%)Fe = 93.90%
11. $m_{CoCl_2 \cdot 6H_2O} = 16.48$ g , $m_{H_2O} = 163.52$ g
12. 3.97 moles
13. $4 \cdot 10^{-4}$ %
14. $V_{water} = 630$ mL, $V_{vinegar} = 270$ mL
15. 19.5 g
16. 11.2 g
18. 1.5 M
19. 2.4 g
20. 0.09 M
21. 4.4 M
22. 0.25 M
23. 60%
24. 0.83 M
25. 58.8 g
26. a) $[Na^+] = [Cl^-] = 0.2$ M
    b) $[Ca^{2+}] = 5$ M
    $[Cl^-] = 10$ M
    c) $[Na^+] = [OH^-] = 0.2$ M

MULTIPLE CHOICE

THE PROPERTIES OF SOLUTIONS

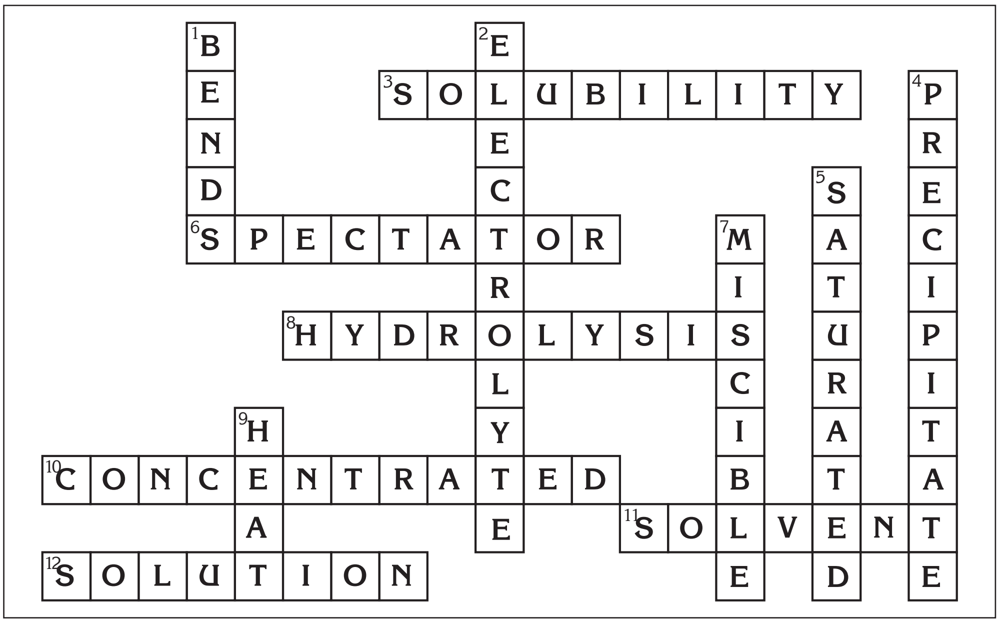
1. D	4. D	7. D	10. D	13. C	16. E	19. E
2. C	5. C	8. A	11. C	14. A	17. B	
3. E	6. A	9. D	12. B	15. E	18. D	

MEASURING CONCENTRATION

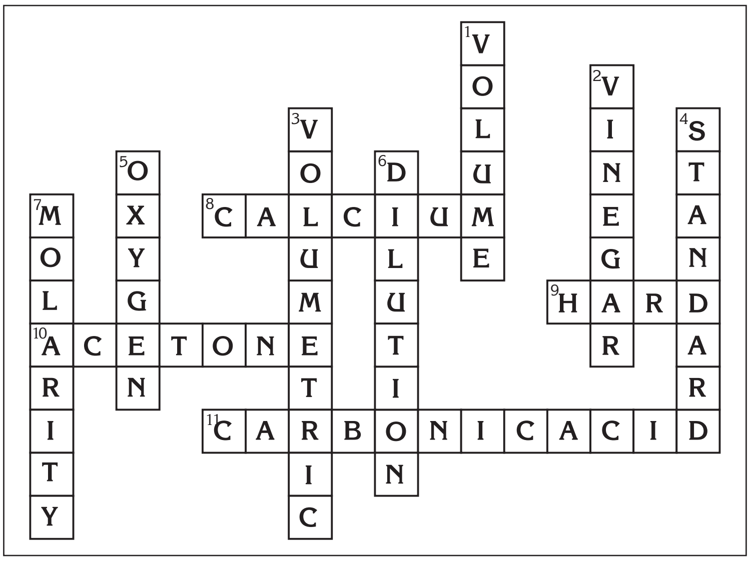
1. B	4. C	7. E	10. B	13. D	16. E	19. C
2. C	5. D	8. C	11. D	14. C	17. C	20. A
3. E	6. C	9. A	12. A	15. D	18. B	

PUZZLE

THE PROPERTIES OF SOLUTIONS



MEASURING CONCENTRATION



1	S	O	L	V	E	N	T
2	A	C	E	T	O	N	E
3	A	C	I	D	I	F	Y
4	V	I	N	E	G	A	R
5	E	T	H	A	N	O	L
6	C	O	N	D	U	C	T
7	G	L	U	C	O	S	E

INDEX

accumulator, 15
acetone, 45
acidic solution, 31, 32
air, 8, 10, 18
alloys, 8, 9, 10
amalgam, 7, 9
antifreezes, 9, 16, 17, 45
aqueous solutions, 8, 27, 28
Arrhenius Svante, 15
artificial honey, 24
basic solution, 31, 32
bends, 23
boiling point, 16, 17
brass, 8, 9
bronze, 10
carbon steel, 9
carbon tetrachloride, 12
carbonated beverages, 9, 10
carbonic acid, 40
chemical dissolution, 11
coin, 7, 9
cold pack, 13
concentrated solution, 17, 40
crystallization, 24
de-icing process, 16
dental amalgam, 9
dental fillings, 7, 9
dilute solution, 17, 40, 42
dilution, 17, 42, 43
discharging, 15
dissolution, 11, 12, 13, 18
double displacement reaction, 29
dry cleaning, 12
electrochemical cell, 12
electrolytes, 14, 15, 18
endothermic process, 13
exothermic process, 13
first aid, 13
fizzing, 23
fluid, 7, 12
fog, 7, 9
freezing point, 16, 17
gas solutions, 8, 9
glucose, 8, 24, 42
hard water, 52
heat pack, 13
heterogeneous mixture, 7
homogeneous mixture, 7
honey, 24
household bleach, 40, 43
hydration, 11
hydrolysis, 31
immiscible, 12
impure, 7, 8, 10
insoluble, 21
ionic equation, 27, 28, 29, 30
ionization, 11
Issyk kul lake, 47
like dissolves like, 12
limestone, 51
liquid solutions, 8, 9, 12
litmus paper, 31
mass percent, 41
metathesis reaction, 29
milk, 52
miscible, 11, 12
mixture, 7
molality, 51
molarity, 47, 48, 49, 50
natural honey, 24
net ionic equation, 27, 28, 29, 30
neutral solution, 31, 32
nonelectrolytes, 14, 18
normality, 51
ocean water, 53
oil, 6, 12
parts per million, 53
perc, 12
percent concentration, 41
physical dissolution, 11
potassium permanganate, 48
precipitate, 27, 30
precipitation, 28, 29
pressure, 23, 25
pure, 8
saturated, 20, 26
slightly soluble, 21
soft drinks, 7
solder, 9, 10
solid solution, 8, 9
solubility, 20, 21, 22, 25
solubility curve, 22
solubility table, 21
soluble, 21
solute, 8, 9, 10
solution, 7, 8, 9, 14, 16
solvation, 11
solvent, 8, 9, 10
spectator ions, 27
stainless steel, 9
standard solution, 50
steel, 9
stock solution, 42, 43
strong electrolytes, 14, 18
supersaturated, 13, 24
tap water, 8, 10, 15
temperature, 22, 25
unsaturated, 20
vinegar, 18, 40
volume percent, 45
volumetric flask, 50
water, 19
water cycle, 19
weak electrolytes, 14, 18

PERIODIC

Name — Symbol
Iron Fe
55.85 26
Atomic mass — Atomic number

IUPAC : 1 American : IA	2 IIA		3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB
Hydrogen H 1.01 1									
Lithium Li 6.94 3	Beryllium Be 9.01 4								
Sodium Na 22.99 11	Magnesium Mg 24.31 12								
Potassium K 39.10 19	Calcium Ca 40.08 20		Scandium Sc 44.96 21	Titanium Ti 47.88 22	Vanadium V 50.94 23	Chromium Cr 52.00 24	Manganese Mn 54.94 25	Iron Fe 55.85 26	Cobalt Co 58.93 27
Rubidium Rb 85.47 37	Strontium Sr 87.62 38		Yttrium Y 88.91 39	Zirconium Zr 91.22 40	Niobium Nb 92.91 41	Molybdenum Mo 95.94 42	Technetium Tc (98) 43	Ruthenium Ru 101.07 44	Rhodium Rh 102.91 45
Cesium Cs 132.91 55	Barium Ba 137.33 56	Lanthanides * 57-70	Lutetium Lu 174.97 71	Hafnium Hf 178.49 72	Tantalum Ta 180.95 73	Tungsten W 183.84 74	Rhenium Re 186.21 75	Osmium Os 190.2 76	Iridium Ir 192.22 77
Francium Fr {223} 87	Radium Ra (226) 88	Actinides ** 89-102	Lawrencium Lr (262) 103	Rutherfordium Rf (261) 104	Dubnium Db (262) 105	Seaborgium Sg (266) 106	Bohrium Bh (264) 107	Hassium Hs (269) 108	Meitnerium Mt (268) 109

* Lanthanides	Lanthanum La 138.91 57	Cerium Ce 140.12 58	Praseodymium Pr 140.91 59	Neodymium Nd 144.24 60	Promethium Pm (145) 61	Samarium Sm 150.36 62
** Actinides	Actinium Ac (227.03) 89	Thorium Th 232.04 90	Protactinium Pa 231.04 91	Uranium U 238.03 92	Neptunium Np (237.05) 93	Plutonium Pu (244) 94

TABLE

10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
								Helium He 4.0026 2
			Boron B 10.81 5	Carbon C 12.01 6	Nitrogen N 14.0067 7	Oxygen O 15.9994 8	Fluorine F 18.9984 9	Neon Ne 20.18 10
			Aluminum Al 26.98 13	Silicon Si 28.09 14	Phosphorus P 30.97 15	Sulfur S 32.07 16	Chlorine Cl 35.45 17	Argon Ar 39.95 18
Nickel Ni 58.69 28	Copper Cu 63.55 29	Zinc Zn 65.41 30	Gallium Ga 69.723 31	Germanium Ge 72.64 32	Arsenic As 74.92 33	Selenium Se 78.96 34	Bromine Br 79.90 35	Krypton Kr 83.80 36
Palladium Pd 106.42 46	Silver Ag 107.87 47	Cadmium Cd 112.41 48	Indium In 114.817 49	Tin Sn 118.71 50	Antimony Sb 121.76 51	Tellurium Te 127.60 52	Iodine I 126.90 53	Xenon Xe 131.29 54
Platinum Pt 195.08 78	Gold Au 196.97 79	Mercury Hg 200.59 80	Thallium Tl 204.383 81	Lead Pb 207.19 82	Bismuth Bi 208.980 83	Polonium Po (209) 84	Astatine At (210) 85	Radon Rn (222) 86
Darmstadium Ds (271) 110	Roentgenium Rg (272) 111	Ununbium Uub (285) 112	113	Ununquadium Uuq 289 114	115	Ununhexium Uuh 292 116	117	118

Europium Eu 151.96 63	Gadolinium Gd 157.25 64	Terbium Tb 158.92 65	Dysprosium Dy 162.50 66	Holmium Ho 164.93 67	Erbium Er 167.26 68	Thulium Tm 168.93 69	Ytterbium Yb 173.04 70
Americium Am (243) 95	Curium Cm (247) 96	Berkelium Bk (247) 97	Californium Cf (251) 98	Einsteinium Es (252) 99	Fermium Fm (257) 100	Mendelevium Md (258) 101	Nobelium No (259) 102

REFERENCES

Çelik, N.; Erdem, A.R.; Gürler, V.; Karabürk, H.; Nazlı, P. **Chemistry 2 as a popular science.** Sürat Publications, İstanbul : 2000

Nazlı, A.; Bayar, M.; Duran, C.; Çelik, N.; Karabürk, H.; Patlı, U.H. **Lise Kimya 1.** Zambak Publishing, İstanbul : 2004

Frank, D.V.; Little, J.G.; Miller, S. **Chemical Interactions.** Prentice Hall Science Explorer, Pearson Education Inc., New Jersey : 2005

Lidin, R.A.; Malochka, V.A.; Andreyeva, L.L. Ximicheski Svoystva Neorganicheckih Veshestv. **Chemical Properties of Inorganic Compounds.** Ximia Publishing, Moscow : 2000

Abbasov, V.M. **Kimya.** TQDK, Abituriyent Bakı : 2005.

Wilson, D.; Bauer, M. **Dynamic Science.** McGraw-Hill., Sydney : 1995

Haire, M.; Kennedy, E.; Lofts, G., Evergreen, M.J. **Core Science 2.** John Wiley & Sons Australia Ltd., Sydney : 1999

Oxtoby, D.W. Nachtrieb, N.H. **Principles of Modern Chemistry 3rd Edition.** Saunders College Publishing, USA : 1996

Sevenair, J.P., Burkett, A.R. **Introductory Chemistry Investigating the Molecular Nature of Matter.** WCB Publishers, USA : 1997

Prescott, C.N., **Chemistry A Course for "O" Level.** Times Media Private Limited, Singapore : 2000

Gallagher. RM.; Ingram, P.; **Modular Science, Chemistry.** Oxford University Press, UK : 2001

Ryan, L. **Chemistry for you.** Stanley Thornes (Publishers) Ltd. UK : 1996

NOTES

NOTES