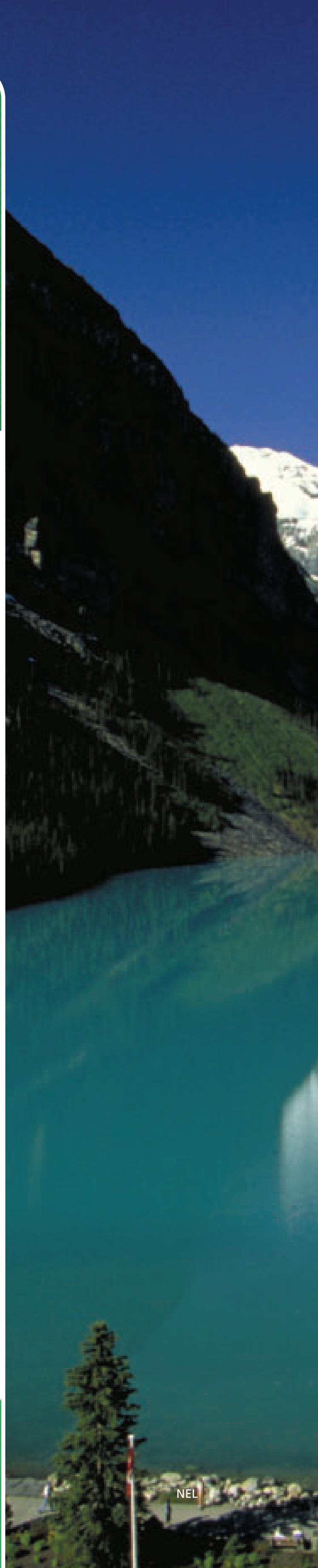
unit 3

Solutions, Acids, and Bases



Solutions, especially of the liquid variety, are everywhere. All fresh water in streams, rivers, and lakes, salt water in the oceans, and even the rain that falls from the sky are examples of solutions. In general, what we call "water" is a solution that is essential to life. The characteristics of natural water can be quite complicated. Various physical, chemical, and biological factors need to be considered. For example, the colour of lakes may be due to dissolved minerals, decomposition of plant materials, and light reflected from suspended solids. In the glacial lake shown in the photograph, the colour is primarily due to the reflection of light from finely ground rock washed down in glacial streams.

Once you include manufactured solutions, there are literally countless examples to consider. Many consumer products, such as liquid cleaners, various drinks, and antiseptics, to name a few types, are solutions. Most industries use solutions in the cleaning, preparation, or treatment of the products they produce. Generally, when we have used or finished with a manufactured solution, we either recycle or dispose of it. Disposal usually means putting the unwanted solution into the environment, with or without some sort of treatment, and assuming that the environment will take care of our wastes. Our view of the environment needs to change from some external entity that can be exploited, to a more holistic view in which we are part of the system. In other words, our attitudes toward the environment need to become more like the traditional attitudes of Aboriginal peoples.

If you want to better understand our natural environment, how we produce solutions from the water in this environment, and the effects of our disposal of wastes, then an empirical and theoretical knowledge of solutions is essential.

As you progress through the unit, think about these focusing questions:

- How can we describe and explain matter as solutions, acids, and bases using empirical and theoretical descriptions?
- Why is an understanding of acid–base and solution chemistry important in our daily lives and in the environment?

GENERAL OUTCOMES

In this unit, you will

- investigate solutions, describing their physical and chemical properties
- describe acid and base solutions qualitatively and quantitatively

Unit 3 Solutions, Acids, and Bases

Prerequisites

Concepts

- classes of matter
- states of matter
- chemical and physical properties
- periodic table
- elements and compounds
- atomic theory
- ions
- chemical reactions
- chemical formulas and equations
- acids and bases

Skills

- WHMIS symbols
- laboratory safety rules
- scientific problem solving
- graphing, and solving linear and exponential equations

You can review prerequisite concepts and skills on the Nelson Web site, in the Chemistry Review unit, and in the Appendices.

A Unit Pre-Test is also available online.

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ARE YOU READY?

These questions will help you find out what you already know, and what you need to review, before you continue with this unit.

Knowledge

1. Copy and complete the classification scheme in **Figure 1**.

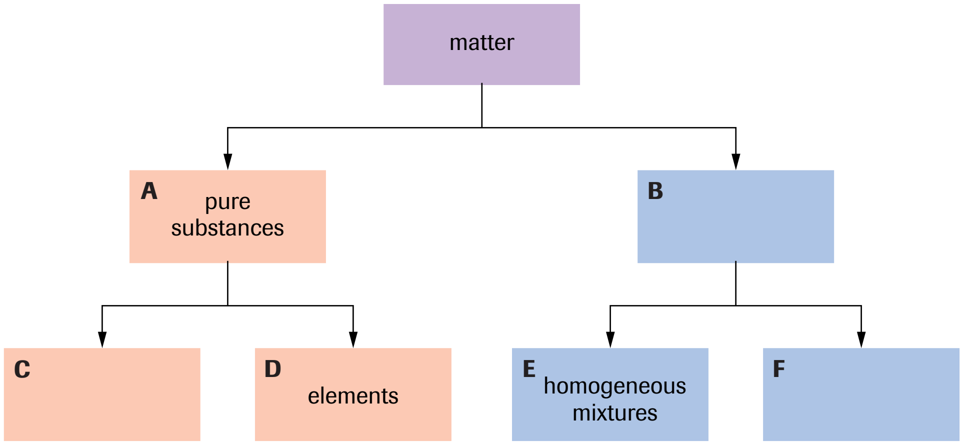

Figure 1
A classification of matter

2. Match each of the substances in **Table 1** to the classification categories illustrated in **Figure 1**.

Table 1 Classification of Substances

Substance	A or B	C or D or E or F
(a) vinegar		
(b) pure water		
(c) sulfur		
(d) air		
(e) milk		

3. Distinguish between ionic and molecular compounds based on their
   (a) chemical name or formula
   (b) empirical (observable) properties

4. In **Table 2**, match each term in column I with its corresponding description in column II.

Table 2 Definitions of Types of Matter

I	II
(a) compound	A. Cannot be broken down into simpler substances
(b) solution	B. Contains two or more visible components
(c) element	C. Can be identified by a single chemical formula
(d) heterogeneous mixture	D. A mixture of two or more pure substances with a single visible component

5. Write the missing words from the following statement in your notebook:

   According to modern atomic theory, an atom contains a number of positively charged ________, determined by the ________ ________ of the element, and an equal number of negatively charged ________.

6. Atoms of the main group (representative) elements generally form predictable ions. Using calcium and fluorine as examples, draw Lewis symbols showing atoms and ions.

7. Draw a diagram to illustrate the model of a small sample (a few particles) of
   (a) sodium chloride
   (b) bromine

8. What is the type of bond between the atoms of a water molecule? What are the types of bonds between the molecules of water in a sample (**Figure 2**)?

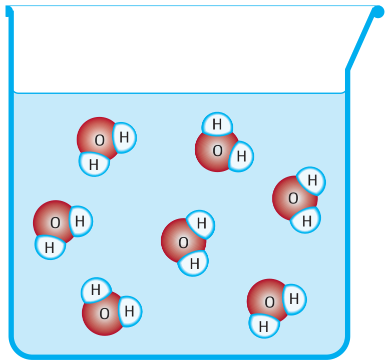

Figure 2
A model of a sample of liquid water

9. Refer to the list of substances in **Table 3** to answer the following questions.
   (a) Which substances have London (dispersion) forces present between the molecules?
   (b) Classify each substance as polar or nonpolar.
   (c) Which substance would be expected to have hydrogen bonding as part of the intermolecular forces between the molecules?
   (d) Distinguish between intermolecular and intramolecular forces.

Table 3 Substances and Their Uses

Substance	Chemical formula	Use
propane	$C_3H_8(g)$	propane barbecues
ethanol	$C_2H_5OH(l)$	in gasohol (gasoline-alcohol fuel)
dichloromethane	$CH_2Cl_2(l)$	paint stripper

10. For each of the following pairs of reactants:
    - write a balanced chemical equation, including states of matter at SATP, for the expected reaction
    - translate the balanced chemical equation into an English sentence, including the coefficients and states of matter
    - state one diagnostic test that could be used to test the predicted reaction

    (a) aqueous iron(III) chloride and aqueous sodium hydroxide
    (b) aqueous silver nitrate and copper metal
    (c) sulfuric acid and aqueous potassium hydroxide
    (d) aqueous chlorine and aqueous sodium bromide

11. Many reactions in solution are single or double replacement reactions.
    (a) Write the word generalization for these two reaction types.
    (b) Classify each of the reactions in question 10 as a single or double replacement reaction.

12. Chemical reactions can also be classified as endothermic or exothermic. What do these terms mean? What diagnostic test would be used for this classification?

Skills

13. In this unit, you will work with many different solutions. What should you do immediately if some solution is spilled on your hand?

14. State the hazard communicated by each of the following WHMIS symbols.
    (a) 
    (b) 
    (c) 

chapter 5

The Nature and Properties of Solutions

In this chapter

Is there such a thing as pure, natural water? Certainly it can't be found in the oceans. Drinking the water of the sea, which is rich in dissolved solutes, can be fatal. Today, seagoing ships carry distillation equipment to convert salt water into drinking water by removing most of those solutes.

Water from lakes and rivers (**Figure 1**), which we depend on for drinking, cooking, irrigation, electric power generation, and recreation, is also impure. Even direct from a spring, fresh water is a solution that contains dissolved minerals and gases. So many substances dissolve in water that it has been called "the universal solvent." Many household products, including soft drinks, fruit juices, vinegar, cleaners, and medicines, are aqueous (water) solutions. ("Aqueous" comes from the Latin *aqua* for "water," as in aquatics.) Our blood plasma is mostly water, and many substances essential to life are dissolved in it, including glucose.

The ability of so many materials to dissolve in water also has some negative implications. Human activities have introduced thousands of unwanted substances into water supplies. These substances include paints, cleaners, industrial waste, insecticides, fertilizers, salt from highways, and other contaminants. Even the atmosphere is contaminated with gases produced when fossil fuels are burned. Rain, falling through these contaminants, may become acidic. From an Aboriginal perspective, we are all connected to water. It flows through us and does not stay in us. If water is contaminated, the contaminates will also flow through us. The effects of water contamination in Walkerton, Ontario were a tragic illustration of this connection and the importance of water to our survival. Learning about aqueous solutions and the limits to purity will help you understand science-related social issues forming around the quality of our water.

STARTING Points

Answer these questions as best you can with your current knowledge. Then, using the concepts and skills you have learned, you will revise your answers at the end of the chapter.

1. What happens when a pure substance dissolves in water? How is dissolving related to any subsequent chemical reactions of the solution?
2. List the different ways you can express the concentration of a solution. Are some ways more useful than others?
3. Is there a limit to how much of a substance dissolves to make a solution? Explain.

Career Connections:
Water and Waste Water Treatment Plant Operator; Toxicologist

Figure 1
The water in this waterfall is clear and clean, but is it pure?

Exploration — Substances in Water (Demonstration)

Even the "universal solvent" does not dissolve all solutes equally well. See what happens when you add different substances to water.

Materials: overhead projector; 5 petri dishes; water; 5 substances, such as fruit drink crystals, a marble chip (calcium carbonate), a sugar cube, a few drops of alcohol (ethanol), a few drops of vegetable oil

- Pour a few millilitres of water into each of five petri dishes, and then carefully add one of the substances to each dish, without stirring. Record your observations.

(a) Which substances dissolved in water?
(b) How certain are you about each substance in (a)? Give your reasons.
(c) Which substances do not appear to dissolve?
(d) How certain are you about your answer in (c)?
(e) Do the mixtures all have the same properties? Other than visible differences, hypothesize how they might differ.
(f) Design some tests for your hypotheses.

- Dispose of any solids into the waste paper basket and pour the liquids down the drain.

5.1 Solutions and Mixtures

Many of the substances that we use every day come dissolved in water. We buy other substances with little or no water, but then mix water with them before use. For example, we may purchase syrup, household ammonia, and pop with water already added, but we mix baking soda, salt, sugar, and powdered drinks with water. Most of the chemical reactions that you see in high school occur in a water environment. Indeed, most of the chemical reactions necessary for life on our planet occur in water.

Science and technology provide us with many useful products and processes that involve substances dissolved in water, such as cleaning solutions and pharmaceuticals. However, as with all technologies, there are risks and benefits arising from the use, misuse, or disposal of these products. A key to understanding the risks and benefits starts with understanding solutions, and in particular, solutions containing water.

Solutions

Solutions are homogeneous mixtures of substances composed of at least one **solute**—a substance that is dissolved, such as salt, NaCl—and one **solvent**—the medium in which a solute is dissolved, such as water. Most liquid-state and gas-state solutions are clear (transparent)—you can see through them; they are not cloudy or murky in appearance. Solutions may be coloured or colourless. Opaque or translucent (cloudy) mixtures, such as milk (Figure 1), contain undissolved particles large enough to block or scatter light waves. These mixtures are considered to be heterogeneous.

It is not immediately obvious whether a clear substance is pure or a mixture, but it is certainly homogeneous. If you were to do a chemical analysis of a sample of a homogeneous mixture (i.e., a solution), you would find that the proportion of each chemical in the sample remains the same, regardless of how small the sample is. This is explained by the idea that there is a uniform mixture of entities (atoms, ions, and/or molecules) in a solution. Empirically, a solution is homogeneous; theoretically, it is uniform at the atomic and molecular levels.

Both solutes and solvents may be gases, liquids, or solids, producing a number of different combinations (**Table 1**). In metal alloys, such as bronze, the dissolving has taken place in liquid form before the solution is used in solid form. Common liquid solutions that have a solvent other than water include varnish, spray furniture polish, and gasoline. Gasoline, for example, is a mixture of many different hydrocarbons and other compounds. These substances form a solution—a uniform mixture at the molecular level. There are many such hydrocarbon solutions, including kerosene (a Canadian-invented fuel for lamps and stoves), and turpentine (used for cleaning paintbrushes). Most greases and oils dissolve in hydrocarbon solvents.

Table 1 Classification of Solutions

Solute in solvent	Example of solution	Source or use
gas in gas	oxygen in nitrogen	air
gas in liquid	oxygen in water	water
gas in solid	oxygen in solid water	ice
liquid in liquid	methanol in water	antifreeze
solid in liquid	sugar in water	syrup
solid in solid	tin in copper	bronze alloy

DID YOU KNOW?

Alternative Meanings

Milk is sometimes labelled as "homogenized," meaning that the cream is equally distributed throughout the milk. This use of the word does not match the chemistry definition of homogeneous. Using the strict chemistry definition, milk is not a homogeneous mixture, but a heterogeneous mixture (**Figure 1**).

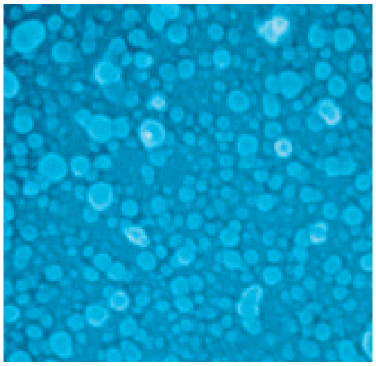

Figure 1
Milk is not a solution. This is quite obvious under magnification.

Other examples of liquids and solids dissolving in solvents other than water include the many chemicals that dissolve in alcohols. For example, solid iodine dissolved in ethanol (an alcohol) is used as an antiseptic (**Figure 2**). Some glues and sealants make use of other solvents: acetic acid is used as a solvent of the components of silicone sealants. You can smell the vinegar odour of acetic acid when sealing around tubs and fish tanks.

The chemical formula representing a solution specifies the solute by using its chemical formula and shows the solvent in parentheses. For example:

$NH_3(aq)$	ammonia gas (solute) dissolved in water (solvent)
$NaCl(aq)$	solid sodium chloride (solute) dissolved in water (solvent)
$I_2(alc)$	solid iodine (solute) dissolved in alcohol (solvent)
$C_2H_5OH(aq)$	liquid ethanol (solute) dissolved in water (solvent)

By far the most numerous and versatile solutions are those in which water is the solvent. Water can dissolve many substances, forming many unique solutions. All *aqueous solutions* have water as the solvent. They may be either coloured or colourless. Although water solutions are all different, they have some similarities and can be classified or described in a number of ways. This chapter deals primarily with the characteristics of aqueous solutions.

Figure 2
Tincture of iodine is a solution of the element iodine and the compound potassium or sodium iodide dissolved in ethanol. It is often found in first aid kits and is used to prevent the infection of minor cuts and scrapes.

mini *Investigation* — *Solutions and Reactions*

For a chemical reaction to occur, does it matter if the reactants are in their pure form or dissolved in a solution when mixed?

Materials: 3 small test tubes with stoppers, vials of $Pb(NO_3)_2(s)$ and $NaI(s)$, wash bottle with pure water, laboratory scoop, test tube rack or small beaker, lead waste container

(a) Describe the appearance of each solid.
- Place a few crystals of $Pb(NO_3)_2(s)$ in one clean, dry test tube.
- Add an equal quantity of $NaI(s)$ to the same test tube. Stopper and shake.

(b) Describe the appearance of the solid mixture.

(c) Is there any evidence of a chemical reaction? Justify your answer.
- Set up two clean, dry test tubes with separate, small quantities (a few crystals) of $Pb(NO_3)_2(s)$ and $NaI(s)$.
- Add pure water to each test tube to a depth of about one-quarter of the test tube. Stopper each test tube and shake to dissolve the solids.

(d) Describe the appearance of each solution.
- Remove the stoppers and pour the contents of one test tube into the other. Stopper and invert to mix.

(e) Is there any evidence of a chemical reaction? Justify your answer.

(f) Compare your answers to (c) and (e). What conclusion can be made?

(g) Suggest a hypothesis to explain your answer to (f).
- Dispose of all materials into the lead waste container.

Properties of Aqueous Solutions

Compounds can be classified as either electrolytes or nonelectrolytes. At this point we will restrict ourselves to compounds in aqueous solutions. Compounds are **electrolytes** if their aqueous solutions conduct electricity. Compounds are **nonelectrolytes** if their aqueous solutions do not conduct electricity. Most household aqueous solutions, such as fruit juices and cleaning solutions, contain electrolytes. The conductivity of a solution

Figure 3
The bulb in this conductivity apparatus lights up if the solute is an electrolyte.

is easily tested with a simple conductivity apparatus (**Figure 3**) or an ohmmeter. This evidence also provides a diagnostic test to determine the class of a solute—electrolyte or nonelectrolyte. This very broad classification of compounds into electrolyte and nonelectrolyte categories can be related to the main types of compounds classified in Chapter 2. Electrolytes are mostly highly soluble ionic compounds (such as $KBr(aq)$), including bases such as ionic hydroxides (for example, sodium hydroxide, $NaOH(aq)$). Most molecular compounds (such as ethanol, $C_2H_5OH(aq)$) are nonelectrolytes, with the exception of acids. Acids (such as nitric acid, $HNO_3(aq)$) are molecular compounds in their pure form but in aqueous solution, conduct electricity.

Another empirical method of classifying solutions uses litmus paper as a test to classify solutes as *acids*, *bases*, or *neutral* substances. Acids form acidic solutions, bases form basic solutions, and most other ionic and molecular compounds form neutral solutions (**Table 2**). These definitions of acids, bases, and neutral substances are empirical, based on the results of the litmus and conductivity tests. Later in this unit, you will encounter theoretical definitions.

Table 2 Properties of Solutes and Their Solutions

Type of solute	Conductivity test
electrolyte	light on conductivity apparatus glows; needle on ohmmeter moves compared to the control
nonelectrolyte	light on conductivity apparatus does not glow; needle on ohmmeter does not move compared to the control
Type of solution	**Litmus test**
acidic	blue litmus turns red
basic	red litmus turns blue
neutral	no change in colour of litmus paper

+ EXTENSION

Fastest Glacier
Water, life's universal solvent, has a very different impact on the environment depending on whether it is in its liquid or solid form. This simple difference is one of the major concerns of environmentalists today. This video shows how the rapid melting of glaciers could have a drastic impact on the environment worldwide.

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INVESTIGATION 5.1 ***Introduction***

Qualitative Chemical Analysis

Solutions have properties determined by the solute that is present. Diagnostic tests based on characteristic properties can be used to identify substances in a qualitative analysis.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ● Design
- ● Materials
- ● Procedure
- ● Evidence
- ● Analysis
- ○ Evaluation

Purpose

The purpose of this investigation is to use known diagnostic tests to distinguish among several pure substances.

Problem

Which of the white solids labelled 1, 2, 3, and 4 is calcium chloride, citric acid, glucose, and calcium hydroxide?

To perform this investigation, turn to page 227.

LAB EXERCISE 5.A

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ● Analysis
- ○ Evaluation

Identifying Solutions

For this investigation, assume that the labels on the four containers have been removed (perhaps washed off). Your task as a laboratory technician is to match the labels to the containers to identify the solutions.

Purpose

The purpose of this investigation is to use diagnostic tests to identify some solutions.

Problem

Which of the solutions labelled 1, 2, 3, and 4 is hydrobromic acid, sodium nitrate, lithium hydroxide, and methanol?

Design

Each solution is tested with both red and blue litmus paper and with conductivity apparatus. The temperature and concentration of the solutions are controlled variables.

Evidence

Table 3 Properties of the Unidentified Solutions

Solution	Red litmus	Blue litmus	Conductivity
1	red	blue	none
2	red	red	high
3	red	blue	high
4	blue	blue	high

Section 5.1 Questions

1. Classify the following mixtures as heterogeneous or homogeneous. Justify your answers.
   (a) fresh-squeezed orange juice
   (b) white vinegar
   (c) an old lead water pipe
   (d) humid air
   (e) a cloud
   (f) a dirty puddle
2. Which of the following substances are solutions?
   (a) milk
   (b) pop
   (c) pure water
   (d) smoke-filled air
   (e) silt-filled water
   (f) rainwater
3. State at least three ways of classifying solutions.
4. (a) Describe an aqueous solution.
   (b) Give at least five examples of aqueous solutions that you can find at home.
5. (a) What types of solutes are electrolytes?
   (b) Write a definition of an electrolyte.
6. Classify each compound as an electrolyte or a nonelectrolyte:
   (a) sodium fluoride (in toothpaste)
   (b) sucrose (table sugar)
   (c) calcium chloride (a road salt)
   (d) ethanol (in wine)
7. Based upon your current knowledge, classify each of the following compounds (**Figure 4**) as forming an acidic, basic, or neutral aqueous solution, and predict the colour of litmus in each solution.
   (a) $HCl(aq)$ (muriatic acid for concrete etching)
   (b) $NaOH(aq)$ (oven and drain cleaner)
   (c) methanol (windshield washer antifreeze)
   (d) sodium hydrogen carbonate (baking soda)

Figure 4
Everyday chemicals form acidic, basic, or neutral solutions.

8. When assessing the risks and benefits of solutions, it is useful to consider multiple perspectives such as scientific, technological, ecological, economic, and political. In a few words, describe the main focus of each of these perspectives.

9. The importance of water can be described from several perspectives. Write a brief statement illustrating the importance of water for each of the following perspectives: technological, economic, ecological, and political.
10. Since grease dissolves in gasoline, some amateur mechanics use gasoline to clean car, bicycle, or motorcycle parts in their basements. Why is this practice unsafe? What precautions would make the use of gasoline for this purpose safer?
11. Electrolytes are lost during physical activity and in hot weather through sweating. The body sweats in order to keep cool—cooling by evaporation of water. Sweating removes water and the substances dissolved in the water, such as salts and other electrolytes. We replace lost electrolytes by eating and drinking. By law, the ingredients of a food item are required to be placed on the label in decreasing order of quantity, as they are in sports drinks.
    (a) Classify the ingredients of the sports drink in **Figure 5** as electrolytes or nonelectrolytes. How does the number and quantity of electrolytes and nonelectrolytes compare?
    (b) Which ingredients contain sodium ions? Which contain potassium ions? Are there more sodium or potassium ions in the drink? Justify your answer.
    (c) Does the most energy in the drink come from proteins, carbohydrates, or fats (oils)?
    (d) What three chemical needs does the drink attempt to satisfy?

Figure 5
Manufacturers recommend sports drinks to athletes, to restore electrolytes to the body.

Extension

12. Aboriginal peoples in North America use many solutions as medicine, both internal and external. Often, the solutions are like a tea, made by placing parts of plants in hot water. Traditional knowledge of the type of plant to use for different ailments also includes where to find the plant, what parts to use, and what procedure to follow.
    (a) Why were rosehips important to Aboriginal peoples of Canada's northwest like the Dene Tha' as well as to the first Europeans who came in contact with them (**Figure 6**)?
    (b) Given that vitamin C is a polar compound while vitamin A is nonpolar, which vitamin would be present in the greatest amount in a solution (tea) of rose hips?

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Figure 6
Aboriginal peoples have traditionally used rosehips, as well as a wide variety of other natural remedies, for their well-being and survival.

Explaining Solutions 5.2

Water is the most important solvent on Earth. The oceans, lakes, rivers, and rain are aqueous solutions containing many different ionic compounds and a few molecular solutes. As you know, there are some ionic compounds that dissolve only very slightly in water, such as limestone (calcium carbonate) and various other rocks and minerals. Nevertheless, many more ionic compounds dissolve in water than in any other known solvent.

Why are ionic compounds so soluble in water? The key to the explanation came from the study of electrolytes. Electrolytes were first explained by Svante Arrhenius who was born in Wijk, Sweden, in 1859. While attending the University of Uppsala, he became intrigued by the problem of how and why some aqueous solutions conduct electricity, but others do not. This problem had puzzled chemists ever since Sir Humphry Davy and Michael Faraday performed experiments over half a century earlier, passing electric currents through chemical substances.

Faraday believed that an electric current produces new charged particles in a solution. He called these electric particles *ions* (a form of the Greek word for "to go"). He could not explain what ions were, or why they did not form in some solutions such as sugar or alcohol dissolved in water.

In 1887, Arrhenius proposed that when a substance dissolves, particles of the substance separate from each other and disperse into the solution. Nonelectrolytes disperse electrically neutral particles throughout the solution. As **Figure 1** shows, molecules of sucrose (a nonelectrolyte) separate from each other and disperse in an aqueous solution as individual molecules of sucrose surrounded by water molecules.

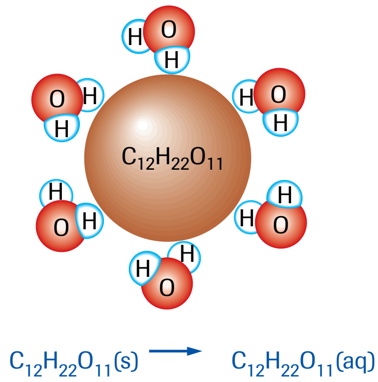

Figure 1
This model illustrates sucrose dissolved in water. The model, showing electrically neutral particles in solution, agrees with the evidence that a sucrose solution does not conduct electricity.

But what about the conductivity of solutions of electrolytes? Arrhenius' explanation for this observation was quite radical. He agreed with the accepted theory that electric current involves the movement of electric charge. Ionic compounds form conducting solutions. Therefore, according to Arrhenius, electrically charged particles must be present in their solutions. We now believe that, when a compound such as table salt dissolves, existing ions from the solid crystal lattice are separated as individual aqueous ions (**Figure 2**). The positive ions are surrounded by the negative ends of the polar water molecules, while the negative ions are surrounded by the positive ends of the polar water molecules (Chapter 3, page 101). **Dissociation** describes the separation of ions that occurs when an ionic compound dissolves in water. Dissociation equations, such as the following examples, show this separation of ions:

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

$$(NH_4)_2SO_4(s) \rightarrow 2\,NH_4^{+}(aq) + SO_4^{2-}(aq)$$

Notice that the formula for the solvent, $H_2O(l)$, does not appear as a reactant in the equation. Although water is necessary for the process of dissociation, it is not consumed and hence is not a reactant. The presence of water molecules surrounding the ions is indicated by (aq).

DID YOU KNOW ?

Restricting Terminology
Dissociation, strictly speaking, is the separation and dispersal of *any* entities that were initially bonded together. For our purposes, we restrict this term to refer to the separation and dispersal of ions as ionic compounds dissolve in water.

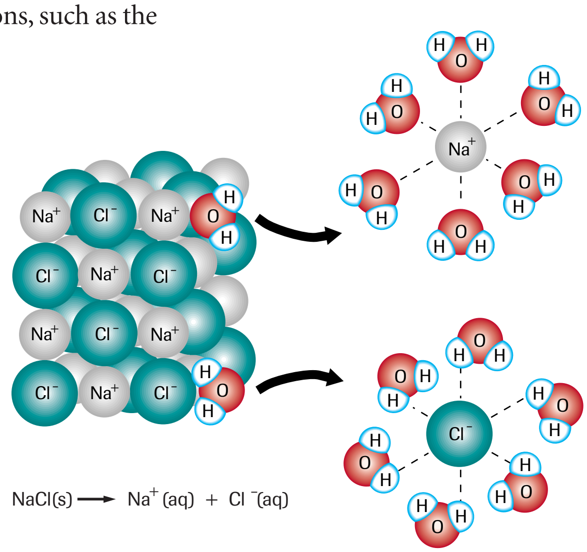

Figure 2
This model represents the dissociation of sodium chloride into positive and negative ions.

Practice

1. Suppose you place a sugar cube (sucrose) and a lump of salt (sodium chloride) into separate glasses of water.
   (a) Predict as many observations as you can about each mixture.
   (b) According to theory, how is the dissolving process similar for both solutes?
   (c) According to theory, how are the final solutions different?
   (d) What theoretical properties of a water molecule help to explain the dissolving of both solutes?
2. Write equations to represent the dissociation of the following ionic compounds when they are placed in water:
   (a) sodium fluoride
   (b) sodium phosphate
   (c) potassium nitrate
   (d) cobalt(II) chloride
   (e) aluminium sulfate
   (f) ammonium hydrogen phosphate

Acids and Bases

Arrhenius eventually extended his theory to explain some of the properties of acids and bases (**Table 1**). According to Arrhenius, bases are ionic hydroxide compounds that dissociate into individual positive ions and negative hydroxide ions in solution. He believed the hydroxide ion was responsible for the properties of basic solutions; for example, turning red litmus paper blue. The dissociation of bases is similar to that of any other ionic compound, as shown in the following dissociation equation for barium hydroxide:

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

Acids, however, are a different story. The properties of acids appear only when compounds containing hydrogen, such as $HCl(g)$ and $H_2SO_4(l)$, dissolve in water. Since acids are electrolytes, the accepted theory is that acidic solutions must contain ions. However, the pure compounds that become acids in solution are usually molecular, so only neutral molecules are initially present. This unique behaviour requires an explanation other than dissociation. According to Arrhenius, acids ionize into positive hydrogen ions and negative ions when dissolved in water.

Ionization is the process by which a neutral atom or molecule is converted to an ion. In the case of acids, Arrhenius assumed that the water solvent somehow causes the acid molecules to ionize, but he did not propose an explanation for this process. The aqueous hydrogen ions are believed to be responsible for changing the colour of litmus in an acidic solution. Hydrogen chloride gas dissolving in water to form hydrochloric acid is a typical example of an acid. An ionization equation, shown below, is used to communicate this process:

$$\underset{\text{neutral molecules}}{HCl(g)} \xrightarrow{\text{ionization}} \underbrace{H^+(aq) + Cl^-(aq)}_{\text{ions}}$$

Arrhenius' theory was a major advance in understanding chemical substances and solutions. Arrhenius also provided the first comprehensive theory of acids and bases. The empirical and theoretical definitions of acids and bases are summarized in Table 1.

DID YOU KNOW?

When is an acid not an acid?
Pure acetic acid, $CH_3COOH(l)$, or a solution of acetic acid in a nonpolar organic solvent such as gasoline, does not conduct electricity or change the colour of litmus. This behaviour is perfectly consistent with molecular substances. As soon as acetic acid is dissolved in water, however, the solution conducts electricity and changes the colour of litmus from blue to red.

EXTENSION

Dissociation vs. Ionization
What is the difference between dissociation and ionization? Both may produce aqueous ions. Dissociation, however, is the separation of ions that already exist before dissolving in water.

$$M^+X^-(s) \rightarrow M^+(aq) + X^-(aq)$$

Ionization involves the production of new ions, specifically hydrogen ions, in the case of acid solutions.

$$HX(aq) \rightarrow H^+(aq) + X^-(aq)$$

You can listen to a detailed discussion of when it is appropriate to use each of these theoretical terms.

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Table 1 Acids, Bases, and Neutral Substances

Type of substance	Empirical definition	Arrhenius' theory
acids	Acids form solutions that • turn blue litmus red and are electrolytes • neutralize bases	• some hydrogen compounds ionize to produce $H^+(aq)$ ions • $H^+(aq)$ ions react with $OH^-(aq)$ ions to produce water
bases	Bases form solutions that • turn red litmus blue and are electrolytes • neutralize acids	• ionic hydroxides dissociate to produce $OH^-(aq)$ ions • $OH^-(aq)$ ions react with $H^+(aq)$ ions to produce water
neutral substances	Neutral substances form solutions that • do not affect litmus • some are electrolytes • some are nonelectrolytes	• no $H^+(aq)$ or $OH^-(aq)$ ions are formed • some are ions in solution • some are molecules in solution

Energy Changes

You already know that chemical reactions can be endothermic (absorb energy from the surroundings) or exothermic (release energy to the surroundings). What about the formation of solutions? Is energy absorbed or released when a substance dissolves in water?

mini *Investigation* — *Hot and Cold Solutions*

In this short investigation, you will determine if changes are endothermic or exothermic simply by feeling the changes in the palm of your hand.

Materials: small plastic bag, plastic spoon, calcium chloride solid, sodium nitrate solid, 10 mL graduated cylinder

- Place about one teaspoon of calcium chloride in a bottom corner of a dry plastic bag.
- Measure about 5 mL of tap water and place into the other corner of the bag, keeping it separate from the solid.
- Lift the bag by the top, shake gently to mix, and immediately place into the palm of your hand.

(a) Is the dissolving of calcium chloride endothermic or exothermic? Justify your answer.

- Dispose of contents into the sink. Rinse and dry the inside of the bag.
- Repeat this procedure using sodium nitrate and dispose of the contents into a waste container.

(b) Is the dissolving of sodium nitrate endothermic or exothermic? Justify your answer.

(c) Suggest some practical applications of this investigation.

Solids may cause irritation to the eyes, skin, and respiratory tract, and are harmful if swallowed. Wear eye protection and avoid direct contact.

At this stage in your chemical education, it is difficult to predict whether the dissolving of a particular solute will be endothermic or exothermic. More extensive investigations have clearly shown that all solution formation involves some energy change. The accepted explanation depends upon two general theoretical principles:

- Breaking existing bonds uses energy.
- Forming new bonds releases energy.

In the dissolving of an ionic solid such as sodium chloride (see Figure 2, page 197), the ionic bond between the positive and negative ions needs to be broken to separate the ions. Energy is also required to overcome the intermolecular forces among the water molecules.

Simultaneously, new bonds between the individual ions and polar water molecules need to form, as shown in Figure 2.

$$\underbrace{NaCl(s) \rightarrow Na^+ + Cl^-}_{\text{uses energy}} \underbrace{\rightarrow Na^+(aq) + Cl^-(aq)}_{\text{releases energy}}$$

$$NaCl(s) \text{ in } H_2O(l) \longrightarrow Na^+(aq) + Cl^-(aq)$$

Bonds broken
- ionic bonds in solid
- intermolecular forces between water molecules

Energy absorbed

Bonds formed
- electrostatic forces between ions and water molecules

Energy released

Net energy change

We can only observe the effects of the net energy change, usually as a change in temperature. Based on the idea of energy changes in bond breaking and forming, we can explain that a dissolving process is either *endothermic* because more energy is absorbed than released or *exothermic* because more energy is released than absorbed. Typical of a simple theory, it is able to explain this process once you know the experimental answer, but is unable to predict the outcome.

When calcium chloride dissolves in water, for example, the evidence (increase in temperature) shows that the overall dissolving process is exothermic. Therefore, more energy must be released when the individual ions bond to water molecules than is used to break the bonds between the ions in the solid calcium chloride and the intermolecular forces between water molecules. For endothermic dissolving, the opposite applies: When sodium nitrate dissolves in water, the temperature of the solution decreases. Therefore, heat is absorbed from the surroundings as the sodium nitrate dissolves. Theoretically, this means that less energy is released when the solute particles bond to water molecules than is used to break the bonds between the particles in the pure substances. You will study endothermic and exothermic changes in more detail in Unit 6.

CAREER CONNECTION

Water and Waste Water Treatment Plant Operator

These operators are essential to everyone's health and safety. They monitor and control our water purification and waste-water treatment facilities. As stewards of our drinking water and freshwater ecosystems, these water specialists use a variety of skills to manage public safety and health. Learn more about the variety of duties and certification requirements for this career direction.

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Substances in Water

The focus in this unit is on substances that dissolve in water. However, many naturally occurring substances, such as limestone (mainly calcium carbonate), do not dissolve in water to an appreciable extent. This property is fortunate for us because limestone is a common building material that would not be very useful if it dissolved when it rained. On the other hand, sodium chloride is a naturally occurring compound found in large underground deposits near Fort Saskatchewan. As you know, sodium chloride easily dissolves in water, and chemical engineers make good use of this property in a process called solution mining—extracting substances by dissolving them in water. Knowledge of the properties of materials is useful for meeting societal needs such as building materials, salt for flavouring, or salt as a raw material to make other useful products. The solubility of ionic compounds (including bases such as ionic hydroxides) can be predicted from a solubility chart (inside back cover of this textbook).

Acids also occur naturally. For example, you have hydrochloric acid in your stomach, carbonic acid is in natural rain, and acetic acid forms during the fermentation process. Acidic solutions vary in their electrical conductivity. Acids that are extremely good conductors are called *strong acids*. Sulfuric acid, nitric acid, and hydrochloric acid are examples of strong acids that are almost completely ionized when in solution. (These strong

BIOLOGY CONNECTION

Ions in Blood

Human blood plasma contains the following ions: $Na^+(aq)$, $K^+(aq)$, $Ca^{2+}(aq)$, $Mg^{2+}(aq)$, $HCO_3^-(aq)$, $Cl^-(aq)$, $HPO_4^{2-}(aq)$, and $SO_4^{2-}(aq)$, as well as many complex acid and protein molecules. A physician can gain information about the state of your health by testing for the quantity of these substances present in a blood sample.

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acids are listed on the inside back cover of this textbook under "Concentrated Reagents.") Most other common acids, such as acetic acid, are *weak acids*. The conductivity of acidic solutions varies a great deal. The accepted explanation, presented in more detail in Chapter 6, is that the degree of ionization of acids varies.

Molecular compounds abound in nature. They include most of the compounds that make up living things, as well as fossil fuels such as oil and natural gas. Like ionic compounds, the solubility of molecular compounds in water varies tremendously, but unlike ionic compounds, there is no simple solubility chart to make specific predictions. According to theories of intermolecular bonding (Chapter 3), however, we can make some general predictions about solubility in water—a polar liquid with hydrogen bonding. Nonpolar molecular compounds generally do not dissolve in water; polar compounds may be slightly soluble in water; and polar compounds with hydrogen bonding are the most likely to be very soluble in water. For efficiency in studying the examples in this textbook, you should memorize the examples in **Table 2**.

To understand the properties of aqueous solutions and the reactions that take place in solutions, it is necessary to know the major entities present when any substance is in a water environment. **Table 3** summarizes this information. The information is based on the solubility and electrical conductivity of substances as determined in the laboratory. Your initial work in chemistry will deal mainly with strong acids and other highly soluble compounds.

Table 2 Solubility of Selected Molecular Compounds

Solubility	Examples
high	ammonia, $NH_3(g)$ hydrogen peroxide, $H_2O_2(l)$ methanol, $CH_3OH(l)$ ethanol, $C_2H_5OH(l)$ glucose, $C_6H_{12}O_6(s)$ sucrose, $C_{12}H_{22}O_{11}(s)$
low	methane, $CH_4(g)$ propane, $C_3H_8(g)$ octane, $C_8H_{18}(l)$

Table 3 Major Entities Present in a Water Environment

Type of substance	Solubility in water	Typical pure substance	Major entities present when substance is placed in water
ionic compounds	high	$NaCl(s)$	$Na^+(aq)$, $Cl^-(aq)$, $H_2O(l)$
	low	$CaCO_3(s)$	$CaCO_3(s)$, $H_2O(l)$
bases	high	$NaOH(s)$	$Na^+(aq)$, $OH^-(aq)$, $H_2O(l)$
	low	$Ca(OH)_2(s)$	$Ca(OH)_2(s)$, $H_2O(l)$
molecular substances	high	$C_{12}H_{22}O_{11}(s)$	$C_{12}H_{22}O_{11}(aq)$, $H_2O(l)$
	low	$C_8H_{18}(l)$	$C_8H_{18}(l)$, $H_2O(l)$
strong acids	high	$HCl(g)$	$H^+(aq)$, $Cl^-(aq)$, $H_2O(l)$
weak acids	high	$CH_3COOH(l)$	$CH_3COOH(aq)$, $H_2O(l)$
elements	low	$Cu(s)$	$Cu(s)$, $H_2O(l)$
		$N_2(g)$	$N_2(g)$, $H_2O(l)$

SUMMARY *Explaining Solutions*

Table 4 Arrhenius' Theory of Solutions

Substance	Process	General equation
molecular	disperse as individual molecules	$XY(s/l/g) \rightarrow XY(aq)$
ionic	dissociate as individual cations and anions	$MX(s) \rightarrow M^+(aq) + X^-(aq)$
base (ionic hydroxide)	dissociate as cations and hydroxide ions	$MOH(s) \rightarrow M^+(aq) + OH^-(aq)$
acid	ionize to form new hydrogen ions and anions	$HX(s/l/g) \rightarrow H^+(aq) + X^-(aq)$

LAB EXERCISE 5.B

Qualitative Analysis

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ● Analysis
- ● Evaluation (1, 3)

In your evaluation, suggest improvements to the Design, using your knowledge of chemicals and **Table 5**.

Table 5 Electrical Conductivity

Class	Solid	Liquid	Aqueous
metal	✓	✓	–
nonmetal	X	X	–
ionic	X	✓	✓
molecular	X	X	X
acid	X	X	✓

Purpose

The purpose of this investigation is to test the diagnostic tests for different classes of substances as a means of qualitative analysis.

Problem

Which of the chemicals numbered 1 to 7 is $KCl(s)$, $Ba(OH)_2(s)$, $Zn(s)$, $C_6H_5COOH(s)$, $Ca_3(PO_4)_2(s)$, $C_{25}H_{52}(s)$ (paraffin wax), and $C_{12}H_{22}O_{11}(s)$?

Design

The chemicals are tested for solubility, conductivity, and effect on litmus paper. Equal amounts of each chemical are added to equal volumes of water.

Evidence

Table 6 Solubility, Conductivity, and Litmus Test Results

Chemical	Solubility in water	Conductivity of solution	Effect of solution on litmus paper
1	high	none	no change
2	high	high	no change
3	none	none	no change
4	high	high	red to blue
5	none	none	no change
6	none	none	no change
7	low	low	blue to red

Section 5.2 Questions

1. In your own words, describe what is believed to happen when an ionic compound dissolves in water.
2. Write dissociation or ionization equations for each of the following pure substances when they dissolve in water:
   (a) $CaCl_2(s)$ (road salt)
   (b) $HF(g)$ (etching glass)
   (c) $(NH_4)_2HPO_4(s)$ (fertilizer)
   (d) $Al_2(SO_4)_3(s)$ (making pickles)
3. Compare dissociation and ionization by listing similarities and differences.
4. According to Arrhenius' theory, what is the explanation for
   (a) an acid turning blue litmus red?
   (b) a base turning red litmus blue?
   (c) neutralization of an acid and a base?
5. List three examples of solutions in consumer products.
6. A key characteristic of science is the goal of explaining natural products and processes. Do we need a theoretical explanation of solutions in order to use them? Answer from consumer and Aboriginal perspectives.
7. Many substances dissolve in water because water is such a polar solvent.
   (a) Are energy changes always involved when substances dissolve in water? Justify your answer.
   (b) Describe a brief experimental design to test your answer to (a).
   (c) What are some limitations that might be encountered if you were to perform this experiment?
8. List the chemical formulas for the major entities present in water for each of the following:
   (a) zinc
   (b) sodium bromide
   (c) oxygen
   (d) nitric acid
   (e) calcium phosphate
   (f) methanol
   (g) aluminium sulfate
   (h) potassium dichromate
   (i) acetic acid
   (j) sulfur
   (k) copper(II) sulfate
   (l) silver chloride
   (m) paraffin wax, $C_{25}H_{52}(s)$
9. Why is water such a good solvent for dissolving many ionic and molecular compounds? How is this property an advantage and a disadvantage?
10. The dissolving of calcium chloride in water is very exothermic compared with dissolving sodium chloride in water. Would calcium chloride be an appropriate substitute for a sidewalk deicer? Identify some positive and negative aspects, including several perspectives.

Extension

11. List some personal values demonstrated by Svante Arrhenius while developing his ideas about solutions.

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12. List some vitamins that are water soluble and some that are fat soluble. Using one example of each, draw a structural formula and explain its solubility. How does the solubility in water and fat relate to how quickly a vitamin is excreted?

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Solution Concentration 5.3

Most aqueous solutions are colourless, so there is no way of knowing, by looking at them, how much of the solute is present in the solution. As we often need to know the quantity of solute in the solution, it is important that solutions be labelled with this information. We use a ratio that compares the quantity of solute to the quantity of the solution. This ratio is called the solution's **concentration**. Chemists describe a solution of a given substance as *dilute* if it has a relatively small quantity of solute per unit volume of solution (**Figure 1**). A *concentrated* solution, on the other hand, has a relatively large quantity of solute per unit volume of solution.

In general, the concentration, c, of any solution is expressed by the ratio

$$\text{concentration} = \frac{\text{quantity of solute}}{\text{quantity of solution}}$$

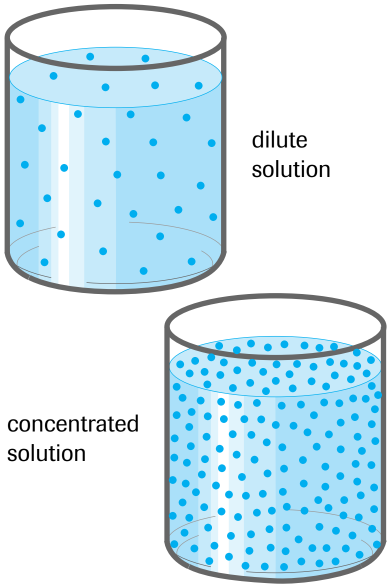

Figure 1
The theoretical model of the dilute solution shows fewer solute entities (particles) per unit volume compared with the model of the concentrated solution.

Percentage Concentration

Many consumer products, such as vinegar (acetic acid), are conveniently labelled with their concentration ratios expressed as percentages (**Figure 2**). A vinegar label listing "5% acetic acid (by volume)" means that there is 5 mL of pure acetic acid dissolved in every 100 mL of the vinegar solution. This type of concentration is often designated as % V/V, percentage volume by volume, or percentage by volume.

$$c_{CH_3COOH} = \frac{5\text{ mL}}{100\text{ mL}} = 5\%\text{ V/V}$$

In general, a percentage by volume concentration may be defined as

$$\boldsymbol{c = \frac{V_{\text{solute}}}{V_{\text{solution}}} \times 100\%}$$

COMMUNICATION *example* 1

A photographic "stop bath" contains 140 mL of pure acetic acid in a 500 mL bottle of solution. What is the percentage by volume concentration of acetic acid?

Solution

$$c_{CH_3COOH} = \frac{140\text{ mL}}{500\text{ mL}} \times 100\%$$

$$= 28.0\%\text{ V/V}$$

The percentage by volume concentration of acetic acid is 28.0%, or the concentration of acetic acid is 28.0% V/V.

Another common concentration ratio used for consumer products is "percentage weight by volume" or % W/V. (In consumer and commercial applications, "weight" is used instead of "mass," which explains the W in the W/V label.) For example, a hydrogen peroxide topical solution used as an antiseptic is 3% W/V (**Figure 2**). This means that 3 g of hydrogen peroxide is dissolved in every 100 mL of solution.

$$c_{H_2O_2} = \frac{3\text{ g}}{100\text{ mL}}$$

$$= 3\%\text{ W/V}$$

Figure 2
The concentrations of different consumer products are usually expressed as a percentage because percentages are generally easy for consumers to understand.

In general, we write a percentage weight by volume concentration as

$$c = \frac{m_{solute}}{V_{solution}} \times 100\%$$

A third concentration ratio is the "percentage weight by weight," or % W/W:

$$c = \frac{m_{solute}}{m_{solution}} \times 100\%$$

CAREER CONNECTION

Toxicologist

Toxicologists are investigators who specialize in detecting poisons and other harmful substances. Toxicologists can specialize in many areas, including analyzing natural substances such as venoms, testing new industrial products, or, in forensic science, testing for poisons in suspicious deaths. Learn more about possibilities of working as a toxicologist, including different specializations, education, and salary.

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COMMUNICATION example 2

A sterling silver ring has a mass of 12.0 g and contains 11.1 g of pure silver. What is the percentage weight by weight concentration of silver in the metal?

Solution

$$\begin{aligned} c_{Ag} &= \frac{11.1\ \text{g}}{12.0\ \text{g}} \times 100\% \\ &= 92.5\%\ \text{W/W} \end{aligned}$$

The percentage weight by weight of silver is 92.5%, or the concentration of silver is 92.5% W/W.

DID YOU KNOW?

Understanding ppm

- A concentration of 1% W/W is equivalent to 10 000 ppm.
- 1 ppm is approximately the concentration obtained by dissolving one grain of salt (about 0.1 mg) in a 100 mL glass of water.

Parts per Million Concentration

In studies of solutions in the environment, we often encounter very low concentrations. For very dilute solutions, we choose a concentration unit to give reasonable numbers for very small quantities of solute. For example, the concentration of toxic substances in the environment, of chlorine in a swimming pool, and of impurities in laboratory chemicals (**Figure 3**) is usually expressed as parts per million (ppm, $1{:}10^6$) or even smaller ratios, such as parts per billion (ppb, $1{:}10^9$) or parts per trillion (ppt, $1{:}10^{12}$). These ratios are a special case of the weight by weight (W/W) ratio. By definition, parts per million means

$$\frac{m_{solute}(\text{g})}{m_{solution}(\text{g})} \times 10^6$$

However, this calculation can be simplified by altering the units to incorporate the factor of 10^6. Therefore, a concentration in parts per million (ppm) can also be expressed as

$$c = \frac{m_{solute}(\text{mg})}{m_{solution}(\text{kg})}$$

which means that 1 ppm = 1 mg/kg.

Because very dilute aqueous solutions are similar to pure water, their densities are considered to be the same: 1 g/mL. Therefore, 1 ppm of chlorine is 1 g in 10^6 g or 10^6 mL (1000 L) of pool water, which is equivalent to 1 mg of chlorine per litre of water. For dilute aqueous solutions only,

$$1\ \text{ppm} = 1\ \text{g}/10^6\ \text{mL} = 1\ \text{mg/L} = 1\ \text{mg/kg}$$

Small concentrations such as ppm, ppb, and ppt are difficult to imagine, but are very important in environmental studies and in the reporting of toxic effects (**Table 1**).

Figure 3
This concentrated sulfuric acid has an amount concentration of 17.8 mol/L. Note the impurities listed in units of parts per million.

COMMUNICATION example 3

Dissolved oxygen in natural waters is an important measure of the health of the ecosystem. In a chemical analysis of 250 mL of water at SATP, 2.2 mg of oxygen was measured. What is the concentration of oxygen in parts per million?

Solution

$$c_{O_2} = \frac{2.2 \text{ mg}}{0.250 \text{ L}}$$
$$= 8.8 \text{ mg/L}$$
$$= 8.8 \text{ ppm}$$

The concentration of dissolved oxygen is 8.8 ppm.

Table 1 Maximum Acceptable Concentration (MAC) of Chemicals in Canadian Drinking Water

Substance	Typical source	MAC (ppm)
cadmium	batteries in landfills	0.005
lead	old plumbing	0.010
nitrates	fertilizers	45.0
cyanides	mining waste	0.2

Amount Concentration

Chemistry is primarily the study of chemical reactions, which we communicate using balanced chemical equations. The coefficients in these equations represent chemical amounts in units of moles. Concentration is therefore communicated using amount concentration. **Amount concentration**, *c*, is the chemical amount of solute dissolved in one litre of solution.

$$\text{amount concentration} = \frac{\text{chemical amount of solute (in moles)}}{\text{volume of solution (in litres)}}$$

$$c = \frac{n}{V}$$

The units of amount concentration (mol/L) come directly from this ratio.

Amount concentration can also be indicated by the use of square brackets. For example, the amount concentration of sodium hydroxide in water could be represented by [NaOH(aq)].

DID YOU KNOW ?

Amount Concentration

According to IUPAC, the molar concentration of X is now officially called the "amount concentration of X" and is denoted by c_X or [X]. Amount concentration replaces the older terms, molar concentration and molarity. Remember that in chemistry, the "amount" of a substance is always a quantity in units of moles.

COMMUNICATION example 4

In a quantitative analysis, a stoichiometry calculation produced 0.186 mol of sodium hydroxide in 0.250 L of solution. Calculate the amount concentration of sodium hydroxide.

Solution

$$c_{NaOH} = \frac{0.186 \text{ mol}}{0.250 \text{ L}}$$
$$= 0.744 \text{ mol/L}$$

The amount concentration of sodium hydroxide is 0.744 mol/L.

Practice

1. Describe the three different systems of expressing the concentration of a solution.
2. Gasohol, a solution of ethanol and gasoline, is considered to be a cleaner fuel than gasoline alone. A typical gasohol mixture available across Canada contains 4.1 L of ethanol in a 55 L tank of fuel. Calculate the percentage by volume concentration of ethanol.
3. Solder flux, available at hardware and craft stores, contains 16 g of zinc chloride in 50 mL of solution. The solvent is aqueous hydrochloric acid. What is the percentage weight by volume of zinc chloride in the solution?

4. Brass is a copper–zinc alloy. If the concentration of zinc is relatively low, the brass has a golden colour and is often used for inexpensive jewellery. If a 35.0 g pendant contains 1.7 g of zinc, what is the percentage weight by weight of zinc in this brass?
5. Formaldehyde, $CH_2O(g)$, an indoor air pollutant that is found in synthetic materials and cigarette smoke, is a carcinogen. If an indoor air sample with a mass of 0.59 kg contained 3.2 mg of formaldehyde, this level would be considered dangerous. What would be the concentration of formaldehyde in parts per million?
6. A plastic dropper bottle for a chemical analysis contains 0.11 mol of calcium chloride in 60 mL of solution. Calculate the amount concentration of calcium chloride.

Figure 4
David Schindler

WWW **WEB Activity**

Canadian Achievers—David Schindler

Dr. David Schindler (**Figure 4**) is a professor of ecology at the University of Alberta, where he specializes in researching land-water interactions. State three specific examples of studies Dr. Schindler has conducted that would involve concentration measurements. Identify some personal values and attitudes that make him both a renowned and a controversial scientist.

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Calculations Involving Concentrations

Solutions are so commonly used in chemistry that calculating concentrations might be the primary reason why chemists pull out their calculators. Chemists and chemical technicians also frequently need to calculate a quantity of solute or solution. Any of these calculations may involve percentage concentrations, ppm concentrations, or amount concentrations. When we know two of these values—quantity of solute, quantity of solution, and concentration of solution—we can calculate the third quantity. Because concentration is a ratio, a simple procedure is to use the concentration ratio (quantity of solute/quantity of solution) as a conversion factor. This approach parallels the one you followed when using molar mass as a conversion factor.

Suppose you are a nurse who needs to calculate the mass of dextrose, $C_6H_{12}O_6(s)$, present in a 1000 mL intravenous feeding of D5W, which is a solution of 5.0% W/V dextrose in water. The conversion factor you need to use is the mass/volume ratio:

$$m_{C_6H_{12}O_6} = 1000\ \cancel{\text{mL}} \times \frac{5.0\ \text{g}}{100\ \cancel{\text{mL}}} = 50\ \text{g}$$

In some calculations, you may want to find the quantity of solution, in which case you will have to invert the ratio to quantity of solution/quantity of solute. This is then the appropriate conversion factor.

For example, what volume of 30.0% W/V hydrogen peroxide solution can be made from 125 g of pure hydrogen peroxide? You know that the answer must be greater than 100 mL because 125 g is greater than 30.0 g (the quantity in 100 mL). Notice how the units cancel to produce the expected volume unit, millilitres, when we use the volume/mass ratio:

$$V_{H_2O_2} = 125\ \cancel{\text{g}} \times \frac{100\ \text{mL}}{30.0\ \cancel{\text{g}}} = 417\ \text{mL}$$

Learning Tip

You can set up calculations involving concentrations as a proportion. For example,

$$\frac{m_{C_6H_{12}O_6}}{1000\ \text{mL}} = \frac{5.0\ \text{g}}{100\ \text{mL}}$$

You can invert the ratios, if it is more convenient. For example,

$$\frac{V_{H_2O_2}}{125\ \text{g}} = \frac{100\ \text{mL}}{30.0\ \text{g}}$$

In both cases, make sure the quantities in the numerator and the denominator on both sides of the equation match.

Thinking about the quantity given and the concentration ratio helps to ensure you are calculating correctly. This method also works for other concentration ratios.

COMMUNICATION *example* 5

A box of apple juice has a fructose (sugar) concentration of 12 g/100 mL (12% W/V) (**Figure 5**). What mass of fructose is present in a 175 mL glass of juice? (The chemical formula for fructose is $C_6H_{12}O_6$.)

Solution

$$m_{C_6H_{12}O_6} = 175\ \cancel{\text{mL}} \times \frac{12\ \text{g}}{100\ \cancel{\text{mL}}}$$
$$= 21\ \text{g}$$

The mass of fructose present in 175 mL of apple juice is 21 g.

COMMUNICATION *example* 6

People with diabetes have to monitor and restrict their sugar intake. What volume of apple juice could a diabetic person drink, if the person's sugar allowance for that beverage was 9.0 g? Assume that the apple juice has a sugar concentration of 12 g/100 mL (12% W/V), and that the sugar in apple juice is fructose.

Solution

$$V_{C_6H_{12}O_6} = 9.0\ \cancel{\text{g}} \times \frac{100\ \text{mL}}{12\ \cancel{\text{g}}}$$
$$= 75\ \text{mL}$$

The volume of apple juice allowed is 75 mL.

Figure 5
The label on a box of apple juice gives the ingredients and some nutritional information, but not the concentration of the various solutes.

When you are given a concentration in parts per million (ppm), it is usually easier to convert the parts per million into units of milligrams per kilogram (1 ppm = 1 mg/kg) before doing your calculation. Remember that 1 ppm = 1 mg/L only applies to aqueous solutions. If, for example, you are given a value of 99 ppm of DDT in a 2 kg gull, what mass of DDT is present? The concentration ratio is 99 mg/kg. Note the cancellation of kilograms.

$$m_{\text{DDT}} = 2\ \cancel{\text{kg}} \times \frac{99\ \text{mg}}{1\ \cancel{\text{kg}}}$$
$$= 0.2\ \text{g (rounded from 198 mg)}$$

COMMUNICATION *example* 7

A sample of well water contains 0.24 ppm of iron(III) sulfate dissolved from the surrounding rocks. What mass of iron(III) sulfate is present in 1.2 L of water in a kettle?

Solution

$$m_{Fe_2(SO_4)_3} = 1.2\ \cancel{\text{L}} \times \frac{0.24\ \text{mg}}{1\ \cancel{\text{L}}}$$
$$= 0.29\ \text{mg}$$

The mass of iron(III) sulfate in 1.2 L is 0.29 mg.

BIOLOGY CONNECTION

Pollutants

The effects of pollutants in the environment, such as toxicity, are an important topic in biology. You will see a much more detailed discussion of this topic if you are taking a biology course.

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Figure 6
Aqueous ammonia is purchased for science laboratories as a concentrated solution.

COMMUNICATION *example* 8

A sample of laboratory ammonia solution has an amount concentration of 14.8 mol/L (**Figure 6**). What chemical amount of ammonia is present in a 2.5 L bottle?

Solution

$$n_{NH_3} = 2.5\ \cancel{L} \times \frac{14.8\ mol}{1\ \cancel{L}}$$

$$= 37\ mol$$

The chemical amount of ammonia in 2.5 L is 37 mol.

You should always check that your answer makes sense. For example, in Communication Example 8, 14.8 mol/L means that there is 14.8 mol of ammonia in 1 L of solution. Therefore, 2.5 L, which is greater than 1 L, must contain a chemical amount greater than 14.8 mol.

In some situations, you may know the amount concentration and need to find either the volume of solution or amount (in moles) of solute. In these situations, use either the volume/amount or amount/volume ratio. Notice that *the units of the quantity you want to find should be the units in the numerator of the conversion factor ratio.*

COMMUNICATION *example* 9

What volume of a 0.25 mol/L salt solution in a laboratory contains 0.10 mol of sodium chloride?

Solution

$$V_{NaCl} = 0.10\ \cancel{mol} \times \frac{1\ L}{0.25\ \cancel{mol}}$$

$$= 0.40\ L$$

The volume of salt solution is 0.40 L.

Figure 7
Hydrogen peroxide solutions are stored in dark bottles to keep out light, which promotes the decomposition of hydrogen peroxide to water and oxygen.

Practice

7. Rubbing alcohol, $C_3H_7OH(l)$, is sold as a 70.0% V/V solution for external use only. What volume of pure $C_3H_7OH(l)$ is present in a 500 mL bottle?

8. Suppose your company makes hydrogen peroxide solution with a generic label for drugstores in your area (**Figure 7**). Calculate the mass of pure hydrogen peroxide needed to make 1000 bottles, each containing 250 mL of 3.0% W/V $H_2O_2(aq)$.

9. Seawater contains approximately 0.055 mol/L of magnesium chloride. Determine the chemical amount of magnesium chloride present in 75 L of seawater.

10. A bottle of 5.0 mol/L hydrochloric acid is opened in the laboratory, and 50 mL of it is poured into a beaker. What chemical amount of acid is in the beaker?

11. A household ammonia solution (e.g., a window-cleaning solution) has an amount concentration of 1.24 mol/L. What volume of this solution would contain 0.500 mol of $NH_3(aq)$?

12. A student needs 0.14 mol of $Na_2SO_4(aq)$ to do a quantitative analysis. The amount concentration of the student's solution is 2.6 mol/L $Na_2SO_4(aq)$. What volume of solution does the student need to measure?

Mass, Volume, and Concentration Calculations

Even though the mole is a very important unit, measurements in a chemistry laboratory are usually of mass (in grams) and of volume (in millilitres). A common chemistry calculation involves the mass of a substance, the volume of a solution, and the amount concentration of that solution. This type of calculation requires the use of two conversion factors—one for molar mass and one for amount concentration. Calculations using molar mass are just like the ones you did in previous units.

SAMPLE problem 5.1

A chemical analysis requires 2.00 L of 0.150 mol/L $AgNO_3(aq)$. What mass of solid silver nitrate is required to prepare this solution?

First determine the chemical amount of silver nitrate needed.

$$n_{AgNO_3} = 2.00\ \cancel{L} \times \frac{0.150\ \text{mol}}{1\ \cancel{L}}$$

$$= 0.300\ \text{mol}$$

Then convert this amount into a mass of silver nitrate by using its molar mass, *M*. The molar mass of silver nitrate is 169.88 g/mol.

$$m_{AgNO_3} = 0.300\ \cancel{\text{mol}} \times \frac{169.88\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$= 51.0\ \text{g}$$

If you clearly understand these two steps, you could combine them into one calculation.

$$m_{AgNO_3} = 2.00\ \cancel{L} \times \frac{0.150\ \text{mol}}{1\ \cancel{L}} \times \frac{169.88\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$= 51.0\ \text{g}$$

Learning Tip

If you prefer to use mathematical formulas, for this sample problem you may use

$n = Vc$

$m = nM$

In order to successfully combine the steps into one operation, as shown above, you need to pay particular attention to the units in the calculation. Cancelling the units will help you to check your procedure.

COMMUNICATION example 10

To study part of the water treatment process in a laboratory, a student requires 1.50 L of 0.12 mol/L aluminium sulfate solution. What mass of aluminium sulfate must she measure for this solution?

Solution

$$n_{Al_2(SO_4)_3} = 1.50\ \cancel{L} \times \frac{0.12\ \text{mol}}{1\ \cancel{L}}$$

$$= 0.180\ \text{mol}$$

$$m_{Al_2(SO_4)_3} = 0.180\ \cancel{\text{mol}} \times \frac{342.14\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$= 61.6\ \text{g}$$

or

$$m_{Al_2(SO_4)_3} = 1.50\ \cancel{L} \times \frac{0.12\ \cancel{\text{mol}}}{1\ \cancel{L}} \times \frac{342.14\ \text{g}}{1\ \cancel{\text{mol}}}$$

$$= 61.6\ \text{g}$$

The mass of aluminium sulfate required is 61.6 g.

Another similar calculation involves the use of a known mass and volume to calculate the amount concentration of a solution. This calculation is similar to the examples given above, using the same conversion factors.

COMMUNICATION example 11

Sodium carbonate is a water softener that is an important part of the detergent used in a washing machine. A student dissolves 5.00 g of solid sodium carbonate to make 250 mL of a solution to study the properties of this component of detergent. What is the amount concentration of the solution?

Solution

$$n_{Na_2CO_3} = 5.00\ \cancel{g} \times \frac{1\ \text{mol}}{105.99\ \cancel{g}} = 0.0472\ \text{mol}$$

$$c_{Na_2CO_3} = \frac{0.0472\ \text{mol}}{0.250\ \text{L}} = 0.189\ \text{mol/L}$$

or

$$c_{Na_2CO_3} = 5.00\ \cancel{g} \times \frac{1\ \text{mol}}{105.99\ \cancel{g}} \times \frac{1}{0.250\ \text{L}} = 0.189\ \text{mol/L}$$

The amount concentration of sodium carbonate is 0.189 mol/L.

Practice

13. A chemical technician needs 3.00 L of 0.125 mol/L sodium hydroxide solution. What mass of solid sodium hydroxide must be measured?
14. Seawater is mostly a solution of sodium chloride in water. The concentration varies, but marine biologists took a sample with an amount concentration of 0.56 mol/L. Calculate the mass of sodium chloride in the biologists' 5.0 L sample.
15. Acid rain may have 355 ppm of dissolved carbon dioxide.
    (a) What mass of carbon dioxide is present in 1.00 L of acid rain?
    (b) Calculate the amount concentration of carbon dioxide in the acid rain sample.
16. A brine (sodium chloride) solution used in pickling contains 235 g of pure sodium chloride dissolved in 3.00 L of solution.
    (a) Determine the percent concentration (% W/V) of sodium chloride.
    (b) What is the amount concentration of sodium chloride?

Concentration of Ions

In solutions of ionic compounds and strong acids, the electrical conductivity suggests the presence of ions in the solution. When these solutes produce aqueous ions, expressing the concentration of individual ions in moles per litre (mol/L) is important. The amount concentrations of the ions in a solution depend on the relative numbers of ions making up the compound: for example, Cl^- ions in NaCl(aq) and $CaCl_2$(aq).

The dissociation or ionization equations for ionic compounds or strong acids allow you to determine the amount concentration of either the ions or the compounds in solution. The ion concentration is always equal to a whole number multiple of the compound concentration. For convenience, square brackets are commonly placed around formulas to indicate the amount concentration of the substance within the brackets. For example, $[NH_3(aq)]$ and $[H^+(aq)]$ indicate the amount concentrations of aqueous ammonia and hydrogen ions respectively.

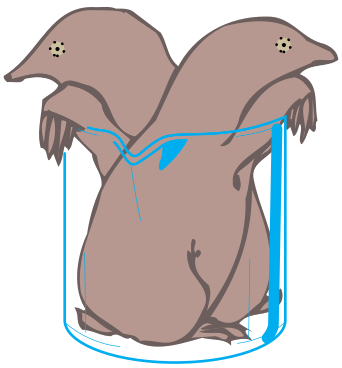


Figure 8
Two moles per litre

When sodium chloride dissolves in water, each mole of sodium chloride produces one mole of sodium ions and one mole of chloride ions.

$$\begin{array}{ccccc} NaCl(aq) & \rightarrow & Na^+(aq) & + & Cl^-(aq) \\ 1\ \text{mol} & & 1\ \text{mol} & & 1\ \text{mol} \end{array}$$

If $[NaCl(aq)] = 1$ mol/L, then $[Na^+(aq)] = 1$ mol/L and $[Cl^-(aq)] = 1$ mol/L because the mole ratio from the dissociation equation is 1:1:1.

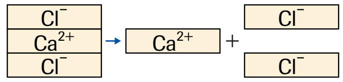
Calcium chloride dissociates in water to produce individual calcium and chloride ions. Each mole of calcium chloride produces one mole of calcium ions and two moles of chloride ions.

$$\begin{array}{lclcl} CaCl_2(aq) & \rightarrow & Ca^{2+}(aq) & + & 2\ Cl^-(aq) \\ 1\ \text{mol} & & 1\ \text{mol} & & 2\ \text{mol} \end{array}$$

If $[CaCl_2(aq)] = 1$ mol/L, then $[Ca^{2+}(aq)] = 1$ mol/L. The $[Cl^-(aq)] = 2$ mol/L because the dissociation equation shows that 2 mol of chloride ions is produced from 1 mol of calcium chloride—a 2:1 mole ratio. Notice that you can easily predict the individual ion concentrations from the concentration of the compound and the subscripts of the ions in the formula of the compound. Even so, it is good practice to write the dissociation or ionization equation prior to calculating concentrations. This practice will help you avoid errors and is good preparation for your later study of stoichiometry in Unit 4.

Learning Tip

Initially it might appear that there is a "lack of conservation" because the dissociation equation shows a greater amount on the product side compared to the reactant side. Here is a simple model for the dissolving of calcium chloride that may help you to understand that no rules are being broken.



1 sheet of paper | cut into | 3 smaller pieces

Although the number of pieces of paper has increased by cutting one sheet into smaller pieces, the total mass of paper has not changed. If you started with five sheets of paper ($CaCl_2$), how many pieces of paper, labelled chloride ion, would you obtain?

COMMUNICATION example 12

What is the amount concentration of aluminium ions and sulfate ions in a 0.40 mol/L solution of $Al_2(SO_4)_3(aq)$?

Solution

$$Al_2(SO_4)_3(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ SO_4^{\ 2-}(aq)$$

$$[Al^{3+}(aq)] = 0.40\ \text{mol/L}\ Al_2(SO_4)_3(aq) \times \frac{2\ \text{mol}\ Al^{3+}(aq)}{1\ \text{mol}\ Al_2(SO_4)_3(aq)} = 0.80\ \text{mol/L}$$

$$[SO_4^{\ 2-}(aq)] = 0.40\ \text{mol/L}\ Al_2(SO_4)_3(aq) \times \frac{3\ \text{mol}\ SO_4^{\ 2-}(aq)}{1\ \text{mol}\ Al_2(SO_4)_3(aq)} = 1.20\ \text{mol/L}$$

The amount concentration of aluminium ions is 0.80 mol/L and of sulfate ions is 1.20 mol/L.

Note that, in Communication Example 12, the chemical formula of the entity you wish to find appears in the numerator of the ratio. The formula of the known entity is in the denominator and cancels with the chemical formula of the known solution concentration. The chemical formulas are omitted from the ratios in the following examples.

COMMUNICATION example 13

Determine the amount concentration of barium and hydroxide ions in a solution made by dissolving 5.48 g of barium hydroxide to make a volume of 250 mL.

Solution

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

$$n_{Ba(OH)_2} = 5.48\ \cancel{g} \times \frac{1\ \text{mol}}{171.35\ \cancel{g}} = 0.0320\ \text{mol}$$

$$[Ba(OH)_2(aq)] = \frac{0.0320\ \text{mol}}{0.250\ \text{L}} = 0.128\ \text{mol/L}$$

$$[Ba^{2+}(aq)] = 0.128\ \text{mol/L} \times \frac{1}{1} = 0.128\ \text{mol/L}$$

$$[OH^-(aq)] = 0.128\ \text{mol/L} \times \frac{2}{1} = 0.256\ \text{mol/L}$$

The amount concentration of barium ions is 0.128 mol/L and of hydroxide ions is 0.256 mol/L.

Learning Tip

Notice that in this example, we are now "going backwards" from the ion concentration to the solute concentration. The first calculation step means that for every two moles of bromide ions, we need only one mole of magnesium bromide.

COMMUNICATION example 14

What mass of magnesium bromide must be dissolved to make 1.50 L of solution with a bromide ion concentration of 0.30 mol/L?

Solution

$$MgBr_2(aq) \rightarrow Mg^{2+}(aq) + 2\,Br^-(aq)$$

$$[MgBr_2(aq)] = 0.30 \text{ mol/L} \times \frac{1}{2} = 0.15 \text{ mol/L}$$

$$n_{MgBr_2} = 1.50\ \cancel{L} \times \frac{0.15 \text{ mol}}{1\ \cancel{L}} = 0.23 \text{ mol}$$

$$m_{MgBr_2} = 0.23\ \cancel{\text{mol}} \times \frac{184.11 \text{ g}}{1\ \cancel{\text{mol}}} = 41 \text{ g}$$

The mass of magnesium bromide required is 41 g.

Practice

17. Find the amount concentration of each ion in the following solutions:

(a) 0.41 mol/L $Na_2S(aq)$

(b) 1.2 mol/L $Sr(NO_3)_2(aq)$

(c) 0.13 mol/L $(NH_4)_3PO_4(aq)$

18. A 250 mL solution is prepared by dissolving 2.01 g of iron(III) chloride in water. What is the amount concentration of each ion in the solution?

19. In order to prepare for a chemical analysis, a lab technician requires 500 mL of each of the following solutions. Calculate the mass of solid required for each solution:

(a) $[Cl^-(aq)] = 0.400$ mol/L from $CaCl_2(s)$

(b) $[CO_3^{2-}(aq)] = 0.35$ mol/L from $Na_2CO_3(s)$

Case Study

Household Chemical Solutions

An amazing number of solutions is available for household use at your local drugstore, hardware store, and supermarket in the form of food products, household cleaners, and health and personal care products (**Figure 9**). They come with a bewildering array of names, instructions, warnings, and concentration labels. For consumer convenience and safety, it is important that household chemical solutions be labelled accurately and honestly. Unfortunately, manufacturers and distributors of household cleaning products are not required by law to list ingredients on their labels. In some cases, you can find the Material Data Safety Sheet (MSDS) for the product on the manufacturer's Web site or phone to request it. Societal concerns about safety and disposal of chemicals have resulted in efforts by chemical manufacturers to promote safe and environmentally sound practices through their Responsible Care® program.

Being able to read information on household product labels is important for personal safety and proper disposal. Hazard symbols and safety warnings on labels are pointless if they go unnoticed, or are not understood. Every year people are injured because they are unaware that bleach (sodium hypochlorite solution) should never be mixed with acids such as vinegar. Although both solutions are effective cleaners for certain stains, when they are combined, they react to produce

Figure 9
Corrosive substances, such as acids and bases, are found in many household products, including cleaning solutions.

a highly toxic gas, chlorine. Trying to use both at once—for example, in cleaning a toilet—has been known to transform a bathroom into a deathtrap.

Concentration is another factor to consider when buying solutions. The labels on consumer products usually give concentration as a percentage, which is easier for the general public to understand than moles per litre. Some consumer products, such as insect repellant, are sold in different concentrations, so it is important to know which one is most suitable for your needs. Another household chemical, isopropyl alcohol, is sold in pure form as a disinfectant, and in 70% concentration as rubbing alcohol. Knowing the key ingredients and their concentrations is useful for safe and proper use and disposal.

Case Study Questions

1. How and why are concentrations for household chemical solutions expressed differently than for laboratory work?
2. According to Health Canada, it is the manufacturer's responsibility to assess and report the hazards associated with a chemical product. What personal or social values would you expect a manufacturer to demonstrate? To what extent do you think these values are demonstrated?
3. Survey your home and list any household solutions you have that are commonly brought to the Household Hazardous Waste Round-Up. When is your local Round-Up? If your location is not listed, phone the local government office to ask why you are not listed and what you should do to dispose of household hazardous materials.

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4. The Canadian Chemical Producers' Association (CCPA) represents over 65 chemical manufacturing industries with over 200 plants across Canada. These industries collectively produce over 90% of all chemicals in Canada. CCPA is the driving force behind the Responsible Care® initiative—a global effort aimed at addressing public concerns about the manufacture, distribution, use, and disposal of chemicals. State their Ethic and list their six Codes of Practice. Do you feel that this program is a suitable replacement for government regulation?

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5. If household products are brought into a workplace, they become restricted products and a MSDS is required. Choose one household product and record the name of one key ingredient. Find the MSDS and identify three pieces of information supplied on this sheet that you think should be on the label.

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WWW **WEB Activity**

Web Quest—Hot Tub Safety

Hot tubs are very popular in private homes, public recreation centres, and commercial hotels (**Figure 10**). How are you protected from infectious diseases transmitted via hot tubs?

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Figure 10
How safe are hot tubs?

SUMMARY Concentration of a Solution

Type	Definition	Units	
percentage by volume	$c = \frac{V_{solute}}{V_{solution}} \times 100\%$	% V/V	(or mL/100 mL)
mass by volume	$c = \frac{m_{solute}}{V_{solution}} \times 100\%$	% W/V	(or g/100 mL)
by mass	$c = \frac{m_{solute}}{m_{solution}} \times 100\%$	% W/W	(or g/100 g)
parts per million	$c = \frac{m_{solute}}{m_{solution}}$	ppm	(or mg/kg)
amount	$c = \frac{n_{solute}}{V_{solution}}$	mol/L	

Section 5.3 Questions

1. What concentration ratio is often found on the labels of consumer products? Why is this unit used?
2. What concentration unit is most useful in the study of chemistry? Briefly describe why this unit is useful.
3. Bags of a D5W intravenous sugar solution used in hospitals contain 50 g of dextrose (glucose) in a 1.00 L bag.
   (a) Calculate the percentage weight by volume concentration of dextrose.
   (b) Suggest a reason why the bags are labelled D5W.
4. An Olympic-bound athlete tested positive for the anabolic steroid nandrolone. The athlete's urine test results showed 0.20 mg of nandrolone in a 10.0 mL urine sample. Convert the test result concentration to parts per million.
5. A 15 mL dose of a cough syrup contains 4.8 mmol of ammonium carbonate, $(NH_4)_2CO_3(aq)$.
   (a) Determine the amount concentration of ammonium carbonate.
   (b) What is the amount concentration of each ion in this solution?
6. The maximum concentration of salt in water at 0 °C is 31.6 g/100 mL. What mass of salt can be dissolved in 250 mL of solution?
7. Bald eagle chicks raised in northern Alberta were found to contain PCBs (polychlorinated biphenyls) at an average concentration of 18.9 ppm. If a chick had a mass of 0.60 kg, predict the mass of PCBs it would contain.
8. An experiment is planned to study the chemistry of a home water-softening process. The brine (sodium chloride solution) used in this process has a concentration of 25 g in every 100 mL of solution. Calculate the amount concentration of this solution.
9. To prepare for an experiment using flame tests, a school lab technician requires 100 mL of 0.10 mol/L solutions of each of the following substances. Calculate the required mass of each solid.
   (a) $NaCl(s)$
   (b) $KCl(s)$
   (c) $CaCl_2(s)$
10. What volume of 0.055 mol/L glucose solution found in a plant contains 2.0 g of glucose, $C_6H_{12}O_6(aq)$?
11. In an experiment, 28.6 g of aluminium chloride is dissolved in 1.50 L of solution.
    (a) Calculate the amount concentration of aluminium chloride.
    (b) Determine the amount concentration of each ion in the final solution.
12. As part of a chemical analysis, a technician requires a 0.25 mol/L bromide ion solution. What mass of magnesium bromide is required to prepare 100 mL of the required solution?
13. How is your report card mark in a subject similar to a concentration? What other ratios have you used that are similar to concentration ratios?
14. List several examples of how solutions and solution concentration are applied in products and processes we use in daily life.
15. Identify the implications of selling medicines in much more concentrated solutions. Present points both in favour and against.
16. Science and technology have both intended and unintended consequences. Illustrate this statement using DDT as your example. Include the role of biomagnification of DDT in the environment.

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17. Very low concentrations of toxic substances sometimes require the use of the parts per billion (ppb) concentration.
    (a) Express parts per billion as a ratio, including the appropriate power of ten.
    (b) How much smaller is 1 ppb than 1 ppm?
    (c) Convert your answer in (a) to a concentration ratio using the appropriate SI prefixes to obtain a mass of solute per kilogram of solution.
    (d) Copper is an essential trace element for animal life. An average adult human requires the equivalent of a litre of water containing 30 ppb of copper a day. What is the mass of copper per kilogram of solution?

Extension

18. Toxicity of substances for animals is usually expressed by a quantity designated as "LD50." Use the Internet to research the use of this quantity. What does LD50 mean? What is the concentration in ppm for a substance considered "extremely toxic" and one considered "slightly toxic"?

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19. Many chemicals that are potentially toxic or harmful to the environment and humans have maximum permissible concentration levels set by government legislation. Nevertheless, some people question the levels that are set and some suggest that the only safe level is zero.
    (a) To what extent should we trust our government agencies to set appropriate levels?
    (b) Outline some risks and benefits, from several perspectives, associated with the use of controversial chemicals such as pesticides.
    (c) What is chemical hormesis? Why might this effect have major implications for government regulatory agencies?
    (d) What does LC50 mean? List some advantages and disadvantages of this method of measuring toxicity.

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Preparation of Solutions 5.4

When you prepare a jug of iced tea using a package of crystals and water, you are preparing a solution from a solid solute (actually, from several solid solutes). However, when you prepare the tea from a container of frozen concentrate, you are preparing a solution by dilution. Scientists use both of these methods to prepare solutions. In this course you will be preparing only aqueous solutions. The knowledge and skills for preparing solutions are necessary to complete some of the more complex laboratory investigations that come later in this course.

Figure 1
An electronic balance is simpler to operate and more efficient than the older mechanical balance. Electronic balances also provide the convenience of taring (Appendix C.3).

Preparation of Standard Solutions from a Solid

Solutions of accurate concentration, called **standard solutions**, are routinely prepared for use in both scientific research laboratories and industrial processes. They are used in chemical analysis as well as for the control of chemical reactions. To prepare a standard solution, good-quality equipment is required to measure the mass of solute and volume of solution. Electronic balances are used for precise and efficient measurement of mass (**Figure 1**). For measuring a precise volume of the final solution, a container called a volumetric flask is used (**Figure 2**).

Figure 2
Volumetric glassware comes in a variety of shapes and sizes. The Erlenmeyer flask on the far left has only approximate volume markings, as does the beaker. The graduated cylinders have much better precision, but for high precision, a volumetric flask (on the right) is used. The volumetric flask shown here, when filled to the line, contains 100.0 mL $\pm$ 0.16 mL at 20 °C. This means that a volume measured in this flask is uncertain by less than 0.2 mL at the specified temperature.

INVESTIGATION 5.2 *Introduction*

A Standard Solution from a Solid

In this investigation, you will practise the skills required to prepare a standard solution from a pure solid (Appendix C.4). You will need these skills in many investigations in this course.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ○ Analysis
- ○ Evaluation

Purpose

The purpose of this investigation is to acquire the skills required to prepare a standard solution starting with a pure solid.

To perform this investigation, turn to page 227.

Figure 3
Hard-water deposits such as calcium carbonate can seriously affect water flow in a pipe.

Figure 4
Solutions of sodium hydroxide in very high concentration are sold as cleaners for clogged drains. The same solution can be made less expensively by dissolving solid lye (a commercial name for sodium hydroxide) in water. The pure chemical is very caustic and the label on the lye container recommends rubber gloves and eye protection.

Practice

1. To test the hardness of water (**Figure 3**), an industrial chemist performs an analysis using 100.0 mL of a 0.250 mol/L standard solution of ammonium oxalate. What mass of ammonium oxalate, $(NH_4)_2C_2O_4(s)$, is needed to make the standard solution?
2. Calculate the mass of solid lye (sodium hydroxide) (**Figure 4**) needed to make 500 mL of a 10.0 mol/L strong cleaning solution.
3. List several examples of solutions that you prepared from solids in the last week.
4. You have been asked to prepare 2.00 L of a 0.100 mol/L aqueous solution of cobalt(II) chloride for an experiment starting with $CoCl_2{\bullet}2H_2O(s)$.
   (a) Show your work for the pre-lab calculation.
   (b) Write a complete specific procedure for preparing this solution, as in Investigation 5.2. Be sure to include all necessary precautions.
5. (a) A technician prepares 500.0 mL of a 0.0750 mol/L solution of potassium permanganate as part of a quality-control analysis in the manufacture of hydrogen peroxide. Calculate the mass of potassium permanganate required to prepare the solution.
   (b) Write a laboratory procedure for preparing the potassium permanganate solution. Follow the conventions of communication for a procedure in a laboratory report.

Preparation of Standard Solutions by Dilution

A second method of preparing solutions is by dilution of an existing solution. You use this process when you add water to concentrated fruit juice, fabric softener, or a cleaning product. Many consumer and commercial products are purchased in concentrated form and then diluted before use. You can save money and help save the environment by diluting concentrated products. Doing so saves on shipping charges and reduces the size of the container, making the product less expensive and more environmentally friendly. Citizens who are comfortable with dilution techniques can live more lightly on Earth.

Because dilution is a simple, quick procedure, it is common scientific practice to begin with a stock solution and to add solvent (usually water) to decrease the concentration to the desired level. A **stock solution** is an initial, usually concentrated, solution from which samples are taken for a dilution. For the most accurate results, the stock solution should be a standard solution.

Even though there are no firm rules, we often describe solutions with an amount concentration of less than 0.1 mol/L as dilute, whereas solutions with a concentration of greater than 1 mol/L may be referred to as concentrated.

Calculating the new concentration after a dilution is straightforward because the quantity of solute is not changed by adding more solvent. Therefore, the mass (or chemical amount) of solute before dilution is the same as the mass (or chemical amount) of solute after dilution.

$$m_i = m_f$$

or

$$n_i = n_f$$

m_i = initial mass of solute
m_f = final mass of solute
n_i = initial chemical amount of solute
n_f = final chemical amount of solute

Using the definitions of solution concentration ($m = Vc$ or $n = Vc$), we can express the constant quantity of solute in terms of the volume and concentration of solution.

$$V_i c_i = V_f c_f$$

This equation means that the concentration is inversely related to the solution's volume. For example, if water is added to 6% hydrogen peroxide disinfectant until the total volume is doubled, the concentration becomes one-half the original value, or 3%.

Any one of the variables in this dilution equation may be calculated for the dilution of a solution, provided the other three values are known. (Note that the dilution calculation for percentage weight by weight (%W/W) will be slightly different because the mass of solution is used: $m_{solute} = m_{solution}c$.)

COMMUNICATION example 1

Water is added to 0.200 L of 2.40 mol/L $NH_3(aq)$ cleaning solution, until the final volume is 1.000 L. Find the amount concentration of the final, diluted solution.

Solution

$$V_i c_i = V_f c_f$$

$$c_f = \frac{V_i c_i}{V_f}$$

$$= \frac{0.200\ \cancel{L} \times 2.40\ \text{mol}/\cancel{L}}{1.000\ \text{L}}$$

$$= 0.480\ \frac{\text{mol}}{\text{L}}$$

The amount concentration of the final, diluted ammonia solution is 0.480 mol/L.

Learning Tip

Alternatively, this problem can be solved another way:

$$n_{NH_3} = 0.200\ \cancel{L} \times \frac{2.40\ \text{mol}}{1.00\ \cancel{L}}$$

$$= 0.480\ \text{mol}$$

$$c_{NH_3} = \frac{0.480\ \text{mol}}{1.000\ \text{L}}$$

$$= 0.480\ \text{mol/L}$$

When diluting all concentrated reagents, especially acids, always add the concentrated reagent, with stirring, to less than the final required quantity of water (**Figure 5**), and then add the rest of the water.

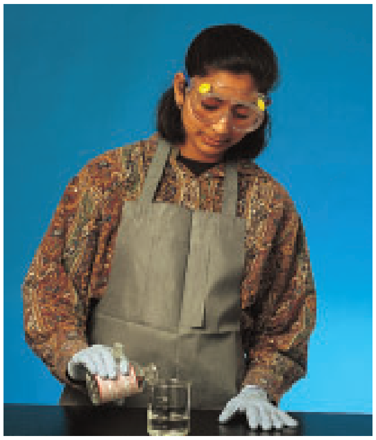

Figure 5
Handling concentrated reagents, especially acids, requires great care. Wear eye protection, a lab apron, and gloves, as shown in the photograph.

COMMUNICATION example 2

A student is instructed to dilute some concentrated HCl(aq) (36%) to make 4.00 L of 10% solution. What volume of hydrochloric acid solution should the student initially measure?

Solution

$$V_i c_i = V_f c_f$$

$$V_i = \frac{V_f c_f}{c_i}$$

$$= \frac{4.00\ \text{L} \times 10\%}{36\%}$$

$$= 1.1\ \text{L}$$

The volume of concentrated hydrochloric acid required is 1.1 L.

You can predict answers to dilution calculations if you understand the dilution process: As the volume increases, the concentration decreases. Use this principle as a useful check on your work. In Communication Example 1, the final concentration must be less than the initial concentration because the solution is being diluted. In Communication Example 2, the initial volume of acid required must be less than the final volume after the dilution.

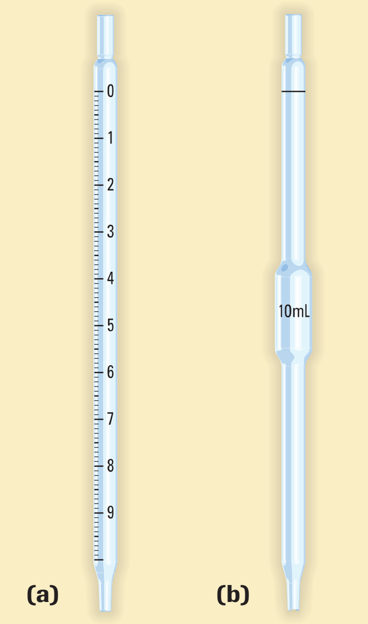

Figure 6
A graduated pipette **(a)** measures a range of volumes, whereas a volumetric pipette **(b)** is calibrated to deliver (TD) a fixed volume.

The dilution technique is especially useful when you need to decrease the concentration of a solution. For example, when doing scientific or technological research, you may want to slow down a reaction that proceeds too rapidly or too violently with a concentrated solution. You could slow down the reaction by lowering the concentration of the solution. In the medical and pharmaceutical industries, prescriptions require not only minute quantities, but also very accurate measurement. If the solutions are diluted before being sold, it is much easier for a patient to take the correct dose. For example, it's easier to accurately measure out 10 mL (two teaspoons) of a cough medicine than it is to measure one-fifth of a teaspoon, which the patient would have to do if the medicine were ten times more concentrated.

The preparation of standard solutions by dilution requires a means of transferring precise and accurate volumes of solution. You know how to use graduated cylinders to measure volumes of solution, but graduated cylinders are not precise enough when working with small volumes. To deliver a precise and accurate, small volume of solution, a laboratory device called a pipette is used. A 10 mL graduated pipette has graduation marks every tenth of a millilitre (**Figure 6**). This type of pipette can transfer any volume from 0.1 mL to 10.0 mL. A volumetric pipette transfers only one specific volume, but has a very high precision and accuracy. For example, a 10 mL volumetric (or delivery) pipette is designed to transfer 10.00 mL of solution with a precision of ± 0.02 mL. The volumetric pipette is often inscribed with TD to indicate that it is calibrated *to deliver* a particular volume with a specified precision. Both kinds of pipettes come in a range of sizes and are used with a pipette bulb. (See Appendix C.4.)

INVESTIGATION 5.3 ***Introduction***

A Standard Solution by Dilution

In this investigation, you will practise the skills required to prepare a standard solution from a more concentrated or stock solution. This laboratory procedure is very common for preparing solutions (Appendix C.4).

Report Checklist

○ Purpose	○ Design	○ Analysis
○ Problem	○ Materials	○ Evaluation
○ Hypothesis	○ Procedure	
○ Prediction	○ Evidence	

Purpose

The purpose of this investigation is to acquire the skills required to prepare a standard solution by diluting a stock solution.

To perform this investigation, turn to page 228.

Practice

6. Radiator antifreeze (ethylene glycol) is diluted with an appropriate quantity of water to prevent freezing of the mixture in the radiator. A 4.00 L container of 94% V/V antifreeze is diluted to 9.00 L. Calculate the concentration of the final solution.

7. Many solutions are prepared in the laboratory from purchased concentrated solutions. Calculate the volume of concentrated 17.8 mol/L stock solution of sulfuric acid a laboratory technician would need to make 2.00 L of 0.200 mol/L solution by dilution of the original concentrated solution.

8. In a study of reaction rates, you need to dilute the copper(II) sulfate solution prepared in Investigation 5.3. You take 5.00 mL of 0.005000 mol/L $CuSO_4$(aq) and dilute it to a final volume of 100.0 mL.
(a) Determine the final concentration of the dilute solution.
(b) What mass of $CuSO_4$(s) is present in 10.0 mL of the final dilute solution?
(c) Can this final dilute solution be prepared directly using the pure solid? Defend your answer.

9. A student tries a reaction and finds that the volume of solution that reacts is too small to be measured with any available equipment. The student takes a 10.00 mL volume of the solution with a pipette, transfers it into a clean 250 mL volumetric flask containing some pure water, adds enough pure water to increase the volume to 250.0 mL, and mixes the solution thoroughly.
   (a) Compare the concentration of the dilute solution to that of the original solution.
   (b) Compare the volume that will react now to the volume that reacted initially.
   (c) Predict the speed or rate of the reaction using the diluted solution compared with the rate using the original solution. Explain your answer.

INVESTIGATION 5.4 *Introduction*

The Iodine Clock Reaction

Technological problem solving often involves a systematic trial-and-error approach that is guided by knowledge and experience. Usually one variable at a time is manipulated, while all other variables are controlled. Variables that may be manipulated include concentration, volume, and temperature. In this investigation, you will compete to see which team is the first to solve the Problem using a reliable process. Create a design to guide your work. Using this design, try several procedures to solve the Problem. The final Analysis will be the materials and procedure that best answer the Problem.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ● Design
- ● Materials
- ● Procedure
- ● Evidence
- ● Analysis
- ○ Evaluation

Purpose

The purpose of this investigation is to find a method for getting a reaction to occur in a specified time period.

Problem

What technological process can be employed to have solution A react with solution B in a reliable time of 20 ± 1 s?

To perform this investigation, turn to page 228.

Section 5.4 Questions

1. List several reasons why scientists make solutions in the course of their work.
2. (a) Briefly describe two different ways of making a solution.
   (b) When should you use each method?
3. In an analysis for sulfate ions in a water treatment plant, a technician needs 100 mL of 0.125 mol/L barium nitrate solution. What mass of pure barium nitrate is required?
4. A 1.00 L bottle of purchased acetic acid is labelled with a concentration of 17.4 mol/L. A technician dilutes this entire bottle of concentrated acid to prepare a 0.400 mol/L solution. Calculate the volume of diluted solution prepared.
5. A 10.00 mL sample of a test solution is diluted in a laboratory to a final volume of 250.0 mL. The concentration of the diluted solution is 0.274 g/L. Determine the concentration of the original test solution.
6. A chemical analysis of silver uses 100 mL of a 0.155 mol/L solution of potassium thiocyanate, $KSCN_{(aq)}$. Write a complete, specific procedure for preparing the solution from the solid. Include all necessary calculations and precautions.
7. A laboratory technician needs 1.00 L of 0.125 mol/L sulfuric acid solution for a quantitative analysis experiment. A commercial 5.00 mol/L sulfuric acid solution is available from a chemical supply company. Write a complete, specific procedure for preparing the solution. Include all necessary calculations and safety precautions.
8. As part of a study of rates of reaction, you are to prepare two aqueous solutions of nickel(II) chloride.
   (a) Calculate the mass of solid nickel(II) chloride that you will need to prepare 100.0 mL of a 0.100 mol/L nickel(II) chloride solution.
   (b) Calculate how to dilute this solution to make 100.0 mL of a 0.0100 mol/L nickel(II) chloride solution.
   (c) Write a list of Materials, and a Procedure for the preparation of the two solutions. Be sure to include all necessary safety precautions and disposal steps.
9. It has been suggested that it is more environmentally friendly to transport chemicals in a highly concentrated state. List arguments for and against this position, including possible intended and unintended consequences.
10. For many years the adage “The solution to pollution is dilution” described the views of some individuals, industries, and governments. They did not realize at that time that chemicals, diluted by water or air, could be concentrated in another system later. What is biomagnification? Describe briefly using a specific chemical as an example. What implications does this effect have for the introduction of new technologies?

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5.5 Solubility

It is easier to handle a great many chemicals when they are in solution, particularly those that are toxic, corrosive, or gaseous. Both in homes and at worksites, transporting, loading, and storing chemicals is more convenient and efficient when the chemicals are in solution rather than in solid or gaseous states. Also, performing a reaction in solution can change the rate (speed), the extent (completeness), and the type (kind of product) of the chemical reaction.

Solutions make it easy to

- handle chemicals—a solid or gas is dissolved in water for ease of use or transportation
- complete reactions—some chemicals do not react until in a solution where there is increased contact between the reacting entities
- control reactions—the rate, extent, and type of reactions are much more easily controlled when one or more reactants are in solution

Figure 1
The low solubility of the soap deposit is overcome by a chemical reaction—a good example of how science and technology provide products useful for daily life.

These three points all apply to the liquid cleaning solution in **Figure 1**. First, the cleaning solution is easy to handle, and the fact that it is sold in a spray bottle adds to its convenience. Spraying a solution is an effective way of handling a chemical that is dissolved in water. Second, the solution allows a reaction to occur between the cleaning chemicals and the dirty deposit, whereas a pure gas or solid would not react well with a solid. Third, the manufacturer can control the rate of the reaction (and thus the safety) by choosing the ideal concentration of the cleaning solution. Having the chemical in solution rather than in its pure state increases our ability to handle and control its use.

Because solutions are very useful at home, in industry, and in scientific research, it is important to consider which substances dissolve easily in solvents such as water and how much of a substance you can dissolve.

mini *Investigation* — Measuring the Dissolving Process

Are there different kinds of salt? How much salt can you dissolve in a given volume of water? What happens to the volume of a solution when a solute is added to it? This quick mini investigation will help you to think about the answers to these questions.

Materials: distilled or deionized water, table salt, coarse pickling salt (pure $NaCl(s)$), a measuring teaspoon (5 mL), two 125 mL Erlenmeyer flasks with stoppers, one 50 mL or 100 mL graduated cylinder

- Place a level teaspoonful of table salt into 25 mL of pure water at room temperature in a 125 mL Erlenmeyer flask. Swirl the flask's contents thoroughly for a minute or two. Record your observations.
- Repeat with pickling salt, again recording your observations.

(a) What does the result, with common table salt as a solute, show about the nature of the substance being used? Compare it with the solution in the second flask.

(b) List the ingredients in common table salt, according to the package label, and explain your observations of the contents of the first flask.

- Add another teaspoon of pickling salt to the second flask, and swirl until the solid is again completely dissolved. Keeping track of how much pickling salt you add, continue to dissolve level teaspoons of salt until no amount of swirling will make all of the solid crystals disappear.

(c) How many level teaspoons of pickling salt (pure $NaCl(s)$) could you get to dissolve in 25 mL of $H_2O(l)$ in the second flask?

(d) What is the final volume of your $NaCl(aq)$ solution in the second flask?

(e) If you dissolve 20.0 mL of $NaCl(s)$ in 100.0 mL of liquid water, what do you suppose the volume of the solution would be? Describe a way to test your supposition. The answer is very interesting.

Solubility of Solids

When you add a small amount of pickling salt (pure sodium chloride) to a jar of water and shake the jar, the salt dissolves and disappears completely. What happens if you continue adding salt and shaking? Eventually, some visible solid salt crystals will remain at the bottom of the jar, despite your efforts to make them dissolve. You have formed a **saturated solution**—a solution in which no more solute will dissolve at a specified temperature. We say it is at maximum solute concentration. If the container is sealed, and the temperature stays the same, no further changes will ever occur in the concentration of this solution. The quantity (mass) of solute that remains undissolved will also stay the same. **Solubility** is the concentration of a saturated solution. The units for solubility are simply units of concentration, such as % W/V or mol/L. You will learn in this section that solubility depends on the temperature, so it is a particular maximum concentration value. Every solubility value must be accompanied by a temperature value. When calculating and using solubility values, we have to make one assumption: The solute is not reacting with the solvent.

Every pure substance has its own unique solubility. Some references provide solubility data for substances in water using units of grams per hundred millilitres of water, not of solution. These units may be convenient for comparing solubilities, but they are not very convenient for calculations. For example, we can find from a reference source, such as the *CRC Handbook of Chemistry and Physics*, that the solubility of sodium sulfate in water at 0 °C is 4.76 g/100 mL H_2O. This means 4.76 g of solute can be dissolved in 100 mL of water—not that you will have 100 mL of solution after dissolving 4.76 g of solute. If more than 4.76 g of this solute is added to 100 mL of water in the container, the excess will not dissolve under the specified conditions (**Figure 2**). The quickest way to see whether you have a saturated solution is to look for the presence of undissolved solids in the solution. There are several experimental designs that can be used to determine the solubility of a solid. For example, the solvent from a measured volume of saturated solution might be removed by evaporation, leaving the crystallized solid solute behind, which can then be collected and measured.

Figure 2
The excess of solid solute in the mixture is visible evidence of a saturated solution.

INVESTIGATION 5.5 *Introduction*

The Solubility of Sodium Chloride in Water

A significant part of the work of science is to test existing theories, laws, and generalizations. You will create a graph from the solubility data provided and use this graph to predict the solubility of sodium chloride in water at a particular temperature. You will then compare the predicted value with a value that you determine experimentally—by crystallization of sodium chloride from a saturated solution.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	○ Procedure	
● Prediction	● Evidence	

Purpose

The purpose of this investigation is to test the known solubility data for a solid in water.

Problem

What is the solubility of sodium chloride, in grams per 100 mL of solution, at room temperature?

Design

A precisely measured volume of a saturated NaCl(aq) solution at room temperature is heated to evaporate the solvent and crystallize the solute. The mass of the dry solute is measured and the concentration of the saturated solution is calculated.

To perform this investigation, turn to page 229.

DID YOU KNOW ?

Generalizations and Indigenous Knowledge

In Western science, a generalization is a statement that summarizes a pattern of empirical properties or trends; for example, solids have a higher solubility in water at higher temperatures. Scientists rely on these generalizations to organize their knowledge and to make predictions. Even though you can find exceptions to all generalizations, they are still very useful.

Aboriginal peoples developed a traditional or indigenous knowledge of their environment. The foundation of this knowledge is empirical and many Aboriginal traditions correspond to generalizations about properties and trends in the natural world. These generalizations allowed Aboriginal peoples to make predictions such as which plants to use for which ailments and weather forecasting.

DID YOU KNOW ?

"The Bends"

When diving underwater using air tanks, a diver breathes air at the same pressure as the surroundings. The increased pressure underwater forces more air to dissolve in the diver's bloodstream. If a diver comes up too quickly, the solubility of air (mostly nitrogen) decreases as the pressure decreases, and nitrogen bubbles form in the blood vessels. These nitrogen bubbles are the cause of a diving danger known as "the bends" (so named because divers typically bend over in agony as they try to relieve the pain). Nitrogen bubbles are especially dangerous if they form in the brain or spinal cord. The bends may be avoided by ascending very slowly or corrected by using a decompression chamber.

Solubility in Water Generalizations

Scientists have carried out a very large number of experiments as they have investigated the effects of temperature on the solubility of various solutes. From the results of their experiments, they have developed several useful generalizations about the solubility of solids, liquids, and gases in water. In all cases, we assume that the solid, liquid, or gas does not react with the solvent, water. The following list outlines how the solubility of various solutes varies with temperature.

Solids

- Solids usually have higher solubility in water at higher temperatures. For example, sucrose has a solubility of about 180 g/100 mL at 0 °C and 487 g/100 mL at 100 °C.

Gases

- Gases always have higher solubility in water at lower temperatures. The solubility of gases decreases as the temperature increases. This inverse relationship is approximately linear.
- Gases always have higher solubility in water at higher pressures.

Liquids

- It is difficult to generalize about the effect of temperature on the solubility of liquids in water. However, for polar liquids in water, the solubility usually increases with temperature. A prediction of the solubility of liquids with temperature will not be as reliable as a prediction for solids and gases.
- Some liquids (mostly nonpolar liquids) do not dissolve in water to any appreciable extent, but form a separate layer. Liquids that behave in this way are said to be *immiscible* with water. For example, benzene, gasoline, and carbon disulfide (which is used in the process of turning wood pulp into rayon or cellophane) are all virtually insoluble in water.
- Some liquids (such as those containing small polar molecules with hydrogen bonding) dissolve completely in water in any proportion. Liquids that behave in this way are said to be *miscible* with water. For example, ethanol (in alcoholic beverages), acetic acid (in vinegar), and ethylene glycol (in antifreeze) all dissolve completely in water, regardless of the quantities mixed.

Elements

- Elements generally have low solubility in water. For example, carbon is used in many water filtration systems to remove organic compounds that cause odours. The carbon does not dissolve in the water passing through it.
- Although the halogens and oxygen dissolve in water to only a very tiny extent, they are so reactive that, even in tiny concentrations, they are often very important in solution reactions.

Solubility Table

A solubility table of ionic compounds (see the inside back cover of this textbook) is best understood by assuming that most substances dissolve in water to some extent. The solubilities of various ionic compounds range from very soluble, like table salt, to slightly soluble, like silver chloride. The classification of compounds into very soluble and slightly soluble categories allows you to predict the state of a compound formed in a reaction in aqueous solution. The cutoff point between very soluble and slightly soluble is arbitrary. A solubility of 0.1 mol/L is commonly used in chemistry as this cutoff point because most ionic compounds have solubilities significantly greater or less than this value, which is a typical concentration for laboratory work. Of course, some compounds seem to be exceptions to the rule. Calcium sulfate, for example, has a solubility close to our arbitrary cutoff point and enough of it will dissolve in water that the solution noticeably conducts electricity.

Practice

1. List three reasons why solutions are useful in a chemistry laboratory or industry.
2. Distinguish between solubility and a saturated solution.
3. Describe in general terms how you would make a saturated solution of a solid in water. How would you know whether the solution is saturated or whether the solute is just very slow in dissolving?
4. For any solute, what important condition must be stated in order to report the solubility?
5. State why you think clothes might be easier to clean in hot water.
6. Sketch a solubility versus temperature graph showing two lines labelled "solids" and "gases." Assume a straight-line relationship and show the generalization for the change in solubility of each type of substance with increasing temperature.
7. Give examples of two liquids that are immiscible and two that are miscible with water.
8. Why do carbonated beverages go "flat" when opened and left at room temperature and pressure?
9. Can more oxygen dissolve in a litre of water in a cold stream or a litre of water in a warm lake? Include your reasoning, according to the kinetic molecular theory.
10. (a) The solubility of oxygen in blood is much greater than its solubility in pure water. Suggest a reason for this observation.
    (b) If the concentration of oxygen in blood were the same as in pure water, how would your life be different?
    (c) Is there an advantage for animals that are cold blooded? Explain briefly.

LAB EXERCISE 5.C

Solubility and Temperature

Purpose
The purpose of this investigation is to test the generalization about the effect of temperature on the solubility of an ionic compound.

Problem
How does temperature affect the solubility of potassium nitrate?

Design
Solid potassium nitrate is added to four flasks of pure water until no more potassium nitrate will dissolve and there is excess solid in each beaker. Each mixture is sealed and stirred at a different temperature until no further changes occur. The same volume of each solution is removed and evaporated to crystallize the solid. The specific relationship of temperature to the solubility of potassium nitrate is determined by graphical analysis. The temperature is the manipulated variable and the solubility is the responding variable.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ● Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ● Analysis
- ● Evaluation (2, 3)

Evidence

Table 1 Solubility of Potassium Nitrate at Various Temperatures

Temperature (°C)	Volume of solution (mL)	Mass of empty beaker (g)	Mass of beaker plus solid (g)
0.0	10.0	92.74	93.99
12.5	10.0	91.75	93.95
23.0	10.0	98.43	101.71
41.5	10.0	93.37	100.15

EXPLORE an issue

Pesticides

Issue Checklist

- ○ Issue
- ○ Resolution
- ○ Design
- ● Evidence
- ● Analysis
- ● Evaluation

The tremendous advances made by science and technology have both intended and unintended consequences for humans and the environment. For example, the development of pesticides has greatly improved crop yields and human health by controlling insect populations. Pesticides are chemicals used to kill pests, including insects and plants. The downside of the use of many pesticides is that they are highly toxic and remain in the environment for years. Some of these pesticides and their residues are part of a group of chemicals known as POPs—persistent organic pollutants that have the potential to harm human health and damage the ecological system on which life depends.

On a more local level, many pesticides are used, particularly in urban areas, to maintain lush green lawns (**Figure 3**). Although these pesticides and their residues contribute slightly to the global problem, they are of more concern locally. The ability of municipalities to ban pesticide use was greatly enhanced by the June 2001 Supreme Court of Canada Ruling that upheld the 1991 pesticide ban in Hudson, Quebec. Since this decision, many other provinces, including Alberta, have passed legislation permitting municipalities to enact by-laws to regulate public health and safety.

Issue

The use of toxic chemicals for the cosmetic appearance of lawns may endanger human health.

Figure 3
Is this your lawn?

Resolution

All municipalities in Alberta should enact a complete ban on lawn pesticides.

Design

Within small groups, research the pros and cons of pesticide use on lawns. Gather information from a wide variety of perspectives.

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Explaining Saturated Solutions

Most substances dissolve in a solvent to a certain extent, and then dissolving appears to stop. If the solution is in a closed system, one in which no substance can enter or leave, then observable properties become constant, or are *in equilibrium* (**Figure 4**).

According to the kinetic molecular theory, particles are always moving and collisions are always occurring in a system, even if no changes are observed. The initial dissolving of sodium chloride in water is thought to be the result of collisions between water molecules and ions that make up the crystals. At equilibrium, water molecules still collide with the ions at the crystal surface. Chemists assume that dissolving of the solid sodium chloride is still occurring at equilibrium. Some of the dissolved sodium and chloride ions must, therefore, be colliding and crystallizing out of the solution to maintain a balance. If both dissolving and crystallizing take place at the same rate, no observable changes would occur in either the concentration of the solution or in the quantity of solid present. The balance that exists when two opposing processes occur at the same rate is known as **dynamic equilibrium** (**Figure 5**).

Figure 4
In a saturated solution of iodine, the concentration of the dissolved solute is constant. According to the theory of dynamic equilibrium, the rate of the dissolving process is equal to the rate of the crystallizing process.

$$I_2(s) \rightleftarrows I_2(aq)$$

Testing the Theory of Dynamic Equilibrium

You can try a simple experiment to illustrate dynamic equilibrium. Dissolve pickling (coarse) salt to make a saturated solution with excess solid in a small jar. Ensure that the lid is firmly in place, and then shake the jar and record the time it takes for the contents to settle so that the solution is clear. Repeat this process once a day for two weeks. Although the same quantity of undissolved salt is present each day, the settling becomes much faster over time because the solid particles in the jar become fewer in number, but larger in size. Chemists usually allow precipitates to digest for a while before filtering them, because larger particles filter more quickly. This evidence supports the idea that both dissolving and crystallizing are occurring simultaneously.

The theory of dynamic equilibrium can be tested by using a saturated solution of iodine in water. Radioactive iodine is used as a marker to follow the movements of some of the molecules in the mixture. To one sample of a saturated solution containing an excess of solid normal iodine, a few crystals of radioactive iodine are added. To a similar second sample, a few millilitres of a saturated solution of radioactive iodine are added (**Figure 6**). The radioactive iodine emits radiation that can be detected by a Geiger counter to show the location of the radioactive iodine. After a few hours, the solution and the solid in both samples clearly show increased radioactivity over the average background readings. Assuming the radioactive iodine molecules are chemically identical to normal iodine, the experimental evidence supports the idea of simultaneous dissolving and crystallizing of iodine molecules in a saturated system.

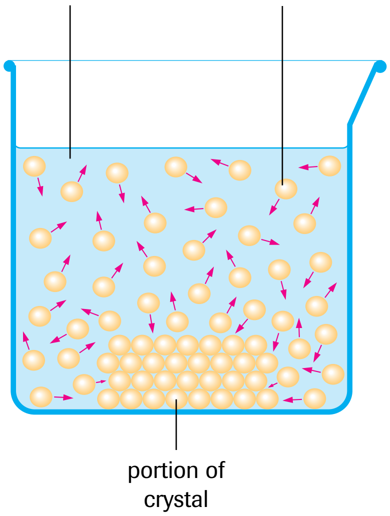
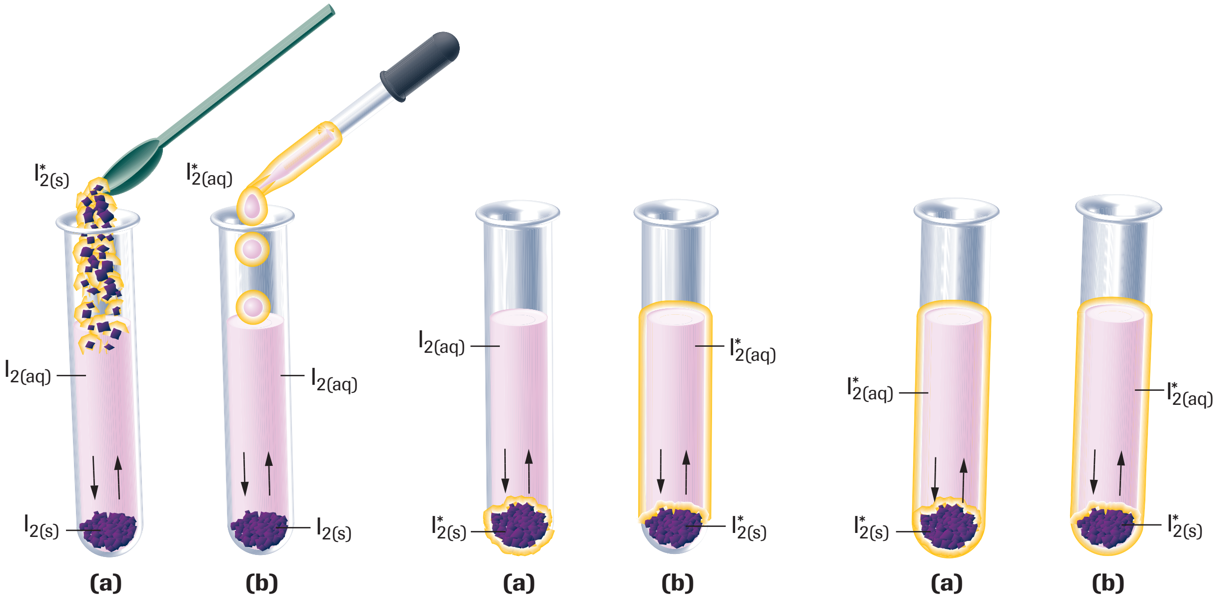


Figure 5
In a saturated solution such as this one, with excess solute present, dissolving and crystallizing occur at the same rate. This situation is known as dynamic equilibrium.



Radioactive iodine crystals are added to sample **(a)**, and a saturated solution of radioactive iodine is added to sample **(b)**.

At first, radioactivity is confined to the solid in **(a)** and the solution in **(b)**.

After several hours, the radioactivity can be detected in both solutions and in both samples of solid iodine.

Figure 6
Radioactive iodine (indicated with an asterisk), added to a saturated solution of normal iodine (I_2), is eventually distributed throughout the mixture. A yellow outline indicates radioactivity.

A solubility equilibrium must contain both dissolved and undissolved solute at the same time. This state can be established by starting with a solute and adding it to a solvent. Consider adding calcium sulfate to water in a large enough quantity that not all will dissolve. We say we have added excess solute. A dissociation equation can be written for a saturated solution established this way:

$$CaSO_4(s) \rightleftarrows Ca^{2+}(aq) + SO_4^{\ 2-}(aq)$$

Now consider a situation where two solutions, containing very high concentrations of calcium and sulfate ions respectively, are mixed. In this situation, the initial rate at which ions combine to form solid crystals is much greater than the rate at which those crystals dissolve, so we observe precipitation until the rates become equal and equilibrium is established.

$$Ca^{2+}(aq) + SO_4^{\ 2-}(aq) \rightleftarrows CaSO_4(s)$$

How the equilibrium is established is not a factor. Viewed this way, most ionic compound precipitation reactions are examples of a dynamic equilibrium just like the equilibrium in a saturated solution.

Section 5.5 Questions

1. Define solubility and state the main factors that affect the solubility of a substance in water.
2. Describe how the solubilities of solids and gases in water depend on temperature.

Use this information to answer questions 3 to 6.

In a chemical analysis experiment, a student notices that a precipitate has formed, and separates this precipitate by filtration. The collected liquid filtrate, which contains aqueous sodium bromide, is set aside in an open beaker. Several days later, some white solid is visible along the top edges of the liquid and at the bottom of the beaker.

3. What does the presence of the solid indicate about the nature of the solution?
4. What interpretation can be made about the concentration of the sodium bromide in the remaining solution? What is the term used for this concentration?
5. Write a brief theoretical explanation for this equilibrium mixture.
6. State two different ways to convert the mixture of the solid and solution into a homogeneous mixture.

7. Burping after drinking pop is common. What gas causes you to burp? Suggest a reason why burping occurs.
8. The purpose of the following investigation is to test the generalization about the effect of temperature on the solubility of an ionic compound known to be slightly soluble. Complete the Prediction and Design sections of the investigation report.

   Problem

   What is the relationship between temperature and the solubility of barium sulfate?

9. Different species of fish are adapted to live in different habitats. Some, such as carp, thrive in relatively warm, still water. Others, such as brook trout, need cold, fast-flowing streams, and will die if moved to the carp's habitat.
   (a) Describe and explain the oxygen conditions in the two habitats.
   (b) Hypothesize about the oxygen requirements of the two species of fish.
   (c) Thermal pollution is the large input of heated water into a lake or slow-moving stream from an industrial plant such as an electric generating station. Predict the effect of thermal pollution on trout in their lakes and streams.
10. Solubility also plays a role in cooking foods. Beans and broccoli should be cooked in water to retain their flavour but asparagus should be cooked in oil (or butter) and not water to best keep its flavour (**Figure 7**).
    (a) Based on this information, classify the solubility in water of the flavour molecules in these foods.
    (b) What interpretations can you make about the nature of the flavour molecules in beans and broccoli versus those in asparagus?

Figure 7
Fat retains the flavour of asparagus better than water does.

INVESTIGATION 5.1

Qualitative Chemical Analysis

Solutions have properties determined by the solute that is present. Diagnostic tests based on characteristic properties can be used to identify substances in a qualitative analysis.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ● Design
- ● Materials
- ● Procedure
- ● Evidence
- ● Analysis
- ○ Evaluation

Purpose
The purpose of this investigation is to use known diagnostic tests to distinguish among several pure substances.

Problem
Which of the white solids labelled 1, 2, 3, and 4 is calcium chloride, citric acid, glucose, and calcium hydroxide?

Calcium hydroxide is corrosive. Do not touch any of the solids. Wear eye protection, gloves, and an apron.

INVESTIGATION 5.2

A Standard Solution from a Solid

In this investigation, you will practise the skills required to prepare a standard solution from a pure solid (Appendix C.4). You will need these skills in many investigations in this course.

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ○ Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ○ Evidence
- ○ Analysis
- ○ Evaluation

Purpose
The purpose of this investigation is to acquire the skills required to prepare a standard solution starting with a pure solid.

Materials
lab apron
eye protection
$CuSO_4 \cdot 5H_2O(s)$, copper(II) sulfate–water (1/5) or copper(II) sulfate pentahydrate
150 mL beaker
centigram balance
laboratory scoop
stirring rod
wash bottle of pure water (distilled or deionized)
100 mL volumetric flask with stopper
small funnel
medicine dropper
meniscus finder

Copper(II) sulfate is harmful if swallowed.

Procedure
1. (Pre-lab) Calculate the mass of $CuSO_4 \cdot 5H_2O(s)$ needed to prepare 100.00 mL of a 0.05000 mol/L solution.
2. Measure the calculated mass of $CuSO_4 \cdot 5H_2O(s)$ in a clean, dry 150 mL beaker. (See Appendix C.3 for tips on using a laboratory balance. See Appendix C.4 and the Nelson Web site for tips on preparing a standard solution from a solid reagent.)

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3. Dissolve the solid in 40 mL to 50 mL of pure water. Use a stirring rod to help dissolve the solid. Be sure to rinse the stirring rod over your beaker of solution.
4. Transfer the solution into a 100 mL volumetric flask. Rinse the beaker two or three times with small quantities of pure water, transferring the rinsings into the volumetric flask.
5. Add pure water to the volumetric flask until the volume is 100.00 mL. Use the dropper and meniscus finder for the final few millilitres to set the bottom of the meniscus on the calibration line.
6. Stopper the flask and mix the contents thoroughly by repeatedly inverting the flask.

Note: Store your solution for the next investigation.

INVESTIGATION 5.3

A Standard Solution by Dilution

In this investigation, you will practise a very common laboratory procedure: preparing a standard solution from a more concentrated or stock solution.

Report Checklist

○ Purpose ○ Problem ○ Hypothesis ○ Prediction ○ Design ○ Materials ○ Procedure ○ Evidence ○ Analysis ○ Evaluation

Purpose

The purpose of this investigation is to acquire the skills required to prepare a standard solution by diluting a stock solution.

Materials

lab apron
eye protection
0.05000 mol/L $CuSO_4$(aq) stock solution
150 mL beaker
10 mL volumetric pipette
pipette bulb
wash bottle of pure water
100 mL volumetric flask with stopper
small funnel
medicine dropper
meniscus finder

Copper(II) sulfate is harmful if swallowed. Wear eye protection and a laboratory apron.

Use a pipette bulb. Do not pipette by mouth.

Procedure

1. (Pre-lab) Calculate the volume of a 0.05000 mol/L stock solution of $CuSO_4$(aq) required to prepare 100.0 mL of a 0.005000 mol/L solution.
2. Measure the required volume of the stock solution using a 10 mL pipette. (See Appendix C.3, Appendix C.4, and the Nelson Web site for tips on pipetting and preparing a standard solution by dilution.)
3. Transfer the required volume of the stock solution into the 100 mL volumetric flask (**Figure 1(a)**).
4. Add pure water until the final volume is reached (**Figure 1(b)**). Use the dropper and meniscus finder for the final few millilitres to set the bottom of the meniscus on the calibration line.
5. Stopper the flask and mix the solution thoroughly.

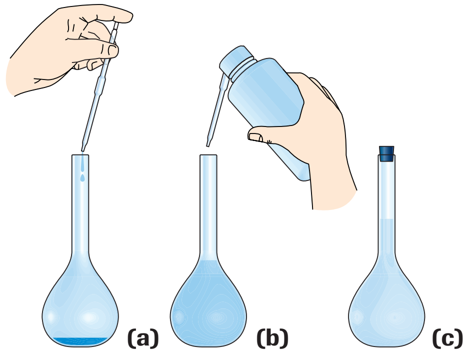

Figure 1
(a) The appropriate volume of $CuSO_4$(aq) is transferred to a volumetric flask.
(b) Water is added to the flask.
(c) In the final dilute solution, the initial amount of copper(II) sulfate is still present, but it is diluted.

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INVESTIGATION 5.4

The Iodine Clock Reaction

Technological problem solving often involves a systematic trial-and-error approach that is guided by knowledge and experience. Usually one variable at a time is manipulated, while all other variables are controlled. Variables that may be manipulated include concentration, volume, and temperature. In this investigation, you will compete to see which team is the first to solve the Problem using a reliable process. Create a design to guide your work. Using this design, try several procedures to solve the Problem. The final Analysis will be the materials and procedure that best answer the Problem.

Report Checklist

○ Purpose ○ Problem ○ Hypothesis ○ Prediction ● Design ● Materials ● Procedure ● Evidence ● Analysis ○ Evaluation

Purpose

The purpose of this investigation is to find a method for getting a reaction to occur in a specified time period.

Problem

What technological process can be employed to have solution A react with solution B in a reliable time of 20 $\pm$ 1 s?

INVESTIGATION 5.5

Report Checklist

- ○ Purpose
- ○ Problem
- ○ Hypothesis
- ● Prediction
- ○ Design
- ○ Materials
- ○ Procedure
- ● Evidence
- ● Analysis
- ● Evaluation (1, 2, 3)

The Solubility of Sodium Chloride in Water

A significant part of the work of science is to test existing theories, laws, and generalizations. You will create a graph from the solubility data (**Table 1**) and use this graph to predict the solubility of sodium chloride in water at a particular temperature. You will then compare the predicted value with a value that you determine experimentally—by crystallization of sodium chloride from a saturated solution.

Table 1 Solubility of Sodium Chloride in Water

Temperature (°C)	Solubility (g /100 mL solution)
0	31.6
40	32.4
70	33.0
100	33.6

Purpose

The purpose of this investigation is to test the known solubility data for a solid in water.

Problem

What is the solubility of sodium chloride, in grams per 100 mL of solution, at room temperature?

Design

A precisely measured volume of a saturated NaCl(aq) solution at room temperature is heated to evaporate the solvent and crystallize the solute. The mass of the dry solute is measured and the concentration of the saturated solution is calculated.

Materials

lab apron
eye protection
oven mitts or heatproof gloves
saturated NaCl(aq) solution
laboratory burner with matches or striker, or hot plate
centigram balance
thermometer or temperature probe
laboratory stand
ring clamp
wire gauze
250 mL beaker
100 mL beaker
10 mL pipette with pipette bulb

When using a laboratory burner, keep long hair tied back and loose clothing secured. If using a hot plate, take all necessary precautions.

Use oven mitts or heatproof gloves to handle hot apparatus.

Procedure

1. Measure and record the mass of a clean, dry 250 mL beaker. (See Appendix C.3 for tips on using a laboratory balance.)
2. Obtain about 40 mL to 50 mL of saturated NaCl(aq) in a 100 mL beaker.
3. Measure and record the temperature of the saturated solution to a precision of 0.2 °C. (See Appendix F.3 for a note on precision of readings.)
4. Pipette a 10.00 mL sample of the saturated solution into the 250 mL beaker. (See Appendix C.3 and the Nelson Web site for tips on pipetting.)

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5. Using a laboratory burner or hot plate, heat the solution evenly in the beaker until all the water boils away, and dry, crystalline NaCl(s) remains. (See Appendix C.3 for tips on using a laboratory burner. Also, see the video on the Nelson Web site.)

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6. Shut off the burner or hot plate, and allow the beaker and contents to cool for at least 5 min.
7. Measure and record the total mass of the beaker and contents.
8. Reheat the beaker and the residue and repeat steps 6 and 7 until two consecutive measurements of the mass give the same value. Record the final mass. (If the mass remains constant, this confirms that the sample is dry.)
9. Dispose of the salt as regular solid waste.

Learning Tip

You will be able to improve the precision of your prediction if you start the vertical axis of your graph at 31 g/100 mL instead of the usual zero value.

Chapter 5 SUMMARY

Outcomes

Knowledge

- explain the nature of solutions and the dissolving process (5.1, 5.2)
- illustrate how dissolving substances in water is often a prerequisite for chemical change (5.1, 5.2)
- differentiate between electrolytes and nonelectrolytes (5.1, 5.2)
- explain dissolving as an endothermic or an exothermic process with regard to breaking and forming of bonds (5.2)
- express concentration in various ways (5.3)
- perform calculations involving concentration, chemical amount, volume, and/or mass (5.3)
- use dissociation equations to calculate ion concentration (5.3)
- describe the procedures and calculations required for preparing solutions from a pure solid and by dilution (5.4)
- define solubility and identify the factors that affect it (5.5)
- explain a saturated solution in terms of equilibrium (5.5)

STS

- illustrate how science and technology are developed to meet societal needs and expand human capabilities (5.1)
- describe interactions of science, technology, and society (5.3, 5.5)
- relate scientific and technological work to personal and social values such as honesty, perseverance, tolerance, open-mindedness, critical-mindedness, creativity, and curiosity (5.1, 5.3, 5.4, 5.5)
- illustrate how science and technology have both intended and unintended consequences (5.3, 5.5)
- evaluate technologies from a variety of perspectives (5.4, 5.5)

Skills

- initiating and planning: design a procedure to identify the type of solution (5.1); design a procedure for determining the concentration of a solution containing a solid solute (5.4); describe procedures for safe handling, storing, and disposal of material used in the laboratory, with reference to WHMIS and consumer product labelling information (5.1, 5.4, 5.5)
- performing and recording: use a conductivity apparatus to classify solutions (5.1); perform an experiment to determine the concentration of a solution (5.4, 5.5); use a balance and volumetric glassware to prepare solutions of specified concentration (5.4); perform an investigation to determine the solubility of a solute in a saturated solution (5.5)
- analyzing and interpreting: use experimental data to determine the concentration of a solution (5.5)
- communication and teamwork: compare personal concentration data with the data of other groups (5.4, 5.5)

Key Terms

5.1
solution
solute
solvent
electrolyte
nonelectrolyte

5.2
dissociation
ionization

5.3
concentration
amount concentration

5.4
standard solution
stock solution

5.5
saturated solution
solubility
dynamic equilibrium

Key Equations

Concentration Types

percentage by volume	$c = \frac{V_{solute}}{V_{solution}} \times 100\%$	% V/V (or mL/100 mL)
mass by volume	$c = \frac{m_{solute}}{V_{solution}} \times 100\%$	% W/V (or g/100 mL)
by mass	$c = \frac{m_{solute}}{m_{solution}} \times 100\%$	% W/W (or g/100 g)
parts per million	$c = \frac{m_{solute}}{m_{solution}}$	ppm (typically mg/kg)
amount	$c = \frac{n_{solute}}{V_{solution}}$	mol/L

Dilution

$V_i c_i = V_f c_f$

MAKE a summary

1. Devise a concept map built around the subject "Solutions" and include all of the Key Terms listed above.
2. Refer back to your answers to the Starting Points questions at the beginning of this chapter. How has your thinking changed?

Go To

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The following components are available on the Nelson Web site. Follow the links for *Nelson Chemistry Alberta 20–30*.

- an interactive Self Quiz for Chapter 5
- additional Diploma Exam-style Review questions
- Illustrated Glossary
- additional IB-related material

There is more information on the Web site wherever you see the Go icon in this chapter.

Chapter 5 REVIEW

Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

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DO NOT WRITE IN THIS TEXTBOOK.

Part 1

1. Rusting of iron occurs extremely slowly in very dry climates. A likely reason for this observation is that
   A. iron is an inert material
   B. there is a lower concentration of oxygen in very dry climates
   C. dissolving substances in water is usually necessary for chemical change
   D. the higher temperatures prevent rusting because it is an exothermic reaction

2. Cold packs (**Figure 1**) contain an ionic compound such as ammonium nitrate and a separate pouch of water that is broken when the cold pack is needed. Which of the following rows indicates the type of change and the process that produces this change?

Row	Type of change	Process
A.	endothermic	ionization
B.	endothermic	dissociation
C.	exothermic	ionization
D.	exothermic	dissociation

Figure 1
A cold pack

3. NR The maximum acceptable concentration of fluoride ions in municipal water supplies corresponds to 0.375 mg of fluoride in a 250 mL glass of water. The concentration of fluoride ions, in ppm, is __________.

Use this information to answer questions 4 to 7.

Hard water contains metal ions, most commonly calcium and magnesium ions. Some moderately hard water is found to contain 200 ppm of calcium hydrogen carbonate.

4. The dissociation equation for calcium hydrogen carbonate in water is
   A. $CaHCO_3(s) \rightarrow CaH^{2+}(aq) + CO_3^{2-}(aq)$
   B. $CaHCO_3(s) \rightarrow Ca^{2+}(aq) + HCO_3^{-}(aq)$
   C. $Ca(HCO_3)_2(s) \rightarrow Ca^{2+}(aq) + 2\,H^{+}(aq) + 2\,CO_3^{2-}(aq)$
   D. $Ca(HCO_3)_2(s) \rightarrow Ca^{2+}(aq) + 2\,HCO_3^{-}(aq)$

5. The ppm concentration of hydrogen carbonate ions in the hard water is
   A. 400 ppm
   B. 300 ppm
   C. 200 ppm
   D. 100 ppm

6. NR The mass of calcium hydrogen carbonate that would be found in 52.0 L of hard water in a bathtub is __________ g.

7. NR The amount concentration of calcium hydrogen carbonate in the hard water is __________ mmol/L.

8. A Web site promoting eco-friendly alternatives to commercial cleaners suggests mixing 125 mL of vinegar with enough water to make 1.0 L of cleaning solution. If the vinegar used contains 5.0% acetic acid (by volume), what is the percentage concentration of acetic acid in the cleaning solution?
   A. 0.025%
   B. 0.13%
   C. 0.63%
   D. 8.0%

9. NR A 500 mL bottle of fireplace window cleaner contains 2.50 mol/L of potassium hydroxide. The mass of KOH(s) contained in the bottle is __________ g.

10. The main piece of laboratory equipment that is used in both procedures for the preparation of a standard solution, from a solid and by dilution, is a/an
    A. Erlenmeyer flask
    B. volumetric pipette
    C. volumetric flask
    D. graduated cylinder

Use this information to answer questions 11 to 13.

The salt tank attached to a water softener contains excess sodium chloride solid, sitting in a fixed quantity of water. This mixture remains for a long period of time before it is used to regenerate the resin in the water softener.

11. The best description of the salt solution is that it is
A. dilute
B. saturated
C. miscible
D. concentrated

12. There are no observable changes in any properties of the mixture because
A. the rate of dissolving equals the rate of crystallizing
B. no change is occurring at the molecular level
C. the rates of dissolving and dissociating are equal
D. there is no more space for any more salt to dissolve

13. The concentration of the salt solution can be increased by
A. stirring vigorously
B. adding more water
C. removing some solution
D. increasing the temperature

Part 2

14. **How** is a homogeneous mixture different from a heterogeneous mixture? Give one example of each.

15. A chemistry student was given the task of identifying four colourless solutions. Complete the **Analysis** of the investigation report.

Problem

Which of the solutions, labelled A, B, C, and D, is calcium hydroxide, glucose, potassium chloride, and sulfuric acid?

Evidence

Table 1 Litmus and Conductivity Tests

Solution*	Red litmus	Blue litmus	Conductivity
A	stays red	blue to red	high
B	stays red	stays blue	none
C	red to blue	stays blue	high
D	stays red	stays blue	high

*same concentration and temperature

16. Scientists have developed a classification system to help organize the study of matter. **Describe** an empirical test that can be used to distinguish between the following classes of matter:
(a) electrolytes and nonelectrolytes
(b) acids, bases, and neutral compounds

17. What is a standard solution, and **why** is such a solution necessary?

18. **Describe** two methods used to prepare standard solutions.

19. Much of the food you eat is converted to glucose in your digestive tract. The glucose dissolves in the blood and circulates throughout your body. Cells use the glucose to produce energy in the process of cellular respiration. State two reasons why it is important for the glucose to be dissolved in a solution rather than remain as a solid.

20. From Mini Investigation: Hot and Cold Solutions (page 199), you know that the dissolving of sodium nitrate is endothermic. What does this mean, empirically and theoretically?

21. **Describe** the ways in which concentrations of solutions are expressed in chemistry laboratories, household products, and environmental studies.

22. A shopper has a choice of yogurt with three different concentrations (% W/W) of milk fat: 5.9%, 2.0%, and 1.2%. If the shopper wants to limit his or her milk fat intake to 3.0 g per serving, **determine** the mass of the largest serving the shopper could have for each type of yogurt.

23. What volume of vinegar contains 15 mL of pure acetic acid (**Figure 2**)?

Figure 2 The label tells us the concentration of acetic acid in vinegar.

24. **Determine** the amount concentration of the following solutions:
(a) 0.35 mol copper(II) nitrate is dissolved in water to make 500 mL of solution.
(b) 10.0 g of sodium hydroxide is dissolved in water to make 2.00 L of solution.
(c) 25 mL of 11.6 mol/L HCl(aq) is diluted to a volume of 145 mL.
(d) A sample of tap water contains 16 ppm of magnesium ions.

25. Standard solutions of sodium oxalate, $Na_2C_2O_4$(aq), are used in a variety of chemical analyses. **Determine** the mass of sodium oxalate required to prepare 250.0 mL of a 0.375 mol/L solution.

26. Phosphoric acid is the active ingredient in many commercial rust-removing solutions. **Determine** the volume of concentrated phosphoric acid (14.6 mol/L) that must be diluted to prepare 500 mL of a 1.25 mol/L solution.

Use this information to answer questions 27 to 29.

For people with diabetes, monitoring blood glucose levels is essential. There are many products available (**Figure 3**) that typically provide the concentration of glucose in units of millimoles per litre and use as little as 1 µL of blood.

Figure 3
A glucose meter

27. A glucose meter shows a normal reading of 7.8 mmol/L for an average adult, two hours after a meal. What mass of glucose is present in 4.7 L of blood of an average adult?

28. DE Glucose meters need to be checked periodically for accuracy. Checking is done using a standard glucose solution, such as one with a concentration of 3.1 mmol/L.
   (a) If you were to prepare 100.0 mL of this standard solution, what mass of solid glucose is required?
   (b) List the materials required to prepare this standard solution. Specify sizes and quantities.
   (c) Write a complete procedure for the preparation of the standard solution.

29. **How** does the glucose meter illustrate the interaction of science, technology, and society?

Use this information to answer questions 30 to 33.

Acids are usually purchased in their pure or concentrated form and then diluted to the concentration required for a particular use. Concentrated 17.8 mol/L sulfuric acid mixed with water can generate localized temperatures in excess of 100 °C. Sulfuric acid is a common example, but you need to be careful when diluting any acid.

30. When concentrated sulfuric acid dissolves in water, is this process endothermic or exothermic? State the evidence.

31. **Describe** the correct procedure for diluting concentrated reagents such as sulfuric acid. **Why** is it recommended that you always follow this procedure?

32. What volume of concentrated sulfuric acid would a technician require to prepare 2.00 L of 0.250 mol/L solution?

33. Write a dissociation equation to explain the electrical conductivity of each of the following chemicals:
   (a) potash: potassium chloride
   (b) Glauber's salt: sodium sulfate
   (c) TSP: trisodium phosphate

34. **Determine** the amount concentration of the cation and the anion in a 0.14 mol/L solution of each of the following chemicals.
   (a) saltpetre: KNO_3
   (b) road salt: calcium chloride
   (c) fertilizer: ammonium phosphate

Extension

35. DE The oil industry is an increasingly important component of Alberta's economy. Part of this industry involves transporting oil products in rail cars. The spill of oil products near Lake Wabamun in August 2005 is a dramatic example of the risks involved in getting products to market. Prepare a fact sheet on the transportation of oil products. Your response should include:
   - the nature of two spilled substances, including names, uses, and properties such as solubility, density, and toxicity
   - the risks and benefits of transporting these products, including several perspectives

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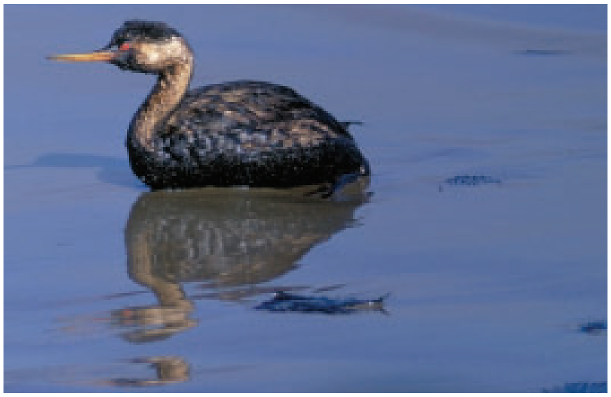

Figure 4
An oil-covered Western Grebe