

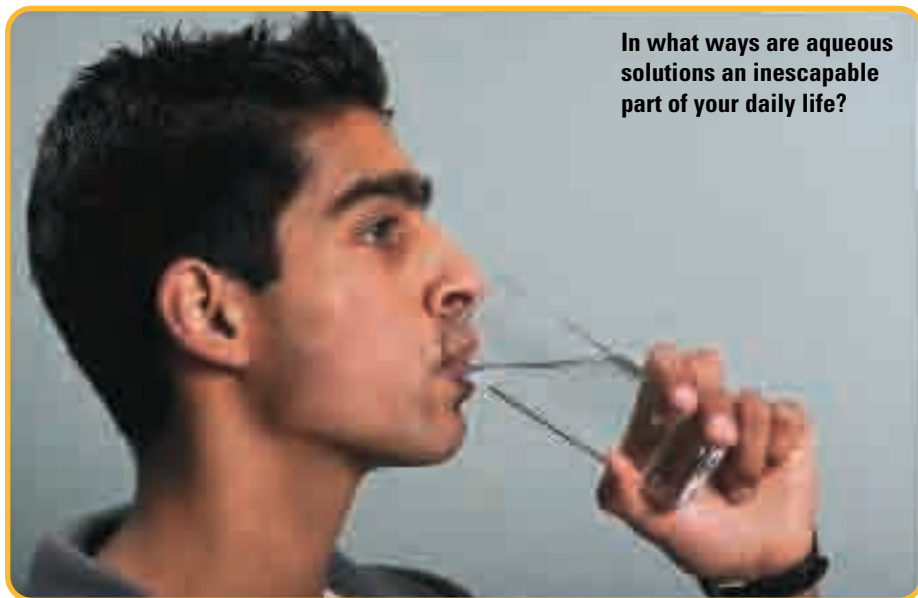
# Aqueous Solutions

**I**n 1966, a Russian scientist who was lecturing in England stunned the scientific community with the report of a new discovery. He said that a colleague had isolated a new form of liquid water. Dubbed “polywater,” this substance was prepared in a laboratory by heating and then condensing  $\text{H}_2\text{O}_{(\text{g})}$  in very narrow glass capillary tubes. Over the next seven years, scientists around the world conducted studies and published over 500 papers about polywater’s properties. For example, its boiling point, density, and viscosity were greater than those of ordinary water. Its freezing point was lower.

Then the popular media took an interest. They turned a simmering scientific curiosity into a boiling concern. Polywater apparently had a powerful capacity for hydrogen bonding. What if it escaped from a laboratory? There were dire predictions that polywater would “take over” Earth’s water resources. The imagined consequences for life on Earth were grim.

Eventually further studies revealed that polywater was simply a concentrated solution of silicon (silicic acid) and several ionic compounds in ordinary water. The glass tubing was the source of these solutes. They had leached into the water.

The polywater event occurred because scientists overlooked, for awhile, water’s remarkable power as a solvent. In this chapter, you will learn how to predict which compounds are soluble in water. As well, you will consider how chemical reactions in aqueous (water) solutions are useful in industry and in protecting the quality of our water supplies.



**In what ways are aqueous solutions an inescapable part of your daily life?**

## Chapter Preview

- 9.1** Making Predictions About Solubility
- 9.2** Reactions in Aqueous Solutions
- 9.3** Stoichiometry in Solution Chemistry
- 9.4** Aqueous Solutions and Water Quality

## Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- identifying factors that affect solubility (Chapter 8, section 8.2)
- performing stoichiometry calculations (Chapter 7, section 7.1)

# 9.1

## Making Predictions About Solubility

### Section Preview/ Specific Expectations

In this section, you will

- **describe** and **identify** combinations of aqueous solutions that result in the formation of precipitates
- **communicate**, using appropriate scientific vocabulary, ideas related to solubility and aqueous solutions
- **communicate** your understanding of the following terms: *precipitate*, *general solubility guidelines*

In Chapter 8, you examined factors that affect the solubility of a compound. As well, you learned that the terms “soluble” and “insoluble” are relative, because no substance is completely insoluble in water. “Soluble” generally means that more than about 1 g of solute will dissolve in 100 mL of water at room temperature. “Insoluble” means that the solubility is less than 0.1 g per 100 mL. While many ionic compounds are soluble in water, many others are not. Cooks, chemists, farmers, pharmacists, and gardeners need to know which compounds are soluble and which are insoluble. (See Figure 9.1.)



**Figure 9.1** Plant food may come in the form of a liquid or a powder. Gardeners dissolve the plant food in water, and then either spray or water the plant with the resulting solution.

### Factors That Affect the Solubility of Ionic Substances

Nearly all alkali metal compounds are soluble in water. Sulfide and phosphate compounds are usually insoluble. How, then, do you account for the fact that sodium sulfide and potassium phosphate are soluble, while iron sulfide and calcium phosphate are insoluble? Why do some ions form soluble compounds, while other ions form insoluble compounds?

#### The Effect of Ion Charge on Solubility

Compounds of ions with small charges tend to be soluble. Compounds of ions with large charges tend to be insoluble. Why? Increasing the charge increases the force that holds the ions together. For example, phosphates (compounds of  $\text{PO}_4^{3-}$ ) tend to be insoluble. On the other hand, the salts of alkali metals are soluble. Alkali metal cations have a single positive charge, so the force that holds the ions together is less.



**CHEM**

**FACT**

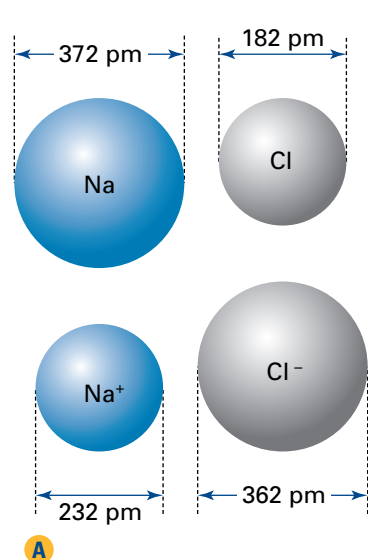
Because the terms “soluble” and “insoluble” are relative, some textbooks give different definitions and units of concentration to describe them. Here are a few examples:

Soluble	Partly or slightly soluble	Insoluble
more than 1 g in 100 mL or greater than 0.1 mol/L	between 1 g and 0.1 g in 100 mL or between 0.1 mol/L and 0.01 mol/L	less than 0.1 g in 100 mL or less than 0.01 mol/L

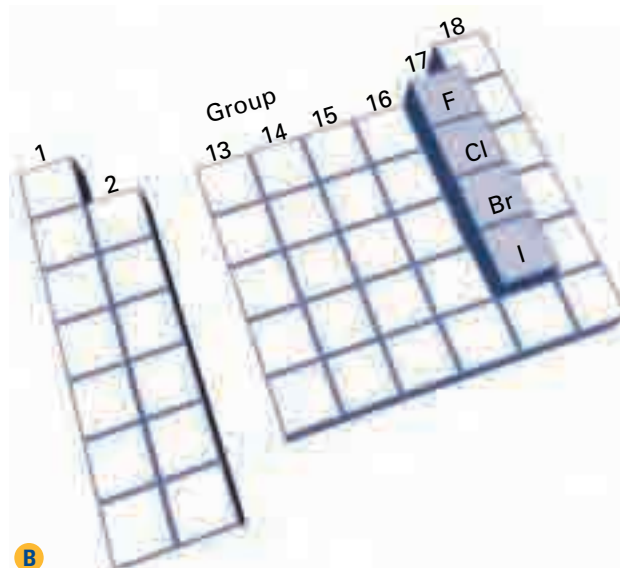
## The Effect of Ion Size on Solubility

When an atom gives up or gains an electron, the size of the ion that results is different from the size of the original atom. In Figure 9.2A, for example, you can see that the sodium ion is smaller than the sodium atom. In general, the ions of metals tend to be smaller than their corresponding neutral atoms. For the ions of non-metals, the reverse is true. The ions of non-metals tend to be larger than their corresponding neutral atoms.

Small ions bond more closely together than large ions. Thus, the bond between small ions is stronger than the bond between large ions with the same charge. As a result, compounds with small ions tend to be less soluble than compounds with large ions. Consider the ions of elements from Group 17 (VIIA), for example. Recall, from Chapters 2 and 3, that the size of the ions increases as you go down a family in the periodic table. (See Figure 9.2B.) Therefore, you would expect that fluoride compounds are less soluble than chloride, bromide, and iodide compounds. This tends to be the case.



**Figure 9.2A** The size of the cation of a metal, such as sodium, tends to be smaller than the size of the atom from which it was formed. The size of the anion of a non-metal, such as phosphorus, tends to be larger than the size of its atom.



**Figure 9.2B** In Group 17 (VIIA), fluoride ions are smaller than chloride, bromide, and iodide ions. How does this periodic trend affect the solubility of compounds that are formed from these elements?

Ionic size		
F <sup>-</sup>		266 pm
Cl <sup>-</sup>		362 pm
Br <sup>-</sup>		392 pm
I <sup>-</sup>		440 pm

## Making Predictions About Solubility

Sulfides (compounds of S<sup>2-</sup>) and oxides (compounds of O<sup>2-</sup>) are influenced by both ion size and ion charge. These compounds tend to be insoluble because their ions have a double charge *and* are relatively small. Even so, a few sulfides and oxides are soluble, as you will discover in Investigation 1-A.

Many interrelated factors affect the solubility of substances in water. This makes it challenging to predict which ionic substances will dissolve in water. By performing experiments, chemists have developed guidelines to help them make predictions about solubility. In Investigation 9-A, you will perform your own experiments to develop guidelines about the solubility of ionic compounds in water.

## mind STRETCH

Sketch an outline of the periodic table. Add labels and arrows to indicate what you think are the trends for ionic size (radius) across a period and down a group. Then suggest reasons for these trends. Specifically, identify the factors that you think are responsible for the differences in the sizes of ions and their “parent” atoms.

# The Solubility of Ionic Compounds

In this investigation, you will work with a set of solutions. You will chemically combine small quantities, two at a time. This will help you determine which combinations react to produce a **precipitate**. A precipitate is an insoluble solid that may result when two aqueous solutions chemically react. *The appearance of a precipitate indicates that an insoluble compound is present.* Then you will compile your data with the data from other groups to develop some guidelines about the solubility of several ionic compounds.

## Problem

How can you develop guidelines to help you predict the solubility of ionic compounds in water?

## Prediction

Read the entire Procedure. Predict which combination of anions and cations will likely be soluble and which combination will likely be insoluble. Justify your prediction by briefly explaining your reasoning.

## Materials

12-well or 24-well plate, or spot plate  
toothpicks  
cotton swabs  
wash bottle with distilled water  
piece of black paper  
piece of white paper  
labelled dropper bottles of aqueous solutions that contain the following cations:  
 $\text{Al}^{3+}$ ,  $\text{NH}_4^+$ ,  $\text{Ba}^{2+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Cu}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Mg}^{2+}$ ,  $\text{Ag}^+$ ,  $\text{Na}^+$ ,  $\text{Zn}^{2+}$   
labelled dropper bottles of aqueous solutions that contain the following anions:  
 $\text{CH}_3\text{COO}^-$ ,  $\text{Br}^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{Cl}^-$ ,  $\text{OH}^-$ ,  $\text{PO}_4^{3-}$ ,  $\text{SO}_4^{2-}$ ,  $\text{S}^{2-}$

## Safety Precautions



- Do not contaminate the dropper bottles. The tip of a dropper should not make contact with either the plate or another solution. Put the cap back on the bottle immediately after use.
- Dispose of solutions as directed by your teacher.
- Make sure that you are working in a well-ventilated area.
- If you accidentally spill any of the solutions on your skin, wash the area immediately with plenty of cool water.

## Procedure

1. Your teacher will give you a set of nine solutions to test. Each solution includes one of five cations or four anions. Design a table to record the results of all the possible combinations of cations with anions in your set of solutions.
2. Decide how to use the well plate or spot plate to test systematically all the combinations of cations with anions in your set. If your plate does not have enough wells, you will need to clean the plate before you can test all the possible combinations. To clean the plate, first discard solutions into the container provided by your teacher. Then rinse the plate with distilled water, and clean the wells using a cotton swab.
3. To test each combination of anion and cation, add one or two drops into your well plate or spot plate. Then stir the mixture using a toothpick. Rinse the toothpick with running water before each stirring. Make sure that you keep track of the combinations of ions in each well or spot.





**Why is it necessary to clean the well or spot plate as described in step 2?**

4. Examine each mixture for evidence of a precipitate. Place the plate on a sheet of white or black paper. (Use whichever colour of paper helps you see a precipitate best.) Any cloudy appearance in the mixture is evidence of a precipitate. Many precipitates are white.
  - If you can see that a precipitate has formed, enter “I” in your table. This indicates that the combination of ions produces an insoluble substance.
  - If you cannot see a precipitate, enter “S” to indicate that the ion you are testing is soluble.
5. Repeat steps 2 to 4 for each cation solution.
6. Discard the solutions and precipitates into the container provided by your teacher. Rinse the plate with water, and clean the wells using a cotton swab.
7. If time permits, your teacher may give you a second set of solutions to test.
8. Add your observations to the class data table. Use your completed copy of the class data table to answer the questions below.

### Analysis

1. Identify any cations that
  - (a) always appear to form soluble compounds
  - (b) always appear to form insoluble compounds

2. Identify any anions that
  - (a) always appear to form soluble compounds
  - (b) always appear to form insoluble compounds
3. Based on your observations, which sulfates are insoluble?
4. Based on your observations, which phosphates are soluble?
5. Explain why each reagent solution you tested must contain both cations and anions.
6. Your teacher prepared the cation solutions using compounds that contain the nitrate ion. For example, the solution marked  $\text{Ca}^{2+}$  was prepared by dissolving  $\text{Ca}(\text{NO}_3)_2$  in water. Why were nitrates used to make these solutions?

### Conclusions

7. Which group in the periodic table most likely forms cations with salts that are usually soluble?
8. Which group in the periodic table most likely forms anions with salts that are usually soluble?
9. Your answers to questions 7 and 8 represent a preliminary set of guidelines for predicting the solubility of the compounds you tested. Many reference books refer to guidelines like these as “solubility rules.” Why might “solubility guidelines” be a better term to use for describing solubility patterns?

### Application

10. Predict another combination of an anion and a cation (not used in this investigation) that you would expect to be soluble. Predict another combination that you would expect to be insoluble. Share your predictions, and your reasons, with the class. Account for any agreement or disagreement.

## Soluble or Insoluble: General Solubility Guidelines

As you have seen, nearly all salts that contain the ammonium ion or an alkali metal are soluble. This observed pattern does not tell you how soluble these salts are, however. As well, it does not tell you whether ammonium chloride is more or less soluble than sodium chloride. Chemists rely on published data for this information. (See Figure 9.3.)

**Figure 9.3** Many web sites on the Internet provide chemical and physical data for tens of thousands of compounds. Print resources, such as *The CRC Handbook of Chemistry and Physics*, provide these data as well.



Many factors affect solubility. Thus, predicting solubility is neither straightforward nor simple. Nevertheless, the **general solubility guidelines** in Table 9.1 are a useful summary of ionic-compound interactions with water. To use Table 9.1, remember that a higher guideline number always takes precedence over a lower guideline number. For example, barium chloride,  $\text{BaCl}_2$ , is a white crystalline powder. The barium ion,  $\text{Ba}^{2+}$ , is listed in guideline 4 as insoluble. The chloride ion,  $\text{Cl}^-$ , is listed in guideline 3 as soluble. The higher guideline number takes precedence. Thus, you would predict that barium chloride is soluble.

You will be referring to the general solubility guidelines often in this chapter and in Chapter 10. They will help you identify salts that are soluble and insoluble in aqueous solutions. Always keep in mind, however, that water is a powerful solvent. Even an “insoluble” salt may dissolve enough to present a serious hazard if it is highly poisonous.

**Table 9.1** General Solubility Guidelines

Guideline	Cations	Anions	Result	Exceptions
1	$\text{Li}^+$ , $\text{Na}^+$ , $\text{K}^+$ , $\text{Rb}^+$ , $\text{Cs}^+$ , $\text{NH}_4^+$	$\text{NO}_3^-$ , $\text{CH}_3\text{COO}^-$ , $\text{ClO}_3^-$	soluble	$\text{Ca}(\text{ClO}_3)_2$ is insoluble
2	$\text{Ag}^+$ , $\text{Pb}^{2+}$ , $\text{Hg}^+$	$\text{CO}_3^{2-}$ , $\text{PO}_4^{3-}$ , $\text{O}^{2-}$ , $\text{S}^{2-}$ , $\text{OH}^-$	insoluble	$\text{BaO}$ and $\text{Ba}(\text{OH})_2$ are soluble. Group 2 sulfides tend to decompose.
3		$\text{Cl}^-$ , $\text{Br}^-$ , $\text{I}^-$	soluble	
4	$\text{Ba}^{2+}$ , $\text{Ca}^{2+}$ , $\text{Sr}^{2+}$		insoluble	
5	$\text{Mg}^{2+}$ , $\text{Cu}^{2+}$ , $\text{Zn}^{2+}$ , $\text{Fe}^{2+}$ , $\text{Fe}^{3+}$ , $\text{Al}^{3+}$	$\text{SO}_4^{2-}$	soluble	

## Practice Problems

1. Decide whether each of the following salts is soluble or insoluble in distilled water. Give reasons for your answer.
  - (a) lead(II) chloride,  $\text{PbCl}_2$  (a white crystalline powder used in paints)
  - (b) zinc oxide,  $\text{ZnO}$  (a white pigment used in paints, cosmetics, and calamine lotion)
  - (c) silver acetate,  $\text{AgCH}_3\text{COO}$  (a whitish powder that is used to help people quit smoking because of the bitter taste it produces)
2. Which of the following compounds are soluble in water? Explain your reasoning for each compound.
  - (a) potassium nitrate,  $\text{KNO}_3$  (used to manufacture gunpowder)
  - (b) lithium carbonate,  $\text{Li}_2\text{CO}_3$  (used to treat people who suffer from depression)
  - (c) lead(II) oxide,  $\text{PbO}$  (used to make crystal glass)
3. Which of the following compounds are insoluble in water?
  - (a) calcium carbonate,  $\text{CaCO}_3$  (present in marble and limestone)
  - (b) magnesium sulfate,  $\text{MgSO}_4$  (found in the hydrated salt,  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ , also known as Epsom salts; used for the relief of aching muscles and as a laxative)
  - (c) aluminum phosphate,  $\text{AlPO}_4$  (found in dental cements)

Web

LINK

[www.school.mcgrawhill.ca/resources/](http://www.school.mcgrawhill.ca/resources/)

Different references present solubility guidelines in different ways. During a 1 h “surf session” on the Internet, a student collected ten different versions. See how many versions you can find. Compare their similarities and differences. Which version(s) do you prefer, and why? To start your search, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

## Section Wrap-up

In Chapter 8, you focussed mainly on physical changes that involve solutions. In the first section of this chapter, you observed that mixing aqueous solutions of ionic compounds may result in either a physical change (dissolving) or a chemical change (a reaction that forms a precipitate). Chemical changes that involve aqueous solutions, especially ionic reactions, are common. They occur in the environment, in your body, and in the bodies of other organisms. In the next section, you will look more closely at reactions that involve aqueous solutions. As well, you will learn how to represent these reactions using a special kind of chemical equation, called an ionic equation.

## Section Review

### Language

### LINK

This poem is by an unknown author. It is not great literature. For decades, however, it has helped many students remember the solubility guidelines. Maybe it will do the same for you. If not, you can try writing your own poem!

Potassium, sodium, and ammonium salts  
Whatever they may be,  
Can always be relied upon  
For solubility.

Every single sulfate  
Is soluble it's said,  
Except barium and calcium  
And strontium and lead.

Most every chloride's soluble,  
That's what we've always read,  
Save silver, mercurous mercury,  
And (slightly) chloride of lead.

When asked about the nitrates,  
The answer's always clear;  
They each and all are soluble,  
That's all we want to hear.

Metallic bases won't dissolve,  
That is, all but three;  
Potassium, sodium, and ammonium  
Dissolve quite readily.

But then you must remember,  
You must surely not forget,  
Calcium and barium  
Dissolve a little bit.

Carbonates are insoluble,  
It's lucky that it's so,  
Or all our marble buildings  
Would melt away like snow.

- 1 (a) **K/U** Name the two factors that affect the solubility of an ionic compound in water.  
(b) Briefly explain how each factor affects solubility.
- 2 **K/U** Which would you expect to be less soluble: sodium fluoride, NaF (used in toothpaste), or sodium iodide, NaI (added to table salt to prevent iodine deficiency in the diet)? Explain your answer.
- 3 **K/U** Which of the following compounds are soluble in water?
  - (a) calcium sulfide, CaS (used in skin products)
  - (b) iron(II) sulfate, FeSO<sub>4</sub> (used as a dietary supplement)
  - (c) magnesium chloride, MgCl<sub>2</sub> (used as a disinfectant and a food tenderizer)
- 4 **MC** Which of the following compounds are insoluble in water? For each compound, relate its solubility to the use described.
  - (a) barium sulfate, BaSO<sub>4</sub> (can be used to obtain images of the stomach and intestines because it is opaque to X-rays)
  - (b) aluminum hydroxide, Al(OH)<sub>3</sub> (found in some antacid tablets)
  - (c) zinc carbonate, ZnCO<sub>3</sub> (used in suntan lotions)
- 5 **C** Calcium nitrate is used in fireworks. Silver nitrate turns dark when exposed to sunlight. When freshly made, both solutions are clear and colourless. Imagine that someone has prepared both solutions but has not labelled them. You do not want to wait for the silver nitrate solution to turn dark in order to identify the solutions. Name a chemical that can be used to precipitate a silver compound with the silver nitrate solution, but will produce no precipitate with the calcium nitrate solution. State the reason for your choice.
- 6 **C** Suppose that you discover four dropper bottles containing clear, colourless liquids in your school laboratory. The following four labels lie nearby:
  - barium, Ba<sup>2+</sup>
  - chloride, Cl<sup>-</sup>
  - silver, Ag<sup>+</sup>
  - sulfate, SO<sub>4</sub><sup>2-</sup>.

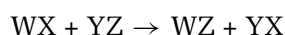
Unfortunately the labels have not been attached to the bottles. You decide to number the bottles 1, 2, 3, and 4. Then you mix the solutions in pairs. Three combinations give white precipitates: bottles 1 and 2, 1 and 4, and 2 and 3. Which ion does each bottle contain?



# Reactions in Aqueous Solutions

## 9.2

When you mix two aqueous ionic compounds together, there are two possible outcomes. Either the compounds will remain in solution without reacting, or one aqueous ionic compound will chemically react with the other. How can you predict which outcome will occur? Figure 9.4 shows what happens when an aqueous solution of lead(II) nitrate is added to an aqueous solution of potassium iodide. As you can see, a yellow solid—a precipitate—is forming. This is a double displacement reaction. Recall, from Chapter 4, that a double displacement reaction is a chemical reaction that involves the exchange of ions to form two new compounds. It has the general equation



In a double displacement reaction, the cations exchange anions. In the reaction shown in Figure 9.4, for example, the lead cation is exchanged with the iodide anion.

You can usually recognize a double displacement reaction by observing one of these possible results:

- the formation of a precipitate (so that ions are removed from solution as an insoluble solid)
- the formation of a gas (so that ions are removed from solution in the form of a gaseous product)
- the formation of water (so that  $H^+$  and  $OH^-$  ions are removed from solution as water)

In this section, you will examine each of these results. At the same time, you will learn how to represent a double displacement reaction using a special kind of chemical equation: an ionic equation.



**Figure 9.4** Lead(II) nitrate and potassium iodide are clear, colourless aqueous solutions. Mixing them causes a double displacement reaction. An insoluble yellow precipitate (lead(II) iodide) and a soluble salt (potassium nitrate) are produced.

### Section Preview/ Specific Expectations

In this section, you will

- **describe** combinations of aqueous solutions that result in the formation of precipitates
- **perform** a qualitative analysis of ions in solutions
- **represent** double displacement reactions by their net ionic equations
- **write** balanced chemical equations and net ionic equations for double displacement reactions
- **communicate** your understanding of the following terms: *spectator ions*, *total ionic equation*, *net ionic equation*, *qualitative analysis*

### Language

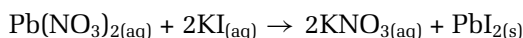
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Double displacement reactions are also called *metathesis reactions*. The word “metathesis” (pronounced with the stress on the second syllable: meh-TATH-e-sis) means “interchange.” In chemistry, a metathesis reaction occurs when ions or atoms are exchanged between different compounds. Chemists are not the only people who use this word. Use a dictionary or encyclopedia to find non-chemistry examples of metathesis.

## Double Displacement Reactions That Produce a Precipitate

A double displacement reaction that results in the formation of an insoluble substance is often called a *precipitation reaction*. Figure 9.4 is a clear example of a precipitation reaction. What if you did not have this photograph, however, and you were unable to do an experiment? Could you have predicted that mixing  $\text{Pb}(\text{NO}_3)_{2(\text{aq})}$  and  $2\text{KI}_{(\text{aq})}$  would result in an insoluble compound? Yes. When you are given (on paper) a pair of solutions to be mixed together, start by thinking about the exchange of ions that may occur. Then use the general solubility guidelines (Table 9.1) to predict which compounds, if any, are insoluble.

For example, consider lead(II) nitrate,  $\text{Pb}(\text{NO}_3)_2$ , and potassium iodide, KI. Lead(II) nitrate contains  $\text{Pb}^{2+}$  cations and  $\text{NO}_3^-$  anions. Potassium iodide contains  $\text{K}^+$  cations and  $\text{I}^-$  anions. Exchanging positive ions results in lead(II) iodide,  $\text{PbI}_2$ , and potassium nitrate,  $\text{KNO}_3$ . From the solubility guidelines, you know that all potassium salts and nitrates are soluble. Thus, potassium nitrate is soluble. The  $\text{Pb}^{2+}$  ion is listed in guideline 2 as an insoluble cation. The  $\text{I}^-$  ion is listed in guideline 3 as a soluble anion. Remember that a higher guideline number takes precedence over a lower guideline number. Thus, you can predict that lead(II) iodide is insoluble. It will form a precipitate when the solutions are mixed. The balanced chemical equation for this reaction is



### Sample Problem

#### Predicting the Formation of a Precipitate

##### Problem

Which of the following pairs of aqueous solutions produce a precipitate when mixed together? Write the balanced chemical equation if you predict a precipitate. Write “NR” if you predict that no reaction takes place.

- (a) potassium carbonate and copper(II) sulfate
- (b) ammonium chloride and zinc sulfate

##### What Is Required?

You need to predict whether or not each pair of aqueous solutions forms an insoluble product (a precipitate). If it does, you need to write a balanced chemical equation.

##### What Is Given?

You know the names of the compounds in each solution.

##### Plan Your Strategy

Start by identifying the ions in each pair of compounds. Then exchange the positive ions in the two compounds. Compare the resulting compounds against the solubility guidelines, and make your prediction.

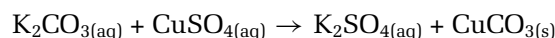
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### Act on Your Strategy

- (a) Potassium carbonate contains  $\text{K}^+$  and  $\text{CO}_3^{2-}$  ions. Copper(II) sulfate contains  $\text{Cu}^{2+}$  and  $\text{SO}_4^{2-}$  ions. Exchanging positive ions results in potassium sulfate,  $\text{K}_2\text{SO}_4$ , and copper(II) carbonate,  $\text{CuCO}_3$ .

All potassium salts are soluble, so these ions remain dissolved in solution.

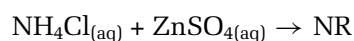
The copper(II) ion is listed in guideline 5 as a soluble cation. The carbonate anion is listed in guideline 2 as insoluble. Because guideline 2 is higher, copper(II) carbonate should be insoluble. So you can predict that a precipitate forms. The balanced chemical equation for this reaction is



- (b) Ammonium chloride contains  $\text{NH}_4^+$  and  $\text{Cl}^-$  ions. Zinc sulfate consists of  $\text{Zn}^{2+}$  and  $\text{SO}_4^{2-}$  ions. Exchanging positive ions results in ammonium sulfate,  $(\text{NH}_4)_2\text{SO}_4$ , and zinc chloride,  $\text{ZnCl}_2$ .

Since all ammonium salts are soluble, the ammonium sulfate stays dissolved in solution.

Zinc chloride consists of a guideline 5 soluble cation and a guideline 3 soluble anion. Because guideline 3 is higher, zinc chloride should be soluble. Thus, you can predict that no precipitate forms.



### Check Your Solution

An experiment is always the best way to check a prediction. If possible, obtain samples of these solutions from your teacher, and mix them together.

### Practice Problems

4. Predict the result of mixing each pair of aqueous solutions. Write a balanced chemical equation if you predict that a precipitate forms. Write “NR” if you predict that no reaction takes place.
- (a) sodium sulfide and iron(II) sulfate
  - (b) sodium hydroxide and barium nitrate
  - (c) cesium phosphate and calcium bromide
  - (d) sodium carbonate and sulfuric acid
  - (e) sodium nitrate and copper(II) sulfate
  - (f) ammonium iodide and silver nitrate
  - (g) potassium carbonate and iron(II) nitrate
  - (h) aluminum nitrate and sodium phosphate
  - (i) potassium chloride and iron(II) nitrate
  - (j) ammonium sulfate and barium chloride
  - (k) sodium sulfide and nickel(II) sulfate
  - (l) lead(II) nitrate and potassium bromide



**Figure 9.5** You can easily identify limestone and marble by their reaction with hydrochloric acid. The gas that is produced by this double displacement reaction is carbon dioxide.

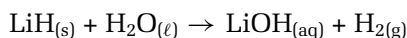
## Double Displacement Reactions That Produce a Gas

Double displacement reactions are responsible for producing a number of gases. (See Figure 9.5.) These gases include

- hydrogen
- hydrogen sulfide (a poisonous gas that smells like rotten eggs)
- sulfur dioxide (a reactant in forming acid rain)
- carbon dioxide
- ammonia

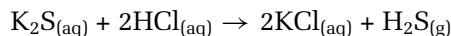
### A Reaction that Produces Hydrogen Gas

The alkali metals form bonds with hydrogen to produce compounds called hydrides. Hydrides react readily with water to produce hydrogen gas. Examine the following equation for the reaction of lithium hydride, LiH, with water. If you have difficulty visualizing the ion exchange that takes place, rewrite the equation for yourself using HOH instead of H<sub>2</sub>O.



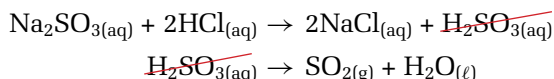
### A Reaction that Produces Hydrogen Sulfide Gas

Sulfides react with certain acids, such as hydrochloric acid, to produce hydrogen sulfide gas.

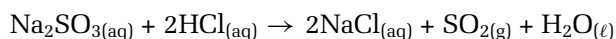


### A Reaction that Produces Sulfur Dioxide Gas

Some reactions produce a compound that, afterward, decomposes into a gas and water. Sodium sulfite is used in photography as a preservative. It reacts with hydrochloric acid to form sulfurous acid. The sulfurous acid then breaks down into sulfur dioxide gas and water. The net reaction is the sum of both changes. If the same compound appears on both sides of an equation (as sulfurous acid, H<sub>2</sub>SO<sub>3</sub>, does here), it can be eliminated. This is just like eliminating terms from an equation in mathematics.

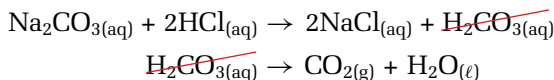


Therefore, the net reaction is

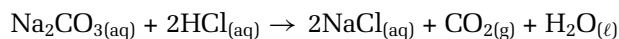


### A Reaction that Produces Carbon Dioxide Gas

The reaction of a carbonate with an acid produces carbonic acid. Carbonic acid decomposes rapidly into carbon dioxide and water.

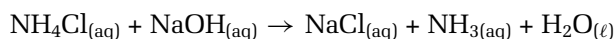


The net reaction is



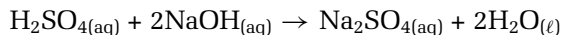
### A Reaction that Produces Ammonia Gas

Ammonia gas is very soluble in water. You can detect it easily, however, by its sharp, pungent smell. Ammonia gas can be prepared by the reaction of an ammonium salt with a base.

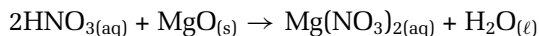


## Double Displacement Reactions That Produce Water

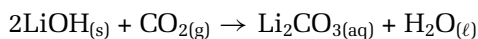
The neutralization reaction between an acid and a base is a very important double displacement reaction. In a neutralization reaction, water results when an  $\text{H}^+$  ion from the acid bonds with an  $\text{OH}^-$  ion from the base.



Most metal oxides are bases. Therefore, a metal oxide will react with an acid in a neutralization reaction to form a salt and water.



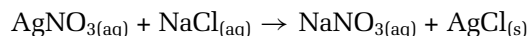
Non-metal oxides are acidic. Therefore, a non-metal oxide will react with a base. This type of reaction is used in the space shuttle. Cabin air is circulated through canisters of lithium hydroxide (a base) to remove the carbon dioxide before it can reach dangerous levels.



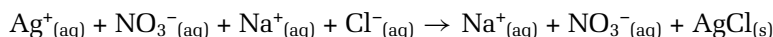
## Representing Aqueous Ionic Reactions with Net Ionic Equations

Mixing a solution that contains silver ions with a solution that contains chloride ions produces a white precipitate of silver chloride. There must have been other ions present in each solution, as well. You know this because it is impossible to have a solution of just a cation or just an anion. Perhaps the solution that contained silver ions was prepared using silver nitrate or silver acetate. Similarly, the solution that contained chloride ions might have been prepared by dissolving  $\text{NaCl}$  in water, or perhaps  $\text{NH}_4\text{Cl}$  or another soluble chloride. Any solution that contains  $\text{Ag}^+_{(\text{aq})}$  will react with any other solution that contains  $\text{Cl}^-_{(\text{aq})}$  to form a precipitate of  $\text{AgCl}_{(\text{s})}$ . The other ions in the solutions are not important to the net result. These ions are like passive onlookers. They are called **spectator ions**.

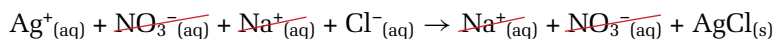
The reaction between silver nitrate and sodium chloride can be represented by the following chemical equation:



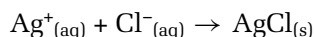
This equation does not show the change that occurs, however. It shows the reactants and products as intact compounds. In reality, soluble ionic compounds dissociate into their respective ions in solution. So chemists often use a **total ionic equation** to show the dissociated ions of the soluble ionic compounds.



Notice that the precipitate,  $\text{AgCl}$ , is still written as an ionic formula. This makes sense because precipitates are insoluble, so they do not dissociate into ions. Also notice that the spectator ions appear on both sides of the equation. Here is the total ionic equation again, with slashes through the spectator ions.



If you eliminate the spectator ions, the equation becomes





An ionic equation that is written this way, without the spectator ions, is called a **net ionic equation**. Before you try writing your own net ionic equations, examine the guidelines in Table 9.2 below.

**Table 9.2** Guidelines for Writing a Net Ionic Equation

1. Include only ions and compounds that have reacted. Do not include spectator ions.
2. Write the soluble ionic compounds as ions. For example, write  $\text{NH}_4^+_{(\text{aq})}$  and  $\text{Cl}^-_{(\text{aq})}$ , instead of  $\text{NH}_4\text{Cl}_{(\text{aq})}$ .
3. Write insoluble ionic compounds as formulas, not ions. For example, zinc sulfide is insoluble, so you write it as  $\text{ZnS}_{(\text{s})}$ , not  $\text{Zn}^{2+}$  and  $\text{S}^{2-}$ .
4. Since covalent compounds do not produce ions in aqueous solution, write their molecular formulas. Water is a common example, because it dissociates only very slightly into ions. When a reaction involves a gas, always include the gas in the net ionic equation.
5. Write strong acids (discussed in the next chapter) in their ionic form. There are six strong acids:
  - hydrochloric acid (write as  $\text{H}^+_{(\text{aq})}$  and  $\text{Cl}^-_{(\text{aq})}$ , not  $\text{HCl}_{(\text{aq})}$ )
  - hydrobromic acid (write as  $\text{H}^+_{(\text{aq})}$  and  $\text{Br}^-_{(\text{aq})}$ )
  - hydroiodic acid (write as  $\text{H}^+_{(\text{aq})}$  and  $\text{I}^-_{(\text{aq})}$ )
  - sulfuric acid (write as  $\text{H}^+_{(\text{aq})}$  and  $\text{SO}_4^{2-}_{(\text{aq})}$ )
  - nitric acid (write as  $\text{H}^+_{(\text{aq})}$  and  $\text{NO}_3^-_{(\text{aq})}$ )
  - perchloric acid (write as  $\text{H}^+_{(\text{aq})}$  and  $\text{ClO}_4^-_{(\text{aq})}$ )All other acids are weak and form few ions. Therefore, write them in their molecular form.
6. Finally, check that the net ionic equation is balanced for charges as well as for atoms.

## Sample Problem

### Writing Net Ionic Equations

#### Problem

A chemical reaction occurs when the following aqueous solutions are mixed: sodium sulfide and iron(II) sulfate. Identify the spectator ions. Then write the balanced net ionic equation.

#### What Is Required?

You need to identify the spectator ions and write a balanced net ionic equation for the reaction between sodium sulfide and iron(II) sulfate.

#### What Is Given?

You know the chemical names of the compounds.

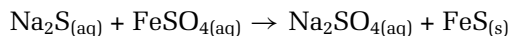
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### Plan Your Strategy

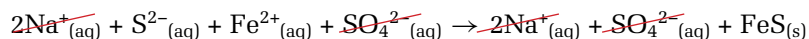
- Step 1** Start by writing the chemical formulas of the given compounds.
- Step 2** Then write the complete chemical equation for the reaction, using your experience in predicting the formation of a precipitate.
- Step 3** Once you have the chemical equation, you can replace the chemical formulas of the soluble ionic compounds with their dissociated ions.
- Step 4** This will give you the total ionic equation. Next you can identify the spectator ions (the ions that appear on both sides of the equation).
- Step 5** Finally, by rewriting the total ionic equation *without* the spectator ions, you will have the net ionic equation.

### Act on Your Strategy

**Steps 1 and 2** The chemical equation for the reaction is

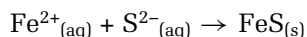


**Step 3** The total ionic equation is



**Step 4** Therefore, the spectator ions are  $\text{Na}^+_{(\text{aq})}$  and  $\text{SO}_4^{2-}_{(\text{aq})}$ .

**Step 5** The net ionic equation is



### Check Your Solution

Take a final look at your net ionic equation to make sure that no ions are on both sides of the equation.

## Practice Problems

- Mixing each pair of aqueous solutions results in a chemical reaction. Identify the spectator ions. Then write the balanced net ionic equation.
  - sodium carbonate and hydrochloric acid
  - sulfuric acid and sodium hydroxide
- Identify the spectator ions for the reaction that takes place when each pair of aqueous solutions is mixed. Then write the balanced net ionic equation.
  - ammonium phosphate and zinc sulfate
  - lithium carbonate and nitric acid
  - sulfuric acid and barium hydroxide

## Identifying Ions in Aqueous Solution

Suppose that you have a sample of water. You want to know what, if any, ions are dissolved in it. Today technological devices, such as the mass spectrometer, make this investigative work fairly simple. Before such devices, however, chemists relied on *wet chemical techniques*: experimental tests, such as submitting a sample to a series of double displacement reactions. Chemists still use wet chemical techniques. With each reaction, insoluble compounds precipitate out of the solution. (See Figure 9.6.) This enables the chemist to determine, eventually, the identity of one or several ions in the solution. This ion-identification process is an example of **qualitative analysis**.

Chemists use a range of techniques for qualitative analysis. For example, the colour of an aqueous solution can help to identify one of the ions that it contains. Examine Table 9.3. However, the intensity of ion colour varies with its concentration in the solution. Also keep in mind that many ions are colourless in aqueous solution. For example, the cations of elements from Groups 1 (IA) and 2 (IIA), as well as aluminum, zinc, and most anions, are colourless. So there are limits to the inferences you can make if you rely on solution colour alone.

Another qualitative analysis technique is a flame test. A dissolved ionic compound is placed in a flame. Table 9.4 lists the flame colours associated with several ions. Notice that all the ions are metallic. The flame test is only useful for identifying metallic ions in aqueous solution.

Qualitative analysis challenges a chemist's creative imagination and chemical understanding. Discover this for yourself in Investigation 9-B.

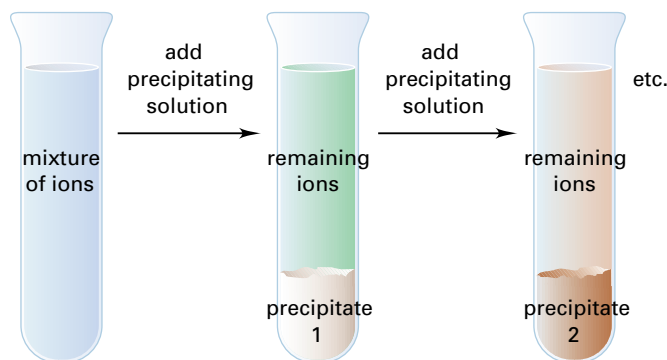
**Table 9.3** The Colour of Some Common Ions in Aqueous Solution

	Ions	Symbol	Colour
Cations	chromium (II) copper(II)	$\text{Cr}^{2+}$ $\text{Cu}^{2+}$	blue
	chromium(III) copper(I) iron(II) nickel(II)	$\text{Cr}^{3+}$ $\text{Cu}^{+}$ $\text{Fe}^{2+}$ $\text{Ni}^{2+}$	green
	iron(III)	$\text{Fe}^{3+}$	pale yellow
	cobalt(II) manganese(II)	$\text{Co}^{2+}$ $\text{Mn}^{2+}$	pink
	chromate	$\text{CrO}_4^{2-}$	yellow
	dichromate	$\text{Cr}_2\text{O}_7^{2-}$	orange
Anions	permanganate	$\text{MnO}_4^-$	purple

**Table 9.4** The Flame Colour of Selected Metallic Ions

Ion	Symbol	Colour
lithium	$\text{Li}^{+}$	red
sodium	$\text{Na}^{+}$	yellow
potassium	$\text{K}^{+}$	violet
cesium	$\text{Cs}^{+}$	violet
calcium	$\text{Ca}^{2+}$	red
strontium	$\text{Sr}^{2+}$	red
barium	$\text{Ba}^{2+}$	yellowish-green
copper	$\text{Cu}^{2+}$	bluish-green
boron	$\text{B}^{2+}$	green
lead	$\text{Pb}^{2+}$	bluish-white

**Figure 9.6** This illustration shows the basic idea behind a qualitative analysis for identifying ions in an aqueous solution. At each stage, the resulting precipitate is removed.



# Qualitative Analysis

In this investigation, you will apply your knowledge of chemical reactions and the general solubility guidelines to identify unknown ions.

## Question

How can you identify ions in solution?

## Predictions

Read the entire Procedure. Can you predict the results of any steps? Write your predictions in your notebook. Justify each prediction.

## Materials

### Part 1

12-well or 24-well plate, or spot plate  
toothpicks  
cotton swabs

unknowns: 4 dropper bottles (labelled A, B, C, and D) of solutions that include  $\text{Na}^+_{(\text{aq})}$ ,  $\text{Ag}^+_{(\text{aq})}$ ,  $\text{Ca}^{2+}_{(\text{aq})}$ , and  $\text{Cu}^{2+}_{(\text{aq})}$

reactants: 2 labelled dropper bottles, containing dilute  $\text{HCl}_{(\text{aq})}$  and dilute  $\text{H}_2\text{SO}_{4(\text{aq})}$

### Part 2

cotton swabs  
Bunsen burner  
heat-resistant pad

unknowns: 4 dropper bottles, containing the same unknowns that were used in Part 1

reactants: 4 labelled dropper bottles, containing  $\text{Na}^+_{(\text{aq})}$ ,  $\text{Ag}^+_{(\text{aq})}$ ,  $\text{Ca}^{2+}_{(\text{aq})}$ , and  $\text{Cu}^{2+}_{(\text{aq})}$

### Part 3

12-well or 24-well plate, or spot plate  
toothpicks  
cotton swabs

unknowns: 3 dropper bottles (labelled X, Y, and Z), containing solutions of  $\text{SO}_4^{2-}_{(\text{aq})}$ ,  $\text{CO}_3^{2-}_{(\text{aq})}$ , and  $\text{I}^-_{(\text{aq})}$

reactants: 3 labelled dropper bottles, containing  $\text{Ba}^{2+}_{(\text{aq})}$ ,  $\text{Ag}^+_{(\text{aq})}$ , and  $\text{HCl}_{(\text{aq})}$

## Safety Precautions



- Be careful not to contaminate the dropper bottles. The tip of a dropper should not make contact with either the plate or another solution. Put the cap back on the bottle immediately after use.
- Hydrochloric acid and sulfuric acid are corrosive. Wash any spills on your skin with plenty of cool water. Inform your teacher immediately.
- Part 2 of this investigation requires an open flame. Tie back long hair, and confine any loose clothing.

## Procedure

### Part 1 Using Acids to Identify Cations

1. Read steps 2 and 3 below. Design a suitable table for recording your observations.
2. Place one or two drops of each unknown solution into four different wells or spots. Add one or two drops of hydrochloric acid to each unknown. Record your observations.
3. Repeat step 2. This time, test each unknown solution with one or two drops of sulfuric acid. Record your observations.
4. Answer Analysis questions 1 to 5.

### Part 2 Using Flame Tests to Identify Cations

**Note:** Your teacher may demonstrate this part or provide you with an alternative version.

1. Design tables to record your observations.
2. Observe the appearance of each known solution. Record your observations. Repeat for each unknown solution. Some cations have a characteristic colour. (Refer to Table 9.4.) If you think that you can identify one of the unknowns, record your identification.

3. Flame tests can identify some cations. Set up the Bunsen burner and heat-resistant pad. Light the burner. Adjust the air supply to produce a hot flame with a blue cone.

4. Place a few drops of solution containing  $\text{Na}^+_{(\text{aq})}$  on one end of a cotton swab.

**CAUTION** Carefully hold the saturated tip so it is just in the Bunsen burner flame, near the blue cone. You may need to hold it in this position for as long as 30 s to allow the solution to vaporize and mix with the flame. Record the colour of the flame.

5. Not all cations give colour to a flame. The sodium ion *does* give a distinctive colour to a flame, however. It is often present in solutions as a contaminant. For a control, repeat step 4 with water and record your observations. You can use the other end of the swab for a second test. Dispose of used swabs in the container your teacher provides.

6. Repeat the flame test for each of the other known solutions. Then test each of the unknown solutions.

7. Answer Analysis question 6.

### Part 3 Identifying Anions

1. Place one or two drops of each unknown solution into three different wells or spots. Add one or two drops of  $\text{Ba}^{2+}_{(\text{aq})}$  to each unknown solution. Stir with a toothpick. Record your observations.

2. Add a drop of hydrochloric acid to any well or spot where you observed a precipitate in step 1. Stir and record your observations.

3. Repeat step 1, adding one or two drops of  $\text{Ag}^+_{(\text{aq})}$  to each unknown solution. Record the colour of any precipitate that forms.

4. Answer Analysis questions 7 to 9.

### Analysis

1. (a) Which of the cations you tested should form a precipitate with hydrochloric acid? Write the net ionic equation.

(b) Did your results support your predictions? Explain.

2. (a) Which cation(s) should form a precipitate when tested with sulfuric acid? Write the net ionic equation.

(b) Did your results support your predictions? Explain.

3. Which cation(s) should form a soluble chloride and a soluble sulfate?

4. Which cation has a solution that is not colourless?

5. Based on your analysis so far, tentatively identify each unknown solution.

6. Use your observations of the flame tests to confirm or refute the identifications you made in question 5. If you are not sure, check your observations and analysis with other students. If necessary, repeat some of your tests.

7. Which anion(s) should form a precipitate with  $\text{Ba}^{2+}$ ? Write the net ionic equation.

8. Which precipitate should react when hydrochloric acid is added? Give reasons for your prediction.

9. Tentatively identify each anion. Check your observations against the results you obtained when you added hydrochloric acid. Were they what you expected? If not, check your observations and analysis with other students. If necessary, repeat some of your tests.

### Conclusion

10. Identify the unknown cations and anions in this investigation. Explain why you do, or do not, have confidence in your decisions. What could you do to be more confident?



## Section Wrap-up

Qualitative analysis helps you identify ions that may be present in a solution. It does not, however, tell you how much of these ions are present. In other words, it does not provide any quantitative information about the quantity or concentration of ions in solution. In the next section, you will find out how to calculate this quantitative information, using techniques you learned in Unit 2.

## Section Review

- 1 C** Briefly compare the relationships among a chemical formula, a total ionic equation, and a net ionic equation. Use sentences or a graphic organizer.
- 2 K/U** Write a net ionic equation for each double displacement reaction in aqueous solution.
  - (a) tin(II) chloride with potassium phosphate
  - (b) nickel(II) chloride with sodium carbonate
  - (c) chromium(III) sulfate with ammonium sulfide
- 3 K/U** For each reaction in question 2, identify the spectator ions.
- 4 K/U** Would you expect a qualitative analysis of a solution to give you the amount of each ion present? Explain why or why not.
- 5 I** A solution of limewater,  $\text{Ca}(\text{OH})_{2(\text{aq})}$ , is basic. It is used to test for the presence of carbon dioxide. Carbon dioxide is weakly acidic and turns limewater milky. Use a chemical equation to explain what happens during the test. What type of reaction occurs?
- 6 K/U** State the name and formula of the precipitate that forms when aqueous solutions of copper(II) sulfate and sodium carbonate are mixed. Write the net ionic equation for the reaction. Identify the spectator ions.
- 7 I** All the solutions in this photograph have the same concentration: 0.1 mol/L. Use Table 9.3 to infer which ion causes the colour in each solution. How much confidence do you have in your inferences? What could you do to increase your confidence?



## 9.3

## Stoichiometry in Solution Chemistry

Section Preview/  
Specific Expectations

In this section, you will

- **represent** double displacement reactions by their net ionic equations
- **write** balanced chemical equations for double displacement reactions
- **solve** stoichiometry problems that involve solutions
- **describe** and **work with** solutions that have known concentrations

Recall that stoichiometry involves calculating the amounts of reactants and products in chemical reactions. If you know the atoms or ions in a formula or a reaction, you can use stoichiometry to determine the amounts of these atoms or ions that react. Solving stoichiometry problems in solution chemistry involves the same strategies you learned in Unit 2. Calculations involving solutions sometimes require a few additional steps, however. For example, if a precipitate forms, the net ionic equation may be easier to use than the chemical equation. Also, some problems may require you to calculate the amount of a reactant, given the volume and concentration of the solution.

Take your time working through the next three Sample Problems. Make sure that you understand how to arrive at the solutions. Then try the Practice Problems on page 352.

## Sample Problem

## The Concentration of Ions

## Problem

Calculate the concentration (in mol/L) of chloride ions in each solution.

- (a) 19.8 g of potassium chloride dissolved in 100 mL of solution
- (b) 26.5 g of calcium chloride dissolved in 150 mL of solution
- (c) a mixture of the two solutions in parts (a) and (b), assuming that the volumes are additive

## What Is Required?

- (a) and (b) You need to find the concentration (in mol/L) of chloride ions in two different solutions.
- (c) You need to find the concentration of chloride ions when the two solutions are mixed.

## What Is Given?

You know that 19.8 g of potassium chloride is dissolved in 100 mL of solution. You also know that 26.5 g of calcium chloride is dissolved in 150 mL of solution.

## Plan Your Strategy

- (a) and (b) For each solution, determine the molar mass. Find the amount (in mol) using the mass and the molar mass. Write equations for the dissociation of the substance. (That is, write the total ionic equation.)

Continued ...

Use the coefficients in the dissociation equation to determine the amount (in mol) of chloride ions present. Calculate the concentration (in mol/L) of chloride ions from the amount and volume of the solution.

- (c) Add the amounts of chloride ions in the two solutions to find the total. Add the volumes of the solutions to find the total volume. Calculate the concentration of chloride ions (in mol/L) using the total amount (in mol) divided by the total volume (in L).

### Act on Your Strategy

(a) and (b)

Solution	KCl	CaCl <sub>2</sub>
Molar mass	39.10 + 35.45 = 74.55 g	40.08 + (2 × 35.45) = 110.98 g
Amount (mol)	$19.8 \text{ g} \times \frac{1 \text{ mol}}{74.55 \text{ g}} = 0.266 \text{ mol}$	$26.5 \text{ g} \times \frac{1 \text{ mol}}{110.98 \text{ g}} = 0.239 \text{ mol}$
Dissociation equation	$\text{KCl}_{(s)} \rightarrow \text{K}^{+}_{(aq)} + \text{Cl}^{-}_{(aq)}$	$\text{CaCl}_{2(s)} \rightarrow \text{Ca}^{2+}_{(aq)} + 2\text{Cl}^{-}_{(aq)}$
Amount of Cl <sup>-</sup>	$0.266 \text{ mol KCl} \times \frac{1 \text{ mol Cl}^{-}}{1 \text{ mol KCl}} = 0.266 \text{ mol}$	$0.239 \text{ mol CaCl}_2 \times \frac{2 \text{ mol Cl}^{-}}{1 \text{ mol CaCl}_2} = 0.478 \text{ mol}$
Concentration of Cl <sup>-</sup>	$\frac{0.266 \text{ mol}}{0.100 \text{ L}} = 2.66 \text{ mol/L}$	$\frac{0.478 \text{ mol}}{0.150 \text{ L}} = 3.19 \text{ mol/L}$

The concentration of chloride ions when 19.8 g of potassium chloride is dissolved in 100 mL of solution is 2.66 mol/L. The concentration of chloride ions when 26.5 g of calcium chloride is dissolved in 150 mL of solution is 3.19 mol/L.

- (c) Total amount of Cl<sup>-</sup><sub>(aq)</sub> = 0.266 + 0.478 mol  
= 0.744 mol

$$\begin{aligned} \text{Total volume of solution} &= 0.100 + 0.150 \text{ L} \\ &= 0.250 \text{ L} \end{aligned}$$

$$\begin{aligned} \text{Total concentration of Cl}^{-}_{(aq)} &= \frac{0.744 \text{ mol}}{0.250 \text{ L}} \\ &= 2.98 \text{ mol/L} \end{aligned}$$

The concentration of chloride ions when the solutions are mixed is 2.98 mol/L.

### Check Your Solution

The units for amount and concentration are correct. The answers appear to be reasonable. When the solutions are mixed, the concentration of the chloride ions is not a simple average of the concentrations of the two solutions. Why? The volumes of the two solutions were different.

## Sample Problem

### The Mass Percent of Ions

#### Problem

The leaves of a rhubarb plant contain a relatively high concentration of oxalate ions,  $\text{C}_2\text{O}_4^{2-}$ . Oxalate ions are poisonous, causing respiratory failure. To determine the percent of oxalate ions, a student measured the mass of some leaves. Then the student ground up the leaves and added excess calcium chloride solution to precipitate calcium oxalate. The student tested 238.6 g of leaves. The dried mass of calcium oxalate was 0.556 g. What was the mass percent of oxalate ions in the leaves?



Although rhubarb stalks are safe to eat, the leaves are poisonous.

#### What Is Required?

You need to find the mass percent of oxalate ions in the leaves.

#### What Is Given?

You know the mass of the leaves is 238.6 g. You also know the mass of dried calcium oxalate is 0.556 g.

#### Plan Your Strategy

Determine the molar mass of calcium oxalate. Use the mass of calcium oxalate and its molar mass to find the amount (in mol) of calcium oxalate. Write the net ionic equation for the formation of calcium oxalate. From the coefficients in the net ionic equation, find the amount of oxalate ions (in mol). Calculate the mass of oxalate ions from the amount of oxalate ions (in mol) and the molar mass. Calculate the mass percent of oxalate ions in rhubarb leaves from the mass of the leaves and the mass of the oxalate ions present.

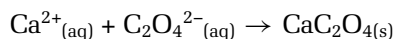
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**Act on Your Strategy**

Molar mass of  $\text{CaC}_2\text{O}_4 = 128.1 \text{ g}$

$$\begin{aligned}\text{Amount of } \text{CaC}_2\text{O}_4 &= 0.556 \text{ g} \times \frac{1 \text{ mol}}{128.1 \text{ g}} \\ &= 0.00434 \text{ mol}\end{aligned}$$

The net ionic equation is



The mole ratio of  $\text{CrO}_4^{2-}$  to  $\text{CaC}_2\text{O}_4$  is 1:1. Therefore, the leaves must have contained 0.00434 mol of  $\text{C}_2\text{O}_4^{2-}$ .

Molar mass of  $\text{C}_2\text{O}_4^{2-} = 88.02 \text{ g}$

$$\begin{aligned}\text{Mass of } \text{C}_2\text{O}_4^{2-} &= 0.00434 \text{ mol} \times \frac{88.02 \text{ g}}{1 \text{ mol}} \\ &= 0.382 \text{ g}\end{aligned}$$

$$\begin{aligned}\text{Mass percent of } \text{C}_2\text{O}_4^{2-} \text{ in the leaves} &= \frac{0.382 \text{ g}}{238.6 \text{ g}} \times 100\% \\ &= 0.160\%\end{aligned}$$

The mass percent of oxalate ions in the leaves is 0.160%.

**Check Your Solution**

Since the units divide out properly, you can be fairly confident that the answer is correct. The final value appears to be reasonable.

**PROBLEM TIP**

Review the method for solving stoichiometry problems you learned in Chapter 7, Section 7.1.

**Sample Problem****Finding the Minimum Volume to Precipitate****Problem**

Aqueous solutions that contain silver ions are usually treated with chloride ions to recover silver chloride. What is the minimum volume of 0.25 mol/L magnesium chloride,  $\text{MgCl}_{2(\text{aq})}$ , needed to precipitate all the silver ions in 60 mL of 0.30 mol/L silver nitrate,  $\text{AgNO}_{3(\text{aq})}$ ? Assume that silver chloride is completely insoluble in water.

**What Is Required?**

You need to find the minimum volume of magnesium chloride that will precipitate all the silver ions.

**What Is Given?**

You know the volumes and concentrations of the silver nitrate (volume = 60 mL; concentration = 0.30 mol/L). The concentration of the magnesium chloride solution is 0.25 mol/L.

**Plan Your Strategy**

Find the amount (in mol) of silver nitrate from the volume and concentration of solution. Write a balanced chemical

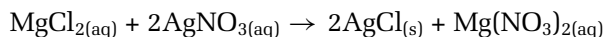
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equation for the reaction. Use mole ratios from the coefficients in the equation to determine the amount (in mol) of magnesium chloride that is needed. Use the amount (in mol) of magnesium chloride and the concentration of solution to find the volume that is needed.

### Act on Your Strategy

$$\begin{aligned}\text{Amount AgNO}_3 &= 0.060 \cancel{\text{L}} \times 0.30 \text{ mol}/\cancel{\text{L}} \\ &= 0.018 \text{ mol}\end{aligned}$$



The mole ratio of  $\text{MgCl}_2$  to  $\text{AgNO}_3$  is 1:2.

$$\begin{aligned}n \text{ mol MgCl}_2 &= 0.018 \text{ mol } \cancel{\text{AgNO}_3} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol } \cancel{\text{AgNO}_3}} \\ &= 0.0090 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Volume of } 0.25 \text{ mol/L MgCl}_2 \text{ needed} &= \frac{0.0090 \cancel{\text{mol}}}{0.25 \cancel{\text{mol/L}}} \\ &= 0.036 \text{ L}\end{aligned}$$

The minimum volume of 0.25 mol/L magnesium chloride that is needed is 36 mL.

### Check Your Solution

The answer is in millilitres, an appropriate unit of volume.  
The amount appears to be reasonable.

## Practice Problems

- Food manufacturers sometimes add calcium acetate to puddings and sweet sauces as a thickening agent. What volume of 0.500 mol/L calcium acetate,  $\text{Ca}(\text{CH}_3\text{COO})_{2(\text{aq})}$ , contains 0.300 mol of acetate ions?
- Ammonium phosphate can be used as a fertilizer. 6.0 g of ammonium phosphate is dissolved in sufficient water to produce 300 mL of solution. What are the concentrations (in mol/L) of the ammonium ions and phosphate ions present?
- An aqueous solution of a certain salt contains chloride ions. A sample of this solution was made by dissolving 17.59 g of the salt in a 1 L volumetric flask. Then 25.00 mL of the solution was treated with excess silver nitrate. The precipitate,  $\text{AgCl}_{(\text{s})}$ , was filtered and dried. If the mass of the dry precipitate was 47.35 g, what was the mass percent of chloride ions in the solution?
- The active ingredient in some rat poisons is thallium(I) sulfate,  $\text{Tl}_2\text{SO}_4$ . A chemist takes a 500 mg sample of thallium(I) sulfate and adds potassium iodide, to precipitate yellow thallium(I) iodide. When the precipitate is dried, its mass is 200 mg. What is the mass percent of  $\text{Tl}_2\text{SO}_4$  in the rat poison?

## Limiting Reactant Problems in Aqueous Solutions

In Chapter 7, you learned how to solve limiting reactant problems. You can always recognize a limiting reactant problem because you are always given the amounts of both reactants. A key step in a limiting reactant problem is determining which one of the two reactants is limiting. In aqueous solutions, this usually means finding the amount of a reactant, given the volume and concentration of the solution.

### Sample Problem

#### Finding the Mass of a Precipitated Compound

##### Problem

Mercury salts have a number of important uses in industry and in chemical analysis. Because mercury compounds are poisonous, however, the mercury ions must be removed from the waste water. Suppose that 25.00 mL of 0.085 mol/L aqueous sodium sulfide is added to 56.5 mL of 0.10 mol/L mercury(II) nitrate. What mass of mercury(II) sulfide,  $\text{HgS}_{(\text{s})}$ , precipitates?

##### What Is Required?

You need to find the mass of mercury(II) sulfide that precipitates.

##### What Is Given?

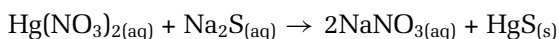
You know the volumes and concentrations of the sodium sulfide and mercury(II) nitrate solutions.

##### Plan Your Strategy

Write a balanced chemical equation for the reaction. Find the amount (in mol) of each reactant, using its volume and concentration. Identify the limiting reactant. Determine the amount (in mol) of mercury(II) sulfide that forms. Calculate the mass of mercury(II) sulfide that precipitates.

##### Act on Your Strategy

The chemical equation is



Calculate the amount (in mol) of each reactant.

$$\begin{aligned}\text{Amount of } \text{Hg}(\text{NO}_3)_2 &= 0.0565 \cancel{\text{L}} \times 0.10 \text{ mol}/\cancel{\text{L}} \\ &= 0.00565 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Amount of } \text{Na}_2\text{S} &= 0.0250 \cancel{\text{L}} \times 0.085 \text{ mol}/\cancel{\text{L}} \\ &= 0.00212 \text{ mol}\end{aligned}$$

The reactants are in a 1:1 ratio. Because  $\text{Na}_2\text{S}$  is present in the smallest amount, it is the limiting reactant.

The equation indicates that each mol of  $\text{Na}_2\text{S}$  reacts to produce the same amount of  $\text{HgS}_{(\text{s})}$  precipitate. This amount is 0.00212 mol.

##### PROBLEM TIP

Chemists solve limiting reactant problems in different ways. The method used here is different from the one you used in Chapter 7, Section 7.2

Continued ...

$$\begin{aligned}\text{Molar mass of HgS} &= 200.6 + 32.1 \\ &= 232.7\end{aligned}$$

$$\begin{aligned}\text{Mass of HgS}_{(s)} &= \text{Amount} \times \text{Molar mass} \\ &= 0.00212 \text{ mol} \times \frac{232.7 \text{ g}}{1 \text{ mol}} \\ &= 0.493 \text{ g}\end{aligned}$$

The mass of mercury(II) sulfide that precipitates is 0.49 g

### Check Your Solution

The answer has appropriate units of mass. This answer appears to be reasonable, given the values in the problem.

## Sample Problem

### Finding the Mass of Another Precipitated Compound

#### Problem

Silver chromate,  $\text{Ag}_2\text{CrO}_4$ , is insoluble. It forms a brick-red precipitate. Calculate the mass of silver chromate that forms when 50.0 mL of 0.100 mol/L silver nitrate reacts with 25.0 mL of 0.150 mol/L sodium chromate.

#### What Is Required?

You need to find the mass of silver chromate that precipitates.

#### What Is Given?

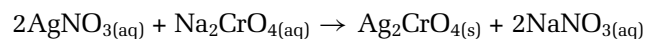
You know the volumes and concentrations of the silver nitrate and sodium chromate solutions.

#### Plan Your Strategy

Write a balanced chemical equation for the reaction. Find the amount (in mol) of each reactant, using its volume and concentration. Identify the limiting reactant. Determine the amount (in mol) of silver chromate that forms. Calculate the mass of silver chromate that precipitates.

#### Act on Your Strategy

The chemical equation is



Calculate the amount (in mol) of each reactant.

$$\begin{aligned}\text{Amount of AgNO}_3 &= 0.0500 \text{ L} \times 0.100 \text{ mol/L} \\ &= 5.00 \times 10^{-3} \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Amount of Na}_2\text{CrO}_4 &= 0.0250 \text{ L} \times 0.150 \text{ mol/L} \\ &= 3.75 \times 10^{-3} \text{ mol}\end{aligned}$$

To identify the limiting reactant, divide by the coefficient in the equation and find the smallest result.

$$\text{AgNO}_3: 5.00 \times \frac{10^{-3} \text{ mol}}{2} = 2.50 \times 10^{-3} \text{ mol}$$

$$\text{Na}_2\text{CrO}_4: 3.75 \times \frac{10^{-3} \text{ mol}}{1} = 3.75 \times 10^{-3} \text{ mol}$$

Since the smallest result is given for  $\text{AgNO}_3$ , this reactant is the limiting reactant.

Using the coefficients in the balanced equation, 2 mol of  $\text{AgNO}_3$  react for each mole of  $\text{Ag}_2\text{CrO}_4$  formed.

$$\begin{aligned} \text{Amount of Ag}_2\text{CrO}_4 &= 5.00 \times 10^{-3} \text{ mol AgNO}_3 \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{2 \text{ mol AgNO}_3} \\ &= 2.50 \times 10^{-3} \text{ mol Ag}_2\text{CrO}_4 \end{aligned}$$

The molar mass of  $\text{Ag}_2\text{CrO}_4$  is 331.7 g/mol.

$$\begin{aligned} \text{Mass of precipitate} &= 2.50 \times 10^{-3} \text{ mol} \times \frac{331.7 \text{ g}}{1 \text{ mol}} \\ &= 0.829 \text{ g} \end{aligned}$$

The mass of silver chromate that precipitates is 0.829 g.

### Check Your Solution

The answer has appropriate units of mass. The answer appears to be reasonable, given the values in the problem.

## Practice Problems

11. 8.76 g of sodium sulfide is added to 350 mL of 0.250 mol/L lead(II) nitrate solution. Calculate the maximum mass of precipitate that can form.
12. 25.0 mL of 0.400 mol/L  $\text{Pb}(\text{NO}_3)_2(\text{aq})$  is mixed with 300 mL of 0.220 mol/L  $\text{KI}(\text{aq})$ . What is the maximum mass of precipitate that can form?
13. A student mixes 15.0 mL of 0.250 mol/L aqueous sodium hydroxide with 20.0 mL of 0.400 mol/L aqueous aluminum nitrate.
  - (a) Write the chemical equation for the reaction.
  - (b) Calculate the maximum mass of precipitate that forms.

## Section Wrap-up

In the last two sections, you have used qualitative and quantitative techniques to investigate ions in aqueous solution. Every drop of water that comes into your home contains a variety of such ions. It also contains other substances, in various concentrations. You have investigated water quality in previous grades. In the next section, you will consider the chemistry of water quality.

## Section Review

- 1 **I** Equal volumes of 0.120 mol/L potassium nitrate and 0.160 mol/L iron(III) nitrate are mixed together. What is the concentration of nitrate ions in the mixture?
- 2 **I** Suppose that you want to remove the barium ions from 120 mL of 0.0500 mol/L aqueous barium nitrate solution. What is the minimum mass of sodium carbonate that you should add?
- 3 **I** An excess of aluminum foil is added to a certain volume of 0.675 mol/L aqueous copper(II) sulfate solution. The mass of solid copper that precipitates is measured and found to be 4.88 g. What was the volume of the copper(II) sulfate solution?
- 4 **I** To generate hydrogen gas, a student adds 5.77 g of mossy zinc to 80.1 mL of 4.00 mol/L hydrochloric acid in an Erlenmeyer flask. When the reaction is over, what is the concentration of aqueous zinc chloride in the flask?
- 5 **I** Copper can be recovered from scrap metal by adding sulfuric acid. Soluble copper sulfate is formed. The copper sulfate then reacts with metallic iron in a single displacement reaction. To simulate this reaction, a student places 1.942 g of iron wool in a beaker that contains 136.3 mL of 0.0750 mol/L aqueous copper(II) sulfate. What mass of copper is formed?
- 6 **I** Your stomach secretes hydrochloric acid to help you digest the food you have eaten. If too much HCl is secreted, however, you may need to take an antacid to neutralize the excess. One antacid product contains the compound magnesium hydroxide,  $\text{Mg}(\text{OH})_2$ .
  - (a) Predict the reaction that takes place when magnesium hydroxide reacts with hydrochloric acid. (**Hint:** This is a double-displacement reaction.)
  - (b) Imagine that you are a chemical analyst testing the effectiveness of antacids. If 0.10 mol/L  $\text{HCl}_{(\text{aq})}$  serves as your model for stomach acid, how many litres will react with an antacid that contains 0.10 g of magnesium hydroxide?
- 7 **I** Even though lead is toxic, many lead compounds are still used as paint pigments (colourings). What volume of 1.50 mol/L lead(II) acetate contains 0.400 mol  $\text{Pb}^{2+}$  ions.

# Aqueous Solutions and Water Quality

## 9.4

Water, as you know, is the most abundant substance on Earth's surface. Unfortunately there is more than enough water in some places, but not enough in others. As well, because water is used and re-used, and because it is such a powerful solvent, it is easily polluted.

Canada, with less than 1% of the world's population, has 22% of its fresh water. So, in most regions of our country, the *quality* of water is usually of greater concern than the quantity available. In this section, you will examine factors that affect water quality. You will also study some of the reactions that make polluted water safe and fit to drink.

### Acceptable Concentrations of Substances in Drinking Water

The federal government assesses the health risks of specific substances in drinking water. Guidelines for acceptable concentrations of each substance are then established in partnership with provincial and territorial governments. See Table 9.5. These governments require water suppliers to meet the guidelines. Corrective action may be taken when guidelines are violated.

**Table 9.5** Acceptable Concentrations of Selected Ions and Compounds in Drinking Water

Ion or compound	Maximum Acceptable Concentration (MAC) (mg/L)	Interim Maximum Acceptable Concentration (mg/L)	Aesthetic Objectives (AO) (mg/L)
aldrin and dieldrin (organic insecticides*)	0.0007		
aluminum		<0.1 or <0.2**	
arsenic		0.025	
benzene (organic component of gasoline)	0.005		
cadmium (component of batteries)	0.005		
chloride			≤ 250
fluoride	1.5		
iron			≤ 0.3
lead	0.010		
malathion (organic insecticide)	0.19		
mercury	0.001		
selenium	0.01		
sulfide (as H <sub>2</sub> S)			≤ 0.05
toluene (organic solvent)			≤ 0.024
uranium	0.1		

\*As you will learn in Unit 5, the term “organic” refers to most compounds structured around the element carbon. Toluene belongs to a large class of petroleum-related compounds called hydrocarbons.

\*\* Health-based guidelines have not yet been established. The concentrations that are listed depend on the method of treatment. They are noted as a precautionary measure.

### Section Preview/ Specific Expectations

In this section, you will

- **explain** the origins of pollutants in natural waters
- **identify** the allowed concentrations of pollutants in drinking water
- **explain** the origins and consequences of water hardness, and **outline** methods for softening hard water
- **describe** the technology involved in purifying drinking water and treating waste water
- **communicate** your understanding of the following terms: *water treatment, hard water, soft water, ion exchange, waste-water treatment*

### CHECKPOINT

You are probably familiar with the water cycle: the natural process that ensures a continual supply of Earth's water.

Sketch, from memory, a schematic drawing of the water cycle. Add labels to indicate the states and energy changes involved. Then infer the main solutes and suspended materials in water at each of the following locations:

- the atmosphere
- Earth's surface (at an urban and a rural location)
- below Earth's surface (that is, in ground water)
- a fresh-water lake
- an ocean

For each location, suggest sources of the solutes and materials you inferred.





Drinking water guidelines exist for microscopic organisms, such as bacteria and protozoans, as well as for chemicals. These guidelines were compromised in the town of Walkerton, Ontario, in 2000. Seven people died and about 2000 people became ill as a result of exposure to *E. coli* bacteria that were present in the town's water supply. The maximum acceptable concentration for *E. coli* bacteria is zero organisms detectable in 100 mL of drinking water. To achieve this target, chlorine is commonly used to kill *E. coli* and other organisms in drinking water. Despite receiving two warnings from the Ministry of the Environment, the town's water manager chose not to increase the concentration of chlorine in the water supply. As a result of this incident, the Ontario government drafted legislation to enforce strict adherence to water quality guidelines.

## Sources That Compromise Water Quality

Pure water does not exist in nature. All water naturally contains dissolved substances or ions. For example, rainwater is naturally acidic. This is because water droplets dissolve atmospheric gases, such as carbon dioxide, to form carbonic acid,  $\text{H}_2\text{CO}_3$ . As water filters through soil and rock, it tends to dissolve (leach) certain ions and compounds, such as  $\text{Ca}^{2+}_{(\text{aq})}$ ,  $\text{Mg}^{2+}_{(\text{aq})}$ ,  $\text{Fe}^{2+}_{(\text{aq})}$ ,  $\text{Fe}^{3+}_{(\text{aq})}$ , and  $\text{SO}_4^{2-}_{(\text{aq})}$ . In general, these substances pose little or no threat to plants, animals, or you. In contrast, many of the substances listed in Table 9.5 may pose a threat. They result mainly from human activities, such as manufacturing, food and materials processing, farming, and garbage disposal. The sources of these substances can be classified in the following three broad categories.

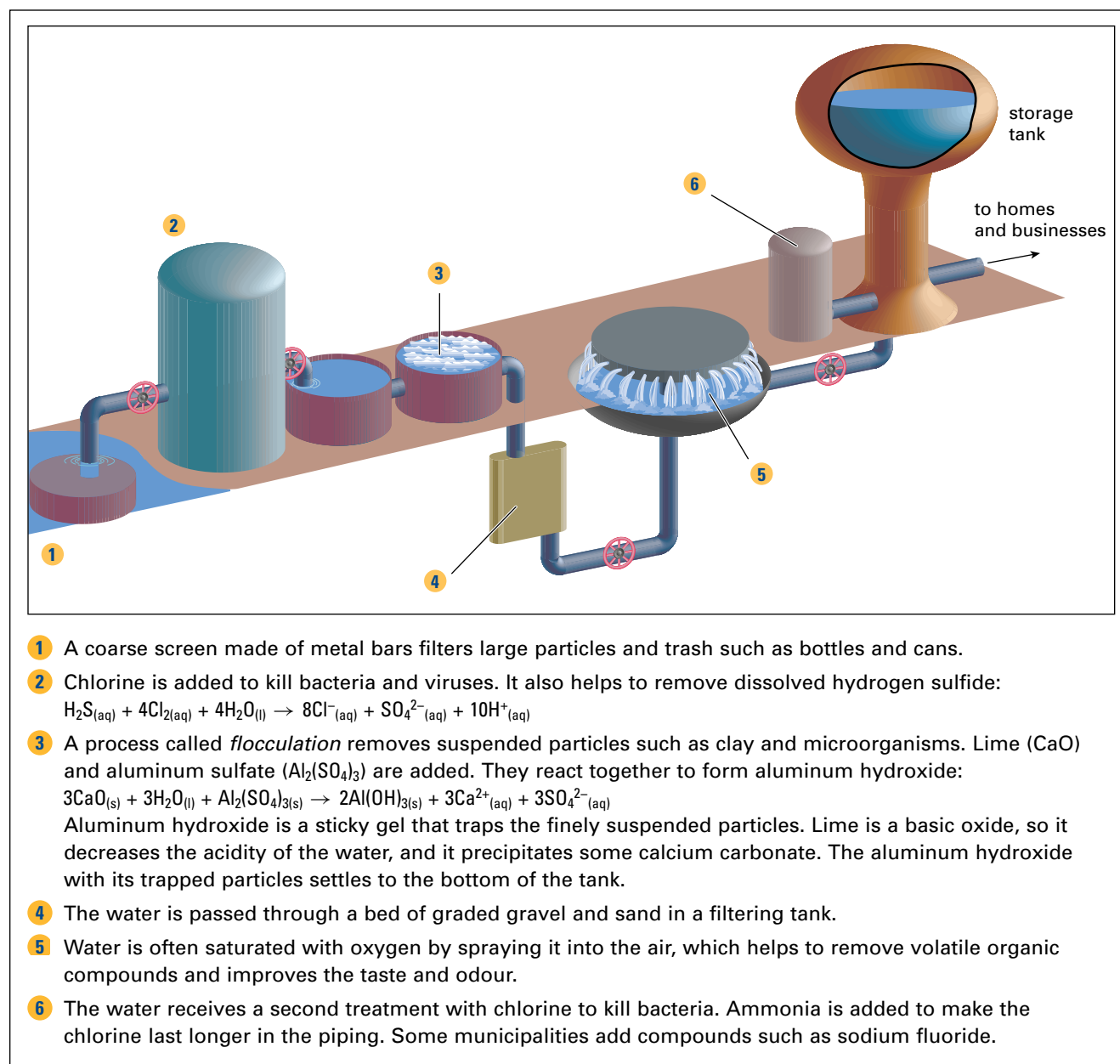
- **Point Sources:** A manufacturing or processing plant that discharges untreated or insufficiently treated waste water into a river or lake is a point source of pollution. Point sources include wrecked tankers that leak oil and factories that discharge metallic ions, organic compounds, acids, and bases. Water itself can be a point source of thermal pollution, when thermal power plants discharge warm water into a lake from their cooling towers. Point sources can spread pollution over huge areas.
- **Diffuse Sources:** Pollution that comes from a wide range of sources, not from a single source, is said to come from a diffuse source. (See Figure 9.7.) A heavy downpour can cause run-off from farm fields to enter rivers and lakes. This run-off often carries undesirable pollutants, including fecal matter, pesticides, and fertilizer compounds, such as nitrates and phosphates.
- **Indirect Sources:** Air, water, and soil can become polluted from a variety of indirect sources. Motor vehicles and factory smokestacks release gases that can indirectly cause many different types of pollution. The acidic gases that are produced (sulfur dioxide and nitrogen oxides) dissolve and contribute to the formation of acid rain. The concentrations of ions in ground water are affected by how acidic the water is. For example, aluminum compounds are commonly found in soil. Acidic water increases the leaching of aluminum ions into rivers and lakes.

**Figure 9.7** Landfill leachate is rainwater that has percolated through landfill wastes. It dissolves numerous compounds, many of them toxic. Modern landfills have containment and treatment procedures to prevent this diffuse source of pollution from contaminating ground water.



## Treating Water for Your Home

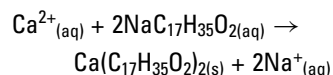
If you live on a farm or in a remote area, you probably obtain your water directly from a well on your property. If you live in an urban community, you probably obtain your water through a municipal or regional water authority. Before the water is made available to you, it is processed at a **water treatment** plant to remove pollutants. Lake, river, or reservoir water enters the treatment plant, where a number of physical and chemical processes take place. Figure 9.8 summarizes these processes.



**Figure 9.8** The following processes take place at a water treatment plant. These processes must ensure that the water meets all the allowable concentrations.



Why is hard water so hard to lather? Ordinary soap is often sodium stearate,  $\text{NaC}_{17}\text{H}_{35}\text{O}_2$ . Sodium stearate consists of the sodium salt of a compound that contains many carbon atoms. The calcium ions in hard water displace the sodium ions, and the new salts are insoluble:



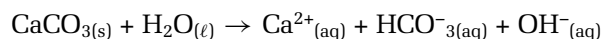
As a result, the effectiveness of the soap is reduced. The insoluble salts form a scum on clothing and a grey slimy deposit in sinks and bathtubs: “bathtub ring.”

## Hard Water and Soft Water

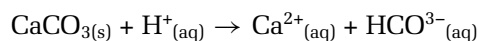
The water that flows through your faucet has been treated to remove, or limit, a large number of pollutants. It is far from pure, however. For example, it still contains dissolved ions, such as  $\text{Ca}^{2+}_{(\text{aq})}$ ,  $\text{Mg}^{2+}_{(\text{aq})}$ ,  $\text{Fe}^{2+}_{(\text{aq})}$ ,  $\text{Fe}^{3+}_{(\text{aq})}$ , and  $\text{SO}_4^{2-}_{(\text{aq})}$ . These ions, especially calcium and magnesium, make it difficult to form lather with soap. Water with high concentrations of these ions is called **hard water**, partly because it is “hard” to lather. Water with relatively low concentrations of these ions lathers well. It is called **soft water**.

Ground water is usually harder than surface water in the same region. The extent of the hardness depends on the types of rocks through which the water flows. It also depends on the length of time that the water is in contact with the rocks. The Rocky Mountains, the Canadian Shield, and most of the Maritimes have very insoluble bedrock. Thus, the water in these regions is usually soft. The sedimentary rocks of the Niagara Peninsula and the Prairies are more soluble, resulting in water that ranges from moderately hard to very hard.

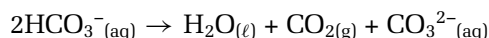
The most common type of rock to cause hard water is limestone (calcium carbonate). Limestone is usually considered insoluble. The small amount that does dissolve forms low concentrations of important ions.



When the water contains dissolved acids (often  $\text{H}_2\text{CO}_3$  from rainwater), the  $\text{H}^{+}_{(\text{aq})}$  ion increases the concentrations of calcium and hydrogen carbonate ions.

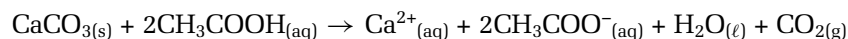


Hydrogen carbonate ions can be economically costly. Solutions that contain these ions decompose when heated to form carbonates.



The carbonate ions recombine with calcium ions to form calcium carbonate deposits. These deposits form a coating on heating elements in kettles and boilers, and build up inside hot water pipes. The coating is commonly called boiler scale. (See Figure 9.9.) It not only reduces the flow of water in pipes, but it also increases the cost of heating the water.

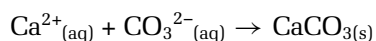
A simple way to remove boiler scale from the inside of a kettle or a coffee maker is to add vinegar. Acetic acid in vinegar reacts with calcium (and magnesium) carbonates to form soluble salts. (Remember that all acetates are soluble.)



**Figure 9.9** This photomicrograph shows the crystalline structure of precipitated boiler scale.

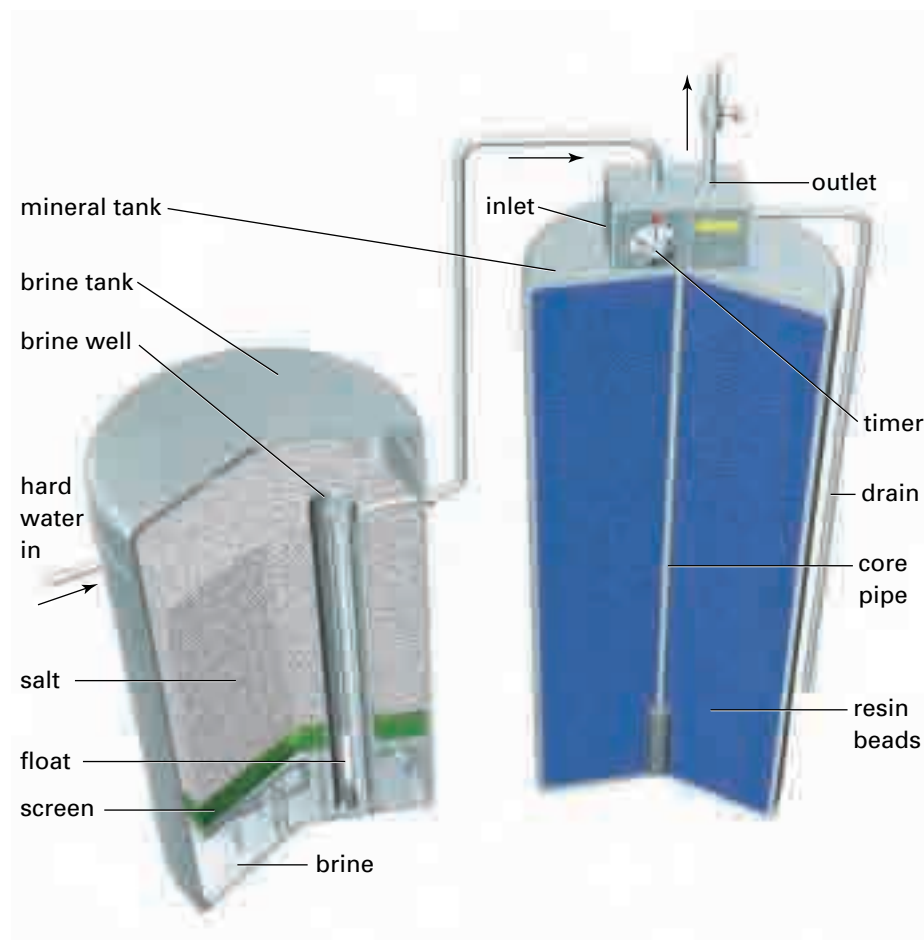
## Treating Water at Home

The ions that cause hard water are not a health hazard but they can be a nuisance. They are not always removed at municipal treatment plants. If you wish, you can remove some of these ions (mainly  $\text{Ca}^{2+}_{(\text{aq})}$ ,  $\text{Mg}^{2+}_{(\text{aq})}$ , and  $\text{Fe}^{2+}_{(\text{aq})}$ ) yourself. For relatively small volumes of water (such as enough to fill a bathtub or a washing machine), you can add sodium carbonate decahydrate,  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ . This compound is commonly called washing soda. It is an inexpensive way to add carbonate ion to the water. The carbonate ions precipitate the unwanted ions.



The sodium ions in the washing soda behave as spectator ions, leaving the water soft.

For the large volume of water needed on a daily basis, people often install an **ion exchange** water softener. This apparatus exchanges one kind of ion for another. (See Figure 9.10.) The hard water passes through a column that is packed with beads. The beads are made from an insoluble plastic material and are coated with sodium salts, often NaCl. (The salt-coated beads are referred to as an ion exchange resin.) As the hard water passes through the column, the ions in the water displace the sodium ions on the resin. After most of the sodium ions have been exchanged for calcium ions (and other hardness-causing ions), the resin is regenerated. This is done by passing a very concentrated solution of sodium chloride (brine) through the column. The calcium ions are flushed out of the system, along with excess sodium chloride solution.



## Biology

## LINK

A vivid red dye is produced from a small insect, called the cochineal, that lives in the Peruvian Andes. Two chemists from Simon Fraser University in British Columbia, Dr. Cam Oehlschlager and Dr. Eva Czyzewska, have developed an improved process for extracting the natural red colouring agent from the insects. The ions that cause hard water change the colour of the dye, however. So the first step in the process uses demineralized hot water. After precipitating the dye, it is dried and used for a variety of products. Some of your classmates may be wearing examples of these products right now. Use the library or the Internet to find out what they are.



### Electronic Learning Partner

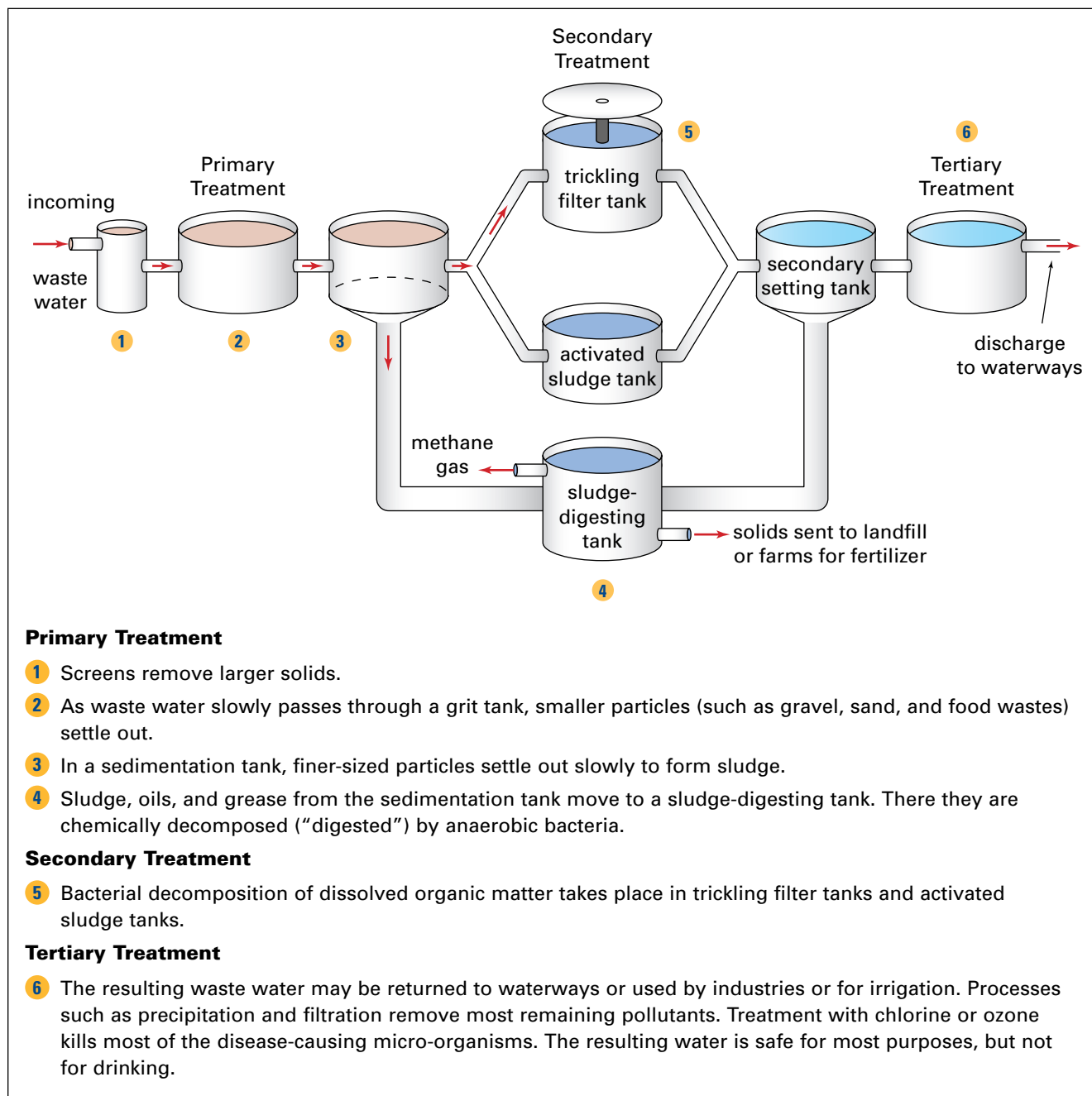
Your Chemistry 11 Electronic Learning Partner has a movie detailing several methods for purifying water.

**Figure 9.10** An ion exchange water softener adds significant amounts of sodium to the water. Thus, it may not be the best choice for drinking water. This is especially true for people on sodium-restricted diets.

## Waste-Water Treatment

The treatment of waste water (sewage) is often divided into three types: primary, secondary, and tertiary treatment. (See Figure 9.11.)

- Primary treatment mainly involves removing solids from waste water physically, using filters and settling tanks.
- Secondary treatment involves using bacteria to chemically decompose dissolved and suspended organic compounds.
- Tertiary treatment involves chemical treatments to remove the majority of remaining ions and disease-causing micro-organisms.



**Figure 9.11** The main steps involved in the treatment of waste water: Municipalities may use one, two, or all three treatment methods, depending on their needs and finances.





Water is hard if it contains high concentrations of calcium and magnesium ions,  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$ . If these ions are present in lower concentrations, the water is considered to be soft. Distilled water has few ions of any kind, and no  $\text{Ca}^{2+}$  or  $\text{Mg}^{2+}$ . Sodium oxalate,  $\text{Na}_2\text{C}_2\text{O}_4$ , is a compound that causes  $\text{Ca}^{2+}$  to precipitate as calcium oxalate,  $\text{CaC}_2\text{O}_4$ , and  $\text{Mg}^{2+}$  to precipitate as magnesium oxalate,  $\text{MgC}_2\text{O}_4$ .

### Procedure

1. A student added 1 mL of hard water to one test tube, 1 mL of soft water to another, and 1 mL of distilled water to a third. The hard water sample contained  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$ . The soft water contained lower concentrations of these ions.
2. Into each test tube, the student put two drops of 0.1 mol/L sodium oxalate solution. Then the student mixed the contents.

### Analysis

1. Infer what the student observed in each test tube. Write a net ionic equation if you predict that a precipitate formed. Write “NR” if you think that no reaction occurred.
2. Imagine that you have the three water samples the student used, but no sodium oxalate solution. How else could you test the validity of your predictions?
3. With your teacher’s permission, test your predictions. **CAUTION** Sodium oxalate is poisonous. Wear an apron and safety goggles. Handle all the solutions carefully.



## Dr. Jiangning Wu: Cleaning the World’s Water



Around the world, researchers are working to reduce the pollutants in our lakes, rivers, and ground water. Both industry and agriculture too often produce waste water that contains everything from animal droppings and excess fertilizer to artificial colours and other chemicals. Yet, without industry and agriculture, we would not have many of the modern comforts we expect and enjoy.

Jiangning Wu was born and raised in Nanjing, China. She came to Canada to complete her studies in chemical engineering at the University of Windsor. After earning her doctorate in 1992, she turned her attention to the use of enzymes for purifying waste water.

Enzymes are naturally occurring substances that speed up chemical reactions without being used up or changed by the reactions. They are basic to living organisms. Thus it is no surprise that they are well-suited to removing organic pollutants from waste water. For instance, an enzyme that is found in horseradish helps to accelerate the oxidation of soluble organic pollutants known as phenols. The resulting compounds are less soluble and more easily removed from the waste water.

In 1993, Dr. Wu found that, with a single additive, the same enzyme was able to remove phenols from waste water. Her strategy worked even for the low concentrations of phenols typical of industrial waste water. Moreover, significantly lower concentrations of the enzyme were needed! Dr. Wu’s discovery brought researchers one step closer to a commercial process, which would mean cleaner waters and less stress on the environment.

Dr. Wu is a professor in the School of Chemical Engineering at Ryerson Polytechnical University in Toronto. She is currently investigating the use of ozone for waste-water treatment and food preservation.



## Section Wrap-up

In this section, you applied your understanding of aqueous solutions to explore the chemistry of water quality. Some of the chemicals you considered belong to a group of compounds that are called acids and bases. In the next chapter, you will investigate the properties and chemical behaviour of these important compounds.



### CHEM

### FACT

Chlorine was first used to disinfect water in Britain in 1904, after a typhoid epidemic. (Typhoid is a water-borne, contagious illness that is caused by a species of *Salmonella* bacteria.) Strict limits are necessary because chlorine is ineffective when its concentration is less than 0.1 mg/L. It gives water an unpleasant taste at concentrations above 1.0 mg/L. Chlorine has a disadvantage, however. It can react with other chemicals in the water to form poisonous compounds, such as chloroform,  $\text{CHCl}_3$ . These chemicals may remain in solution even after the entire treatment process.

## Section Review

- 1** **K/U** In what ways can water in the environment become polluted? Give at least two examples for each.
- 2** **K/U** Which chemical in Table 9.5 has the lowest acceptable concentration? Which chemical has the highest acceptable concentration? Rearrange the chemicals in the table so they are organized by concentration, from lowest to highest, rather than alphabetically by name.
- 3** **MC** What is the source of the water you use at home? Based on your experience with this water, which ions do you think it contains? Explain your answer.
- 4** **C** Use a graphic organizer to outline the main steps involved in treating drinking water and waste water.
- 5** **MC** An alternative to ion exchange water softening is a process called *reverse osmosis*. Use your knowledge of osmosis to infer, in general terms, how this method might work. Consult print or electronic resources to modify or expand on your ideas. Then use suitable software to communicate your findings as a graphic organizer or a virtual slide presentation.
- 6** **MC** One component of the waste water that enters a waste water plant for treatment is urea,  $(\text{NH}_2)_2\text{CO}$ .
  - (a)** Use reference materials to find out how urea is chemically changed to nitrate ions. (There are several steps.)
  - (b)** Write chemical equations to represent the steps involved in this process.
  - (c)** What happens to aqueous nitrates when they are discharged to waterways in the environment?

### Reflecting on Chapter 9

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Predict combinations of aqueous solutions that result in the formation of precipitates.
- Describe your experiences with a qualitative analysis of ions in solution.
- Represent a double-displacement reaction using its net ionic equation.
- Write balanced chemical equations and net ionic equations for double-displacement reactions.
- Apply your understanding of stoichiometry to solve quantitative problems involving solutions.
- Identify the origins of pollutants in drinking water, and the allowed concentrations of some of these pollutants.
- Examine the causes and effects of water hardness. Considered several methods for softening hard water.
- Compare the chemistry and the technology of water treatment and waste-water treatment.

### Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

general solubility guidelines

hard water

net ionic equation

qualitative analysis

spectator ions

waste-water treatment

ion exchange

precipitate

soft water

total ionic equation

water treatment

### Knowledge/Understanding

- In your own words, define the terms “spectator ion” and “net ionic equation.”
- Identify the spectator ions in the following skeleton equation. Then write the balanced ionic equation for the reaction.  

$$\text{Al}(\text{NO}_3)_3(\text{aq}) + \text{NH}_4\text{OH}(\text{aq}) \rightarrow$$

$$\text{Al}(\text{OH})_3(\text{s}) + \text{NH}_4\text{NO}_3(\text{aq})$$
- Hydrogen sulfide gas can be prepared by the reaction of sulfuric acid with sodium sulfide.  

$$\text{H}_2\text{SO}_4(\text{aq}) + \text{Na}_2\text{S}(\text{aq}) \rightarrow \text{H}_2\text{S}(\text{g}) + \text{Na}_2\text{SO}_4(\text{aq})$$

Write the net ionic equation for this reaction.
- Each of the following combinations of reagents results in a double displacement reaction. In your notebook, complete the chemical equation. Then identify the spectator ions, and write the net ionic equation.
  - copper(II) chloride<sub>(aq)</sub> + ammonium phosphate<sub>(aq)</sub>  $\rightarrow$
  - aluminum nitrate<sub>(aq)</sub> + barium hydroxide<sub>(aq)</sub>  $\rightarrow$
  - sodium hydroxide<sub>(aq)</sub> + magnesium chloride<sub>(aq)</sub>  $\rightarrow$
- Use the general solubility guidelines to name three reagents that will combine with each ion below to form a precipitate. Assume that the reactions take place in aqueous solution. For each reaction, write the net ionic equation.
  - bromide ion
  - carbonate ion
  - lead(II) ion
  - iron(III) ion
- The transition metals form insoluble sulfides, often with a characteristic colour. Write the net ionic equation for the precipitation of each ion by the addition of an aqueous solution of sodium sulfide.
  - $\text{Cr}^{3+}(\text{aq})$  (**Note:**  $\text{Cr}_2\text{S}_3(\text{s})$  is brown-black.)
  - $\text{Ni}^{2+}(\text{aq})$  (**Note:**  $\text{NiS}(\text{s})$  is black.)
  - $\text{Mn}^{4+}(\text{aq})$  (**Note:**  $\text{MnS}_2(\text{s})$  is green or red, depending on the arrangement of ions in the solid.)
- Identify three cations and three anions that are commonly found in ground water. Suggest at least one likely source for each.
- Briefly describe two steps in the primary treatment of waste water, one involving a physical change and the other involving a chemical change.
- Many liquid antacids contain magnesium hydroxide,  $\text{Mg}(\text{OH})_2$ . Why must the bottle be shaken before a dose is poured?
  - Stomach acid contains hydrochloric acid. Excess acid that backs up into the esophagus is the cause of “heartburn.” Write the chemical equation and the net ionic equation for the reaction that takes place when someone with heartburn swallows a dose of liquid antacid.

10. Aqueous solutions of iron(III) chloride and ammonium sulfide react in a double displacement reaction.
- Write the name and formula of the substance that precipitates.
  - Write the chemical equation for the reaction.
  - Write the net ionic equation.

## Inquiry

11. A reference book states that the solubility of silver sulfate is 0.57 g in 100 mL of cold water. You decide to check this by measuring the mass of a silver salt precipitated from a known volume of saturated silver sulfate solution. Solubility data show that silver chloride is much less soluble than silver nitrate. Explain why you should not use barium chloride to precipitate the silver ions. Suggest a different reagent, and write the net ionic equation for the reaction.
12. The presence of copper(II) ions in solution can be tested by adding an aqueous solution of sodium sulfide. The appearance of a black precipitate indicates that the test is positive. A solution of copper(II) bromide is tested this way. What precipitate is formed? Write the net ionic equation for the reaction.
13. An old home-gardening “recipe” for fertilizer suggests adding 15 g of Epsom salts (magnesium sulfate heptahydrate,  $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ ) to 4 L of water. What will be the concentration of magnesium ions?
14. Calculate the concentration (in mol/L) of each aqueous solution.
- 7.37 g of table sugar,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ , dissolved in 125 mL of solution
  - 15.5 g of ammonium phosphate,  $(\text{NH}_4)_3\text{PO}_4$ , dissolved in 180 mL of solution
  - 76.7 g of glycerol,  $\text{C}_3\text{H}_8\text{O}_3$ , dissolved in 1.20 L of solution
15. 50.0 mL of 0.200 mol/L  $\text{Ca}(\text{NO}_3)_2(\text{aq})$  is mixed with 200 mL of 0.180 mol/L  $\text{K}_2\text{SO}_4(\text{aq})$ . What is the concentration of sulfate ions in the final solution?
16. Suppose that 1.00 L of 0.200 mol/L  $\text{KNO}_3(\text{aq})$  is mixed with 2.00 L of 0.100 mol/L  $\text{Ca}(\text{NO}_3)_2(\text{aq})$ . Determine the concentrations of the major ions in the solution.
17. Equal masses of each of the following salts are dissolved in equal volumes of water: sodium chloride, calcium chloride, and iron(III) chloride. Which salt produces the largest concentration of chloride ions?
18. Imagine that you are the chemist at a cement factory. You are responsible for analyzing the factory’s waste water. If a 50.0 mL sample of waste water contains 0.090 g of  $\text{Ca}^{2+}(\text{aq})$  and 0.029 g of  $\text{Mg}^{2+}(\text{aq})$ , calculate
- the concentration of each ion in mol/L
  - the concentration of each ion in ppm
19. The concentration of calcium ions,  $\text{Ca}^{2+}$ , in blood plasma is about  $2.5 \times 10^{-3}$  mol/L. Calcium ions are important in muscle contraction and in regulating heartbeat. If the concentration of calcium ions falls too low, death is inevitable. In a television drama, a patient is brought to hospital after being accidentally splashed with hydrofluoric acid. The acid readily penetrates the skin, and the fluoride ions combine with the calcium ions in the blood. If the patient’s volume of blood plasma is 2.8 L, what amount (in mol) of fluoride ions would completely combine with all the calcium ions in the patient’s blood?
20. A double displacement reaction occurs in aqueous solution when magnesium phosphate reacts with lead(II) nitrate. If 20.0 mL of 0.750 mol/L magnesium phosphate reacts, what is the maximum mass of precipitate that can be formed?

## Communication

21. Phosphate ions act as a fertilizer. They promote the growth of algae in rivers and lakes. They can enter rivers and lakes from fields that are improperly fertilized or from untreated waste water that contains phosphate detergents. How can the water be treated to remove the phosphate ions?
22. A chemist analyzes the sulfate salt of an unknown alkaline earth metal. The chemist adds 1.273 g of the salt to excess barium chloride solution. After filtering and drying, the mass of precipitate is found to be 2.468 g.
- Use the formula  $\text{MSO}_4$  to represent the unknown salt. Write the molecular and net ionic equations for the reaction.

- (b) Calculate the amount (in mol) of  $\text{MSO}_4$  used in the reaction.
- (c) Determine the molar mass of the unknown salt.
- (d) What is the likely identity of the unknown metal cation? What test might the chemist perform to help confirm this conclusion?
23. The same volume of solution is made using the same masses of two salts: rubidium carbonate and calcium carbonate. Which salt gives the larger concentration of aqueous carbonate ions?
24. A prospector asks you to analyze a bag of silver ore. You measure the mass of the ore and add excess nitric acid to it. Then you add excess sodium chloride solution. You filter and dry the precipitate. The mass of the ore is 856.1 g, and 1.092 g of silver chloride is collected.
- (a) Why did you first treat the ore with excess nitric acid?
- (b) Calculate the mass percent of silver in the ore. The ore that is extracted at a silver mine typically contains about 0.085% silver by mass. Should the prospector keep looking or begin celebrating?

## Making Connections

25. Think about your activities yesterday. Which activity required the most use of water? Estimate the volume of water you used. Would it make sense to have two supplies of water to your home, one for drinking and a second of lower purity for every other activity that uses water? Give reasons for your answer.
26. List three different household wastes that are commonly discarded and have the potential to contaminate ground water if rain leaches through your local landfill site. What chemical(s) does each contain? For each waste, identify an alternative to dumping it in a landfill site.
27. Water is essential for crops. Improper irrigation over a number of years, however, can result in farmed land becoming laden with toxic chemical compounds. Research how this happens. Find out whether this is a concern to farmers near where you live.

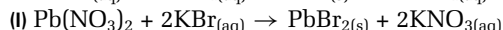
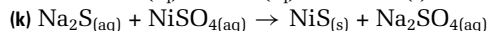
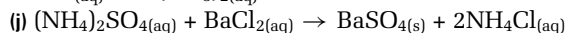
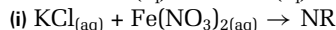
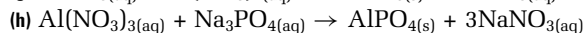
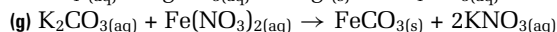
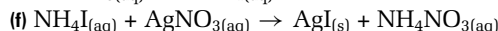
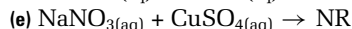
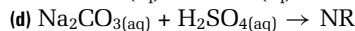
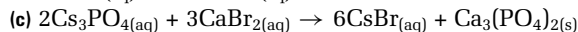
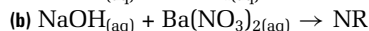
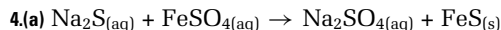
## Answers to Practice Problems and

### Short Answers to Section Review Questions:

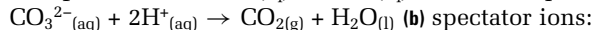
**Practice Problems:** 1.(a) insoluble (b) insoluble

(c) soluble 2.(a) soluble (b) soluble (c) insoluble

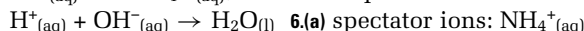
3.(a) insoluble (b) soluble (c) insoluble



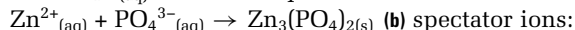
5.(a) spectator ions:  $\text{Na}^+_{(\text{aq})}$  and  $\text{Cl}^-_{(\text{aq})}$ ; net ionic equation:



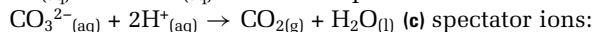
$\text{Na}^+_{(\text{aq})}$  and  $\text{SO}_4^{2-}_{(\text{aq})}$ ; net ionic equation:



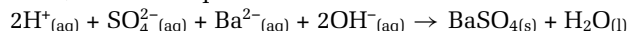
and  $\text{SO}_4^{2-}_{(\text{aq})}$ ; net ionic equation:



$\text{Li}^+_{(\text{aq})}$  and  $\text{NO}_3^-_{(\text{aq})}$ ; net ionic equation:



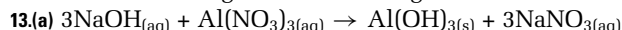
none; net ionic equation:



7. 300 mL 8. concentration of  $\text{NH}_4^+_{(\text{aq})}$  is 0.40 mol/L;

concentration of  $\text{PO}_4^{3-}_{(\text{aq})}$  is 0.13 mol/L 9. 66.57%

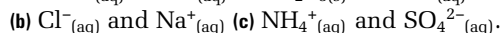
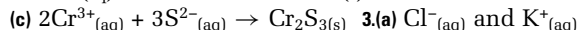
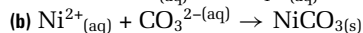
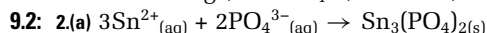
10. 30.47% 11. 20.9 g of PbS 12. 4.61 g PbI<sub>2</sub>



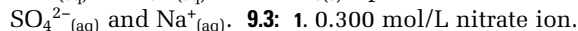
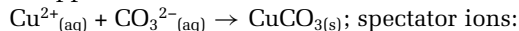
(b) 0.0975 g

**Section Review: 9.1:** 2. NaF less soluble, because  $\text{F}^-$  is smaller than  $\text{I}^-$ . 3.(a) insoluble (b) soluble (c) soluble.

4. all insoluble. 5. Any reagent containing  $\text{Cl}^-$ ,  $\text{Br}^-$ , or  $\text{I}^-$  will precipitate silver ion but leave calcium ion in solution. 6. 1 =  $\text{Ag}^+$ , 2 =  $\text{SO}_4^{2+}$ , 3 =  $\text{Ba}^{2+}$ , 4 =  $\text{Cl}^-$ .



6. copper(II) carbonate,  $\text{CuCO}_3$ ;



2. 0.636 g  $\text{Na}_2\text{CO}_{3(\text{s})}$ . 3. 114 mL. 4. 1.09 mol/L  $\text{ZnCl}_2$ .

5. 0.650 g Cu 6.  $3.4 \times 10^{-2}$  L 7. 0.267 L



