Chapter

7

Solubility and Reactions

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

Solutions make it easy to

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

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

In this chapter we will examine several concepts, including the extent to which one substance will dissolve in another, and the effect of temperature on the extent of dissolving. In chemical reactions we often wish to produce a product that is more or less soluble than the reactants. In fact, this is one of the most common techniques used for separating chemical substances. This chapter will help you understand how to do this.

Reflect on your Learning

1. A liquid you are using at home or in the laboratory may be a solution or a pure substance. Which is most likely? How would you test the liquid to determine whether it is a solution or a pure substance?

2. When dissolving a chemical in water, the rate of dissolving is often confused with the extent of dissolving. How do you speed up the rate of dissolving of a solute? How do you know if no more solute will dissolve?

3. Reactions in solution are common. If two aqueous solutions are mixed, what evidence would indicate that a reaction has occurred? How would you know if the change was just a physical change rather than a chemical change?

Try This
Activity

Measuring the Dissolving Process

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

Materials: distilled or deionized water, table salt, coarse pickling salt (pure $NaCl_{(s)}$), a measuring teaspoon (5 mL), two 250-mL Erlenmeyer flasks, with stoppers, one 100-mL graduated cylinder or measuring cup

- Place a level teaspoonful of table salt into 100 mL of pure water at room temperature in a 250-mL Erlenmeyer flask. Swirl the flask's contents thoroughly for a minute or two. Record your observations.
- Repeat with pickling salt, again recording your observations.

  (a) What does the result, with common table salt as a solute, show about the nature of the substance being used? Compare it with the solution in the second flask.
  (b) List the ingredients in common table salt, according to the package label, and explain your observations of the contents of the first flask.

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

  (c) How many level teaspoons of pickling salt (pure $NaCl_{(s)}$) could you get to dissolve in 100 mL of $H_2O_{(l)}$ in the second flask?
  (d) What is the final volume of your $NaCl_{(aq)}$ solution in the second flask?
  (e) If you dissolve 20.0 mL of $NaCl_{(s)}$ in 100.0 mL of liquid water, what do you suppose the volume of the solution would be? Can you think of a way to test this? The answer is very interesting.

Figure 1
The excess of solid solute in the mixture is visible evidence for a saturated solution.

INQUIRY SKILLS

○ Questioning   ● Recording
○ Hypothesizing   ● Analyzing
○ Predicting   ● Evaluating
○ Planning   ● Communicating
● Conducting

7.1   Solubility

When you add a small amount of pickling salt (pure sodium chloride) to a jar of water and shake the jar, the salt dissolves and disappears completely. What happens if you continue adding salt and shaking? Eventually, some visible solid salt crystals will remain at the bottom of the jar, despite your efforts to make them dissolve. You have formed a **saturated solution**—a solution in which no more solute will dissolve. We say it is at maximum solute concentration. If the container is sealed, and the temperature stays the same, no further changes will ever occur in the concentration of this solution. The quantity (mass) of solute that remains undissolved will also stay the same. The **solubility** of sodium chloride in water is the concentration of your saturated solution. The units for solubility and maximum concentration are therefore the same: usually grams of solute per 100 mL of solvent. You will learn in this chapter that solubility depends on the temperature, so it is a *particular* maximum concentration value. Every solubility value must be accompanied by a temperature value.

When calculating and using solubility values we have to make one assumption: the solute is not reacting with the solvent.

Solubility of Solids

Every pure substance has its own unique solubility. For example, we can find from a reference source, such as the *CRC Handbook of Chemistry and Physics*, that the solubility of sodium sulfate in water at 0°C is 4.76 g/100 mL. Remember that this means 4.76 g of solute can be dissolved in 100 mL of water; *not* that you will have 100 mL of solution after dissolving 4.76 g of solute. If more than 4.76 g of this solute is added to 100 mL of water in the container, the excess will not dissolve under the specified conditions (**Figure 1**). The quickest way to see whether you have a saturated solution is to look for the presence of undissolved solids in the solution. There are several experimental designs that can be used to determine the solubility of a solid. For example, the solvent from a measured volume of saturated solution might be removed by evaporation, leaving the crystallized solid solute behind—which can then be collected and measured.

Investigation 7.1.1

Solubility Curve of a Solid

Because the solubility of a substance changes with the temperature of the saturated solution, it is sometimes useful to plot graphs of the relationship between these two variables. A graph gives an instant visual picture of the relationship and can then be used to determine the solubility at any temperature. A graph of solubility and temperature of the solution is called a *solubility curve*.

In this investigation, you will measure the temperature at which the solute precipitates out of solution, indicating that the solute has reached its maximum concentration. You will do this for several different concentrations. You will then use your evidence to create a solubility curve. Complete the **Analysis** and **Evaluation** sections of the lab report.

Question

What is the relationship between the solubility of potassium nitrate and the temperature of its solution?

Experimental Design

A known mass of potassium nitrate is dissolved in warm water. As the solution cools, the temperature is recorded when the *first* sign of crystal formation occurs. This Procedure is repeated several times with different volumes of water (i.e., different solution concentrations).

(a) Identify dependent, independent, and controlled variables in this design.

Materials

eye protection	medicine dropper
laboratory apron	laboratory scoop
oven mitts	small test tube
250-mL beaker	10-mL glass graduated cylinder
thermometer	pure (distilled) water
centigram balance	$KNO_{3(s)}$ (about 3 g)
hot plate	

Be careful with the hot plate; it looks the same whether hot or cold. Use oven mitts when handling hot apparatus.

Wear eye protection and a laboratory apron.

Procedure

1. Set up a hot-water bath by placing the 250-mL beaker, about two-thirds full of water, on the hot plate. Heat the water to about 85°C.

2. Add about 3 g of $KNO_{3(s)}$ to a clean, dry test tube and record the actual mass used.

3. Add 10.0 mL pure water to the graduated cylinder.

4. Use the medicine dropper to remove 2.0 mL of pure water from the graduated cylinder. Transfer the 2.0 mL of water to the test tube.

5. Insert the thermometer into the test tube and then place the test tube in the hot-water bath.

6. Carefully stir the contents of the test tube with the thermometer until the solid has completely dissolved. Do not leave the test tube in the water bath any longer than necessary to dissolve the solid.

7. Remove the test tube from the bath, stir carefully, and record the temperature when the *first crystals appear*. It may help to hold the test tube up to a light.

8. Repeat steps 4 to 7 four more times with the same test tube, adding 1.0 mL more water (instead of 2.0 mL) each time.

9. Empty the contents of the test tube into the labelled beaker provided to recycle the potassium nitrate.

Analysis

(b) For each trial, calculate the solubility in grams of $KNO_{3(s)}$ per 100 mL of water.
(c) Plot a graph of solubility (in g/100 mL) against temperature.
(d) Write a sentence to answer the Question: What is the relationship between the solubility of potassium nitrate and the temperature of its solution?

Evaluation

(e) Evaluate the Evidence by judging the Experimental Design, Materials, Procedure, and your skills. Identify sources of experimental uncertainty or error. Note any flaws and potential improvements.
(f) Suggest a different experimental design that could be used to determine the solubility of a solid such as potassium nitrate.

Synthesis

(g) What property of the solution is illustrated at the moment the crystals start forming? Explain briefly.

(h) According to your graph, which of the following mixtures is a saturated solution and which is an unsaturated solution?

    (i)   100 g of $KNO_{3(s)}$ in 100 mL of $H_2O_{(l)}$ at 40°C

    (ii) 50 g of $KNO_{3(s)}$ in 100 mL of $H_2O_{(l)}$ at 70°C

    (iii) 120 g of $KNO_{3(s)}$ in 200 mL of $H_2O_{(l)}$ at 60°C

(i) According to your graph, what mass of potassium nitrate will dissolve in 100 mL of water at

    (i) 20°C?   (ii) 50°C?

Solubility Curves

Graphs of solubility (maximum concentration) against temperature allow quick and easy reference, and are very useful for a wide variety of questions involving solution concentrations. **Figure 2** shows some solubility curves for ionic compounds in aqueous solution.

Solubility of Some Ionic Compounds in Water
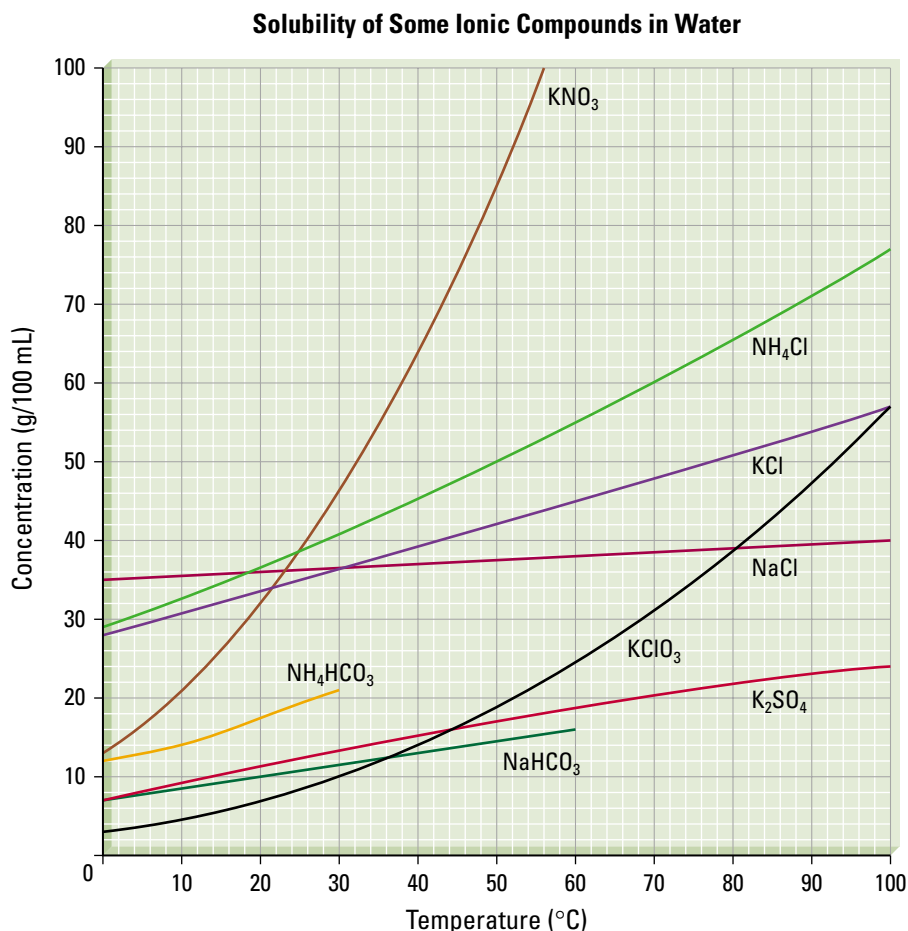
Figure 2
The solubility of ionic compounds in water is related to temperature.

Practice

Understanding Concepts

1. Refer to **Figure 2**.

    (a) What mass of K_2SO_4 can be dissolved in 100 mL of water at 70°C?

(b) At what temperature are the solubilities of KNO_3 and KCl approximately equal?

(c) Why is it not possible to determine the molar concentration of a saturated solution of $KClO_3$ at 25°C from its solubility curve? What additional information would be required?

(d) Which substance forms a saturated solution in which the mass ratio of solute to solvent is 1:1, and at what temperature?

2. A solution containing equal masses of NaCL and KCL is slowly cooled.

(a) At what temperatures might sodium chloride precipitate first, without precipitating potassium chloride?

(b) At what temperatures might potassium chloride precipitate first, without precipitating sodium chloride?

Applying Inquiry Skills

3. If you completed Investigation 7.1.1, you generated a solubility curve for $KNO_{3(aq)}$. Compare your curve to the curve in **Figure 2**. Are there differences? How would you account for any differences?

Solubility of Gases

Although we are not so aware of gases dissolving in liquids, they do indeed do so. Swimming pools contain dissolved chlorine, rivers and streams contain dissolved oxygen, and cans of pop contain dissolved carbon dioxide.

Try This
Activity

Gas Solubility

This activity illustrates the various effects of temperature (and even some effects of pressure) on solubility. It also highlights an interesting point: air dissolves in water, and water dissolves in air!

Materials: clear, colourless container (e.g., beaker or drinking glass), a long match, any clear, light-coloured carbonated beverage

• Open the carbonated beverage, and fill your container half full. Observe the rapidity of formation, source, and size of the bubbles.

• Let the bubbling drink stand a few moments in a draft-free area, then light a match and lower it into the top half of the container. Observe what happens to the flame.

(a) Describe and explain your observation.

• Empty and rinse your container.

• Run a tap until the water is quite hot and then fill your container. Normally, the water will at first be clouded or milky in appearance. Let it stand for a minute or two as you observe closely what happens. Record your observations. Dispose of the hot water.

• Run a tap until the water is very cold and then fill your container. Normally, the water will at first be nicely clear in appearance. Let it stand for a few minutes as you observe closely what happens to the outside of the container. Record your observations.

(b) In terms of solubility, explain what causes the condensation here, and describe how you could test qualitatively what the liquid is.

• Let your container of cold water stand for another half-hour. Occasionally observe and record what happens inside of container.

(c) In terms of solubility, explain your observations of the container of cold water. What causes the bubbles to form? Discuss what might cause the bubbles to be larger than they were in the hot water.

INQUIRY SKILLS

○ Questioning   ○ Recording
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○ Planning   ○ Communicating
○ Conducting

Lab Exercise 7.1.1

Solubility of a Gas

We have seen that the solubility of a solid increases with increasing temperature. How does the solubility of a gas change as the temperature changes? Using the evidence provided in this lab exercise, create a generalization for the effect of temperature on the solubility of a gas. Complete the **Experimental Design, Analysis,** and **Synthesis** sections of the report.

Question

What effect does temperature have on the solubility of various gases in water?

Experimental Design

(a) Based on **Table 1**, identify the independent, dependent, and controlled variables.

Evidence

Table 1: Solubility of Gases at Different Temperatures

Gas	Solubility (g/100 mL) at specified temperature		
	0°C	20°C	50°C
$N_{2(g)}$	0.0029	0.0019	0.0012
$O_{2(g)}$	0.0069	0.0043	0.0027
$CO_{2(g)}$	0.335	0.169	0.076
$NH_{3(g)}$	89.9	51.8	28.4

Analysis

(b) Use the Evidence to answer the Question. Write your response as a generalization.

Synthesis

(c) What assumption must we make about the dissolving of a gas in water for the solubility evidence to be valid?

(d) Use your knowledge of the kinetic molecular theory to explain the generalization for the solubility of gas as the temperature increases.

Practice

Understanding Concepts

4. When you open a can of pop, which is more likely to fizz and spray: a can at room temperature or a cold can from the refrigerator? Explain why using your understanding of gas solubility.

Reflecting

5. (a) The solubility of oxygen in blood is much greater than its solubility in pure water. Suggest a reason for this observation.

   (b) If the solubility of the oxygen in blood were the same as in pure water, how would your life be different?

Solubility in Water—Generalizations and Examples

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

- Solids usually have higher solubility in water at higher temperatures. **Figure 2** shows the solubility curves of many ionic compounds. This trend is generally the same for soluble molecular compounds. For example, sucrose has a solubility of about 180 g/100 mL at 0°C and 487 g/100 mL at 100°C.
- Gases always have higher solubility in water at lower temperatures. The solubility of gases decreases as the temperature increases. The relationship between solubility and temperature is inverse and approximately linear.
- It is difficult to generalize about the effect of temperature on the solubility of liquids in water. However, for polar liquids in water, the solubility usually increases with temperature. A prediction of the solubility of liquids with temperature will not be as reliable as a prediction for solids and gases.
- Some liquids (mostly nonpolar liquids) do not dissolve in water to any appreciable extent, but form a separate layer. Liquids that behave in this way are said to be **immiscible** with water. For example, benzene, gasoline, and carbon disulfide (which is used in the process of turning wood pulp into rayon or cellophane) are all virtually insoluble in water.
- Some liquids (such as those containing small polar molecules with hydrogen bonding) dissolve completely in water in any proportion. Liquids that behave in this way are said to be **miscible** with water. For example, ethanol (in alcoholic beverages), acetic acid (in vinegar), and ethylene glycol (in antifreeze) all dissolve completely in water.
- Elements generally have low solubility in water. For example, some people put a lump of sulfur in their dog's water bowl, to keep the water "fresh." The same piece of sulfur can last, apparently unchanged, for years. Carbon is used in many water filtration systems to remove organic compounds that cause odours. The carbon does not dissolve in the water passing through it.
- Although the halogens and oxygen dissolve in water to only a very tiny extent, they are so reactive that even in tiny concentrations they are often very important in solution reactions.

Of course, there are exceptions to all generalizations. For example, the solubility of lithium carbonate in water actually decreases as the temperature increases.

immiscible: two liquids that form separate layers instead of dissolving

miscible: liquids that mix in all proportions and have no maximum concentration

Practice

Understanding Concepts

6. For any solute, what important condition must be stated in order to report the solubility?

7. Sketch a solubility graph showing two lines labelled "solids" and "gases." Assume a straight-line relationship and show the generalization for the change in solubility of each type of substance with increasing temperature.

8. A common diagnostic test for $CO_{2(g)}$ involves the use of a saturated solution of calcium hydroxide, $Ca(OH)_{2(aq)}$, historically called lime-water. Calcium hydroxide has low solubility—but is actually soluble enough to have a noticeable effect. The solution is easily prepared by placing a spoonful of solid $Ca(OH)_{2(s)}$ in pure water, and stirring or shaking for a while. A small amount of the solid dissolves. The excess solute settles to the bottom, leaving a clear, saturated solution. Bubbling carbon dioxide gas through this saturated solution causes reactions that finally precipitate calcium carbonate, which is roughly 10 times less soluble than calcium hydroxide.

   (a) The first reaction is the apparent "dissolving" of a tiny amount of carbon dioxide in water. What we believe really occurs is that some carbon dioxide gas reacts with water to form carbonic acid in solution. Write a balanced chemical equation to represent this reaction.

   (b) The acid reacts with dissolved calcium hydroxide to form calcium carbonate precipitate. Write a balanced chemical equation to represent this reaction.

   (c) Use the reaction equations to explain why this is a test for the presence of $CO_{2(g)}$.

   (d) If the calcium hydroxide solution were not saturated, or nearly so, would this diagnostic test still work? Explain the reasoning behind your answer.

   (e) Plot a solubility curve for $Ca(OH)_2$, using the data in **Table 2**.

   (f) Assuming you wished to prepare 1.0 L of saturated limewater solution to be stored at a normal room temperature of 22°C, what minimum mass of solid calcium hydroxide would you require? Why should you actually use considerably more than the minimum required?

   (g) Does the solubility curve of this compound fit the generalization about the solubility of ionic solids at different temperatures? Does this mean that the generalization is not valid or useful? Discuss briefly.

Applying Inquiry Skills

9. An experiment is conducted to test the generalization for the temperature dependence of the solubility of solids. Complete the **Prediction** and **Analysis** sections of the following report.

   Question

   How does temperature affect the solubility of potassium chlorate?

   Prediction

   (a) Use the generalization for the solubility of solids to answer the Question and state your reasons.

   Evidence

Table 2: Solubility of Calcium Hydroxide at Various Temperatures

Temperature (°C)	Solubility (g /100 mL)
0	0.18
20	0.17
40	0.14
60	0.11
80	0.09
100	0.08

Table 3: Solubility of Potassium Chlorate

Temperature (°C)	Solubility (g /100 mL)
0	5.0
20	8.5
40	16.3
60	27.5
80	42.5
100	59.5

Analysis

(b) Plot a solubility curve for potassium chlorate.

(c) What is the answer to the Question, according to the Evidence?

Making Connections

10. Some industries, particularly electric power generating stations, get rid of waste hot water by releasing it into a nearby lake or river. Use your knowledge of solubility and temperature to describe why this thermal pollution is detrimental to most fish.

Crystallization

Have you ever opened a jar of liquid honey only to find that the clear, golden syrup is mixed with hard, white particles (**Figure 3**)? This is an example of what happens to a saturated sugar solution (honey is mostly sugar and water) when some of the water evaporates. The white particles are sugar crystals that have come out of solution in a process called crystallization. This crystallization process also occurs naturally (and extremely slowly) during the formation of stalactites and stalagmites in caves (**Figure 4**).

Crystallization can be artificially speeded up by heating a solution to evaporate off the solvent. With less solvent present, the concentration of the solute quickly exceeds the solubility, so the excess solute crystallizes out. This evaporation method is used industrially to isolate the solid solute from many solutions. For example, table salt and table sugar are produced industrially by heating saturated solutions of the substances. Rapid evaporation forms lots of small crystals, which are then screened for uniformity of size. Crystals of unwanted sizes are just redissolved and cycled through the process again. Large crystals can be formed by slowly evaporating the solvent (**Figure 5**).

There are several experimental designs that we can use to determine the solubility (maximum concentration) of a solid. In Unit 2 you used precipitation and filtration as an experimental design to determine the quantity of a solute present in a solution. Later in this unit you will use a titration procedure to determine the concentration of a solution. At this point you are going to use a procedure called crystallization. Crystallization involves removing the solvent from a solution by evaporation, leaving behind the solid solute—which can then be collected and measured.

Figure 3
Liquid honey normally crystallizes over time. The label instructions usually suggest that you just place the container in warm water for a while if you want the sugar crystals to redissolve.

Figure 4
Stalactites and stalagmites form in caves when calcium carbonate crystallizes from ground water solutions. The solvent evaporates.

Figure 5
You can create your own stalactite–stalagmite system with a saturated solution of Epsom salts, two glasses, a plate, and some heavy string—and it won't take centuries to form!

Wear eye protection and a laboratory apron.

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

When heating anything in a container, reduce the danger of splattering by keeping the container loosely covered.

Use oven mitts or heat-proof gloves to handle hot apparatus.

Table 4: Solubility of Sodium Chloride in Water

Temperature (°C)	Solubility (g /100 mL)
0	35.7
20	35.9
40	36.4
60	37.1
80	38.0
100	39.2

Investigation 7.1.2

The Solubility of Sodium Chloride in Water

A significant part of the work of science is to test existing theories, laws, and generalizations. Your purpose in this investigation is to test the solubility curve for an ionic solid. To do this, you will create a graph from the solubility data (**Table 4**), and use this graph to predict the solubility of sodium chloride in water at a particular temperature. You will then compare the predicted value with a value that you determine experimentally—by crystallization of sodium chloride from a saturated solution.

Complete the **Prediction**, **Evidence**, **Analysis**, and **Evaluation** sections of this report.

Question

What is the solubility of sodium chloride at room temperature?

Prediction

(a) Create a graph of the data in **Table 4**, and use it to predict the solubility of sodium chloride in water at the measured room temperature. (You will be able to improve the precision of your prediction if you start the vertical axis at 35 g/100 mL instead of the usual zero value.)

Experimental Design

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

Materials

lab apron
eye protection
oven mitts or heatproof gloves
saturated $NaCl_{(aq)}$ solution
laboratory burner with matches or striker, or hot plate
centigram balance
thermometer
laboratory stand
ring clamp
wire gauze
250-mL beaker
100-mL beaker
10-mL pipet with pipet bulb to fit

Procedure

1. Measure and record the total mass of a clean, dry 250-mL beaker (plus watch glass cover, if used).

2. Obtain about 40 mL to 50 mL of saturated $NaCl_{(aq)}$ in a 100-mL beaker.

3. Measure and record the temperature of the saturated solution to a precision of 0.2°C.

4. Pipet a 10.00-mL sample of the saturated solution into the 250-mL beaker.

5. Using a laboratory burner or hot plate, heat the solution evenly in the beaker until all the water boils away, and dry, crystalline $NaCl_{(s)}$ remains (**Figure 6**).

6. Shut off the burner or hot plate, and allow the beaker (plus any cover) and contents to cool for at least 5 min.

7. Measure and record the total mass of the beaker, cover, and contents.

8. Reheat the beaker and the residue and repeat steps 6 and 7 until two consecutive measurements of the mass give the same value. Record the final mass. (If the mass remains constant, this confirms that the sample is dry.)

9. Dispose of the salt as regular solid waste.

Analysis

(b) Using the evidence you collected, answer the Question. Write the solubility of sodium chloride in grams per 100 mL.

Evaluation

(c) Evaluate the Evidence by judging the Experimental Design, Materials, Procedure, and skills used. Note any flaws and suggest improvements. List some sources of experimental uncertainty or error.

(d) Considering your answers to (c), how certain are you about the experimental answer that you obtained?

(e) Determine the accuracy of your result by calculating the percentage difference.

(f) Use the accuracy you calculated and your answer to (c) to evaluate the Prediction you made based on the solubility curve.

(g) Based on your Evaluation of the Prediction, is the authority you used (the solubility data and curve) acceptable?

Solubility Categories

People have been using the property of solubility for thousands, perhaps millions, of years. Over the last couple of hundred years, experimenters have been investigating solubility in a more quantitative fashion, and have developed tables of solubilities of various substances under a variety of conditions. Measurements of extremely high precision show that all substances are soluble in water to some extent—so the question then just becomes, to what extent? The solubilities of various simple ionic compounds range from very **high solubility**, like that of ammonium chloride, to extremely **low solubility**, like that of silver chloride. Strictly speaking, nothing is absolutely insoluble in water, but we use the term **insoluble** to mean negligible solubility—where the effect of the quantity that will dissolve is not easily detectable. This means that, although your drinking glass may dissolve to the extent of a few hundred molecules every time you wash it, you won't have to worry about the effect for a few million years.

As you probably found in Chapter 3, reading solubility tables takes a little practice. You will notice that **Table 5**, page 324, shows only the anions as headings. We can safely assume that most compounds containing those anions have similar solubilities. It is useful to classify simple ionic compounds into categories of high and low aqueous solubility. The classification allows you to predict the state of many compounds formed in single and double displacement reactions. Of course, any defined cutoff point between high and low solubility is entirely arbitrary. We usually assign a solubility of 0.1 mol/L as the cutoff point, because most ionic compounds have solubilities either significantly greater or significantly less than

Figure 6
A saturated sodium chloride solution is heated to evaporate the water and crystallize the solute.

high solubility: with a maximum concentration at SATP (standard ambient temperature and pressure) of greater than or equal to 0.1 mol/L

low solubility: with a maximum concentration at SATP of less than 0.1 mol/L

insoluble: a substance that has a negligible solubility at SATP

Table 5: Solubility of Ionic Compounds at SATP

		Anions						
		Cl⁻, Br⁻, I⁻	**S²⁻**	**OH⁻**	**SO₄²⁻**	**CO₃²⁻, PO₄³⁻, SO₃²⁻**	**C₂H₃O₂⁻**	**NO₃⁻**
Cations	High solubility (aq) ≥0.1 mol/L (at SATP)	most	Group 1, NH_4^+ Group 2	Group 1, NH_4^+ Sr^{2+}, Ba^{2+}, Tl^+	most	Group 1, NH_4^+	most	all
		All Group 1 compounds, including acids, and all ammonium compounds are assumed to have high solubility in water.						
	Low solubility (s) <0.1 mol/L (at SATP)	Ag^+, Pb^{2+}, Tl^+, Hg_2^{2+}, (Hg^+), Cu^+	most	most	Ag^+, Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}	most	Ag^+	none

this value. As well, we often work with solution concentrations of 0.01 mol/L to 1.0 mol/L in school laboratories. In other words, we choose this 0.1 mol/L cutoff because it usually works for our purposes.

As with all simple generalizations that *usually* work, there are exceptions. There are some compounds with solubilities that are low, but close enough to our cutoff to cause problems in investigation. Calcium sulfate has a low solubility by our definition (about 0.02 mol/L at SATP), and when a spoonful is stirred into water it doesn't appear to dissolve. However, a conductivity test shows that enough does dissolve for the saturated solution to noticeably conduct electric current (**Figure 7**). This also means that, if you washed a sample of $CaSO_{4(s)}$ in a filter paper with water, you would be losing a little bit of your sample every time you rinsed it, slowly ruining your experiment's accuracy.

Figure 7
If we rely on visual evidence, calcium sulfate appears insoluble, but other evidence (i.e., conductivity) indicates that the solubility is very low but not negligible.

precipitate: (*verb*) form a low solubility solid from a solution; (*noun*) the solid formed in a chemical reaction or by decreased solubility

There is another familiar compound that is only slightly soluble in water: calcium hydroxide, $Ca(OH_{2(s)})$, used to make limewater, which is used in a common diagnostic test for $CO_{2(g)}$. This compound is defined as having low solubility by our arbitrary cutoff point, but it is actually soluble enough to have a noticeable effect. Bubbling carbon dioxide gas through this saturated solution causes reactions that finally **precipitate** calcium carbonate, which has *very* low solubility. A precipitate is a pure, solid substance that appears in a solution, either as a result of a reaction or because of decreasing solubility.

One final example of a substance that is defined as having low solubility, but that in fact does dissolve very slightly, is oxygen. The solubility of oxygen gas in

home aquariums is, for example, approximately 0.002 mol/L at 20°C—about 50 times lower than our "low solubility" cutoff—but obviously the fish don't care. This is a sufficient concentration for their purposes.

Practice

Understanding Concepts

11. The solubility table (**Table 5**) summarizes many solubility generalizations for common ionic compounds in water at SATP.
    (a) What cutoff point is used to separate low and high solubility substances in water?
    (b) State two reasons for the choice of this concentration value.
    (c) Describe, in your own words, how to use this table to find out whether an ionic compound has low or high solubility.
    (d) In a chemistry class, how do the terms "soluble" and "insoluble" relate to this table?

12. Classify the following compounds as having high or low solubilities:
    (a) $NaOH_{(s)}$ (oven cleaner)
    (b) $MnCl_{2(s)}$ (used in the dyeing industry)
    (c) $Al(OH)_{3(s)}$ (water purification)
    (d) $Ca_3(PO_4)_{2(s)}$ (raw material in production of fertilizers)
    (e) $CuSO_4 \cdot 5H_2O_{(s)}$ (bluestone, a fungicide)

13. Specialty sugar, made up of crystals as large as 5 mm long, is sold for sweetening coffee.
    (a) Suggest how this sugar might have been produced.
    (b) Predict the solubility of this sugar, compared with that of regular white sugar.
    (c) Predict the rate of dissolving of this sugar, compared with that of regular white sugar.

Applying Inquiry Skills

14. Describe two experimental designs that can be used to determine the solubility of a solid in water.

Making Connections

15. Some pollutants in natural waters, such as heavy metals and organic compounds, would be classified as having low solubility. What are some origins of pollutants in natural waters? If some are low solubility compounds, why are these a problem?

The Solvay Process—An Effect of Solubility

In Chapter 5 you studied the sequence of chemical reactions called the Solvay process. This commercial process was developed to produce sodium carbonate, $Na_2CO_{3(s)}$, a compound historically called soda ash and now commonly known as washing soda. The overall reaction for the Solvay process is

$$CaCO_{3(s)} + 2\,NaCl_{(aq)} \rightarrow Na_2CO_{3(aq)} + CaCl_{2(aq)}$$
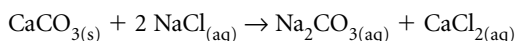
The key to the process, however, wasn't really a chemical change. Ernest Solvay used a solubility effect to separate chemicals in solution, and found a way to produce the product at one-third of the previous cost. This made Solvay both wealthy and famous, as you can imagine.

The reactants in the crucial step of the overall process are $NH_4HCO_{3(aq)}$ and $NaCl_{(aq)}$. Using our solubility chart, we find this reaction improbable.

$$NH_4HCO_{3(aq)} + NaCl_{(aq)} \rightarrow NH_4Cl_{(aq)} + NaHCO_{3(s)}$$
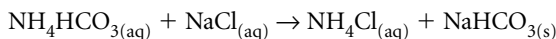
From the generalizations you have learned, all the possible products are high solubility compounds, and should not precipitate in water solution. According to our solubility table (**Table 5**), none of the possible ion combinations has low solubility.

However, the solubility table gives solubilities at SATP. Solvay's effect was made possible by Solvay's realization that, in cold water, the solubility of $NaHCO_3$ (7 g/100 mL at 0°C) is lower than the solubility of NH_4HCO_3 (12 g/100 mL at 0°C). It is also much lower than the solubilities of either NH_4Cl or $NaCl$ (**Figure 2**). This meant that the solution, at room temperature, would already be nearly saturated with sodium hydrogen carbonate. When he decreased the temperature, the $Na^+_{(aq)}$ and $HCO_3^-_{(aq)}$ combination reached its solubility limit and crystallized out of the solution as $NaHCO_{3(s)}$, leaving the other more soluble ions behind.

Practice

Understanding Concepts

16. Sketch a solubility graph (qualitative only) to illustrate the crucial step for the Solvay process in which sodium hydrogen carbonate was precipitated from an aqueous mixture of sodium hydrogen carbonate and ammonium chloride.

Making Connections

17. How is the development of the Solvay process an example of science leading to a new technology?

Section 7.1 Questions

Understanding Concepts

1. In a chemical analysis experiment, a student notices that a precipitate has formed, and separates this precipitate by filtration. The collected liquid filtrate, which contains aqueous sodium bromide, is set aside in an open beaker. Several days later, some white solid is visible along the top edges of the liquid and at the bottom of the beaker.
   (a) What does the presence of the solid indicate about the nature of the solution?
   (b) What interpretation can be made about the concentration of the sodium bromide in the remaining solution? What is the term used for this concentration?
   (c) State two different ways to convert the mixture of the solid and solution into a homogeneous mixture.

2. Describe how the solubilities of solids and gases in water depend on temperature.

Applying Inquiry Skills

3. The following investigation is carried out to test the generalization about the effect of temperature on the solubility of an ionic

compound classified as having low solubility. Complete the **Prediction**, **Experimental Design**, **Analysis**, and **Evaluation** sections of the investigation report.

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

Prediction
(a) Answer the Question, including your reasoning.

Experimental Design
Pure barium sulfate is added to three flasks of pure water until no more will dissolve and there is excess solid in each beaker. The flasks are sealed, and each is stirred at a different temperature until no further changes occur. The same volume of each solution is removed and evaporated to crystallize the solid.
(b) Identify the independent, dependent, and controlled variables.

Evidence

Solubility of Barium Sulfate

Temperature (°C)	Mass of $BaSO_{4(s)}$ (mg/100 mL)
0	0.19
20	0.25
50	0.34

Analysis
(c) Plot a solubility curve for barium sulfate.
(d) Write a sentence to answer the Question, based on the Evidence collected.

Evaluation
(e) Suggest a significant improvement that could be made in this experiment.
(f) Was the prediction verified? State your reasons.
(g) Does the solubility generalization appear acceptable for low solubility ionic compounds? Is one example enough of a test? Discuss briefly.

Making Connections

4. In "dry" cleaning, a non-aqueous solvent is used to remove grease stains from clothing. Suggest a reason (hypothesis) why a non-aqueous solvent is used, rather than water and detergent.

5. Different species of fish are adapted to live in different habitats. Some, such as carp, can survive perfectly well in relatively warm, still water. Others, such as brook trout, need cold, fast-flowing streams, and will die if moved to the carp's habitat.
   (a) Describe and explain the oxygen conditions in the two habitats.
   (b) Hypothesize about the oxygen requirements of the two species of fish.
   (c) Predict the effect of thermal pollution on trout in their streams.

7.2 Hard Water Treatment

When using soap in the bath, do you find that a floating scum forms on the surface of the bath water? Is there a hard, greyish-white scale inside your kettle and pipes (**Figure 1**)? When you wash your hands, does your soap fail to produce a good lather? This is all evidence that you have **hard water**. When detergents were initially developed they were hailed as a laundry revolution. They were chemicals formulated to work even in hard water, without producing the precipitates common when using soap.

hard water: water containing an appreciable concentration of calcium and magnesium ions

Table 1: Water Hardness

Hardness index (mg/L or ppm)	Water classification
< 50	soft
50–200	slightly hard
200–400	moderately hard
400–600	hard
> 600	very hard

Figure 1
Scale deposits caused by hard water can reduce the efficiency of some kitchen appliances.

When water travels through soil and rock it dissolves some of the minerals from the rock. Some minerals will always dissolve in ground water: The amount that dissolves depends on the rock type. Some rocks, such as granite, contain mostly low-solubility minerals. Other rocks, such as limestone, contain large quantities of minerals that dissolve relatively easily in water. The ground water that flows through or over such rock will then contain metal cations, such as calcium, $Ca^{2+}_{(aq)}$, magnesium, $Mg^{2+}_{(aq)}$, iron(II), $Fe^{2+}_{(aq)}$, iron(III), $Fe^{3+}_{(aq)}$, and manganese, $Mn^{2+}_{(aq)}$.

Water hardness is primarily due to calcium and magnesium ions. The hardness index used by chemists is the total concentration of these two ions, expressed in milligrams per litre (which is equivalent to parts per million). There is no agreed-upon standard system for grading water hardness, but **Table 1** shows an approximate classification.

Water Softening

Water is "softened" by removing the calcium and magnesium ions. This may be done during large-scale water treatment, or within a home or business by various kinds of commercial water-softening devices (**Figure 2**).

In some parts of Ontario the hardness of the water is a problem. As you learned earlier in your study of freshwater treatment, municipalities in these areas often use the **soda-lime process** to soften the water. Both sodium carbonate (the soda) and calcium hydroxide (called hydrated lime or slaked lime) are added to the hard water. After a series of reactions, the calcium and magnesium ions in the hard water are precipitated out as the respective carbonates, softening the water. This process depends entirely on differences in solubility, as both calcium and magnesium carbonates are much less soluble than the respective hydrogen carbonate (bicarbonate) compounds found in water.
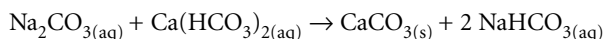

Figure 2
The use of home water-softening units is increasing steadily.

soda-lime process: a water-softening process involving sodium carbonate and calcium hydroxide, in which calcium carbonate and magnesium carbonate are precipitated out

$$Na_2CO_{3(aq)} + Ca(HCO_3)_{2(aq)} \rightarrow CaCO_{3(s)} + 2\,NaHCO_{3(aq)}$$

Home water softening usually uses an ion exchange process. In the softener tank is a bed of small resin (plastic) grains, through which the water is passed. You can think of the molecules of resin as extremely long chains, along which are attached many charged sulfonate groups: SO_3^- (**Figure 3**). The resin begins with sodium ions attached to each sulfonate group. When hard water moves through the resin, calcium and magnesium ions attach to the sulfonate groups, displacing the sodium ions into the water (**Figure 3**). Water emerging from the resin is thus "softened": Sodium ions have been exchanged for the water's calcium and magnesium ions.

Eventually, there will be no more sodium ions left on the resin. The water softener must then be regenerated or the disposable cartridge replaced. Regeneration involves washing the resin with a saturated brine (sodium chloride) solution from a salt tank attached to the water softener. This causes sodium ions to replace the calcium and magnesium ions on the resin. The hard water ions are flushed down the drain, and the softener is ready to work for another cycle.

Home softeners may regenerate by clock setting, usually during the night. More advanced models measure water usage, and regenerate only when the resin's capacity to exchange ions is nearly exhausted. This means that less salt is wasted, but these models are more expensive. Water softened in this way becomes higher in sodium ions, which may be a concern for people on low-sodium diets. Potassium chloride is sometimes used in place of sodium chloride in the brine tank. Marketers claim that since plants require potassium, this may be better for the environment, as well as for people on low-sodium diets. Potassium ions do not work quite as well as sodium ions, though, and $KCl_{(s)}$ is more expensive to buy than $NaCl_{(s)}$.

Practice

Understanding Concepts

1. Hard water is not a problem when cleaning laundry with modern laundry detergents, but it can still cause other problems.
   (a) State two observations that are evidence of hard water.
   (b) Explain why ground water is hard in some areas.

2. The soda-lime process uses two common chemicals to react with and remove certain metal ions from water.
   (a) Write the chemical formulas for the chemicals used to remove the ions.
   (b) Explain how this reaction solves the hard-water problem.

3. In municipal water treatment to soften water, calcium ions are mostly removed from the water by precipitation of calcium carbonate, which has a low solubility of 7.1×10^{-5} mol/L.
   (a) Calculate the concentration of calcium carbonate in units of milligrams per litre in the treated water.
   (b) What is the concentration of calcium carbonate in units of parts per million?
   (c) Ignoring other ions that may be present, classify the hardness of the treated water.

4. Why does a home water softener unit have to be regenerated?

Making Connections

5. Suggest reasons why an ion exchange water softener would not be appropriate for large-scale municipal water softening.

(a)

(b)
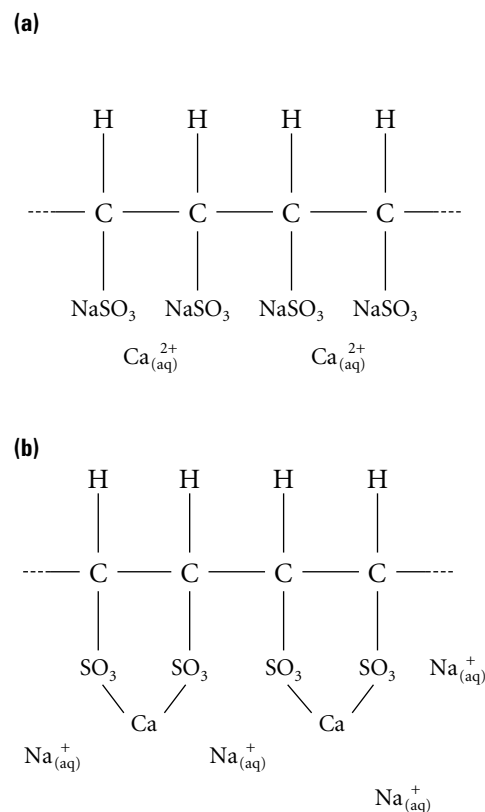
Figure 3
(a) Each resin molecule is actually hundreds of thousands of these units long. Each sulfonate group acts as a site where positive (metal) ions may be attached and held.
(b) Two aqueous calcium ions have been attached and held by the resin surface—exchanged for four aqueous sodium ions.

Answer
3. (a) 7.1 mg/L

6. "When detergents were initially developed they were hailed as a laundry revolution." Comment on this statement, giving examples of their positive and negative effects.

7. When homes in rural areas have water softeners, does the soft water go to all of the taps in the house? Use the Internet to find out. Report on what you find. What are some health concerns of drinking water with high sodium ion content?

Follow the links for Nelson Chemistry 11, 7.2.

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Section 7.2 Questions

Understanding Concepts

1. Why is water in some parts of Ontario soft, and in other parts hard?

2. What are the two metal ions most responsible for the hardness of water?

3. If you have hard water at home, what are the consequences
   (a) that you would likely notice first?
   (b) that are not easily visible? What serious problems may result?

4. In municipal water softening, calcium ions in water are precipitated as solid calcium carbonate. Assuming that the hard water contains aqueous calcium hydrogen carbonate, write the complete balanced chemical equation for the reaction of the hard water with aqueous sodium carbonate.

5. Explain the following features of ion-exchange resins in a home water softener.
   (a) Negatively charged sulfonate "sites" on resin molecules attract $Ca^{2+}_{(aq)}$ ions better than $Na^+_{(aq)}$ ions. (Consider the ion charges in your answer.)
   (b) Sodium ions from saturated $NaCl_{(aq)}$ replace calcium ions during regeneration. (Consider the concentration and collision theory in your answer.)
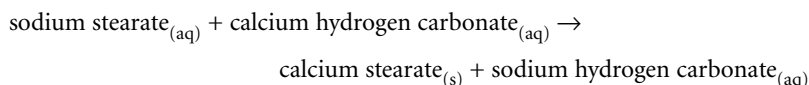
Applying Inquiry Skills

6. According to many news media reports, a researcher in the United Kingdom claimed that magnets could be used to soften water flowing through a pipe.
   (a) Write a specific question for a possible investigation and design a controlled experiment to answer the question.
   (b) The claim of water-softening magnets was, in fact, erroneously reported by most news media. The researcher actually claimed that a magnet attached to a water pipe containing flowing water makes it easier to remove the hard-water scale once it has formed. Evaluate your design in (a). What improvements can you now suggest?

Making Connections

7. Summarize the stages of ion-exchange water softening as a circular flow chart. Note the benefits inside the circle. Note the environmental or health concerns of each stage outside the circle.

7.3 Reactions in Solution

At the beginning of this chapter, we mentioned the removal of soap deposits in, for example, a bathtub or shower enclosure. The deposits (precipitates) form when soluble soap chemicals react with hard-water ions. The resulting calcium stearate, being insoluble, precipitates out on the walls of the shower.

sodium stearate$_{(aq)}$ + calcium hydrogen carbonate$_{(aq)}$ →

calcium stearate$_{(s)}$ + sodium hydrogen carbonate$_{(aq)}$

To remove these deposits, we could initiate a chemical reaction between the calcium stearate and the cleaning agent to produce soluble products that would rinse away.

calcium stearate$_{(s)}$ + cleaning agent$_{(aq)}$ → products$_{(aq)}$
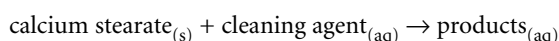
Your solubility table is useful in predicting solubilities, but there are many more low- and high-solubility ionic compounds than are presented in the table. From a scientific perspective, choosing the right chemical for the cleaning agent is a matter of choosing the right cation and anion. In this section, we look at chemical reactions that will either produce precipitates in the laboratory or remove precipitates at home.

According to the collision theory, all chemical reactions must involve collisions between atoms, ions, or molecules. If there are more collisions, there will be more reactions. Imagine a reaction between a solid and a gas. Collisions can occur only on the relatively small surface area of the solid. If the same quantity of solid is broken up, it will have far more surface area, so more collisions will occur and the substances will react more quickly. Chopping firewood into kindling to make it easy to light is a simple example of this. Dissolving a solid in a solution is the ultimate breaking-up of that solid, since the substance is reduced to the smallest possible particles: separate atoms, ions, or molecules. This creates the largest possible surface area and, therefore, the greatest number of collisions. This situation is only possible in gas- or liquid-phase solutions, but is very common.

Investigation 7.3.1

Precipitation Reactions in Solution

The purpose of this investigation is to test the solubility generalizations contained in a solubility table. To do this, you will use the solubility table (**Table 5**, page 324, or Appendix C) to predict possible precipitates.

Make a **Prediction**, carry out the **Procedure**, then complete the **Analysis** and **Evaluation** sections of the investigation report.

Question

Which aqueous solutions of cobalt(II) chloride, silver nitrate, and lead(II) nitrate react with each of sodium hydroxide, potassium iodide, and sodium carbonate?

Prediction

(a) Use the solubility table (**Table 5**) to predict the precipitates formed from each of the nine combinations of solutions.

INQUIRY SKILLS

○ Questioning	◉ Recording
○ Hypothesizing	◉ Analyzing
◉ Predicting	◉ Evaluating
○ Planning	◉ Communicating
◉ Conducting	

Experimental Design

A few drops of each pair of solutions are combined and evidence of a precipitate is noted. The volumes, concentrations, and temperatures of the solutions are all controlled.

Materials

lab apron
eye protection
rubber gloves
spot or well plate
2 medicine droppers
dropper bottles containing 0.10 mol/L aqueous solutions of
   cobalt(II) chloride
   silver nitrate
   lead(II) nitrate
   sodium hydroxide
   potassium iodide
   sodium carbonate
wash bottle of distilled water
waste containers for waste silver, lead, and cobalt compounds

Procedure

1. Place the spot or well plate on a sheet of paper and label, on the paper, the first three columns with the chemical formulas for sodium hydroxide, potassium iodide, and sodium carbonate.

2. Label the first three rows with the chemical formulas for cobalt(II) chloride, silver nitrate, and lead(II) nitrate.

3. At the intersection of each row and column, combine two drops of the solutions listed. Record specific observations for each combination of solutions.

4. Deposit all waste solutions in the containers provided by your teacher labelled "silver recycle" and "heavy metal waste."

Analysis

(b) Answer the Question by identifying which combinations did and did not produce precipitates, according to your evidence.

Evaluation

(c) Evaluate the evidence gathered by considering the Experimental Design, Materials, and Procedure. Did you gather the kind of evidence necessary to test your Prediction? If not, what evidence would be necessary?

(d) If you are confident in your evidence, compare it with your Prediction. Does your evidence support or cast doubt on the generalizations that form the basis of the solubility table?

Practice

Understanding Concepts

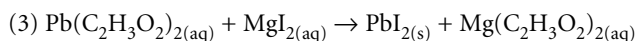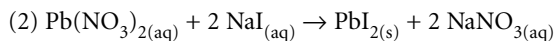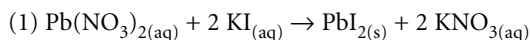1. Classify each of the following substances as having high or low solubility:

(a) silver sulfide
(b) magnesium nitrate
(c) zinc carbonate

2. Which combinations of the reactants listed below will produce a precipitate? For each precipitate predicted, write the chemical formula.
   (a) aqueous strontium nitrate and aqueous potassium sulfate
   (b) aqueous sodium acetate and aqueous strontium chloride
   (c) aqueous copper(II) bromide and aqueous sodium sulfite

Net Ionic Equations

Two students, performing Investigation 7.3.1, were curious about the yellow precipitate that formed when they mixed lead(II) nitrate with potassium iodide. They looked in the solubility table, and predicted that the precipitate is lead(II) iodide. They set out to design an experiment to provide evidence in support of their prediction. They decided to mix aqueous lead(II) nitrate with aqueous sodium iodide, and aqueous lead(II) acetate with aqueous magnesium iodide to see if they got the same yellow precipitate as they observed for lead(II) nitrate with potassium iodide.

The first student mixed aqueous solutions of lead(II) nitrate and sodium iodide and observed the instant formation of a bright yellow precipitate (**Figure 1**). The student's lab partner recorded the same observation after mixing aqueous solutions of lead(II) acetate and magnesium iodide. The balanced chemical equations for these reactions show some similarities and some differences.

(1) $Pb(NO_3)_{2(aq)} + 2\ KI_{(aq)} \rightarrow PbI_{2(s)} + 2\ KNO_{3(aq)}$

(2) $Pb(NO_3)_{2(aq)} + 2\ NaI_{(aq)} \rightarrow PbI_{2(s)} + 2\ NaNO_{3(aq)}$

(3) $Pb(C_2H_3O_2)_{2(aq)} + MgI_{2(aq)} \rightarrow PbI_{2(s)} + Mg(C_2H_3O_2)_{2(aq)}$

The evidence gathered from these reactions is identical: yellow precipitates. This supports the students' prediction that the yellow precipitate is lead(II) iodide. The results of this test increase our confidence in both the double displacement reaction generalization and the solubility generalizations.

Let's now see if we can develop a theoretical description for reactions 1, 2, and 3 above.

Using the Arrhenius theory of dissociation, we can describe these reactions in more detail. We believe that each of the high-solubility ionic reactants dissociates in aqueous solution to form separate cations and anions. For example, reaction (1) could be written as follows:

$Pb^{2+}_{(aq)} + 2\ NO^-_{3(aq)} + 2\ K^+_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} + 2\ K^+_{(aq)} + 2\ NO^-_{3(aq)}$
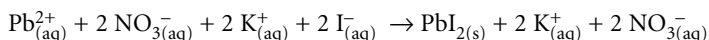
Notice that the precipitate, $PbI_{2(s)}$, is not dissociated because lead(II) iodide has a low solubility. If an ionic compound does not dissolve, it cannot dissociate into individual ions. The chemical equation above is called a **total ionic equation**. It shows what entities are present in the reaction. In this total ionic equation, the nitrate and potassium ions present on the reactant side also appear unchanged among the products.

Chemical reaction equations normally do not show chemical substances that do not change in any way. We can extend this practice to entities in solution—in this case, the ions that do not change. Any ion (or other entity, such as molecules or atoms) present in a reaction system that does not change during the course of

Figure 1
A solution of sodium iodide reacts immediately, when mixed with a solution of lead(II) nitrate, to form a bright yellow precipitate.
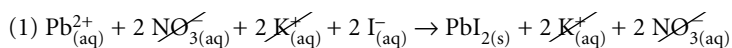
total ionic equation: a chemical equation that shows all high-solubility ionic compounds in their dissociated form
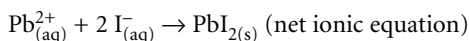
spectator: an entity such as an ion, molecule, or ionic solid that does not change or take part in a chemical reaction

net ionic equation: a way of representing a reaction by writing only those ions or neutral substances specifically involved in an overall chemical reaction
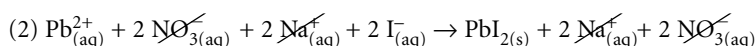
the chemical reaction is called a **spectator**. A spectator ion is like a spectator at a hockey game—present for the action but not taking part in the game. By ignoring these ions and rewriting this reaction equation showing only the entities that change, you write a **net ionic equation** for the reaction. When modifying a total ionic equation to give a net ionic equation, be careful to cancel only identical amounts of identical entities appearing on both reactant and product sides.
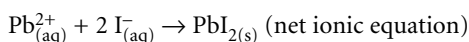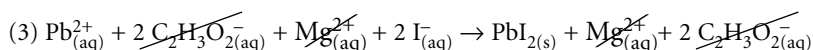
$$\text{(1)} \ Pb^{2+}_{(aq)} + 2\ \cancel{NO_{3(aq)}} + 2\ \cancel{K^+_{(aq)}} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} + 2\ \cancel{K^+_{(aq)}} + 2\ \cancel{NO_{3(aq)}}$$

can be shortened to

$$Pb^{2+}_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} \text{ (net ionic equation)}$$

Similarly, for reaction (2),

$$\text{(2)} \ Pb^{2+}_{(aq)} + 2\ \cancel{NO_{3(aq)}} + 2\ \cancel{Na^+_{(aq)}} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} + 2\ \cancel{Na^+_{(aq)}} + 2\ \cancel{NO_{3(aq)}}$$

becomes

$$Pb^{2+}_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} \text{ (net ionic equation)}$$

If we apply the same procedure to reaction (3),

$$\text{(3)} \ Pb^{2+}_{(aq)} + 2\ \cancel{C_2H_3O_{2(aq)}^-} + \cancel{Mg^{2+}_{(aq)}} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} + \cancel{Mg^{2+}_{(aq)}} + 2\ \cancel{C_2H_3O_{2(aq)}^-}$$

is written

$$Pb^{2+}_{(aq)} + 2\ I^-_{(aq)} \rightarrow PbI_{2(s)} \text{ (net ionic equation)}$$
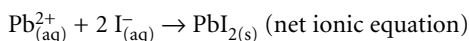
Notice that the same observation made from three apparently different chemical reactions gives the same net ionic equation. Although solutions of different substances were combined in these two cases, the net ionic equation makes it clear that the same chemical reaction occurs. Apparently, bright yellow lead(II) iodide will precipitate when we mix any two solutions containing both lead(II) ions and iodide ions.

SUMMARY   **Writing Net Ionic Equations**

1. Write the balanced chemical equation with full chemical formulas for all reactants and products.
2. Using solubility information, such as the table in Appendix C, rewrite the formulas for all high-solubility ionic compounds as dissociated ions, to show the total ionic equation.
3. Cancel identical amounts of identical entities appearing on both reactant and product sides.
4. Write the net ionic equation, reducing coefficients if necessary.

Sample Problem 1

Write the net ionic equation for the reaction of aqueous barium chloride and aqueous sodium sulfate.

Solution

$BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2\ NaCl_{(aq)}$

$Ba^{2+}_{(aq)} + 2\ \cancel{Cl^-_{(aq)}} + 2\ \cancel{Na^+_{(aq)}} + SO^{2-}_{4(aq)} \rightarrow BaSO_{4(s)} + 2\ \cancel{Na^+_{(aq)}} + 2\ \cancel{Cl^-_{(aq)}}$

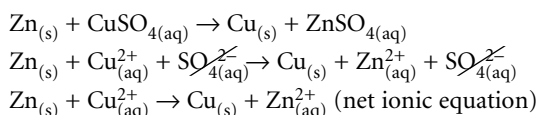$Ba^{2+}_{(aq)} + SO^{2-}_{4(aq)} \rightarrow BaSO_{4(s)}$ (net ionic equation)
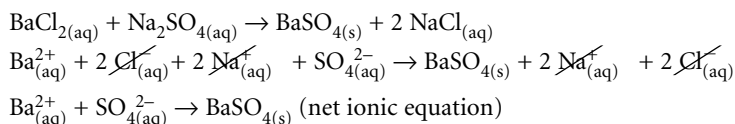
Net ionic equations are useful for more than double displacement reactions that produce precipitates. We can also use them for communicating other reactions. The following single displacement reaction is a good example.

Sample Problem 2

Write the net ionic equation for the reaction of zinc metal and aqueous copper(II) sulfate, and then write a statement to communicate the meaning of the net ionic equation.

Solution

$Zn_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + ZnSO_{4(aq)}$

$Zn_{(s)} + Cu^{2+}_{(aq)} + \cancel{SO^{2-}_{4(aq)}} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)} + \cancel{SO^{2-}_{4(aq)}}$

$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$ (net ionic equation)

Placing solid zinc in any aqueous solution containing copper(II) ions will produce solid copper and aqueous zinc ions.

Practice

Understanding Concepts

3. Strontium compounds are often used in flares because their flame colour is bright red. One industrial process to produce low-solubility strontium compounds (that are less affected by getting wet) involves the reaction of aqueous solutions of strontium nitrate and sodium carbonate. Write the balanced chemical equation, total ionic equation, and net ionic equation for this reaction.

4. Placing aluminum foil in any solution containing aqueous copper(II) ions will result in a reaction. The reaction is slow to begin with, but then proceeds rapidly.
   (a) Referring to the solubility table (Appendix C), name at least four ionic compounds that could be dissolved in water to make a solution containing aqueous copper(II) ions.
   (b) Write a balanced chemical equation for the reaction of aluminum with one of the compounds you suggested in (a).
   (c) Write the total ionic equation for the reaction.
   (d) Write the net ionic equation for the reaction.

5. One industrial method of producing bromine is to react seawater, containing a low concentration of sodium bromide, with chlorine gas. The chlorine gas is bubbled through the seawater in a specially designed vessel. Write the net ionic equation for this reaction.

6. In a hard-water analysis, sodium oxalate solution reacts with calcium hydrogen carbonate (in the hard water) to precipitate a calcium compound. Write the net ionic equation for this reaction.

7. In a laboratory test of the metal activity series, a student places a strip of lead metal into aqueous silver nitrate. Write the net ionic equation for this reaction.

Making Connections

8. Some natural waters contain iron ions that affect the taste of the water and cause rust stains. Aeration converts any iron(II) ions into iron(III) ions. A basic solution (containing hydroxide ions) is added to produce a precipitate.
   (a) Write the net ionic equation for the reaction of aqueous iron(III) ions and aqueous hydroxide ions.
   (b) What separation method is most likely to be used during this water treatment process?

Explore an
Issue

Debate: Producing Photographs

DECISION-MAKING SKILLS

○ Define the Issue ○ Analyze the Issue
○ Identify Alternatives ○ Defend a Decision
○ Research ○ Evaluate

Photographers and the photographic developing industry use a great many resources. You are probably aware of some of the drawbacks of using film. There are disposable cameras. Film is bought in plastic canisters. Every day developers and printers use huge quantities of chemicals in solution, some of which are toxic, and others of which are costly to produce. All of these aspects of film photography have an environmental impact. But what about digital photography? Think about the resources required to create and distribute digital photographs. Is this technology as environmentally clean as it first appears?

(a) In small groups, prepare for a debate on the proposition, "Digital cameras are more environmentally friendly than film cameras."

(b) Your group will be either defending or opposing the proposition. Brainstorm and research arguments in support of your position. Collect scientific evidence wherever possible. Try to go to primary sources as much as possible for your information.

(c) Assemble your evidence into separate, logical subtopics, and prepare to debate. (In the debate, each subtopic could perhaps be presented by a different member of the group.)

(d) After the debate, discuss in your group how you could have improved on your group's performance. Which (if any) aspects of your preparation could have been done better?

Follow the links for Nelson Chemistry 11, 7.3.

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Section 7.3 Questions

Understanding Concepts

1. A common method for the disposal of soluble lead waste is to precipitate the lead as the low-solubility lead(II) silicate. Write the net ionic equation for the reaction of aqueous lead(II) nitrate and aqueous sodium silicate.

2. In a water treatment plant, sodium phosphate is added to remove calcium ions from the water. Write the net ionic equation for the reaction of aqueous calcium chloride and aqueous sodium phosphate.

3. As part of a recycling process, silver metal is recovered from a silver nitrate solution by reacting it with copper metal. Write the net ionic equation for this reaction.

Making Connections

4. To reduce the amount of poisonous substances dumped in our landfill sites, many municipalities have hazardous waste disposal depots. Find out whether there is one in your area, and research the skills and qualifications of the people who work there, and the types of substances they handle.

Follow the links for Nelson Chemistry 11, 7.3.

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7.4 Waste Water Treatment

In Chapter 6 we looked at the treatment of drinking water, and the ways in which we make sure that it is safe. Then, earlier in this chapter we explored the technologies used to reduce the effects of hard water. In this section we will look at another aspect of water treatment: the treatment of waste water (**Figure 1**).

Ideally, all used water returned to the environment would be clean enough to be drinkable, whatever kind of treatment was necessary. In reality, though, some municipalities still permit raw sewage to be discharged into our rivers, lakes, and coastal waters.

Sewage is anything that gets flushed out of our homes through the sewer pipes, and includes a considerable amount of organic waste. There are three major problems with organic waste.

First is the issue of disease transmission. Untreated sewage may contain bodily wastes from people with a wide range of infectious diseases, many of which can be transmitted by contact with feces. If the disease-causing organisms are not killed, they could easily cause infection in anyone who drinks, swims in, or washes with the contaminated water.

The second problem is that the decomposition (rotting) of organic wastes uses up a lot of oxygen. Biologists measure the use of dissolved oxygen by bacteria, and call it the BOD—the biological oxygen demand. This is defined as the quantity of oxygen used up by bacteria over five days at 20°C to decompose any organic matter that is present in (or added to) water. A high BOD reading indicates that there is a lot of organic matter in the water, and that a considerable quantity of oxygen is being removed from the water in order to decompose those organics. Oxygen is really only very slightly soluble in water. Under natural conditions, oxygen is present in water to about a 0.010 g/L (10 ppm) level at a temperature of 10°C, and about 8 ppm at 25°C. Many fish and other aquatic organisms require a minimum of about 4–5 ppm of oxygen to survive. Bacteria, growing as a result of raw sewage discharge into a river, can therefore cause a critical lowering in oxygen levels for a long way downstream. This effect is even more serious if the river flows slowly and so the water is not naturally aerated. The low oxygen level can be a serious problem for anything living in the water.

Figure 1
Modern sewage treatment technology allows waste treatment plants to be less offensive to neighbours. The resulting sludge has a wide range of uses, and is even useful in some industrial processes.

DID YOU KNOW ?

Natural Oxygen Levels
The dissolved oxygen content in rivers and streams is considerably less than the maximum solubility of oxygen in water at the same temperature. There are many reasons for this, including the small surface area in contact with the air (in deep, smoothly flowing water) and the use by organisms in the water of the dissolved oxygen during respiration.

2. In a water treatment plant, sodium phosphate is added to remove calcium ions from the water. Write the net ionic equation for the reaction of aqueous calcium chloride and aqueous sodium phosphate.

3. As part of a recycling process, silver metal is recovered from a silver nitrate solution by reacting it with copper metal. Write the net ionic equation for this reaction.

Making Connections

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Figure 1
Modern sewage treatment technology allows waste treatment plants to be less offensive to neighbours. The resulting sludge has a wide range of uses, and is even useful in some industrial processes.
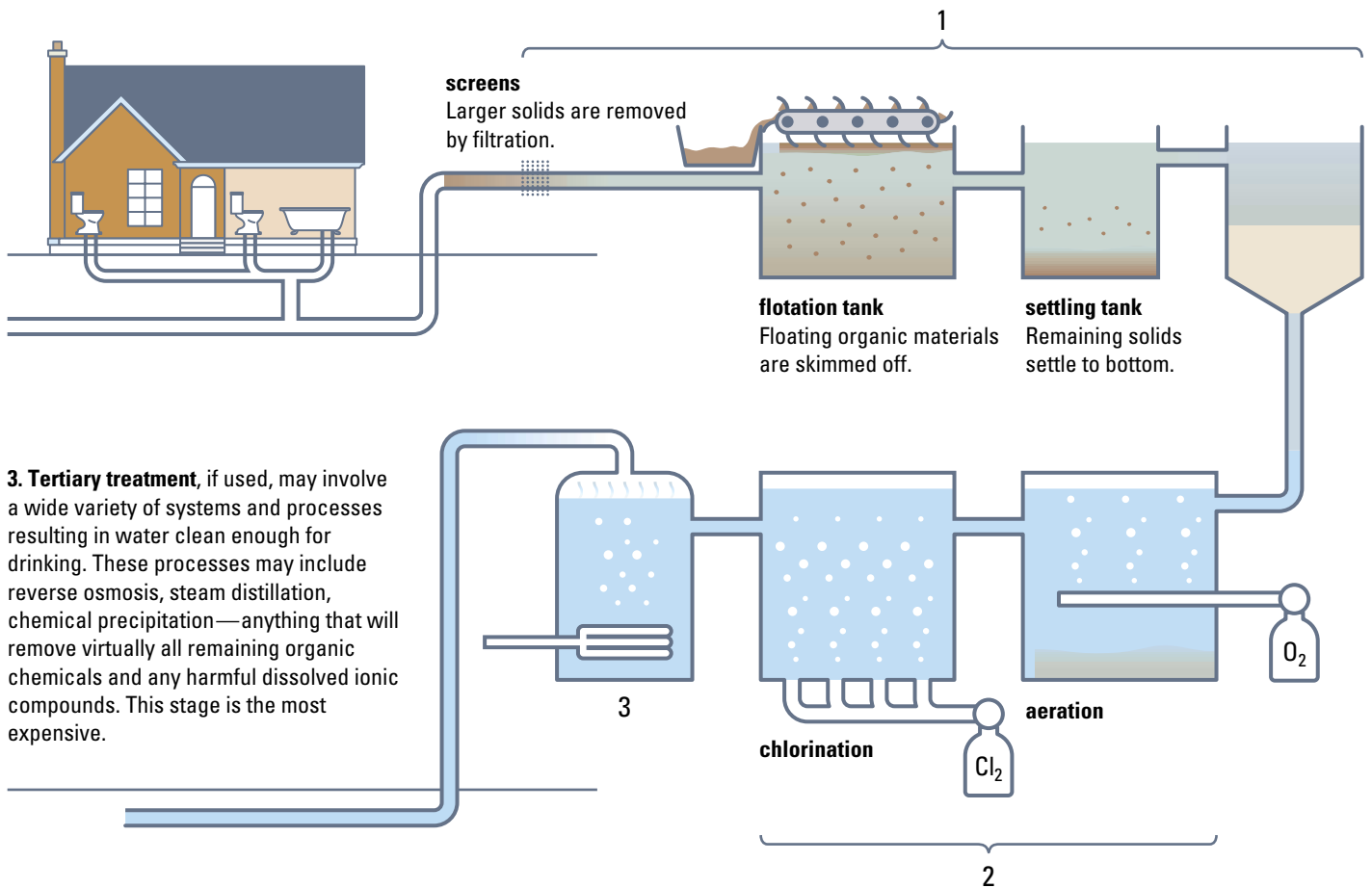
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The dissolved oxygen content in rivers and streams is considerably less than the maximum solubility of oxygen in water at the same temperature. There are many reasons for this, including the small surface area in contact with the air (in deep, smoothly flowing water) and the use by organisms in the water of the dissolved oxygen during respiration.

The third problem is that the decomposition of organic matter releases chemicals, including nitrates and phosphates, that stimulate the rapid growth of aquatic plants and algae. As the surface plants grow in much greater abundance than is usual, they block out light to deeper water, killing the bottom-dwelling plants. Later, as the artificially fertilized "bloom" dies away, it becomes another load of organic material to be decomposed—and to use up dissolved oxygen.

Along with organic waste, many other materials are delivered into the sewers: cleaning solutions (e.g., phosphates), paint (e.g., oil-based paints), garden chemicals (e.g., fertilizers and pesticides), insoluble materials (e.g., calcium and magnesium carbonates), minerals from water softeners (e.g., sodium and chloride ions), and so on. All of these would be harmful if discharged

1. Primary treatment involves screening, flotation, settling, and filtering out of solid particles. It has no effect on dissolved materials or microorganisms. This treatment removes about 40% of BOD.

screens
Larger solids are removed by filtration.

flotation tank
Floating organic materials are skimmed off.

settling tank
Remaining solids settle to bottom.

3. Tertiary treatment, if used, may involve a wide variety of systems and processes resulting in water clean enough for drinking. These processes may include reverse osmosis, steam distillation, chemical precipitation—anything that will remove virtually all remaining organic chemicals and any harmful dissolved ionic compounds. This stage is the most expensive.

chlorination

aeration

2. Secondary treatment is a two-step process. The first step usually involves aerating the water to support oxygen-using organisms, which react with dissolved organic substances to produce a sludge precipitate. The second step is chlorination, which further purifies the water. The resulting sludge may be used as landfill or fertilizer, although any heavy metal pollutants present are not removed. About 90% of BOD is removed from the water by the end of this stage. After this treatment, the water could be returned to the environment as it is suitable for most non-drinking purposes.

Figure 2
Three stages of a waste water treatment system

directly into the natural environment. Fortunately, provided that they are only present in small quantities, the waste water treatment system can deal with them.

Treatment of waste water is becoming more and more important as urban populations grow, and high-quality drinking water becomes more scarce. The general sequence for treatment of municipal sewage is shown in **Figure 2**.

In Canada there is public pressure to place the responsibility for waste water treatment on the user. The idea is that waste water should be cleaned and treated to a point that is not polluting. Not only is the treatment beneficial to life forms in the river, but also to anyone downstream who wants to use that river's water as a municipal supply of drinking water.

There are many parts of Ontario where homes and businesses are not connected to municipal sewage treatment facilities. In most cases, this is not a problem. The sewage is either stored in a holding tank until it is periodically pumped out and trucked away, or it is piped directly into a septic system (**Figure 3**). Here the solid matter is decomposed by bacteria, and the liquid is piped off to a leaching bed, where it gradually trickles through layers of sand and gravel and eventually runs down into the water table. Although septic systems generally remove most of the organic material and bacteria from the sewage, they cannot remove the other substances, such as phosphates (from cleaning products) and nonaqueous liquids.

Figure 3
Many rural homes and cottages have septic systems to purify their waste water.
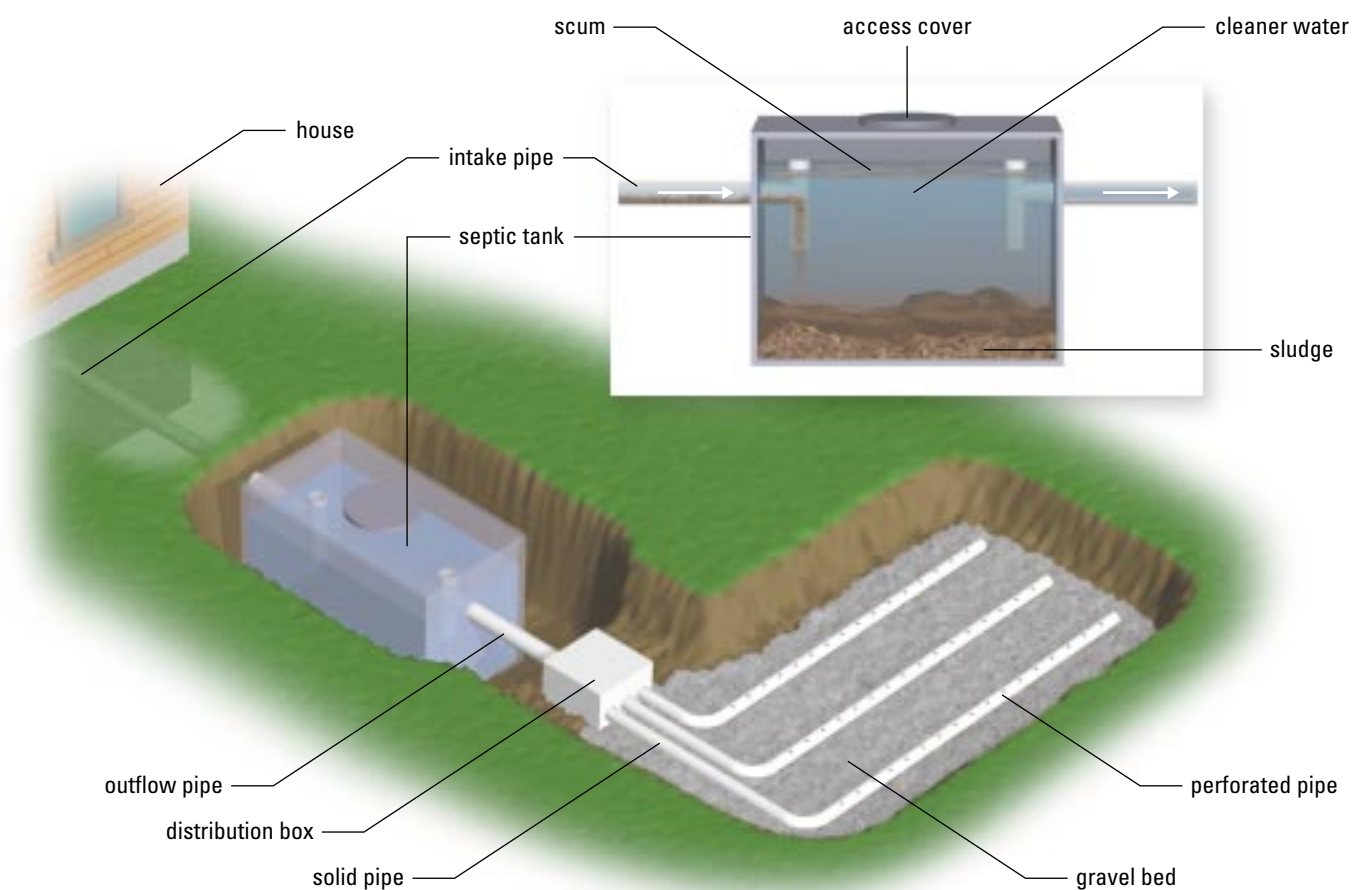
The natural environment can tolerate a certain amount of pollution. For example, the combination of sunlight with a healthy community of organisms living in large bodies of fresh water is able to break down some of the pollutants from domestic sewage, agricultural fertilizers, and industrial waste. Small quantities of these chemicals—nitrates, phosphates, and organic compounds—can be broken down by chemical reactions or by natural bacteria into simple, non-polluting substances such as carbon dioxide and nitrogen.

Practice

Understanding Concepts

1. Outline the problems that are likely to result from releasing untreated sewage into the environment.
2. Briefly outline the causes and effects of a high BOD.
3. Create a diagram indicating the stages of waste water treatment.
4. Explain why rural homeowners should be especially careful about what they flush down their drains.

Reflecting

5. Will the information in this section affect your attitude toward the materials that you pour down the sink at home? at school?

Section 7.4 Questions

Understanding Concepts

1. Create a flow chart diagram indicating the ways in which water is treated during its cycle from raw, untreated water to drinkable water in our homes, and then returned to the environment as treated waste water.

Applying Inquiry Skills

2. A team of environmental scientists discovers many dead fish in a river downstream from an industrial town. The scientists need to find the reason for this observation.
   (a) Write two possible hypotheses that they might test.
   (b) Write a prediction and experimental design for each of the hypotheses.

Making Connections

3. To reduce the quantity of garbage going to landfill sites, some people have suggested that food waste be disposed of in a garburator. These household devices grind up the waste and flush it, with lots of water, into the sewer. What are some advantages and disadvantages of this suggestion? What would be the effect of this material on the sewage treatment system?

4. How is sewage treated after it leaves your home? Research to find out, and draw your own conclusions (with further research if necessary) about whether this treatment is adequate. Outline your position (supporting it with evidence) in a communication to the person or organization responsible for your water treatment.

Follow the links for Nelson Chemistry 11, 7.4.

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7.5 Qualitative Chemical Analysis

Chemical analysis of an unknown chemical sample can include both

- **qualitative** analysis—the identification of the specific substances present, and
- **quantitative** analysis—the measurement of the quantity of a substance present.

You have probably heard about at least one widely used chemical analysis: the blood test for alcohol. Technicians test a blood sample to determine whether ethanol (from alcoholic beverages) is present, and if so, how much is present. As another example, water quality analysts are continually checking our drinking water for a wide variety of dissolved substances, some potentially harmful, some beneficial.

You have already seen many examples of qualitative analysis, such as the colour reaction of litmus paper to identify the presence of an acid or a base. The conductivity test for electrolytes, the limewater test for carbon dioxide, the explosion ("pop") test for hydrogen gas, and the glowing splint test for oxygen are all qualitative analyses. As you can see, qualitative analysis is often used for diagnostic testing. We can plan diagnostic tests using the format, "If [procedure], and [evidence], then [analysis]." For example, if cobalt(II) chloride paper is exposed to a liquid or vapour [procedure], and the paper turns from blue to pink [evidence], then water is likely present [analysis].

Flame tests, such as those that you may have conducted in Investigation 1.4.2, are also qualitative analyses.

qualitative: describes a quality or change in matter that has no numerical value expressed

quantitative: describes a quantity of matter or degree of change of matter

Qualitative Analysis by Colour

You have already learned in Unit 1 that some ions have a characteristic colour in a solution, a flame, or a gas discharge tube.

First, let's look at the colours of aqueous solutions. Most aqueous solutions are colourless, as **Table 1** shows: ions of elements in Group 1, 2, and 17 have no colour at all. Other ions, not listed in the table, are also colourless. However, many solutions containing monatomic and polyatomic ions of the transition elements do have a colour in solution.

Table 1: Colours of Solutions

Ion	Solution colour
Groups 1, 2, 17	colourless
$Cr^{2+}_{(aq)}$	blue
$Cr^{3+}_{(aq)}$	green
$Co^{2+}_{(aq)}$	pink
$Cu^{+}_{(aq)}$	green
$Cu^{2+}_{(aq)}$	blue
$Fe^{2+}_{(aq)}$	pale green
$Fe^{3+}_{(aq)}$	yellow-brown
$Mn^{2+}_{(aq)}$	pale pink
$Ni^{2+}_{(aq)}$	green
$CrO_4^{2-}_{(aq)}$	yellow
$Cr_2O_7^{2-}_{(aq)}$	orange
$MnO_4^{-}_{(aq)}$	purple

Sample Problem 1

According to the evidence in **Table 2** and **Figure 1**, which of the numbered solutions is potassium dichromate, sodium chloride, sodium chromate, potassium permanganate, nickel(II) nitrate, and copper(II) sulfate? (Refer to **Table 1**.)

Table 2: Colours of the Unknown Solutions

Solution	1	2	3	4	5	6
Colour	purple	colourless	green	yellow	blue	orange

Solution

According to **Table 2** and **Table 1**, the solutions are (1) potassium permanganate, (2) sodium chloride, (3) nickel(II) nitrate, (4) sodium chromate, (5) copper(II) sulfate, and (6) potassium dichromate.

Figure 1
Which solution is which?

Table 3: Colours of Flames

Ion	Flame colour
H^+	colourless
Li^+	bright red
Na^+	yellow
K^+	violet
Ca^{2+}	yellow-red
Sr^{2+}	bright red
Ba^{2+}	yellow-green
Cu^{2+}	blue (halides) green (others)
Pb^{2+}	light blue-grey
Zn^{2+}	whitish green

You may be familiar with the blue colour of solutions containing copper(II) ions, such as copper(II) sulfate, but did you know that copper(II) ions produce a green flame (**Figure 2**)?

We can use flame tests to indicate the presence of several metal ions, such as copper(II), calcium, and sodium (**Table 3**). In a flame test, a clean platinum or nichrome wire is dipped into a test solution and then held in a nearly colourless flame (**Figure 2**). The wire is cleaned by dipping it alternately into hydrochloric acid and then into the flame, until very little flame colour is produced. There are other ways to conduct flame tests: you could dip a wood splint in the aqueous solution and then hold it close to a flame; you could hold a tiny solid sample of a substance in the flame; or you could spray the aqueous solution into the flame.

Robert Wilhelm Bunsen and Gustav Robert Kirchhoff took the idea of the flame test and developed it into a technique called spectroscopy (**Figure 3**). Bunsen had previously invented an efficient gas laboratory burner that produced an easily adjustable, hot, nearly colourless flame. Bunsen's burner made better research possible—a classic example of technology leading science—and made his name famous. Bunsen and Kirchhoff soon discovered two new elements, cesium and rubidium, by examining the spectra produced by passing the light from flame tests through a prism.

Flame tests are still used for identification today. Of course, the technology has become much more sophisticated. The equipment used is called a spectrophotometer (**Figure 4**). It analyzes the light produced by samples vaporized in a flame. It can even detect light not visible to humans. The spectrophotometer can detect minute quantities of substances, in concentrations as tiny as parts per billion. By measuring the quantity of light emitted, this device can also do quantitative analysis—measuring the concentrations of various elements precisely and accurately. Forensic scientists may use this technique of high-tech qualitative analysis when gathering evidence for criminal investigations. Similar technology is used in a completely different branch of science: astronomy. Astronomers study the light spectra from distant stars to find out what elements the stars are composed of.

Figure 2

Copper(II) ions usually impart a green colour to a flame. This green flame, and the characteristic blue colour in aqueous solution, can be used as diagnostic tests for copper(II) ions.

Practice

Understanding Concepts

1. What is the expected colour of solutions that contain the following? (Refer to **Table 1**.)
   (a) $Na^+_{(aq)}$
   (b) $Cu^{2+}_{(aq)}$
   (c) $Fe^{3+}_{(aq)}$
   (d) $Cr_2O_7{}^{2-}_{(aq)}$
   (e) $Cl^-_{(aq)}$
   (f) $Ni^{2+}_{(aq)}$

2. What colour is imparted to a flame by the following ions in a flame test?
   (a) calcium
   (b) copper(II) (halide)
   (c) Na^+
   (d) K^+
   (e) H^+

3. Flame tests on solids produce the same results as flame tests on solutions. These tests may be used as additional evidence to support

the identification of precipitates. What colour would the following precipitates give to a flame?

(a) $CaCO_{3(s)}$
(b) $PbCl_{2(s)}$
(c) $SrSO_{4(s)}$
(d) $Cu(OH)_{2(s)}$

Applying Inquiry Skills

4. Complete the **Analysis** and **Evaluation** sections of the following report.

   Question

   What ions are present in the solutions provided?

   Experimental Design

   The solution colour is noted and a flame test is conducted on each solution.

   Evidence

   Table 4: Solution and Flame Colours

Solution	Solution colour	Flame colour
A	colourless	violet
B	blue	green
C	colourless	yellow
D	colourless	yellow-red
E	colourless	bright red

   Analysis
   (a) Which ions are present in solutions A–E?

   Evaluation
   (b) Critique the Experimental Design.

Figure 3
In 1860 Bunsen and Kirchhoff developed techniques of spectroscopy for analysis.

Figure 4
The spectrophotometer is a valuable tool, essential for precise qualitative and quantitative analyses in many areas of science.

qualitative chemical analysis: the identification of substances present in a sample; may involve several diagnostic tests
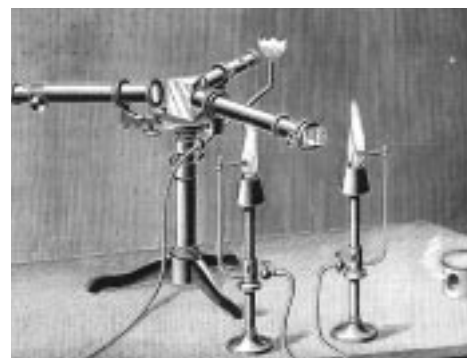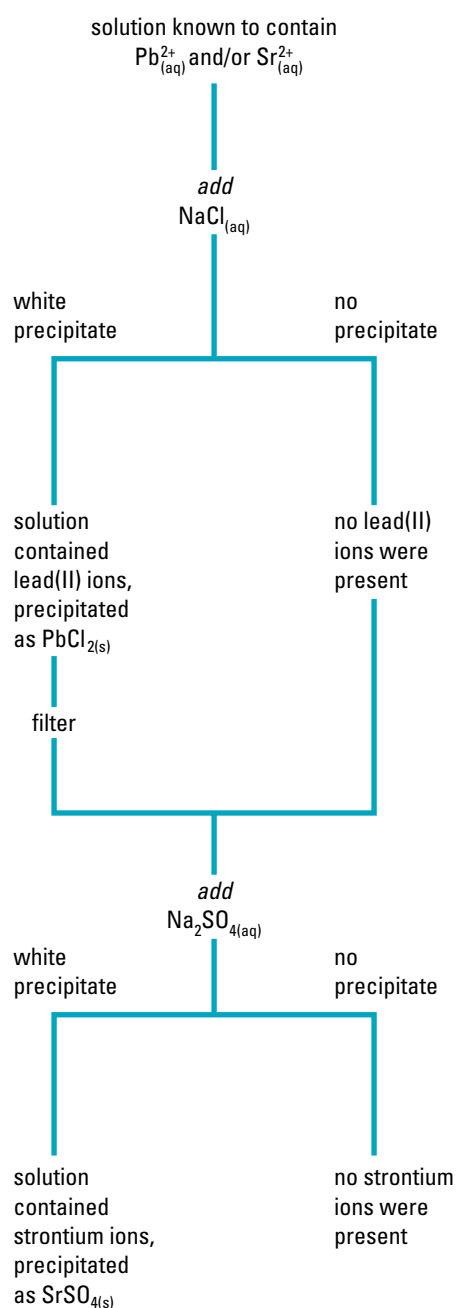
# Sequential Qualitative Chemical Analysis

In addition to colours of solutions and flames, analytical chemists have created very specific qualitative tests for ions in aqueous solution. These specific tests use the precipitation of low-solubility products. The chemist plans a double displacement reaction involving one unknown solution and one known solution, and predicts that, if a precipitate forms, then a certain ion must have been present in the unknown solution. Solubility tables help the chemist (and us) to choose reactants that will produce precipitates as evidence of the presence of specific ions. This type of test is called a **qualitative chemical analysis**.

Suppose you were given a solution that might contain either lead(II) ions or strontium ions, or possibly both or neither. How could you determine which ions, if any, were present? You could perform a qualitative analysis. For this, you must design an experiment that involves two diagnostic tests, one for each ion—that is, tests that definitely identify a particular substance. Of course, it is hard to identify an ion in a solution unless it forms a compound with low solubility. In other words, unless it forms a precipitate. So you choose reactants that, if lead(II) ions or strontium ions were present, would form a precipitate.

Test for the Presence of Lead(II) Ions

Let's test for the possible presence of lead(II) ions in a solution, before going on to test for the presence of strontium ions (**Figure 5**, page 344). Chloride ions form a low-solubility compound with lead(II) ions, so if lead(II) ions are present

solution known to contain
$Pb^{2+}_{(aq)}$ and/or $Sr^{2+}_{(aq)}$

add
$NaCl_{(aq)}$

white
precipitate

no
precipitate

solution
contained
lead(II) ions,
precipitated
as $PbCl_{2(s)}$

no lead(II)
ions were
present

filter

add
$Na_2SO_{4(aq)}$

white
precipitate

no
precipitate

solution
contained
strontium ions,
precipitated
as $SrSO_{4(s)}$

no strontium
ions were
present

Figure 5
Reading down, we see that this is one experimental design for analyzing a solution for lead(II) and/or strontium ions. In this example, the two tests could not be done in reverse order, because both lead(II) and strontium ions precipitate with sulfate ions.

limiting reagent: a reactant that is completely consumed in a chemical reaction

excess reagent: the reactant that is present in more than the required amount for complete reaction

in a solution, a precipitate forms when chloride ions are added. An appropriate source of chloride ions would be a solution of sodium chloride—a high-solubility chloride compound. Refer to the solubility table (Appendix C) to confirm that strontium ions do not precipitate with chloride ions, but that lead(II) ions do.

If sodium chloride solution is added to a sample of the test solution and a precipitate forms, then lead(II) ions are likely present and the following reaction has taken place.

$$Pb^{2+}_{(aq)} + 2\ Cl^-_{(aq)} \rightarrow PbCl_{2(s)}$$
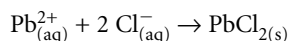
We add an excess of sodium chloride so that there are sufficient chloride ions to precipitate all of the lead(II) ions, leaving none in solution. This makes certain that there are no leftover lead(II) ions to interfere with a subsequent test for strontium ions.

When a sample under investigation (e.g., our unknown solution) is combined with an excess quantity of another reactant (e.g., sodium chloride), all of the sample reacts. The reactant that is completely used up (in this case, lead(II) ions) is called the **limiting reagent**. The reactant that is present in more than the required quantity (the chloride ions) is called the **excess reagent**.

Our precipitate indicates that lead(II) ions were present in the solution. Now we need to test our solution for strontium ions. First, however, we should remove the lead(II) chloride precipitate by filtering the solution. The remaining filtrate can then be tested for strontium ions.

Test for the Presence of Strontium Ions

What strontium-containing compound would form a precipitate? A look at the solubility table shows us that a compound containing sulfate ions and strontium ions is relatively insoluble. If a highly soluble sulfate solution (e.g., sodium sulfate) is added to the filtrate, and a precipitate forms, then strontium ions are likely to be present. (In the unlikely event that there are still some lead(II) ions in solution, they will react with the sulfate ions and would precipitate also.) If $Sr^{2+}_{(aq)}$ is present, then

$$Sr^{2+}_{(aq)} + SO_{4(aq)}^{2-} \rightarrow SrSO_{4(s)}$$
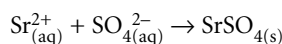
If we see a precipitate form, it is likely to be strontium sulfate. This indicates the presence of strontium ions in the original solution. We could also confirm the presence of strontium ions with a flame test.

In our example, both the unknown solution and the diagnostic-test solutions contained dissociated ions. If a precipitate forms, collisions must have occurred between two kinds of ions to form a low-solubility solid. The chloride and sodium ions present in the diagnostic-test solutions are spectator ions.

By carefully planning an experimental design for qualitative chemical analysis, you can do a sequence of diagnostic tests to detect many different ions, beginning with only one sample of the unknown solution.

We can use our experimental design as a guide for planning diagnostic tests for many ions (**Figure 6**, page 346).

SUMMARY

To complete a sequential analysis involving solubility, follow these steps. (These steps are written for cation (metal ion) analysis. For anion analysis, reverse the words cations and anions.)

1. Locate the possible cations on the solubility table.

2. Determine which anions precipitate the possible cations.

3. Plan a sequence of precipitation reactions that uses anions to precipitate a single cation at a time.

4. Use filtration between steps to remove cation precipitates that might interfere with subsequent additions of anions.

5. Draw a flow chart to assist your testing and communication.

Recognize that the absence of a precipitate is an indicator that the ion being tested for is not present. Realize that sometimes an experimental design can be created where parallel tests rather than sequential steps are used. For example, you can test for the presence or absence of calcium and mercury(I) ions by adding sulfate and chloride ions to separate samples of a solution.

Investigation 7.5.1

Sequential Chemical Analysis in Solution

A security guard noticed a trickle of clear liquid dripping from a drainage channel in the warehouse of a chemical plant. The guard alerted the chemists. To find out which of the containers of chemical solutions was leaking, the chemists needed to find out what ions were in the leaked solution. They decided to test for the most likely ions first.

In this investigation you are one of the chemists identifying the spill. Your purpose is to test a sample of the spilled solution for the presence of silver, barium, and zinc ions. You have to plan and carry out a qualitative chemical analysis using precipitation reactions, and analyze the evidence gathered.

Complete the **Materials**, **Experimental Design**, **Procedure**, and **Analysis** sections of the investigation report.

Question

Are there any silver, barium, and/or zinc ions present in the leaked solution?

Experimental Design

(a) Plan a sequential qualitative analysis involving precipitation and filtration at each step. Use sodium chloride, sodium sulfate, and sodium carbonate to test for silver, barium, and zinc ions. Draw a flow chart to communicate the experimental design and procedure.

(b) Suggest flame tests on the precipitates that would increase your confidence in the qualitative analysis.

(c) Create a list of Materials and write out your Procedure. Include any necessary safety precautions and special disposal arrangements.

Procedure

1. With your teacher's approval, carry out your investigation. Record your observations at every step.

Analysis

(d) Analyze your evidence to answer the Question.

INQUIRY SKILLS

○ Questioning	● Recording
○ Hypothesizing	● Analyzing
○ Predicting	○ Evaluating
● Planning	● Communicating
● Conducting	

See Handling Chemicals Safely in Appendix B.

Barium solutions are toxic and should be treated with care. Avoid swallowing and avoid contact with the skin. Wear gloves.
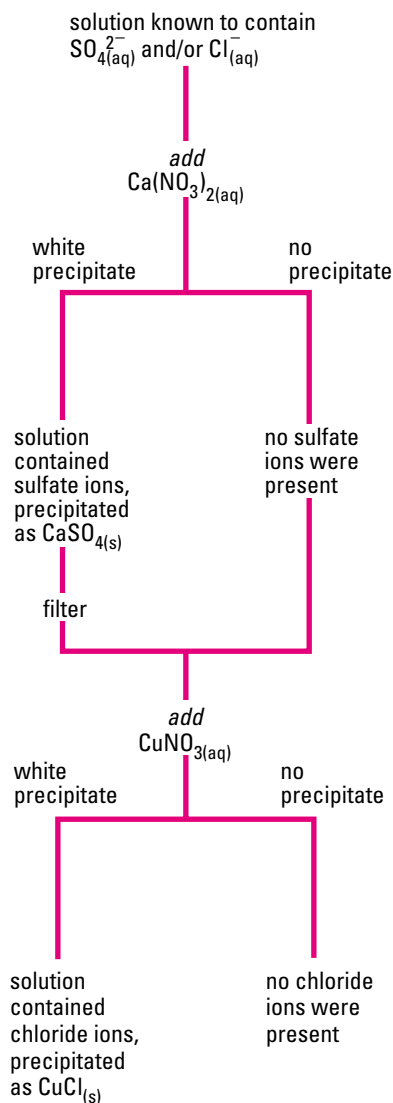
Silver solutions are corrosive and must be kept away from the eyes and skin, and must not be swallowed. Wear eye protection and do not rub your eyes.

Zinc nitrate is harmful if swallowed.

Dispose of all waste substances in a special container labelled "Heavy Metal Waste."

solution known to contain
$SO_{4(aq)}^{2-}$ and/or $Cl_{(aq)}^{-}$

add
$Ca(NO_3)_{2(aq)}$

white precipitate | no precipitate

solution contained sulfate ions, precipitated as $CaSO_{4(s)}$ | no sulfate ions were present

filter

add
$CuNO_{3(aq)}$

white precipitate | no precipitate

solution contained chloride ions, precipitated as $CuCl_{(s)}$ | no chloride ions were present

Figure 6

Reading down, we see that this is one experimental design for analyzing a solution for sulfate and/or chloride ions. In this particular example, the calcium nitrate and copper(II) nitrate tests could be done in reverse order, but this is usually not true in sequential analyses.

Practice

Understanding Concepts

5. Distinguish between qualitative and quantitative analysis.

6. What three parts are included in the "If..., and...,then..." statement of a diagnostic test? Provide an example to illustrate this.

7. Use **Table 5**: Solubility of Ionic Compounds at SATP (page 324) to suggest ion(s) that could be used to precipitate the ion listed.
   (a) $Mg_{(aq)}^{2+}$
   (b) $Ba_{(aq)}^{2+}$
   (c) $I_{(aq)}^{-}$
   (d) $SO_{4(aq)}^{2-}$

Applying Inquiry Skills

8. (a) Design a qualitative analyis for carbonate ions, using a reactant that would not precipitate sulfide ions in the sample. Refer to **Table 5**: Solubility of Ionic Compounds at SATP (page 324).
   (b) Write a net ionic equation for the precipitation reaction.

9. You are given a solution that may contain acetate and carbonate ions. Design an experiment to analyze a single sample of this solution to find out which of the ions (if any) it contains.

Making Connections

10. Provide some examples of qualitative chemical analyses that are important in our society. Use the Internet in your research.

Follow the links for Nelson Chemistry 11, 7.5.

GO TO **www.science.nelson.com**

Section 7.5 Questions

Understanding Concepts

1. Predict which of the following combinations of aqueous chemicals produce a precipitate. Write a net ionic equation (including states of matter) for the formation of any precipitate.
   (a) lead(II) nitrate and calcium chloride
   (b) ammonium sulfide and zinc bromide
   (c) potassium iodide and sodium nitrate
   (d) silver sulfate and ammonium acetate
   (e) barium nitrate and ammonium phosphate
   (f) sodium hydroxide and calcium nitrate

2. Predict the colour of aqueous solutions containing the following ions.
   (a) iron(III)
   (b) sodium
   (c) $Cu_{(aq)}^{2+}$
   (d) $Ni_{(aq)}^{2+}$
   (e) $Cl_{(aq)}^{-}$

3. (a) Which of the ions, lithium, calcium, and strontium, in separate unlabelled solutions, can be distinguished using a simple flame test with a laboratory burner?

(b) Suggest a different method to distinguish experimentally between the remaining ions.

Applying Inquiry Skills

4. (a) Design an experiment to determine whether potassium and/or strontium ions are present in a solution.
   (b) Evaluate your design.

5. Design an experiment to analyze a single sample of a solution for any or all of the $Tl^+_{(aq)}$, $Ba^{2+}_{(aq)}$, and $Ca^{2+}_{(aq)}$ ions.

6. An investigation is planned to test for the presence of copper(II) and/or calcium ions in a single sample of a solution.

   Experimental Design

   A sample of the solution is examined for colour, and is tested using a flame test.

   (a) Critique the Experimental Design.
   (b) Write an alternative Experimental Design to test for the presence of the two ions in one solution.

7. Provide an example of a household solution that contains a gas as a solute. Identify the gas and state a diagnostic test for this gas.

8. Provide an example of a household solution that contains a solid as a solute. State diagnostic tests for the presence of the ions of this solid in the solution.

9. To be transported and used in your body, minerals must exist in solution. Oxalic acid in foods such as rhubarb, Swiss chard, spinach, cocoa, and tea may remove the minerals through precipitation. Design an investigation to determine which metal-ion minerals are precipitated by oxalic acid.

10. You have been told, by an unreliable source, that a product marketed as a water-softening agent is sodium carbonate.
    (a) Design an investigation to test this hypothesis.
    (b) With your teacher's approval, conduct your investigation.

Making Connections

11. Using the Internet, research the career of forensic chemist. Find out how a spectrophotometer is used in this line of work, and give examples of the kinds of substances analyzed. Why must the amounts of substances be so precisely known?
    Follow the links for Nelson Chemistry 11, 7.5.

    GO TO www.science.nelson.com

7.6 Quantitative Analysis

Quantitative analyses are carried out for many reasons. They can help chemical engineers find out whether an industrial chemical process is operating optimally. They can also be used by water quality analysts to establish whether a contaminant is concentrated enough to be toxic. And medical laboratory technicians use quantitative analyses to decide whether the concentration of a substance in human blood is healthy or legal (**Figure 1**).

Figure 7

Figure 1
Human blood plasma contains the following ions: $Na^+_{(aq)}$, $K^+_{(aq)}$, $Ca^{2+}_{(aq)}$, $Mg^{2+}_{(aq)}$, $HCO_3^-$$_{(aq)}$, $Cl^-_{(aq)}$, $HPO_4^{2-}_{(aq)}$, and $SO_4^{2-}_{(aq)}$, as well as many complex acid and protein molecules. The concentrations of these various substances can indicate, to a trained medical professional, the state of your health.

(b) Suggest a different method to distinguish experimentally between the remaining ions.

Applying Inquiry Skills

4. (a) Design an experiment to determine whether potassium and/or strontium ions are present in a solution.
   (b) Evaluate your design.

5. Design an experiment to analyze a single sample of a solution for any or all of the $Tl^+_{(aq)}$, $Ba^{2+}_{(aq)}$, and $Ca^{2+}_{(aq)}$ ions.

6. An investigation is planned to test for the presence of copper(II) and/or calcium ions in a single sample of a solution.

   Experimental Design

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8. Provide an example of a household solution that contains a solid as a solute. State diagnostic tests for the presence of the ions of this solid in the solution.

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Figure 7

Figure 1

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Quantitative analysis quite often follows qualitative analysis. Many of us can detect the presence of alcohol by its odour. This is a crude (but direct) form of qualitative analysis. If the police suspect someone of drinking and driving, they may decide to use a breathalyzer. Breathalyzers can indirectly determine the level of blood alcohol quantitatively. Exhaled air (which may contain alcohol) is blown through a bright orange solution containing dichromate ions. If there is alcohol in the breath, it reacts with the dichromate ions to produce chromium(III) ions. The new solution is pale green in colour. This change in colour is detected when light of a certain wavelength is shone through the solution. The light that passes through is converted into an electric current by a photocell. The size of the electric current is indicated on a current meter, which is labelled in terms of blood alcohol content.

There is a more direct (and therefore more certain) way of measuring blood alcohol content. This method involves taking a blood sample and reacting it for several hours with the same kind of solution as in the breathalyzer. A laboratory technologist then performs a procedure called titration to determine the concentration of alcohol in the blood. The legal limit for blood alcohol varies from 0 to 0.080 g/100 mL (800 ppm), depending on what kind of driving licence the suspect holds. If the suspect's blood alcohol concentration is greater than the legal limit, the police may decide to press impaired driving charges.

There are various experimental designs used to determine blood alcohol content. Most of the current designs require chemical reactions and stoichiometry for quantitative analysis. In this section you will learn how to complete stoichiometric calculations involving reactions in solution.

Solution Stoichiometry

In Unit 2 you saw the usefulness of gravimetric (mass-to-mass) stoichiometry for both prediction and analysis. However, most reactions in research and industry take place in aqueous solutions. This is because chemicals in solution are easy to manipulate, and reaction rates and extents in solution are relatively easy to control—not to mention the fact that many chemical reactions only occur if the reagents are in solution. We therefore need to learn how to find the concentration of reactants in solutions. Scientists have developed a concept to help them do this, a concept known as **solution stoichiometry**.

solution stoichiometry: a method of calculating the concentration of substances in a chemical reaction by measuring the volumes of solutions that react completely; sometimes called volumetric stoichiometry

When using the concept of solution stoichiometry, we often know some of the information, such as the mass or concentration of one of the products of a reaction. We can then use this information to determine another quantity, such as the molar concentration of a reactant.

SUMMARY   **Stoichiometry Calculations**

When performing any kind of stoichiometry calculation by any method, follow these general steps.

1. Write a balanced equation for the reaction, to obtain the mole ratios.

2. Convert the given value to an amount in moles using the appropriate conversion factor.

3. Convert the given amount in moles to the required amount in moles, using the mole ratio from the balanced equation.

4. Convert the required amount in moles to the required value using the appropriate conversion factor.

Lab Exercise 7.6.1

Quantitative Analysis in Solution

INQUIRY SKILLS

○ Questioning   ○ Recording
○ Hypothesizing   ● Analyzing
○ Predicting   ○ Evaluating
○ Planning   ○ Communicating
○ Conducting

It is more financially viable to recycle metals if they are in fairly concentrated solutions, so recycling companies will pay more for those solutions than for more dilute solutions. How do companies find out how much silver, for example, is in a solution? Technicians carry out a reaction that involves removing all the silver from a known volume of the solution, drying it, and measuring its mass. Knowing the mass of silver and the volume of solution, they can calculate the molar concentration of silver in the solution.

The purpose of this investigation is to use the stoichiometric concept to find the unknown concentration of a known volume of aqueous silver nitrate: to extend gravimetric stoichiometry into solution stoichiometry.

Although you will not be conducting this investigation, the investigation report models the type of report that you will be expected to produce in later investigations. Here the evidence is provided. You are to complete the **Analysis** section.

Question

What is the molar concentration of silver nitrate in the aqueous solution provided?

Experimental Design

A precisely measured volume of aqueous silver nitrate solution, $AgNO_{3(aq)}$, is completely reacted with excess copper metal, $Cu_{(s)}$. The silver metal product, $Ag_{(s)}$, is separated by filtration and dried, and the mass of silver measured to the precision of the balance. The concentration of the initial solution is calculated from the mass of product by the stoichiometric method.

Materials

>100 mL $AgNO_{3(aq)}$ of unknown concentration
centigram balance
#16–#20 gauge solid (not braided) copper wire
fine steel wool
wash bottle of pure water
wash bottle of pure acetone, $CH_3COCH_{3(l)}$
filtration apparatus
filter paper
250-mL beaker, with watch glass to fit
400-mL waste beaker for acetone
400-mL waste beaker for filtrate
100-mL graduated cylinder
stirring rod

Procedure

Day 1

1. Using a graduated cylinder, obtain 100 mL of silver nitrate solution and pour it into a 250-mL beaker.

2. Clean about 30 cm of solid copper wire with fine steel wool, and form about 20 cm of it into a coil with a 10-cm handle, so the coiled section will

Figure 2

The blue solution that appears verifies that the most common ion of copper, $Cu^{2+}_{(aq)}$, is being slowly formed in this reaction. Silver crystals and excess copper are also visible.

be submerged when placed in the beaker (**Figure 2**) of silver nitrate solution.

3. Record any immediate evidence of chemical reaction, cover with a watch glass, and set aside until the next day.

Day 2

4. Check for completeness of reaction. If the coil is intact, with unreacted (excess) copper remaining, the reaction is complete; proceed to step 6. If all of the copper from Day 1 has reacted away, proceed to step 5.

5. Add another coil of copper wire, cover with a watch glass, and set aside until the next day.

6. When the reaction is complete, remove the wire coil, using shaking, washing, and rubbing with a stirring rod to make sure that all solid silver particles remain in the beaker.

7. Measure and record the mass of a piece of filter paper.

8. Filter the beaker contents to separate the solid silver from the filtrate.

9. Do the final three washes of the solid silver and filter paper with acetone from a wash bottle. Catch the rinsing acetone in a waste beaker.

10. Place the unfolded filter paper and contents on a paper towel to dry for a few minutes.

11. Measure and record the mass of the dry silver plus filter paper.

12. Dispose of the remaining substances by placing solids in the garbage (or recycling them), and rinsing aqueous solutions (not acetone) down the drain with plenty of water. Transfer the acetone to a recycle container.

Evidence

The following observations were recorded:
volume of silver nitrate solution = 100.0 mL
mass of filter paper = 0.93 g
mass of filter paper plus silver = 2.73 g
The copper wire in the silver nitrate solution turned black, grey, and then silver, with spikes of silvery crystals growing from the copper wire.
The solution gradually went from colourless to blue.
There was an excess of copper wire at the end of the reaction time.
After cleaning and washing, the copper wire was noticeably thinner.

Analysis

(a) Analyze the Evidence and use it to answer the Question.

Evaluation

(b) Why was the copper wire cleaned with steel wool before starting the reaction?

(c) Why did the Procedure specify to use acetone for the final three rinses?

(d) What alternative is there to the acetone rinses?

(e) What other test could be done to determine whether all of the silver nitrate in solution had reacted?

(f) If some copper were to break off the wire and join the silver product, how could you separate the two metals?

Using Molar Concentration

In Lab Exercise 7.6.1 you calculated the molar concentration of silver nitrate from the mass of silver, the molar mass of silver, and the volume of the solution. This was a stoichiometric calculation.

Some solution stoichiometry problems are a little more complicated. Sometimes a problem may involve determining the amount (in moles) of a substance by considering how much of that substance reacts with, or is produced by, a known quantity of another substance. In other words, solving the problem is dependent upon writing a balanced chemical equation for the reaction, and converting from concentration (or volume) to amount in moles, and back again.

When you did gravimetric stoichiometry calculations in Chapter 5, you used molar mass as the conversion factor to convert from mass to amount in moles, and back to mass. In solution (volumetric) stoichiometry you use molar concentration as a conversion factor.
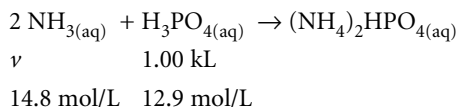
However, the general stoichiometric method remains the same. Whatever value you initially measure, you have to first convert it to an amount in moles by using a conversion factor. Second, use the mole ratio between the chemicals with given and required values. Finally, use an appropriate conversion factor to convert the required amount in moles into the requested value with appropriate units.

Let us now work through a typical stoichiometry calculation involving molar concentrations of aqueous solutions.

Many popular chemical fertilizers include ammonium hydrogen phosphate (**Figure 3**). This compound is made commercially by reacting concentrated aqueous solutions of ammonia and phosphoric acid. The reaction is most efficient if the right proportions of reactants are used. Of course, the amounts required will depend on the concentrations of the reactant solutions. What volume of 14.8 mol/L $NH_{3(aq)}$ would be needed to react completely with each 1.00 kL (1.00 m^3) of 12.9 mol/L $H_3PO_{4(aq)}$ to produce fertilizer in a commercial operation?

Step 1: Write a balanced chemical equation to find the relationship between the amount of ammonia (in moles) and the amount of phosphoric acid (in moles).

Beneath the equation, list both the given and the required measurements, and the conversion factors, in that order.

$$2\ NH_{3(aq)}\ +\ H_3PO_{4(aq)}\ \rightarrow\ (NH_4)_2HPO_{4(aq)}$$

v	1.00 kL	
14.8 mol/L	12.9 mol/L	

m is mass in, for example, grams (g).
v is volume in, for example, millilitres (mL).
n is amount in, for example, moles (mol).
C is molar concentration in moles per litre (mol/L).
M is molar mass in, for example, grams per mole (g/mol).

$$v_{H_3PO_4} = 1.00\ kL$$
$$C_{H_3PO_4} = 12.9\ mol/L$$
$$C_{NH_3} = 14.8\ mol/L$$

Step 2: Convert the information given for the reactant of known volume (phosphoric acid) to an amount in moles.

Figure 3
Fertilizers can have a dramatic effect on plant growth. The plants on the left were fertilized with ammonium hydrogen phosphate. Compare them with the unfertilized plants on the right.

$$n_{\text{H}_3\text{PO}_4} = 1.00 \text{ k\cancel{L}} \times \frac{12.9 \text{ mol}}{1 \text{ \cancel{L}}}$$

$$n_{\text{H}_3\text{PO}_4} = 12.9 \text{ kmol}$$

Step 3: Use the mole ratio to calculate the amount of the required substance (ammonia). According to the balanced chemical equation in Step 1, two moles of ammonia react for every one mole of phosphoric acid.

$$n_{\text{NH}_3} = 12.9 \text{ kmol} \times \frac{2}{1}$$

$$n_{\text{NH}_3} = 25.8 \text{ kmol}$$

Step 4: Convert the amount of ammonia to the quantity requested in the question; in this example, volume. To obtain the volume of ammonia, the molar concentration (which is the amount/volume ratio) is used to convert the amount in moles to the solution volume. Note that the prefix "kilo" does *not* cancel, while the unit "mol" does.

$$v_{\text{NH}_3} = 25.8 \text{ k\cancel{mol}} \times \frac{1 \text{ L}}{14.8 \text{ \cancel{mol}}}$$

$$v_{\text{NH}_3} = 1.74 \text{ kL}$$

According to the stoichiometric method, the required volume of ammonia solution is 1.74 kL.

We can summarize steps 2 to 4 into one step, shown below. Note that the order of use of the conversion factors is the same as above. Again, in the quantity 1.00 kL, note that only the unit cancels.

$$v_{\text{NH}_3} = 1.00 \text{ k\cancel{L} H}_3\text{PO}_4 \times \frac{12.9 \text{ \cancel{mol} H}_3\text{PO}_4}{1 \text{ \cancel{L} H}_3\text{PO}_4} \times \frac{2 \text{ \cancel{mol} NH}_3}{1 \text{ \cancel{mol} H}_3\text{PO}_4} \times \frac{1 \text{ L NH}_3}{14.8 \text{ \cancel{mol} NH}_3}$$
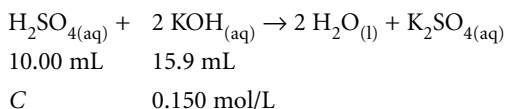
$$v_{\text{NH}_3} = 1.74 \text{ kL}$$

The following sample problem shows how to solve and communicate a solution to another stoichiometric problem involving solutions.

Sample Problem 1

Chemical technologists work in the laboratories of chemical industries. One of their jobs is to monitor the concentrations of solutions in the process stream. For example, sulfuric acid is a reactant in the production of sulfates (e.g., ammonium sulfate) in fertilizer plants. A technician needs to determine the concentration of the sulfuric acid solution. In the experiment, a 10.00-mL sample of sulfuric acid reacts completely with 15.9 mL of 0.150 mol/L potassium hydroxide solution. Calculate the molar concentration of the sulfuric acid.

Solution

$$\text{H}_2\text{SO}_{4(\text{aq})} + 2 \text{ KOH}_{(\text{aq})} \rightarrow 2 \text{ H}_2\text{O}_{(\text{l})} + \text{K}_2\text{SO}_{4(\text{aq})}$$

10.00 mL   15.9 mL

C 0.150 mol/L

$$n_{KOH} = 15.9 \text{ mL} \times \frac{0.150 \text{ mol}}{1 \text{ L}} = 2.39 \text{ mmol}$$

$$n_{H_2SO_4} = 2.39 \text{ mmol} \times \frac{1}{2} = 1.19 \text{ mmol}$$

$$C_{H_2SO_4} = \frac{1.19 \text{ mmol}}{10.00 \text{ mL}}$$

$$C_{H_2SO_4} = 0.119 \text{ mol/L}$$

or $C_{H_2SO_4} = 15.9 \text{ mL KOH} \times \dfrac{0.150 \text{ mol KOH}}{1 \text{ L KOH}} \times \dfrac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol KOH}} \times \dfrac{1}{10.00 \text{ mL}}$

$$C_{H_2SO_4} = 0.119 \text{ mol/L}$$

According to the stoichiometric method, the concentration of the sulfuric acid solution is 0.119 mol/L.

Practice

Understanding Concepts

1. Ammonium sulfate is a "high-nitrogen" fertilizer. It is manufactured by reacting sulfuric acid with ammonia. In a laboratory study of this process, 50.0 mL of sulfuric acid reacts with 24.4 mL of a 2.20 mol/L ammonia solution to yield the product ammonium sulfate in solution. From this evidence, calculate the molar concentration of the sulfuric acid at this stage in the process.

2. Slaked lime is sometimes used in water treatment plants to clarify water for residential use. The lime is added to an aluminum sulfate solution in the water. Fine particles in the water stick to the floc precipitate produced, and settle out with it. Calculate the volume of 0.0250 mol/L calcium hydroxide solution that can be completely reacted with 25.0 mL of 0.125 mol/L aluminum sulfate solution.

3. In designing a solution stoichiometry experiment for her class to perform, a chemistry teacher wants 75.0 mL of 0.200 mol/L iron(III) chloride solution to react completely with an excess of 0.250 mol/L sodium carbonate solution.
   (a) What is the *minimum* volume of this sodium carbonate solution needed?
   (b) What would be a *reasonable* volume of this sodium carbonate solution to use in this experiment? Provide your reasoning.

Applying Inquiry Skills

4. Every concept introduced in science goes through a create-test-use cycle. Even after a scientific concept has been used for some time, scientists continue to test it in new conditions, using new technologies and new experimental designs. This is why scientific knowledge is considered trustworthy. Concepts that have been around for a long time have withstood the tests of time. Their certainty is high.

   The purpose of this investigation is to test solution stoichiometry. Complete the **Prediction**, the **Analysis**, and the **Evaluation**.

   Question
   What is the mass of precipitate produced by the reaction of 20.0 mL of a 2.50 mol/L stock solution of sodium hydroxide with an excess of zinc chloride solution?

Answers
1. 0.537 mol/L
2. 375 mL
3. (a) 90.0 mL
   (b) 100 mL

Soluble barium compounds, such as barium chloride, are toxic and must not be swallowed. Wear gloves and wash hands thoroughly after handling the barium ion solution.

Wear eye protection and a laboratory apron.

Figure 4
Once the precipitate settles and the top layer becomes clear, you can test for the completeness of a reaction. Carefully run a drop or two of the excess reagent down the side of the beaker and watch for any additional precipitation, which would indicate that some of the limiting reagent remains in the filtrate solution. No precipitate formation indicates that the reaction of the limiting reagent is complete.

Prediction

(a) Use the stoichiometric method to predict the mass of precipitate.

Evidence

volume of $NaOH_{(aq)}$ = 20.0 mL
mass of filter paper = 0.91 g
mass of filter paper plus precipitate = 3.30 g
Litmus tests of the filtrate showed no change in colour.

Analysis

(b) Use the Evidence to answer the Question.

Evaluation

(c) Use your answer in (b) to evaluate your Prediction.
(d) Use your evaluation of the Prediction to evaluate the stoichiometry concept.

Investigation 7.6.1

Percentage Yield of Barium Sulfate

Barium sulfate is a white, odourless, tasteless powder that has a variety of different uses: as a weighting mud in oil drilling; in the manufacture of paper, paints, and inks; and taken internally for gastrointestinal X-ray analysis. It is so insoluble that it is non-toxic, and is therefore very safe to handle.

The reaction studied in this investigation is similar to the one used in the industrial manufacture of barium sulfate. The purpose of this investigation is to evaluate this procedure for producing barium sulfate by comparing the experimental to the expected yield in a percentage yield calculation. To do this, you will first have to decide which is the limiting and which the excess reagent, and then assuming 100% yield, predict the expected mass of product by the stoichiometric method.

Complete the **Question**, **Prediction**, **Materials**, **Procedure**, **Analysis**, and **Evaluation** sections of the investigation report.

Question

(a) Write a Question for this investigation.

Prediction

(b) From the information in the Experimental Design, decide which is the limiting and which the excess reagent. Use solution stoichiometry to predict the expected yield.

Experimental Design

A 40.0-mL sample of 0.15 mol/L sodium sulfate solution is mixed with 50.0 mL of 0.100 mol/L barium chloride solution. A diagnostic test is used to test the filtrate for excess reagent (**Figure 4**). The actual mass produced is compared with the predicted mass to assess the completeness of this reaction.

(c) Write a numbered step-by-step Procedure, including all necessary safety and disposal precautions.
(d) Write a list of Materials.

Procedure

1. With your teacher's approval, carry out your experiment and record your observations.

Analysis

(e) Analyze your evidence to calculate the experimental yield.

Evaluation

(f) Evaluate the method of production by comparing the experimental yield with the expected yield by calculating a percentage yield.

Section 7.6 Questions

Applying Inquiry Skills

1. A student wishes to precipitate all the lead(II) ions from 2.0 L of solution containing, among other substances, 0.34 mol/L $Pb(NO_3)_{2(aq)}$. The purpose of this reaction is to make the filtrate solution non-toxic. If the student intends to precipitate lead(II) sulfate, suggest and calculate an appropriate solute, and calculate the required mass of this solute.

2. When designing an experiment to determine the concentration of an aqueous solution, two students decide to carry out a precipitation reaction, followed by a crystallization. They precipitate one of the ions of the solution, boil the water away, and then measure the mass of solid remaining. Critique this experimental design.

3. The purpose of the following investigation is to test solution stoichiometry. Complete the Prediction, Analysis, and Evaluation of the investigation report.

   Question

   What mass of precipitate is produced by the reaction of 20.0 mL of 0.210 mol/L sodium sulfide with an excess quantity of aluminum nitrate solution?

   Prediction

   (a) Use the stoichiometric concept to predict an expected yield.

   Experimental Design

   The two solutions provided react with each other and the resulting precipitate is separated by filtration and dried. The mass of the dried precipitate is determined.

   Evidence

   A precipitate was formed very rapidly when the solutions were mixed.
   mass of filter paper = 0.97 g
   mass of dried filter paper plus precipitate = 1.17 g
   A few additional drops of the aluminum nitrate solution added to the filtrate produced no precipitate.
   A few additional drops of the sodium sulfide solution added to the filtrate produced a yellow precipitate.

   Analysis

   (b) Use the Evidence to answer the Question.

   (continued)

Evaluation

(c) Evaluate the experimental design, the prediction, and the stoichiometric method.

4. Once a scientific concept has passed several tests, it can be used in industry. Many industries recycle valuable byproducts, such as silver nitrate solution. Suppose you are a technician in an industry that needs to determine the molar concentration of a solution of silver nitrate. Complete the **Analysis** of the investigation report.

Question

What is the molar concentration of silver nitrate in the solution to be recycled?

Experimental Design

A sample of the silver nitrate solution to be recycled reacts with an excess quantity of sodium sulfate in solution. The precipitate formed is filtered and the mass of dried precipitate is measured.

Evidence

A white precipitate was formed in the reaction.
A similar precipitate formed when a few drops of silver nitrate were added to the filtrate.
No precipitate formed when a few drops of sodium sulfate were added to the filtrate.
volume of silver nitrate solution = 100 mL
mass of filter paper = 1.27 g
mass of dried filter paper plus precipitate = 6.74 g

Analysis

(a) Analyze the Evidence to answer the Question.

Making Connections

5. Antifreeze is used in car cooling systems. It is usually an aqueous solution of ethylene glycol, $C_2H_4(OH)_{2(aq)}$. The concentration of this solution is usually 50% by volume. What is the most common way to determine the concentration of antifreeze solutions at home and in the garage? What other solutions in an automobile have their concentrations determined in a similar way? Describe how the quantitative test works and what units of measure are used to express the concentration of the solutions.

Follow the links for Nelson Chemistry 11, 7.6.

Key Expectations

Throughout this chapter, you have had the opportunity to do the following:

- Describe, in words and by graphs, the dependence on temperature of the solubility of solids, liquids, and gases in water. (7.1)
- Explain hardness of water, its consequences, and water-softening methods. (7.2)
- Predict common combinations of aqueous solutions that produce precipitates and represent these reactions using net ionic equations. (7.3, 7.5)
- Describe the technology and the major steps involved in the treatment of waste water. (7.4)
- Perform qualitative and quantitative analyses of solutions. (7.5, 7.6)
- Solve solution stoichiometry problems. (7.6)
- Use the Key Terms for this chapter to communicate clearly. (all sections)

Key Terms

excess reagent	qualitative
hard water	qualitive chemical analysis
high solubility	quantitative
immiscible	saturated solution
insoluble	soda-lime process
limiting reagent	solubility
low solubility	solution stoichiometry
miscible	spectator
net ionic equation	total ionic equation
precipitate	

Make a Summary

Make a concept map showing how an understanding of solubility is connected to, and is central to, each of the other topics discussed in this chapter. Use as many of the Key Terms as you can.

Reflect on your Learning

Revisit your answers to the Reflect on Your Learning questions at the begining of this chapter.

- How has your thinking changed?
- What new questions do you have?

Understanding Concepts

1. Equal volumes of 1.0 mol/L solutions of each of the following pairs of solutions are mixed. Predict which combinations will form a precipitate and write net ionic equations for the predicted reactions.
    (a) $CuSO_{4(aq)}$ and $NaOH_{(aq)}$
    (b) $H_2SO_{4(aq)}$ and $NaOH_{(aq)}$
    (c) $Na_3PO_{4(aq)}$ and $CaCl_{2(aq)}$
    (d) $AgNO_{3(aq)}$ and $KCl_{(aq)}$
    (e) $MgSO_{4(aq)}$ and $LiBr_{(aq)}$
    (f) $CuNO_{3(aq)}$ and $NaCl_{(aq)}$

2. A lab technician uses 1.0 mol/L $Na_2CO_{3(aq)}$ to precipitate metal ions from waste solutions. The resulting filtered solids can be disposed of more easily than large volumes of solution. Write net ionic equations for the reaction between the $Na_2CO_{3(aq)}$ and each of the following waste solutions.
    (a) $Zn(NO_3)_{2(aq)}$
    (b) $Pb(NO_3)_{2(aq)}$
    (c) $Fe(NO_3)_{3(aq)}$
    (d) $CuSO_{4(aq)}$
    (e) $AgNO_{3(aq)}$
    (f) $NiCl_{2(aq)}$
    (g) Defend the technician's choice of $Na_2CO_{3(aq)}$ as the excess reagent.

3. The purification of water can involve several precipitation reactions. Write balanced net ionic equations to represent the reactions described below.
    (a) aqueous aluminum sulfate reacts with aqueous calcium hydroxide
    (b) aqueous sodium phosphate reacts with dissolved calcium bicarbonate
    (c) dissolved magnesium bicarbonate reacts with aqueous calcium hydroxide
    (d) aqueous calcium hydroxide reacts with dissolved iron(III) sulfate

4. When a flame test is done on a blue solution, a green flame is produced (**Figure 1**). Which aqueous cation could account for these observations?

5. List the ions whose compounds are assumed to have high solubility in water. (Refer to the solubility table on the inside back cover.)

6. A flame test on an unknown colourless solution produces a violet flame. When $HgNO_{3(aq)}$ is added to the unknown solution, a precipitate is formed. Name one possible solute for the unknown solution.

Figure 1

7. How can colour be used to distinguish between the following pairs of ions?
    (a) $Cu^+_{(aq)}$ and $Cu^{2+}_{(aq)}$
    (b) $Fe^{2+}_{(aq)}$ and $Fe^{3+}_{(aq)}$
    (c) $CrO_4^{2-}{}_{(aq)}$ and $Cr_2O_7^{2-}{}_{(aq)}$

8. Which ions are primarily responsible for "hard" water?

9. Copper(II) ions can be precipitated from waste solutions by adding aqueous sodium carbonate.
    (a) What is the minimum volume of 1.25 mol/L $Na_2CO_{3(aq)}$ needed to precipitate all the copper(II) ions in 4.54 L of 0.0875 mol/L $CuSO_{4(aq)}$?
    (b) Suggest a suitable volume to use for this reaction.

10. A 24.89-g piece of zinc is placed into a beaker containing 350 mL of hydrochloric acid. The next day the remaining zinc is removed, dried, weighed, and found to have a mass of 21.62 g. Determine the concentration of zinc chloride in the beaker.

Applying Inquiry Skills

11. A sample of drinking water from a well turns cloudy when $Ba(NO_3)_{2(aq)}$ is added to it. List the anions that could produce the cloudiness. Suggest further tests that could narrow down the possibilities.

12. A solution is known to contain sodium sulfate and/or lithium bromide. Design an experiment to test for the presence of these two compounds.

13. Nitrogen and its compounds are an essential part of the growth and decomposition of plants, and they find their way into soil and surface water. Design an experiment to distinguish among dilute aqueous solutions of the following nitrogen-containing substances: ammonium hydroxide, nitrogen gas, nitrous acid, and potassium nitrate.

14. Seawater contains traces of every chemical compound found on land. A student uses a filtration design to determine the concentration of sodium chloride, the most abundant solute in seawater. Complete the

Materials and Analysis sections, and evaluate the Experimental Design in the Evaluation section.

Question

What is the molar concentration of sodium chloride in a sample of seawater?

Experimental Design

The sodium chloride in a test sample of deep seawater reacts with an excess volume of a 1.00 mol/L $Pb(NO_3)_{2(aq)}$ solution to form a precipitate, which is then filtered and dried.

Materials

(a) What materials would be needed to conduct this investigation?

Evidence

A white precipitate formed when the solutions were mixed.

    volume of seawater = 50.0 mL
    mass of filter paper = 0.91 g
    mass of dried filter paper plus precipitate = 4.58 g

Several drops of potassium iodide solution added to the filtrate produced a yellow precipitate.

Analysis

(b) According to the Evidence, what is the molar concentration of sodium chloride in seawater?

Evaluation

(c) Critique the Experimental Design. What flaws can you see in the student's plan?

15. Copper(II) sulfate is very toxic to algae (tiny water plants). Sometimes copper(II) sulfate is added to swimming pools and water reservoirs to kill algae. A lab technician uses a filtration design to determine the concentration of copper(II) sulfate in a solution prepared for use in a water reservoir. Complete the **Materials, Analysis,** and **Evaluation** sections of the report.

Question

What is the molar concentration of copper(II) sulfate in a solution?

Experimental Design

The copper(II) sulfate solution reacts with an excess volume of a 0.750 mol/L $NaOH_{(aq)}$ solution to form a precipitate, which is then filtered and dried.

Materials

(a) What materials are needed to conduct this investigation?

Evidence

A gelatinous precipitate formed when the solutions were mixed.

    volume of copper(II) sulfate solution = 25.0 mL
    mass of filter paper = 0.88 g
    mass of dried filter paper plus precipitate = 2.83 g

The filtrate is colourless; the dried precipitate is blue.

Analysis

(b) According to the Evidence, what is the molar concentration of copper(II) sulfate in the solution?

Evaluation

(c) Critique the Experimental Design. What flaws can you see in the lab technician's plan?

Making Connections

16. What is the level of hardness in the water source for your community? Find out whether the water is treated for hardness. If so, what treatment is used?

Follow the links for Nelson Chemistry 11, Chapter 7 Review.

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17. What waste treatment process is used in your community or in a cottage that you know? What are the stages in the process? How old are the process and the equipment—is this of concern?

18. The maximum quantity of oxygen that dissolves in water at 0°C is 14.7 ppm, and at 25°C is 8.7 ppm.
    (a) Calculate the difference in the mass of oxygen that can be dissolved in 50 L of water at the two temperatures.
    (b) If you were a fish, which temperature might you prefer? Explain your answer.