3.1 Recognizing and Describing Chemical Reactions

Figure 3.1 Iron in this truck reacted with water and oxygen to form rust.

A **chemical reaction** occurs when one or more substances change to form different substances. A chemical reaction is also known as a **chemical change**. For example, a chemical reaction occurs when carbon, C, and oxygen, O_2, react to form carbon dioxide, CO_2. The substances that undergo a chemical reaction are called the **reactants**. The substances formed in a chemical reaction are called **products**. For example, in the reaction just described, carbon and oxygen are the reactants, while carbon dioxide is the product. What are the reactants and products in Figure 3.1?

Recognizing Chemical Reactions

Suppose you are baking a cake. You mix baking soda with other ingredients. When the batter cooks, the cake rises. Has there been a chemical reaction? How do you know? You need to look for evidence. Often, the clearest evidence of a chemical reaction is a release or absorption of energy. *All chemical reactions involve a change in energy.* A change in energy, however, may also accompany a physical change. Evidence of the formation of new substances suggests that a chemical reaction has taken place. For example, there may be an odour change or a colour change.

Examine the types of evidence shown in Table 3.1. Then examine the photographs in Figure 3.2 on the next page. What evidence of a chemical reaction is there in each photograph?

Chemical reactions in which light is given off are called *chemiluminescent* reactions. What is bioluminescence? Use an online or print dictionary to investigate the answer. Compare the two terms, and give an example of each one.

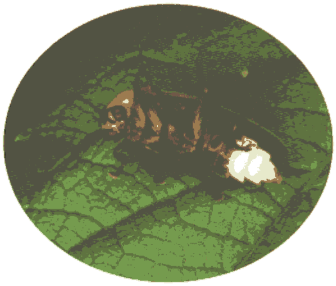

Table 3.1 Evidence of Chemical Reactions

Energy change	All chemical reactions involve a change in energy. Energy changes can be observed in the following ways: • temperature change (thermal energy released or absorbed) • emission of light (light energy released) • emission of sound • electrical energy
Odour change	In some chemical reactions, the odour of the products is different from the odour of the reactants.
Colour change	In some chemical reactions, the colour of the products is different from the colour of the reactants.
Formation of a gas	Gases are formed in some chemical reactions. When the reaction takes place in solution, you can observe the gas being formed as bubbles in the solution.
Formation of a solid (precipitate) in solution	In some reactions that take place in solution, a substance is formed that is not soluble in the solution. The substance comes out of solution as a solid (precipitate).

A Potassium perchlorate, sulfur, and other compounds react in highway flares. Energy is released as light and thermal energy.

B Baking powder in batter reacts at high temperatures. Bubbles of gas form. This is why batter rises when placed in the oven.

C The odour, colour, and taste of food change when you cook it because substances in the food undergo chemical reactions.

D Substances in bleach react with coloured compounds to form colourless compounds. This reaction is useful when you want to remove a stain from white fabric, but an accidental bleach spill can damage a pair of jeans.

E The combustion of fuel such as gasoline produces energy in the form of heat, and produces gases. The expanding gases drive the pistons in the engine of the car.

F The characteristic smell of lightning storms comes from ozone and other substances. These substances are formed when the electrical energy of lightning causes chemical reactions.

Figure 3.2 Each of these photographs illustrates a common chemical reaction.

Solubility and Chemical Reactions

One of the pieces of evidence for a chemical reaction, as listed in Table 3.1, is the formation of a solid, or **precipitate**. A precipitate is an insoluble solid that is formed in a chemical reaction that occurs in solution. Such reactions often involve ionic compounds.

Recall from Chapter 2 that many ionic compounds are soluble in water. When an ionic compound dissolves in water, the ions break away from the crystal lattice structure. The cations and anions that made up the solid crystal now move freely in the water. Some ionic compounds, however, do *not* dissolve well in water.

Why does the formation of a precipitate signal that a chemical reaction has taken place? Suppose two solutions of ionic compounds are mixed. The anions and cations may react to form a compound that is insoluble in water. This compound comes out of solution as a precipitate. The solution turns cloudy and particles of the precipitate may collect at the surface or bottom of the liquid. Figure 3.3 shows an example of precipitation that you may have seen.

In Inquiry Investigation 3-A, you can mix solutions of ions to determine whether a precipitate will form. You can use your observations to identify ionic compounds that are soluble or insoluble in water.

Figure 3.3 *Hard water* contains dissolved ions that react with compounds in soap. Insoluble compounds form as products and appear as a precipitate called soap scum.

INQUIRY

INVESTIGATION 3-A

SKILLCHECK

Initiating and Planning

Performing and Recording

Analyzing and Interpreting

Communication and Teamwork

Recognizing Reactions of Ionic Compounds in Solution

When two solutions of ionic compounds are mixed, the positive and negative ions mix and may react. When a positive ion and a negative ion react to form an insoluble compound, a precipitate forms. In this investigation, you will observe the formation of a precipitate and attempt to identify the ions that have combined.

Question

How can you test whether a reaction occurs between dissolved ionic compounds?

Safety Precautions

- Wash your hands thoroughly at the end of this investigation.
- Dispose of materials as your teacher directs. Do not pour any chemicals down the drain.

Apparatus

test tube rack
6 test tubes

Materials

6 labels
sodium nitrate solution
calcium chloride solution
sodium phosphate solution

Calcium chloride is a more expensive, but less corrosive, alternative to sodium chloride (road salt) for melting ice on streets and sidewalks.

Procedure

1. Label three test tubes with the names of the solutions to be tested. One test tube should be labelled "sodium nitrate solution," another should be labelled "calcium chloride solution," and the third test tube, "sodium phosphate solution."
2. Pour 10 mL of each solution into the test tube with the appropriate label.
3. Carefully pour about half the sodium phosphate solution into another, empty test tube. To this test tube, add about half the calcium chloride solution. **Record your observations.** Make a label to identify the contents of this test tube.
4. Pour about half the sodium nitrate solution into another test tube, and add the remaining calcium chloride solution. **Record your observations**, and make a label to identify the contents of the test tube.

5 Pour the remaining sodium nitrate solution and sodium phosphate solution into the last empty test tube. **Record your observations,** and make a label to identify the contents of the test tube.

6 Dispose of the solutions as your teacher directs. Do not pour any solutions down the drain. Wash your hands thoroughly.

Analyze

1. In step 3, you mixed solutions of sodium phosphate and calcium chloride. Write the names and formulas of the ions in the sodium phosphate solution. Then do the same for the calcium chloride solution.

2. Sodium nitrate and calcium chloride are soluble in water. Expain how your observations confirm this statement. Use your observations in step 4 to identify two other ionic compounds that are soluble in water. Explain your reasoning.

3. Write the names and formulas of the ions in the solutions you mixed in step 3. What two substances, other than sodium phosphate and calcium chloride, could form from these ions? What is the name and formula of the precipitate formed in this step? Explain your reasoning.

Conclude and Apply

4. When you mixed solutions of sodium nitrate and sodium phosphate in step 5, did new compounds form? Explain your answer.

5. Potassium nitrate, $KNO_{3(s)}$, is highly soluble in water. If you mixed each of the following pairs of solutions, would a precipitate form in each case? Explain your answers.
   (a) potassium phosphate and calcium nitrate
   (b) potassium chloride and sodium nitrate
   (c) potassium nitrate and sodium nitrate

6. **(a)** List all the sodium compounds you encountered in this investigation. Describe each compound as soluble or insoluble according to your observations and previous experience.
   (b) What generalization can you suggest about the solubility of compounds containing sodium ions? How could you test this proposed generalization?
   (c) Repeat steps (a) and (b) for the calcium compounds you encountered in this investigation.

Extend Your Skills

7. From your teacher, obtain another set of three solutions of ionic compounds. Two of the ions in these solutions combine to form a precipitate. Design a set of tests to identify the precipitate, and perform them. **CAUTION:** One of these solutions, silver nitrate solution, can stain your skin. Handle it very carefully.

Predicting Solubility

How do you know whether a certain pair of positive and negative ions will combine to form a precipitate in an aqueous solution? Scientists have tested a wide range of ionic compounds to see how well they dissolve in water. Table 3.2 outlines their results. The table gives combinations of common positive and negative ions and shows whether ionic compounds containing two types of ions are soluble in water.

Table 3.2 Solubility of Ionic Compounds in Water at 25°C

Ions	NH_4^+	NO_3^- ClO_3^- ClO_4^-	CH_3COO^-	Cl^- Br^- I^-	SO_4^{2-}	S^{2-}	OH^-	CO_3^{2-} PO_4^{3-} SO_3^{2-}
High solubility	all	all	most	most	most	Group 1 Group 2 NH_4^+	Group 1 NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group 1 NH_4^+
Low solubility	none	none	Ag^+ Hg^+	Ag^+ Pb^{2+} Cu^+ Hg^+ Tl^+	Ag^+ Pb^{2+} Ca^{2+} Ba^{2+} Sr^{2+} Ra^{2+}	most	most	most

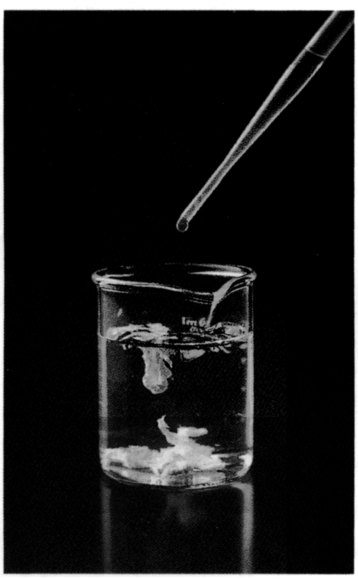

Figure 3.4 The solution in this beaker contained dissolved sodium chloride, $NaCl_{(s)}$. The chemist added silver nitrate solution, $AgNO_{3(aq)}$, from a dropper. What compound is the white precipitate? How do you know?

You can predict the solubility of ionic compounds in water by using the information in the table. For example, suppose you wish to determine whether barium hydroxide, $Ba(OH)_{2(s)}$, is soluble in water. First, locate one of its ions (Ba^{2+} or OH^-) in the first row of Table 3.2. The negative ion, OH^-, is in the seventh column of the first row. Looking down the column, locate Ba^{2+}. You will find it in the row with the heading "High solubility." This heading means that $Ba(OH)_{2(s)}$ is soluble in water.

How soluble is silver sulfide, $Ag_2S_{(s)}$? The sulfide ion is in the seventh column of the first row. Silver does not appear specifically in that column; therefore, it is included among "most" ions in the "Low solubility" category. Silver sulfide therefore has a relatively low solubility in water.

Knowing the solubility of different ionic compounds can help chemists identify the solutes in solutions using a precipitation test. For example, Figure 3.4 shows a test for chloride ions, Cl^-. The chemist adds silver nitrate solution to the solution being tested. If a white precipitate forms, the chemist infers that there are likely chloride ions in the solution. Use Table 3.2 to justify this inference. Suggest one other inference the chemist could make.

How do you predict whether a precipitate will form when you mix two solutions together? Sometimes it is useful to draw diagrams in order to visualize what is happening when the solutions are mixed. Work through the Model Problem on the following page to see how to solve solubility problems. Practise your skills by solving the Practice Problems on page 90.

Model Problem 1

A chemist mixes a solution of lead nitrate, $Pb(NO_3)_{2(aq)}$, and potassium iodide, $KI_{(aq)}$. Does a precipitate form? If so, what is the precipitate?

Solution

- First determine which ions are present in each solution (Figure 3.5 A). The lead nitrate solution contains Pb^{2+} ions and NO_3^- ions. The potassium iodide solution contains K^+ ions and I^- ions.
- When the solutions are first mixed (Figure 3.5 B), the mixture contains all four ions: Pb^{2+}, NO_3^-, K^+, and I^-. Using your knowledge of how ionic compounds form, list the possible compounds that can form from these ions. They are:

  $Pb(NO_3)_2$ KI KNO_3 PbI_2
- Decide whether any of the compounds are insoluble. You know that $Pb(NO_3)_{2(s)}$ and $KI_{(s)}$ are both soluble, since they were the two compounds initially in solution.
- According to the solubility chart, all ionic compounds containing the nitrate ion, NO_3^-, are soluble. Therefore, $KNO_{3(s)}$ is soluble.
- According to the solubility chart, most ionic compounds containing the iodide ion, I^-, are soluble. One exception occurs when the compound contains Pb^{2+}, as is the case with $PbI_{2(s)}$. Therefore, $PbI_{2(s)}$ is insoluble.
- PbI_2 is the only insoluble compound that can form from the four ions. Figures 3.5 C and D shows what happens when the solution $KI_{(aq)}$ is added to the solution $Pb(NO_3)_{2(aq)}$.

When lead(II) nitrate, $Pb(NO_3)_{2(aq)}$, and potassium iodide, $KI_{(aq)}$, are mixed, a precipitate forms. The precipitate is $PbI_{2(s)}$.

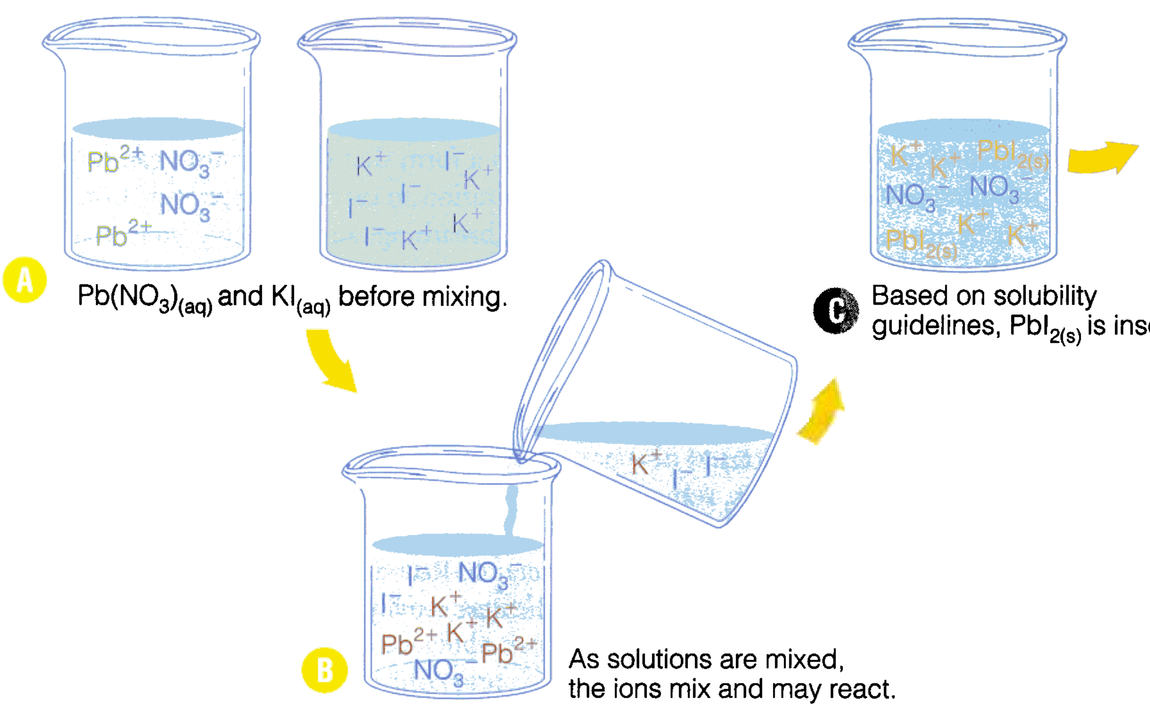

Figure 3.5 The bright yellow precipitate shown in (D) is lead(II) iodide.

Practice Problems

1. Determine whether the following ionic compounds have high solubility or low solubility in water.
   (a) $KCl_{(s)}$ (b) $SrSO_{4(s)}$ (c) $NH_4OH_{(s)}$ (d) $CuCl_{(s)}$

2. Determine whether the following compounds have high solubility or low solubility in water.
   (a) $PbI_{2(s)}$ (b) $Rb_2S_{(s)}$ (c) $Ba(OH)_{2(s)}$ (d) $CaSO_{4(s)}$

3. When solutions of $KI_{(aq)}$ and $AgNO_{3(aq)}$ are mixed, a precipitate forms.
   (a) What four types of ions are in solution when $KI_{(aq)}$ and $AgNO_{3(aq)}$ are first mixed?
   (b) Other than $KI_{(aq)}$ and $AgNO_{3(aq)}$, what two possible compounds can form from these ions?
   (c) Which one of these compounds is the precipitate? Briefly explain your answer.

4. (a) Which polyatomic cation forms soluble compounds with all anions?
   (b) List the other cations that form soluble compounds with all anions.
   (c) List the anions that form soluble compounds with all cations.

Figure 3.6 Plants, such as these trees, absorb the greenhouse gas carbon dioxide and produce oxygen through the endothermic process of photosynthesis.

Chemical Reactions and Energy Changes

All chemical reactions release or absorb energy. For example, consider the process of photosynthesis. In photosynthesis, plants convert carbon dioxide and water to glucose (food) and oxygen. Energy from the Sun is required for the reaction to proceed. What evidence is there that photosynthesis requires energy in the form of light? Plants do not produce oxygen at night. Also, a plant kept in a darkened room will eventually die, because it cannot produce food for itself.

The reverse process of photosynthesis is cellular respiration. In cellular respiration, plants and animals convert glucose and oxygen into carbon dioxide and water. This process releases energy in a form that the organism can use. Your body uses energy from cellular respiration to carry out every function it performs. Running, solving a problem, breathing, and digesting your lunch all require energy from this process.

For many chemical reactions, *thermal* energy is the most significant form of energy involved. For this reason, reactions that release energy in any form are called **exo*therm*ic reactions.** Cellular respiration, an explosion, the combustion of gasoline in an engine, and the rusting of iron are all examples of exothermic reactions.

Like photosynthesis, some reactions need the continuous addition of energy to cause a chemical change. For example, when you cook food the energy you add is absorbed by the reactants. Reactions that absorb energy are called **endothermic reactions.** In the next activity, decide whether a reaction is endothermic or exothermic.

Find Out ACTIVITY

Exothermic or Endothermic?

(Demonstration)

All chemical reactions involve a change in energy. This change may be quite noticeable, or it may be so slight that it can be detected only with sensitive instruments. Note any evidence of a change in energy as your teacher demonstrates two reactions.

Materials

iron powder, $Fe_{(s)}$
sodium chloride, $NaCl_{(s)}$
vermiculite
resealable plastic bag
barium hydroxide octahydrate, $Ba(OH)_2•8H_2O_{(s)}$
ammonium thiocyanate, $NH_4SCN_{(s)}$
Erlenmeyer flask and stopper
block of wood
water

Safety Precautions

- Your teacher will demonstrate these reactions.
- Do not handle any of the materials unless you are wearing gloves.
- Barium hydroxide octahydrate and ammonium thiocyanate are poisonous. One of the products of the second reaction is ammonia gas, which is a respiratory irritant.

Procedure • Performing and Recording

1. Place about 25 g of iron powder, $Fe_{(s)}$, and 1 g of sodium chloride, $NaCl_{(s)}$, in a resealable plastic bag.
2. Add about 30 g of vermiculite to the bag. Seal the bag. Gently squeeze and shake the contents to mix them.
3. Pass the bag around the class. Hold the bag between your hands and note any changes in temperature.
4. Pour a very thin layer of water on the wood block.
5. Add 20 g barium hydroxide octahydrate and 10 g ammonium thiocyanate to the Erlenmeyer flask.
6. Insert the stopper and shake the dry mixture until it begins to resemble slush. Place the flask on the block for a few minutes.
7. Lift up the flask.

What Did You Find Out? • Analyzing and Interpreting

1. Classify the first and second reactions as endothermic or exothermic. Explain your reasoning.
2. Besides a change in energy, what other evidence, if any, showed that a chemical reaction was taking place in each case?

Chemical Reactions and the Law of Conservation of Energy

Like all natural events, chemical changes obey the **law of conservation of energy**. This law states that *energy can be converted from one form to another, but the total energy of the universe remains constant.* In other words, energy cannot be destroyed or created. Figure 3.7 on the next page shows the scientists who are credited with developing the law.

At first, the law of conservation of energy may seem puzzling when you think of chemical reactions. For example, burning wood seems to create thermal energy. You might conclude that energy has not been conserved. This conclusion, however, is inaccurate. It overlooks the fact that energy is

Pause & Reflect

A change in energy does not always mean that a chemical change has taken place. Physical changes also involve energy changes. Give three examples of physical changes. Classify them as endothermic or exothermic and justify your classification. What questions would you ask to decide whether a change is physical or chemical?

present in the bonds of the reactants and products of every chemical reaction. The term "total energy" takes this chemical energy into account.

Energy is always required to break chemical bonds. As well, energy is always released when new bonds form. In a chemical reaction, chemical bonds in the reactants break, and new bonds form in the products. For example, think about what happens when water reacts to form hydrogen and oxygen. The bonds within the water molecules break (this requires energy) as bonds within the new hydrogen and oxygen molecules form (this releases energy).

- Breaking chemical bonds is an endothermic process.
- Forming new chemical bonds is an exothermic process.

The energy released or absorbed in the reaction is equal to the difference between the energy required to break bonds and the energy released when bonds are formed. For example, the electrolysis of water to form hydrogen and oxygen is endothermic. This means that less energy is released by the formation of new bonds in the hydrogen and oxygen molecules than is required to break the bonds in the water molecules.

1. When more energy is required to break bonds than is released when new bonds form, the reaction is endothermic. For example,

energy + water → hydrogen + oxygen

When oxygen and hydrogen react to form water, energy is released.

2. When less energy is required to break bonds than is released when new bonds form, the reaction is exothermic. For example,

hydrogen + oxygen → water + energy

Figure 3.8 summarizes these ideas.

Figure 3.7 The law of conservation of energy as we know it today was formulated by Hermann Helmholtz (A), based in part on the experimental data of James Joule (B) and the theories of Julius Robert Mayer (C).

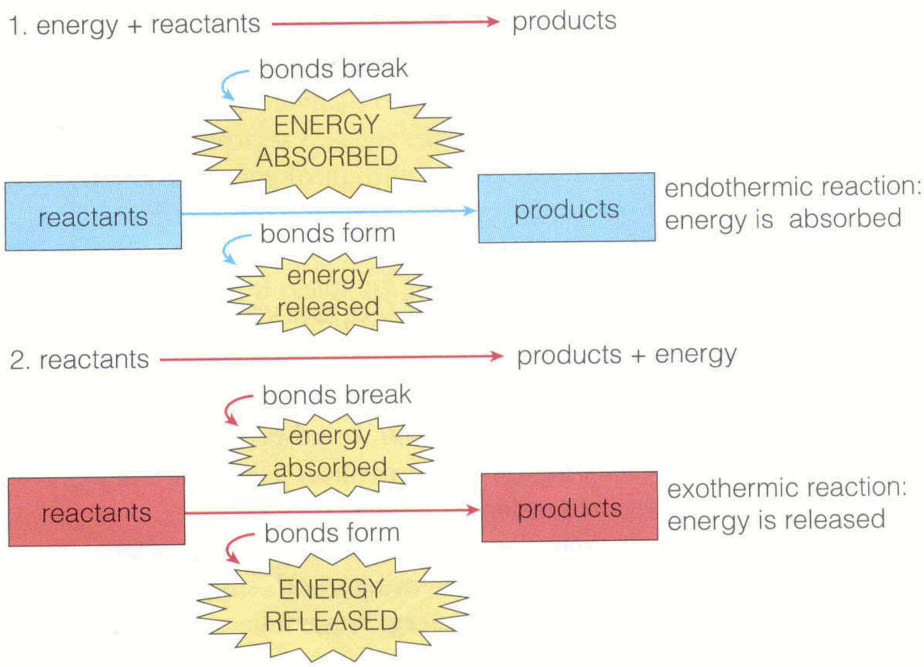

Figure 3.8 The difference in energy between the reactants and the products is released or absorbed in the chemical reaction.

Section 3.1 Summary

In this section, you reviewed the evidence of chemical reactions. The formation of a precipitate and a change in energy both indicate that a chemical reaction has taken place. In the next section, you will learn how to represent chemical reactions using chemical equations.

Check Your Understanding

1. Describe five types of evidence of a chemical reaction. For each type of evidence, give an example that has not been used in this section.
2. Distinguish between an exothermic and an endothermic reaction, using examples to illustrate your answers. In the laboratory, which type of reaction is more likely to be dangerous? Why?
3. Name each of the following ionic compounds and determine whether or not each one is relatively soluble in water.
   (a) $AgNO_{3(s)}$
   (b) $CaS_{(s)}$
   (c) $PbI_{2(s)}$
   (d) $Na_2SO_{4(s)}$
   (e) $(NH_4)_2SO_{4(s)}$
   (f) $Li_2CO_{3(s)}$
4. **Apply** Classify each of the following changes as exothermic or endothermic. Also classify each change as a chemical reaction or a physical change. Explain your reasoning for each.
   (a) Ice melts.
   (b) A match burns after it is struck against a rough surface.
   (c) Dissolving ammonium nitrate in water decreases the temperature of the solution.
   (d) Heating a frying pan that contains a raw egg cooks the egg. When the frying pan is removed from the burner, the cooking stops.
5. **Thinking Critically** A student mixes aqueous potassium sulfate and aqueous barium nitrate. Will the student observe the formation of a precipitate? If so, what is the name and chemical formula of the precipitate? If not, explain why not.
6. **Thinking Critically** A student has two beakers containing clear, colourless liquids. One of the beakers contains aqueous sodium bromide, $NaBr_{(aq)}$. The other beaker contains pure water.
   (a) How could the student use aqueous copper(I) nitrate, $CuNO_{3(aq)}$, to determine which beaker contains aqueous sodium bromide? Explain your answer in detail.
   (b) Suggest a different soluble ionic compound that the student could have used to test for the presence of sodium bromide, $NaBr_{(aq)}$.

DidYouKnow?

Many cooks use sodium hydrogencarbonate, $NaHCO_{3(s)}$ (commonly called sodium bicarbonate or baking soda), to put out cooking fires. When heated, sodium hydrogencarbonate forms carbon dioxide gas, which does not support combustion (burning). The solid residue of sodium carbonate helps to smother the fire. Why is it usually a bad idea to use water to extinguish a cooking fire?