PRE-PUBLICATION MATERIAL
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Feature:
Unit 2: Chemical Reactions
Chapter 4: The Effects of Chemical Reactions

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The challenge in the Unit Task is to reclaim toxic copper ions from a dump site that contains mine tailings. How will you recover the copper and treat the remaining solution so that it is safe to return to the environment?

The Unit Task is described in detail on page xxx. As you work through the unit, look for Unit Task Bookmarks to see how information in the section relates to the Unit Task.
FOCUS ON STSE

GLEENING CHEMICAL PROCESSES

Look around you. Almost all the consumer products you see have been manufactured using processes that involve chemicals. It is difficult to imagine our lives without the plastics, fabrics, metals, medications, and household chemical products that the chemical industry provides. Chemical reactions play a key role in the processes that manufacture these products. However, many of these reactions have adverse effects on the environment or human health.

We are starting to better understand the hazards of the chemicals used in chemical processes. Scientists and engineers are incorporating this understanding into the manufacturing processes. Industries are now paying greater attention to the hazards of the raw materials they use and the by-products they release. Consequently, industries are adapting their processes to reduce their environmental impact. Switching to less toxic reactants and becoming more energy efficient are just two ways in which this transformation is made possible. As a result, industries that use chemicals are becoming “greener.” As they do, many are finding that they are also becoming more profitable.

“Green chemistry” is a movement to make processes involving chemicals more environmentally friendly and sustainable. Green chemistry is based on the concept, “Why pollute if there is a greener alternative?”

Developing a green alternative begins with considering the hazards of the chemicals used in a process as well as their properties. Engineers then develop a manufacturing process in which every stage of product development is environmentally safe—from the raw materials to what happens to the final product at the end of its useful life. In other words, the entire process is “benign by design.”

A green chemistry solution may involve using safer chemicals. Liquid carbon dioxide, for example, is gradually replacing toxic organic solvents used in dry cleaning. Greening a chemical process can also involve making the process more efficient by reducing the number of steps involved. For example, a new process to manufacture ibuprofen, an important pain reliever, has been developed that reduces the number of steps involved by half. The result is a process that is safer, generates less waste, uses less energy, and is more profitable!

Not all chemical processes can realistically be made totally green. However, as you progress through this unit, you will see how the principles of green chemistry can reduce the environmental footprint of these processes.

Questions
1. How do engineers attempt to make chemical processes “greener”?  
2. Identify at least one chemical process that you are familiar with. What impact does it have on the environment? How can it be made greener?  
3. What manufacturing industries do you know about that affect the environment or human health in a negative way? What aspects of these industries could perhaps be improved? Brainstorm possible solutions to these problems.
CONCEPTS
- chemical and physical properties of a substance
- the law of conservation of mass
- balanced chemical equations
- environmental effects of chemical reactions
- properties of acids and bases

SKILLS
- name or write the chemical formula of an ionic or molecular compound
- write a balanced chemical equation for a given reaction
- research and collect information
- plan and conduct experiments
- communicate scientific information clearly and accurately

Concepts Review
1. (a) What evidence of chemical change can you observe in Figure 1(a)?
(b) Hydrogen peroxide, \( \text{H}_2\text{O}_2 \), decomposes rapidly when in contact with some vegetables. What is the evidence of a chemical change in Figure 1(b)?
(c) The gas produced in Figure 1(b) causes a glowing splint to relight. Identify the gas.
(d) Write a balanced chemical equation for the decomposition of hydrogen peroxide, \( \text{H}_2\text{O}_2 \). Assume that the other reaction product is water.
(e) Write a chemical equation for the reaction occurring in Figure 1(a).
(f) What other evidence suggests that chemical change is occurring in other reactions?

2. Figure 2 shows models of the combustion of methane, \( \text{CH}_4 \).

\[
\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

Figure 2 Oxygen atoms are represented by red spheres, carbon by black, and hydrogen by white.
(a) Write a balanced chemical equation for this reaction.
(b) Distinguish between coefficients and subscripts in the chemical equation.
(c) Define the law of conservation of mass.
(d) Explain why the balanced equation for the combustion of methane obeys the law of conservation of mass.
(e) When 16 g of methane reacts with 64 g of oxygen, 44 g of carbon dioxide is produced. Predict what mass of water is produced. Explain your prediction.

3. A shiny, silver-coloured element X is a good conductor of electricity. Reacting X in chlorine produces a white solid.
(a) Classify X as a metal, non-metal, or metalloid.
(b) If X acquires an ionic charge of +2 when it reacts, what is the chemical formula of the chloride compound that it forms?
(c) If X has an atomic mass of less than 40 u, to which group of the periodic table does it belong?
Skills Review

4. Write the chemical formula for each of the following compounds. Use the periodic table on the inside back cover of the textbook and the table of ion charges in Section 2.5.

(a) potassium sulfide
(b) ammonium chlorate
(c) sodium sulfite
(d) iron(III) nitrite
(e) calcium phosphate
(f) carbon disulfide
(g) potassium dihydrogen phosphate
(h) barium hypochlorite
(i) manganese sulfate
(j) ammonium perchlorate

5. Name the following compounds.

(a) AlCl3
(b) Zn(ClO3)2
(c) PbO2
(d) (NH4)2CO3
(e) Na3PO3
(f) Ca(HCO3)2
(g) Zn(ClO3)2
(h) SO2
(i) N2O
(j) Fe(OH)3

6. Balance these chemical equations.

(a) K + O2 → K2O
(b) P4 + Cl2 → PCl3
(c) C7H8 + O2 → CO2 + H2O
(d) Fe + H2SO4 → Fe2(SO4)3 + H2
(e) HBr + Mg(OH)2 → MgBr2 + H2O
(f) Na3PO4 + CaF2 → NaF + Ca3(PO4)2
(g) (NH4)3PO4 + Sn(NO3)4 → Sn2(PO4)3 + NH4NO3

7. What safety precautions should be taken in each of the following situations?

(a) Chemicals or glassware are used in an experiment.
(b) Open flames are used.
(c) A flask containing a solution is heated on a hot plate.
(d) A flammable solvent is brought into the laboratory.
(e) A substance in a test tube is heated over a Bunsen burner.
(f) You are asked to smell a sample of a chemical.
(g) A solution is provided with a label bearing the symbol shown in Figure 3.

8. Figure 4 shows two samples of rainwater. Universal indicator has been added to both samples.

(a) Which sample is more acidic? How do you know?
(b) Which of the following chemicals could be added to make these samples less acidic: HCl(aq), NaHCO3, or NaCl?
(c) What effect would this addition have on the pH?

9. The cleanup of an acid spill often requires two steps: containment of the spill followed by treatment.

(a) What properties should the substances have that are used to contain an acid spill?
(b) Which substance would you recommend to treat an acid spill at home: sodium hydroxide, NaOH(aq) (in drain-opening products), or sodium hydrogen carbonate, NaHCO3(aq) (baking soda)? Why?

10. Compounds can be broken down or decomposed into simpler substances; elements cannot. How does this difference affect the disposal of hazardous substances?
How Do Chemical Reactions Affect Us and the Environment?

Only a few years ago, an environmentally friendly car or “eco car” was typically slow and not much to look at. Not anymore! The latest eco cars are sleek, green, and attractive. Also, going green no longer means sacrificing performance. The performance of many eco cars is now comparable to that of their conventional cousins. For example, the Tesla Roadster—an electric car—can accelerate from 0 to 100 km/h in less than 4 s and has a range of about 320 km between recharges, according to the manufacturer.

Most vehicles rely on chemical reactions as their energy source. For example, a conventional car burns gasoline in an internal combustion engine. The problem with this technology is that it pollutes the environment and is very inefficient. The electrical energy used to power an electric car comes from reversible chemical reactions inside its batteries. Electric cars are far more efficient at converting stored energy into motion. Hybrid vehicles use both gasoline and electrical motors. Other eco cars generate energy by burning an alternative fuel such as hydrogen or ethanol.

Advances in our understanding of chemical reactions have also produced the high-tech materials used in eco cars. Strong, yet lightweight materials such as carbon fibre are replacing steel as the materials of choice for automotive bodies. This not only improves their performance but also makes them virtually corrosion free. Chemical reactions also provide the means to recycle and reuse these materials once the car is scrapped.

Before deciding that an eco car is for you, consider all the pros and cons. For example, how “clean” is its energy source? The sales pitch of “zero emissions” is not a guarantee that electric cars run without producing any pollution. These vehicles are only as “clean” as their source of electricity. Another consideration is the impact on society of producing the energy source. For example, the alternative fuel ethanol is primarily made from corn. Diverting a large percentage of the annual corn crop to ethanol production affects the cost of food. And, finally, new technology is always expensive. Be prepared to pay more for the eco car. Is the extra cost worth it?

Almost every product and service that we buy has an impact on society and the environment. This chapter will help you to be more aware of the chemical reactions that are involved in these impacts so that you can make better-informed choices.
Mini Investigation

Back and Forth

Skills: Predicting, Performing, Observing, Analyzing, Evaluating, Communicating

Many chemical changes, such as the rusting of steel on a car, cannot be reversed. However, as you will see in this investigation, some chemical changes can be reversed.

Equipment and Materials: eye protection; lab apron; test tube; test-tube rack or small beaker; dropper bottles containing dilute solutions of hydrochloric acid, HCl(aq), sodium hydroxide, NaOH(aq), and iron(II) sulfate, FeSO₄(aq) ☢️

(Teacher Note: The concentration of each of the three solutions is 0.1 mol/L.)

1. Put on your eye protection and lab apron.
2. Add a solution of iron(II) sulfate, FeSO₄(aq), to a depth of about 2 cm in a test tube.
3. Add the same volume of sodium hydroxide solution, NaOH(aq), to the same test tube. Record your observations.
4. Slowly add a small volume of hydrochloric acid, HCl(aq), to the test tube until the solution is clear again.
5. Predict how you could reverse the effects of Step 4. Test your prediction by trying it. Record your observations.

A. What evidence of chemical change did you observe? ☐
B. Evaluate your prediction in Step 5. ☐
C. Suggest an explanation for your observations in Step 5. ☐

☢️ The hydrochloric acid and sodium hydroxide solution used in this activity are irritants. Wash any spills on skin or clothing immediately with plenty of cold water. Report any spills to your teacher.
Chemistry is the study of substances and the changes these substances undergo. These changes can be either physical or chemical.

During a physical change, the properties of a substance may change, but its chemical identity remains the same. For example, the rapid expansion of compressed butane gas as it is released from its container is an example of a physical change. The volume of the gas increases, but the gas is still butane. Other common physical changes are changes of state and dissolving.

Chemical changes result in the formation of new substances. For example, when butane reaches the hot nozzle of a blowtorch, it ignites, producing carbon dioxide gas and water vapour (Figure 1). It is no longer butane. This chemical change also releases the energy needed to caramelize the sugar in the dessert. The process in which one or more substances undergo a chemical change to produce one or more new substances is called a chemical reaction.

### Evidence of Chemical Reactions

All chemical reactions result in the formation of new substances. How can you tell that a new substance has been produced? Any one of the following clues may indicate that a chemical change has produced one or more new substances:

- There is a change in colour.
- Energy is released or absorbed.
- A gas is produced.
- A precipitate forms.

Two of these clues apply to the work of the chef in Figure 1. Butane is a colourless gas. However, as it burns, it produces an intense blue flame that radiates a great deal of thermal energy. Both of these observations suggest that the combustion of butane is a chemical reaction. Similarly, the darkening of the surface of the dessert indicates that caramelizing sugar is likely a chemical reaction.

A precipitate is a solid produced when two liquids are combined (Figure 2).

None of these four clues are conclusive proof that a chemical reaction has taken place. For example, heating water until it boils produces a gas, but no new substance is produced, so this is not a chemical change. The only way to be certain that a chemical change has occurred is to test the product to show that it is a new substance.

### Atoms in Chemical Reactions

We can use an example to explore what happens at the molecular level during a chemical reaction. Hydrogen peroxide, \( \text{H}_2\text{O}_2 \), decomposes quickly in the presence of manganese(IV) oxide to produce water and a colourless gas (Figure 3(a)).

A test can help to identify the gas that bubbles out of the liquid. A glowing splint inserted in the gas relights, suggesting that it is oxygen.

Manganese(IV) oxide is not used up or produced in the reaction; it is neither a reactant nor a product. Instead, it is a catalyst. A catalyst is a substance that makes a reaction occur faster without being used up in the reaction.

For every two molecules of hydrogen peroxide that react, one molecule of oxygen and two molecules of water are produced (Figure 3(b)). For this to occur, the bonds within the hydrogen peroxide molecules must first break, allowing the atoms to rearrange and form new bonds. All the atoms that were present at the start of the reaction must also be present at the end. Therefore, the total mass of the reactant (hydrogen peroxide) must equal the total mass of the products (water and oxygen). This is true for all chemical reactions and is called the law of conservation of mass.
Figure 3 (a) Hydrogen peroxide breaks down in the presence of black manganese dioxide. The products are water and oxygen. (b) All the atoms in the reactants are accounted for in the products. Notice the subscripts, indicating the numbers of atoms in each molecule.

**Mini Investigation**

**Elephant Toothpaste**

Skills: Controlling Variables, Performing, Observing, Analyzing, Evaluating, Communicating

Hydrogen peroxide can be purchased in a variety of concentrations, depending on their application. Drugstores sell 3% hydrogen peroxide as a disinfectant for cuts. Beauty supply stores sell 6% and 12% hydrogen peroxide for use with hair-coloring products. In this activity, you will compare the effect of the concentration of a hydrogen peroxide solution on its reactivity.

**Equipment and Materials:** eye protection; lab apron; 2 large narrow test tubes; 2 stirring rods; small beaker; scoopula; candle; 3% hydrogen peroxide solution; 6% hydrogen peroxide solution; masking tape; marker; liquid dish detergent; dry yeast; wooden splint

1. Put on your eye protection and lab apron.
2. Pour 3% hydrogen peroxide into one test tube to a depth of about 3 cm and the same amount of 6% hydrogen peroxide into the other test tube. Label each test tube appropriately.
3. Add about 5 drops of liquid dish detergent to each test tube.
4. Use a different stirring rod to mix the contents of each test tube.
5. Place the test tubes in the beaker.
6. Place the beaker and its contents in the sink to catch any spills.
7. Add enough yeast to cover the tip of the scoopula to each test tube.
8. Light your splint from the lit candle. Test the gas produced with a glowing splint.

**LEARNING TIP**

An open flame is used. Tie back long hair and secure loose clothing. Never leave the flame unattended.

9. Follow your teacher’s instructions for proper disposal.

A. What evidence did you observe that chemical reactions were occurring?
B. What variables were controlled in this activity? What variables were changed?
C. Why do you think hair salons use hydrogen peroxide solutions that are more concentrated than 3%? What precautions should the technicians in salons follow when using these products?

**Describing Chemical Reactions**

Chemists use both word equations and chemical equations to describe chemical reactions. Both types of equations list reactants on the left of an arrow and products on the right. A word equation gives the names of the reactants and products. A chemical equation, however, provides far more detail: it gives the chemical formulas of the reactants and products, their state (Table 1), and specific conditions required for the reaction to occur. The chemical equation also gives the ratio in which the chemicals react. This is done through coefficients placed before each chemical formula. A coefficient of “1” is implied if no coefficient is written. For example, the word and chemical equations for the decomposition of hydrogen peroxide are

**Word equation:** hydrogen peroxide $\xrightarrow{\text{(catalyst)}}$ water + oxygen + energy

**Chemical equation:** $2 \text{H}_2\text{O}_2(aq) \xrightarrow{\text{MnO}_2, \text{(catalyst)}} 2 \text{H}_2\text{O}(l) + \text{O}_2(g) + \text{energy}

This chemical equation is “balanced” because the total number of atoms of each type in the reactants is the same as in the products. In particular, there are four atoms of hydrogen and four atoms of oxygen on both sides of the arrow.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
</tr>
</thead>
<tbody>
<tr>
<td>(s)</td>
<td>solid</td>
</tr>
<tr>
<td>(l)</td>
<td>liquid</td>
</tr>
<tr>
<td>(g)</td>
<td>gas</td>
</tr>
<tr>
<td>(aq)</td>
<td>aqueous (dissolved in water)</td>
</tr>
</tbody>
</table>

**LEARNING TIP**

Balanced Chemical Equations

A chemical equation is balanced if the total number of atoms of each type is the same on both sides of the equation.
Tutorial 1 Balancing Chemical Equations

Follow these useful strategies when balancing chemical equations. For simplicity, the state symbols have been omitted from these examples.

- Start by writing a skeleton equation. This is an equation that includes only the chemical formulas, without coefficients.
- Balance atoms that appear only once on each side of the equation.
- Leave atoms that appear more than once to the end.
- Treat polyatomic ions as one unit, rather than as individual atoms, providing that they do not change during the reaction.
- Check that the final equation is balanced.

Sample Problem 1: Balancing a Chemical Equation

Write the balanced chemical equation for the reaction of sodium in oxygen to produce sodium oxide (Figure 4).

Step 1. Write a skeleton chemical equation for the reaction.

\[ \text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O} \]

Step 2. Count the number of atoms of each type on either side of the arrow.

\[ \begin{align*}
\text{Na} & : 1 \text{ Na atom} \\
\text{O}_2 & : 2 \text{ O atoms} \\
\text{Na}_2\text{O} & : 2 \text{ Na atoms} 1 \text{ O atom}
\end{align*} \]

Step 3. Multiply the formulas by an appropriate coefficient until all the atoms are balanced. Keep checking whether the numbers of each type of atom on both sides are balanced.

Na\(_2\)O (on the right) must be multiplied by 2 to balance the two oxygen atoms on the left side.

\[ \text{Na} + \text{O}_2 \rightarrow 2 \text{Na}_2\text{O} \]

Na (on the left) must be multiplied by 4 to balance Na on the right side.

Step 4. Write the final chemical equation.

\[ 4 \text{ Na} + \text{O}_2 \rightarrow 2 \text{Na}_2\text{O} \]

Sample Problem 2: Balancing a Chemical Equation Involving Polyatomic Ions

Write the balanced chemical equation for the reaction of copper in silver nitrate to produce copper(II) nitrate and silver.

Step 1. Write a skeleton chemical equation for the reaction.

\[ \text{Cu} + \text{AgNO}_3 \rightarrow \text{Cu(NO}_3\text{)}_2 + \text{Ag} \]

Step 2. Count the number of atoms of each type on either side of the arrow.

\[ \begin{align*}
\text{Cu} & : 1 \text{ Cu atom} \\
\text{Ag}^{+} & : 1 \text{ Ag atom} \\
\text{NO}_3^{-} & : 1 \text{ NO}_3^{-} \\
\text{Cu}^{2+} & : 1 \text{ Cu}^{2+} \\
\text{Ag}^{+} & : 2 \text{ Ag atom} \\
\text{NO}_3^{-} & : 2 \text{ NO}_3^{-}
\end{align*} \]

Step 3. Check whether the numbers of atoms and ions are the same on both sides. Since NO\(_3^{-}\) remains intact, it can be counted as a unit. AgNO\(_3\) must be multiplied by 2 to balance nitrates. Therefore, silver must also be multiplied by 2.

\[ \begin{align*}
\text{Cu} & : 1 \text{ Cu}^{2+} \\
\text{Ag}^{+} & : 2 \text{ Ag atom} \\
\text{NO}_3^{-} & : 2 \text{ NO}_3^{-}
\end{align*} \]
Step 4. Write the final chemical equation.
\[ \text{Cu} + 2 \text{AgNO}_3 \rightarrow \text{Cu(NO}_3)_2 + 2 \text{Ag} \]

Practice

1. Balance the following chemical equations.
   (a) \( \text{P} + \text{O}_2 \rightarrow \text{P}_2\text{O}_5 \)
   (b) \( \text{K}_2\text{O} + \text{H}_2\text{O} \rightarrow \text{KOH} \)
   (c) \( \text{AlBr}_3 + \text{K}_2\text{SO}_4 \rightarrow \text{KBr} + \text{Al}_2(\text{SO}_4)_3 \)
   (d) \( \text{FeCl}_3 + \text{NaOH} \rightarrow \text{Fe(OH)}_3 + \text{NaCl} \)
   (e) \( \text{AgNO}_3 + \text{H}_2\text{S} \rightarrow \text{Ag}_2\text{S} + \text{HNO}_3 \)
   (f) \( (\text{NH}_4)_2\text{CO}_3 \rightarrow \text{NH}_3 + \text{H}_2\text{O} + \text{CO}_2 \)

UNIT TASK BOOKMARK

You will need to balance chemical equations in the Unit Task, described on page XXX.

4.1 Summary

- Evidence of a chemical reaction includes colour change; absorption or release of energy; production of a gas (except evaporating or boiling of a liquid); and formation of a precipitate.
- During a chemical reaction, reactant atoms rearrange to form products.
- Chemical reactions are described using either word equations or chemical equations.
- A balanced chemical equation gives the correct proportions of chemicals in a chemical reaction. As a result, it obeys the law of conservation of mass.

4.1 Questions

(a) Write a word equation for this reaction.
(b) Write a chemical equation, including all state symbols.
(c) What is the subscript for oxygen in nitric acid?
(d) What is the coefficient of nitric acid in the equation?
(e) How many atoms of oxygen appear on the left-hand side of the equation?
(f) Distinguish between the symbols (l) and (aq) used in this equation.

4. Why are the coefficients, but not the subscripts, sometimes changed when balancing a chemical equation?

5. Balance the following chemical equations.
   (a) \( \text{S}_8 + \text{O}_2 \rightarrow \text{SO}_2 \)
   (b) \( \text{N}_2 + \text{H}_2 \rightarrow \text{NH}_3 \)
   (c) \( \text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2 \)
   (d) \( \text{Li} + \text{AlCl}_3 \rightarrow \text{LiCl} + \text{Al} \)
   (e) \( \text{C}_2\text{H}_10 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)
   (f) \( \text{N}_2 + \text{O}_2 \rightarrow \text{N}_2\text{O}_5 \)
   (g) \( \text{Li} + \text{B}_2\text{O}_3 \rightarrow \text{Li}_2\text{O} + \text{B} \)
   (h) \( \text{Fe}_2\text{O}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Fe}_2(\text{SO}_4)_3 + \text{H}_2\text{O} \)
   (i) \( \text{H}_3\text{PO}_4 + \text{Ca(OH)}_2 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + \text{H}_2\text{O} \)
   (j) \( \text{NH}_3 + \text{O}_2 \rightarrow \text{N}_2 + \text{H}_2\text{O} \)
   (k) \( \text{Ca}_3(\text{PO}_4)_2 + \text{SiO}_2 + \text{C} \rightarrow \text{CaSiO}_3 + \text{CO} + \text{P} \)
   (l) \( \text{C}_6\text{H}_6 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \)

Figure 5 (a) Copper metal reacts in a silver nitrate solution to produce impure silver and dissolved copper nitrate. Dissolved copper(II) ions are blue. (b) A match flares as it catches fire.
4.2

Synthesis and Decomposition Reactions

In the previous unit, you learned that the periodic table is a useful tool to predict the properties of elements. Because the properties of elements are predictable, many of the chemical reactions that these elements undergo are predictable as well. For example, Figure 1 shows the reaction of two metals—sodium and potassium—with chlorine. Because sodium and potassium are both alkali metals, they undergo similar chemical reactions to form chloride compounds. The chemical equations for these reactions are also similar: two elements react to form one compound.

Figure 1 Alkali metals are highly reactive elements. They react with chlorine to form very stable compounds.

We can use this pattern to predict the products of other reactions involving a metal and chlorine. Magnesium, for example, also reacts with chlorine. The product of this reaction is the compound magnesium chloride. Magnesium has an ionic charge of $+2$, so the chemical formula of this compound is MgCl$_2$. Therefore, the balanced chemical equation for this reaction is

$$\text{Mg}(s) + \text{Cl}_2(g) \rightarrow \text{MgCl}_2(s)$$

There are millions of known reactions. Chemists group them into patterns to organize them. Grouping reactions also makes the prediction of reaction products much simpler. In the next few sections, you will learn to recognize five patterns that chemists use to classify chemical reactions. Note, however, that there are many ways in which reactions can be classified. Using these patterns is just one of the ways.

**Synthesis Reactions**

The reactions in Figure 1 are examples of synthesis reactions. In a **synthesis reaction**, two simple reactants combine to form one larger or more complex product. For this to occur, the reactants must first collide, break existing bonds between their atoms, and form new bonds. The chemical equations for synthesis reactions fit the general pattern:

$$\text{A} + \text{B} \rightarrow \text{AB}$$

**Figure 2** A model of a typical synthesis reaction

Can we use this pattern to predict the products of a synthesis reaction?
All the reactions that we have considered so far involve metals reacting with non-metals to form ionic compounds. Some reactions, however, involve only non-metals.

**Synthesis Reactions of Non-Metals**
We can apply the pattern of synthesis reactions to this situation. We expect that the product will be one or more molecular compounds since non-metals are combining.

**SYNTHESIS REACTIONS INVOLVING HYDROGEN**
Hydrogen is unusual for a Group 1 element. Most Group 1 elements form ionic compounds when they bond with other elements. Hydrogen, however, usually forms molecular compounds. Molecular compounds are made up of atoms, not ions. However, ionic charges are still useful in predicting the products of synthesis reactions involving hydrogen (Figure 4). For example, the reaction between hydrogen, \( \text{H}_2 \), and chlorine, \( \text{Cl}_2 \), follows the familiar pattern of synthesis reactions:

\[
\text{A} + \text{B} \rightarrow \text{AB}
\]

\[
\text{H}_2(g) + \text{Cl}_2(g) \rightarrow \text{H}_2\text{Cl}_2
\]

If we apply the ionic charges of hydrogen (+1) and chlorine (−1), the product of the reaction is hydrogen chloride, \( \text{HCl} \). Therefore, the balanced chemical equation for this reaction is

\[
\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2 \text{HCl}(g)
\]
The reaction between hydrogen and oxygen also follows the pattern of synthesis reactions:

\[ A + B \rightarrow AB \]

\[ \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}_y \]

If the ionic charges of hydrogen (+1) and oxygen (−2) are applied, we can see that two hydrogen atoms are needed to balance the ionic charge of one oxygen atom. The product of the reaction is \( \text{H}_2\text{O} \). Therefore, the balanced chemical equation for this reaction is

\[ 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g}) + \text{energy} \]

The reaction of hydrogen with oxygen releases a great deal of energy, as well as water. This reaction has the potential to be an important source of “green energy.” In fact, a fleet of hydrogen buses was used in Vancouver during the 2010 Winter Olympics (Figure 5). However, since hydrogen does not occur naturally as an element, it has to be extracted from compounds that contain hydrogen. This, of course, requires energy. Therefore, hydrogen fuel can truly be considered “green” only if it is generated from environmentally friendly raw materials using a renewable energy source. Water would be a suitable raw material. Solar power would be a suitable energy source.

**SYNTHESIS REACTIONS NOT INVOLVING HYDROGEN**

The products of synthesis reactions of non-metals other than hydrogen are difficult to predict. The products of these reactions often depend on the reaction conditions. For example, two products are possible for the reaction of carbon with oxygen, depending on the availability of oxygen:

\[ \text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \]

\[ 2 \text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2 \text{CO}(\text{g}) \]

In these cases, the only way to identify the reaction products is by conducting chemical tests.

**Synthesis Reactions Involving Compounds**

So far, we have only considered synthesis reactions involving elements as reactants. However, compounds also can participate in synthesis reactions. For example, bubbling carbon dioxide into a mixture of water and bromothymol blue indicator causes the indicator to change colour from blue to yellow (Figure 6).

Bromothymol blue is an acid–base indicator. It is blue when in a basic solution and yellow when in an acidic solution. The colour change in Figure 6 indicates that an acid has formed. In this case, the combination of water and carbon dioxide produces carbonic acid:

\[ \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) \rightarrow \text{H}_2\text{CO}_3(\text{aq}) \]

Air contains a small percentage of carbon dioxide. Whenever water is in contact with air, some of the carbon dioxide dissolves in the water and reacts to produce carbonic acid. The synthesis of carbonic acid makes “normal” rain slightly acidic. The same process also acidifies seawater. The recent increase in carbon dioxide in the atmosphere has produced a significant increase in the acidity of the world’s oceans. This increased acidity could have devastating effects (Figure 7). Coral is made mostly of calcium carbonate. This calcium carbonate will react with the acid in the ocean to produce soluble compounds. Coral reefs will be “eaten away.”
Sulfur trioxide is another compound that is affecting the environment because of a synthesis reaction. It is produced during the combustion of fossil fuels. Water in the environment reacts with sulfur trioxide in a synthesis reaction to form sulfuric acid:

\[ \text{H}_2\text{O}(l) + \text{SO}_3(g) \rightarrow \text{H}_2\text{SO}_4(aq) \]

The product of each of the previous two examples could be predicted simply by combining the two reactant chemical formulas. Unfortunately, we cannot predict the products of all synthesis reactions so simply. The products that actually form depend on the reaction conditions.

### Decomposition Reactions

Special effects are an important part of any action movie. Many effects are computer generated, but others are the result of chemical reactions. The reactants are explosives—placed in just the right spot to produce the desired effect—and oxygen (Figure 8). The science of handling explosives is called pyrotechnics and, as you would expect, requires highly specialized training. Explosives experts do not only work on Hollywood movie sets. Explosives are also used for clearing rock for road construction and mining.

![Figure 8](image)

**Figure 8** Carefully placed explosives produce spectacular movie effects such as vehicle rollovers. An explosion is a rapid chemical reaction.

Many explosives are molecular compounds of four non-metal elements: carbon, hydrogen, nitrogen, and oxygen. TNT, for example, is more correctly called trinitrotoluene, \( \text{C}_7\text{H}_5(\text{NO}_2)_3 \). It is a common industrial explosive. The detonation of this compound releases a great deal of energy and several reaction products:

\[ 2 \text{C}_7\text{H}_5(\text{NO}_2)_3(s) \rightarrow 7 \text{C}(s) + 7 \text{CO}(g) + 3 \text{N}_2(g) + 5 \text{H}_2\text{O}(g) + \text{energy} \]

A chemical reaction in which one large compound breaks down or decomposes into two or more smaller substances is called a **decomposition reaction**. The general pattern for a decomposition reaction is

\[ \text{AB} \rightarrow \text{A} + \text{B} \]

![Figure 9](image)

**Figure 9** Decomposition reactions produce two or more smaller substances.
Decomposition reactions usually need energy to get started—even for reactions that release a lot of energy. In the case of TNT, a small electric current starts the reaction. It is not only the release of energy that makes TNT so dangerous, however. The source of TNT's destructive power becomes clear when you consider the states of its decomposition products. Most of these substances are gases. Gases occupy far more space than solids and expand very quickly when they are heated. The thermal energy released by exploding TNT makes these gases expand very quickly, creating a powerful destructive force a fraction of a second after detonation.

**PREDICTING THE PRODUCTS OF DECOMPOSITION REACTIONS**

Simple ionic compounds such as potassium chloride, KCl, can be made to decompose into their elements. For example, potassium metal and chlorine gas are made industrially by passing electricity through molten potassium chloride:

\[ 2 \text{KCl(s)} \xrightarrow{\text{electricity}} 2 \text{K(s)} + \text{Cl}_2(\text{g}) \]

However, predicting the decomposition products of most molecular compounds or ionic compounds involving polyatomic ions is difficult. For example, you might expect potassium chlorate, KClO₃, to decompose into its elements when heated (Figure 10). Instead, the decomposition products are potassium chloride and a colourless gas. A glowing splint relights when placed in the gas, identifying it as oxygen. The chemical equation for the decomposition of potassium chlorate is therefore

\[ 2 \text{KClO}_3(\text{s}) \xrightarrow{\text{heat}} 2 \text{KCl(s)} + 3 \text{O}_2(\text{g}) \]

Some decomposition reactions produce only compounds. For example, the intense heat of a Bunsen burner flame causes calcium carbonate, CaCO₃, to decompose into calcium oxide, CaO, and an invisible gas. Bubbling this gas into limewater produces a precipitate, indicating that the gas is carbon dioxide. Therefore, the chemical equation for the decomposition of calcium carbonate is

\[ \text{CaCO}_3(\text{s}) \xrightarrow{\text{heat}} \text{CaO(s)} + \text{CO}_2(\text{g}) \]

Calcium oxide, also known as lime, is a key ingredient in cement. Industrially, kilns operating above 1400 °C are used to decompose limestone (CaCO₃) to produce calcium oxide (Figure 11). Kilns consume huge quantities of fossil fuels to maintain this temperature. Fossil fuel combustion and the decomposition of calcium carbonate both generate carbon dioxide—a greenhouse gas. Not surprisingly, cement kilns are a significant source of greenhouse gas emissions.

**Laughing Gas**

In 1799, Humphrey Davy first produced nitrous oxide, N₂O(g), by decomposing ammonium nitrate:

\[ \text{NH}_4\text{NO}_3(\text{s}) \rightarrow 2 \text{H}_2\text{O}(\text{g}) + \text{N}_2\text{O}(\text{g}) \]

Davy found that inhaling this gas made people giddy and laugh. Davy was one of the most prominent scientists of his time. He popularized chemistry by putting on lavish shows for the British upper class where his audience could test the effects of “laughing gas.” Today, dentists use laughing gas to relax their patients.

**Figure 10** The head of a match contains potassium chlorate, among other chemicals. Oxygen, a product of the decomposition reaction, makes the match burn faster.

**Figure 11** Cement kilns must operate at high temperatures to decompose limestone (calcium carbonate).
4.2 Summary

- Chemical reactions can be grouped according to the pattern of reactants and products in their chemical equations.
- In a synthesis reaction, two simple reactants combine to make a larger or more complex compound and follow the general pattern
  \[ A_1 + B \rightarrow AB \]
- In a decomposition reaction, a complex compound breaks down into two or more simpler products and follows the general pattern
  \[ AB \rightarrow A_1 + B \]

4.2 Questions

1. Classify these reactions as synthesis or decomposition. Justify your choice.
   - (a) \( 2 \text{Al} + 3 \text{Br}_2 \rightarrow 2 \text{AlBr}_3 \)
   - (b) \( 2 \text{HCl} \rightarrow \text{H}_2 + \text{Cl}_2 \)
   - (c) \( \text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 \)
   - (d) \( \text{P}_4 + 5 \text{O}_2 \rightarrow 2 \text{P}_2\text{O}_5 \)

2. Predict the products of these synthesis or decomposition reactions. Write a balanced chemical equation to represent each reaction.
   - (a) \( Z + S \rightarrow \)
   - (b) \( \text{CaCl}_2 \rightarrow \)
   - (c) \( \text{NH}_3 + \text{HCl} \rightarrow \)

3. Hydrogen peroxide forms gas bubbles when it is added to blood. The other reaction product is water. Inserting a glowing splint into a sample of this gas causes the splint to relight. Identify the gas.

4. Write a balanced chemical equation for each of these reactions.
   - (a) Aluminum metal readily reacts in air to form a hard protective coating of aluminum oxide.
   - (b) Copper(II) oxide and carbon dioxide are produced when copper(II) carbonate is heated.
   - (c) Solid nitrogen triiodide is a shock-sensitive explosive that is stable when wet and explosive when dry. This compound decomposes rapidly to produce gases when detonated.

5. Photosynthesis is a chemical process that occurs in green plants in which solar energy is converted into stored chemical energy. This process is the basis of life on Earth. The overall chemical equation for photosynthesis is
  \[ 6 \text{CO}_2(g) + 6 \text{H}_2\text{O}(l) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(s) + 6 \text{O}_2(g) \]
  Compare photosynthesis to the synthesis reactions discussed in this section.

6. Most fossil fuels contain traces of sulfur. When the fuel is burned, the sulfur impurities become a major cause of acid precipitation. Three chemical reactions are involved. Write a balanced chemical equation for each reaction.
   - (a) Solid sulfur reacts with oxygen to form gaseous sulfur dioxide.
   - (b) The product from (a) then further reacts with oxygen to form gaseous sulfur trioxide.
   - (c) The product from (b) then reacts with droplets of water in the atmosphere to form sulfuric acid.

7. Sodium hydrogen carbonate (baking soda) easily decomposes to release sodium carbonate, water, and a gas that makes limewater cloudy. Cake batter rises because of the decomposition of sodium hydrogen carbonate. What gas is produced? Write a balanced chemical equation for this reaction.

8. A red solid is heated in a test tube. A colourless gas that relights a glowing splint is collected from the test tube. After the test tube cools, a silver-coloured liquid metal remains. Further testing shows that the liquid cannot be decomposed. What type of chemical reaction occurred in the test tube? Why? Identify the gaseous product. Identify the metal. Name and write the chemical formula for the red solid. Write a balanced chemical equation for this reaction.

9. The decomposition of solid sodium azide, NaN\(_3\), is the chemical reaction that inflates automobile airbags during a collision. Write a chemical equation for this reaction. Assume that the first products formed are both gases.

10. Cement kiln operators sometimes supplement the fuel used in cement kilns with old car tires. Research the potential environmental benefits and threats of this practice. Communicate your findings in a Pros and Cons chart. In your opinion, should the practice be continued or banned?
Ontario’s garbage is mounting. Despite aggressive recycling and composting campaigns, we are producing more garbage than we know what to do with. So far, most garbage that is not recycled is dumped in landfill sites (Figure 1). However, existing landfill sites are rapidly filling up and the opposition to opening new sites is intense. We need to find an alternative way of getting rid of waste.

Many municipalities are investigating gasification as a possible solution to their garbage dilemma. Gasification is not simply burning the garbage. Gasification involves heating waste to temperatures high enough to cause their molecules to decompose into simpler substances. This process is already being used to convert toxic industrial and hospital waste into safer substances.

Plasma gasification is a particular gasification technology that involves heating waste in a sealed chamber in which there is almost no oxygen (Figure 2). Under these conditions, large carbon-based molecules, such as those in plastics, decompose into “syngas.” Syngas is a mixture composed mostly of hydrogen and carbon monoxide. Chemical treatment or “scrubbing” removes contaminants from the syngas.

Cleaned syngas is an energy-rich mixture that burns like a fossil fuel and is used to generate electricity. The amount of electricity that it generates far exceeds the energy requirements of the gasification plant. The excess energy would be available to be sold.

Alternatively, the components in syngas can be used to make other carbon-based industrial chemicals, such as ethanol, C₂H₅OH.

Plasma gasification uses high-energy electrical sparks called plasmas to achieve temperatures as high as 10 000 °C—hotter than the surface of the Sun. The energy is so intense that it literally rips molecules apart and forces atoms to ionize. It is somewhat like placing waste into a continuous lightning bolt (Figure 3).

**Figure 1** Hundreds of truckloads of Ontario garbage arrive at landfill sites each day. The dilemma of what to do with this mountain of mess remains a topic of hot debate in many municipalities.

**Figure 2** Plasma gasification furnace

Plasma gasification uses high-energy electrical sparks called plasmas to achieve temperatures as high as 10 000 °C—hotter than the surface of the Sun. The energy is so intense that it literally rips molecules apart and forces atoms to ionize. It is somewhat like placing waste into a continuous lightning bolt (Figure 3).

**Figure 3** Plasma is a high-energy mixture of ionized particles and electrons. In plasma gasification, the energy of the plasma decomposes the molecules in waste into simpler substances.
Critics of gasification are concerned about its emissions. Previous attempts to burn waste produced toxic by-products that escaped into the environment. In particular, dioxin is a highly toxic compound produced when plastics containing chlorine are burned. Critics also argue that the convenience of “zapping away” garbage undermines progress Ontarians have made in recycling. The “buy today and destroy tomorrow” mentality is not sustainable, given that our planet has finite resources. Finally, critics maintain that there are always risks associated with implementing a new technology. Are the potential benefits worth it?

The Issue
The Ontario city of Greenborough is considering adding a gasification plant to its waste management strategy. Recycling programs have been well received and are successful. However, every year the city still generates about 80,000 t of waste that cannot be composted or recycled. Furthermore, landfill sites are almost full. Attempts to incinerate waste 20 years ago failed miserably. The stench and fumes from that incinerator still linger in citizens' memories.

Your environmental consulting team has been asked to prepare a report outlining the advantages and disadvantages of including gasification as part of the city's waste management strategy. Your report will be presented to Greenborough City Council.

Goal
To decide whether a gasification facility should be built in the city

Research
Work in pairs or in small groups to learn more about waste gasification. Research these questions:
- Compare plasma gasification to one other gasification technology.
- What toxic emissions or waste products does gasification generate?
- How successful are gasification technologies at treating their emissions and waste?
- What other disadvantages are associated with gasification?
- What other cities are currently using this technology and how successful has it been?

Identify Solutions
Is waste gasification a viable solution to deal with a portion of Greenborough's garbage? If so, which type of gasification process should be selected? If not, what alternative should be investigated?

Make a Decision
Which solution does your firm recommend? Outline the rationale you used to reach your decision.

Communicate
Complete a presentation of your findings that will be shared with City Council. State your recommendation and your reasons for it. Remember that only a few of the members of your target audience are scientists or engineers. However, these people will make the final decision on whether to proceed with your recommendation.

Plan for Action
Inform yourself about the garbage situation in your region. How does your municipality currently dispose of garbage? Is there a garbage crisis in your region? If so, what possible solutions have been proposed?
4.4 Single Displacement Reactions

Many of us would not want to be without our cellphones. But for an increasing number of Canadians, cellphones are becoming literally a “pain in the ear.” These people suffer from nickel contact dermatitis—a skin allergy to nickel (Figure 1). The source of the nickel is the shiny metal parts of the cellphone that are in frequent contact with the skin: the buttons and outer rim. This metal contains nickel. Nickel ions gradually leave the cellphone and transfer to the skin. Over time, constant exposure to nickel may make the skin sensitive to it. Once sensitized, the next exposure to nickel can result in an allergic reaction.

Figure 1 (a) and (b) Cellphones are sometimes responsible for serious skin problems. (c) A chemical test shows whether nickel ions are present.

All skin naturally produces secretions, such as sweat, that contain acidic compounds like lactic acid and amino acids. These compounds make sweat mildly acidic. The acidity of skin is sufficient to slowly corrode nickel from the cellphone. We can use the general formula HA(aq) to represent these acidic compounds. Therefore, the chemical equation for this reaction is

\[ \text{Ni(s)} + 2 \text{HA(aq)} \rightarrow \text{NiA}_2(aq) + \text{H}_2(g) \]

**Elements Changing Places**

The reaction of nickel with an acid follows the same pattern as other reactions. These reactions are called single displacement reactions. In a **single displacement reaction**, one element displaces or replaces an element in a compound (Figure 2). The general pattern for this type of reaction is

\[ \text{A + BC} \rightarrow \text{AC} + \text{B} \]

Figure 2 In a single displacement reaction, one element, A, displaces another element, B, from a compound, BC. The products are a new compound, AC, and the displaced element, B.

Single displacement reactions are useful when you need to recover a dissolved metal from solution. For example, copper(II) sulfate is a common chemical used in high school labs. Each year, schools generate a considerable volume of waste copper(II) sulfate. Unfortunately, dissolved copper ions are toxic and must be treated before disposal. One treatment option is to add a piece of reactive metal, such as zinc, to the solution. A single displacement reaction occurs in which zinc displaces copper from the solution (Figure 3). The equation for this reaction is

\[ \text{Zn(s)} + \text{CuSO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{Cu(s)} \]

**LEARNING TIP**

Chemistry Shorthand
Chemists sometimes use abbreviations to represent groups of compounds or even specific compounds. For example, HA means “any acid” and HC means “any hydrocarbon.” Do not confuse these abbreviations with chemical symbols.

**single displacement reaction** a reaction in which an element displaces another element in a compound, producing a new compound and the new element; general pattern: \( \text{A + BC} \rightarrow \text{AC} + \text{B} \)

Figure 3 The reaction of zinc with a solution of copper(II) sulfate. The original blue colour of the copper(II) sulfate solution fades as copper ions are displaced out of the solution to form copper metal.
Zinc metal also reacts with solutions of many other metal compounds. In other words, zinc is a reactive metal. Gold, on the other hand, rarely reacts when placed in solutions. This property explains why gold is one of the few metals that are found as elements in nature. Most other metals exist only as compounds. Chemists use the word “activity” to describe the reactivity of a metal. The more reactive the metal, the greater its activity is. Zinc is a more reactive metal than gold.

The Activity Series of Metals

In Investigation 1.4.1, The Search for Patterns, you compared the activity of different metals. Chemists have used the results of experiments like this to develop an activity series: a ranking of the reactivity of metals relative to each other (Figure 4). The most reactive metals—lithium and potassium—are at the top of the activity series. The least reactive metals are at the bottom. Hydrogen is included in the series, even though it is not a metal, because it forms a positively charged ion like a metal. Including hydrogen also makes it easier to predict which metals react in acids and water. The activity series is based on two important generalizations:

1. One element can displace elements below it from compounds in solution but cannot displace elements above it.
2. The farther apart two elements are, the more likely it is that the displacement reaction will occur quickly.

Sample Problem 1: Reactions Involving a Metal and an Ionic Compound

Using the activity series, write a balanced chemical equation for the reaction that occurs when magnesium metal is placed in a solution of copper(II) sulfate. If you predict that no reaction occurs, write no reaction.

Step 1. Write the chemical formulas of the reactants.

\[
\text{Mg(s)} + \text{CuSO}_4(\text{aq}) \rightarrow \text{A} + \text{BC} \rightarrow \text{AC} + \text{B}
\]

Step 2. Identify which metal is higher on the activity series. If the higher (more active) metal is the element, the reaction proceeds. If the higher metal is in the compound, no reaction occurs.

The metal element, Mg, is higher than the metal in the compound, Cu. Therefore, a displacement reaction will occur.

Step 3. Determine the chemical formula of the new compound AC.

\[
\text{MgSO}_4
\]

Step 4. Write the balanced chemical equation for the reaction.

\[
\text{Mg(s)} + \text{CuSO}_4(\text{aq}) \rightarrow \text{MgSO}_4(\text{aq}) + \text{Cu(s)}
\]

Sample Problem 2: Reactions Involving a Metal and an Ionic Compound

Using the activity series, write a balanced chemical equation for the reaction that occurs when lead is placed in a zinc nitrate solution. If you predict that no reaction occurs, write no reaction.
Step 1. Write the chemical formulas of the reactants.

\[ \text{Pb(s)} \quad \text{Zn(NO}_3\text{)}_2(\text{aq}) \]

\[ \text{A} + \text{BC} \rightarrow \text{AC} + \text{B} \]

Step 2. Identify which metal is higher on the activity series. If the higher metal is the element, the reaction proceeds. If the higher metal is in the compound, no reaction occurs.

The metal in the compound, Zn, is higher than the metal element, Pb. Therefore, no displacement reaction occurs.

No reaction

Practice

1. Use the activity series to write a chemical equation for the following reactions that actually occur. Write no reaction if you predict that no reaction occurs.

   (a) \( \text{Mg(s)} + \text{AgNO}_3(\text{aq}) \rightarrow \)
   (b) \( \text{Zn(s)} + \text{FeCl}_2(\text{aq}) \rightarrow \)
   (c) \( \text{Ni(s)} + \text{Al(NO}_3\text{)}_3 \rightarrow \)

Reactions Involving Metals and Water or Acids

Not all single displacement reactions involve a metal in an aqueous solution of an ionic compound. Sometimes the reactants are a metal and a compound that contains hydrogen. Among the most familiar hydrogen compounds are water and acids (Figure 5). The same procedure outlined above can be used to predict displacement reactions involving acids and water. These tips will help:

1. Treat acids as if they were ionic compounds, e.g.,

   \[ \text{HCl(aq)} = \text{H}^+ + \text{Cl}^- \]
   \[ \text{H}_2\text{SO}_4(\text{aq}) = \text{H}^+ + \text{H}^+ + \text{SO}_4^{2-} \]

2. Treat water as if it were an ionic compound:

   \[ \text{H}_2\text{O} = \text{H}^+ + \text{OH}^- \]

Sample Problem 3: Reactions Involving a Metal and an Acid

Use the activity series to write a balanced chemical equation for the reaction that occurs when lithium is added to sulfuric acid. If you predict that no reaction occurs, write no reaction.

Step 1. Write the chemical formulas of the reactants.

\[ \text{Li(s)} \quad \text{H}_2\text{SO}_4(\text{aq}) \]

\[ \text{Li(s)} \quad \text{H}^+ + \text{H}^+ + \text{SO}_4^{2-} (\text{aq}) \]

\[ \text{A} + \text{BC} \rightarrow \text{AC} + \text{B} \]

Step 2. Identify which is higher on the activity series: the element or hydrogen. If the element is higher, the reaction proceeds. If the hydrogen is higher, no reaction occurs.

Li is higher than H. Therefore, a displacement reaction does occur.

Step 3. Determine the chemical formula of the new compound AC.

\[ \text{Li}^+ \quad \text{SO}_4^{2-} \]

\[ \text{Li}_2\text{SO}_4 \]

Step 4. Write the balanced chemical equation for the reaction.

\[ \text{Li(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Li}_2\text{SO}_4(\text{aq}) + \text{H}_2(\text{g}) \]
The Halogen Activity Series

The elements of the halogen family also have their own activity series based on displacement reactions. For example, Figure 7 shows the reaction that occurs when chlorine gas is bubbled into a colourless solution of potassium iodide. The reddish orange colour in the test tube is the characteristic colour of iodine, I₂(s).

Note that the single displacement pattern for halogens is slightly different from the pattern used for metals. This is because a negative ion (anion) is being displaced rather than a positive ion (cation).

Bromine also reacts with potassium iodide solution. In this case, the balanced chemical equation is

\[ \text{Br}_2(l) + 2 \text{KI}(aq) \rightarrow 2 \text{KBr}(aq) + \text{I}_2(aq) \]

Because of this evidence, you would expect iodine to be below both bromine and chlorine on a halogen activity series. You would therefore not expect a reaction when iodine is added to a potassium chloride solution (Figure 8).

The evidence from similar experiments is the basis for the activity series of halogens (Figure 9). Just like the activity series of metals, it helps us to predict which elements will displace others from compounds and which will not. Fluorine, for example, displaces all other halogens from compounds.
Single Displacement Reactions of Solids

All the examples we have discussed so far involve aqueous solutions. Some important industrial reactions are single displacement reactions that involve only solid reactants. For example, a key step in making iron for steel involves reacting iron ore (mostly Fe₂O₃) with coke (a form of carbon) in a blast furnace. Temperatures within the furnace reach 2000 °C. This is hot enough to melt iron. The chemical equation for this reaction is

\[3 \text{C}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow 3 \text{CO}(g) + 2 \text{Fe}(l)\]

A + BC → AC + B

**An Activity Series of Ions**

This investigation establishes an activity series for metal ions and compares it with the activity series of solids.

**Ironing Out Rust**

Skills: Questioning, Performing, Observing, Analyzing, Communicating

Iron ore is made up of several compounds, including iron(III) oxide. In this activity, you will recreate, on a small scale, the chemical reaction that converts iron(III) oxide to iron (Figure 10). Instead of a blast furnace, you will use a Bunsen burner. A burning match will be the source of carbon.

**Equipment and Materials:**
- eye protection
- lab apron
- 2 Petri dishes
- magnet
- Bunsen burner
- spark lighter
- tongs
- ceramic dish
- scoopula
- vial of iron(III) oxide
- match

1. Put on your eye protection and lab apron.
2. Place a small amount of iron(III) oxide powder in the Petri dish.
3. Place a magnet on the underside of the dish and test the powder for magnetism.
4. Your teacher will light a Bunsen burner and adjust its flame until it is blue.
5. Your teacher will dip a match head into a vial of iron(III) oxide, coating the match head well with the compound.
6. The match, held with tongs, will be positioned so that the coated end is in the centre of the blue flame. Record your observations.
7. The match will be removed from the flame and allowed to burn for 3–5 s. Your teacher will place the match in the ceramic dish. Allow the flame to extinguish.
8. Once the contents of the dish have cooled, crush them with the end of the scoopula and transfer them to the second Petri dish.
9. Test the contents of the dish from underneath with the magnet.
10. Dispose of the materials as directed by your teacher.

A. What evidence suggests that pure iron was produced in this reaction?

B. Based on this experiment, where would carbon fit on an activity series, relative to iron? Why?
4.4 Questions

1. Compare the types of reactants in synthesis and single displacement reactions.

2. Use the activity series of metals to predict whether each of these reactions occurs. Write a balanced chemical equation for each reaction that does occur. Write no reaction for those that do not.
   (a) \( \text{Al}(s) + \text{AgNO}_3(aq) \rightarrow \)
   (b) \( \text{Zn}(s) + \text{Pb(NO}_3)_2(aq) \rightarrow \)
   (c) \( \text{Au}(s) + \text{H}_2\text{O}(l) \rightarrow \)
   (d) \( \text{Mg}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \)
   (e) \( \text{Ca}(s) + \text{H}_2\text{O}(l) \rightarrow \)
   (f) \( \text{Al}(s) + \text{HCl}(aq) \rightarrow \)

3. Use the halogen activity series to predict whether each of these reactions occurs. Write a balanced chemical equation for each reaction that does occur. Write no reaction for those that do not.
   (a) \( \text{Br}_2(l) + \text{NaI}(aq) \rightarrow \)
   (b) \( \text{Cl}_2(g) + \text{KF}(aq) \rightarrow \)
   (c) \( \text{F}_2(g) + \text{CaBr}_2(aq) \rightarrow \)

4. Use the activity series to rank these reactions in order from slowest to fastest. Justify your prediction.
   (a) \( \text{Zn}(s) + \text{AgNO}_3(aq) \rightarrow \)
   (b) \( \text{Mg}(s) + \text{AgNO}_3(aq) \rightarrow \)
   (c) \( \text{K}(s) + \text{AgNO}_3(aq) \rightarrow \)
   (d) \( \text{Au}(s) + \text{AgNO}_3(aq) \rightarrow \)

5. Some pieces of jewellery corrode, leaving stains on your skin. Others do not. Use the activity series of metals to explain this observation.

6. Figure 11 shows magnesium metal burning brightly inside a block of dry ice (solid carbon dioxide). One of the products of this reaction is elemental carbon.

   Figure 11 Magnesium burning in dry ice

(a) Write a chemical equation for this reaction.
(b) Fires involving reactive metals such as magnesium are difficult to extinguish. Why is pouring sand on the fire better than attempting to use a carbon dioxide fire extinguisher?

7. Some pipes supplying water to old homes are made of lead. Use the activity series to explain why many homeowners are replacing these pipes with copper pipes.

8. Table 1 summarizes the reaction of metals A, B, and C in different temperatures of water. Based on this evidence, describe the position of these metals relative to each other and to hydrogen on the activity series. Justify your prediction.

   Table 1 Activity of Three Metals in Water of Different Temperatures

<table>
<thead>
<tr>
<th>Temperature of water</th>
<th>Metals</th>
</tr>
</thead>
<tbody>
<tr>
<td>cold</td>
<td>NR</td>
</tr>
<tr>
<td>room temperature</td>
<td>NR</td>
</tr>
<tr>
<td>hot</td>
<td>NR</td>
</tr>
</tbody>
</table>

9. Figure 12 shows three metals, A, B, and C, in hydrochloric acid.
   (a) Compare the relative position of these metals and hydrogen on the activity series.
   (b) Which reactive metal is highest in the activity series? Justify your answer.
   (c) If the three metals were copper, magnesium, and zinc, which letter would correspond to which metal?

   Figure 12 Three metals in acid

10. Most metals exist naturally as compounds rather than as elements. Extracting metals is far more complicated than simply decomposing their compounds. The Goldschmidt process was developed to extract important metals from their ores. Research this process and how it produces iron. Write a chemical equation for the reaction.
The Mystery of the Missing Mercury

ABSTRACT
Mercury and ozone are two industrial pollutants that accumulate in the Arctic air during winter. Each spring, both substances mysteriously vanish. Mercury is a particular concern because of its toxicity and ability to accumulate in food chains. At first, these depletion events were thought to be unrelated. Recent discoveries suggest that they are caused by bromine released from seawater. The chemical reactions involved and the ultimate fate of the lost mercury are still the subject of ongoing research.

Introduction
Each spring, as the first rays of the Sun glisten across the Arctic ice and snow, a strange disappearing act occurs. All the mercury in the Arctic air suddenly vanishes. Several months later, as the dark of the Arctic winter returns, mercury reappears. Canadian scientists first noticed this peculiar cycling of mercury in 1982 and have been studying it ever since (Figure 1). Researchers led by Jan Bottenheim of Environment Canada want to get to the bottom of why mercury vanishes from the atmosphere. Perhaps even more importantly, they want to find out where it goes.

Sources of Mercury
Like gold, mercury is one of the few metals that are found in nature as pure elements. It is the only metal that is liquid at room temperature (Figure 2). Like any liquid, mercury gradually evaporates into the air, becoming a gas even at low temperatures. You may have seen mercury in thermometers, although this type of thermometer is being phased out because of the danger if it breaks. Mercury is also used in fluorescent lights.

Some mercury enters the atmosphere as a result of natural processes such as volcanic activity. However, as much as two-thirds of atmospheric mercury is released by human activities such as the combustion of coal. Canada and the United States burn vast quantities of coal to generate power, refine metals, and incinerate waste. Coal is mostly carbon, but it contains some impurities: traces of mercury, sulfur, and other elements.

Clues Emerge
The first clue to solving the missing mercury mystery came from an unlikely source. Ground-level ozone, O_3(g), is another air pollutant. Like mercury, ozone is produced largely by human activities in the south and redistributed by air currents (wind) throughout the atmosphere. By coincidence, scientists noticed that ozone also vanishes from the Arctic air each spring. Is sunlight providing the energy...
needed for ozone and mercury to react and form new compounds? Or is another factor involved? Bottenheim’s team knew that ozone depletion is usually caused by halogens or their compounds. For example, chlorofluorocarbons, or CFCs, were largely responsible for the depletion of the ozone layer in the 20th century. But CFCs have been banned for decades. Could there be a natural source of halogens in the Arctic?

**A Salty Solution to the Mystery**

Bottenheim accidentally found the answer to this question while examining data collected by another Arctic research team. Sure enough, Arctic air contains bromine—a halogen! Furthermore, the concentration of bromine in the air rises each spring, just when levels of ozone and mercury fall. The fluctuations appeared to be connected. According to Bottenheim, “this started the whole ball rolling.”

The team soon determined that the source of the atmospheric bromine was seawater. Seawater contains dissolved bromide ions, Br\textsuperscript{−}(aq). Bromide is a stable ion because it has eight valence electrons (Figure 3(a)). Scientists believe that sunlight causes bromide ions to lose one electron to form highly reactive bromine atoms, Br\textsuperscript{0} (Figure 3(b)).

![Lewis structure of a bromide ion and a bromine atom](image)

Bromine atoms are very reactive, due to their incomplete valence electron orbits. Consequently, these atoms promptly react with oxygen in the air to form highly reactive compounds such as bromine monoxide, BrO (Figure 4).

**Figure 4** Lewis structure of a bromine monoxide molecule. This molecule is reactive because of the single electron on the oxygen atom.

Scientists theorize that both bromine and bromine monoxide react with elemental mercury in the air, pulling away two of its electrons. This leaves a mercury(II) ion, Hg\textsuperscript{2+}. The ion then attaches itself to snowflakes and falls to the ice surface. Thus, almost all of the mercury is removed from the air and trapped in the ice and snow.

The Arctic Sun fuels further photochemical reactions, causing the ice-bound mercury ions to regain two electrons. They now no longer have a positive charge and lose their attraction to the snow and ice crystals.

Some of this mercury is absorbed by micro-organisms and converted to compounds such as methyl mercury, CH\textsubscript{3}Hg. This compound is dangerous because it readily accumulates in the fat tissue of larger organisms as they eat these micro-organisms. However, this process does not account for all the mercury that was in the air. The ultimate fate of the missing mercury remains unclear, and the research continues.

**Further Reading**


**4.5 Questions**

1. Mercury is an industrial pollutant. Why is mercury showing up in the Arctic where there is little or no heavy industry? 

2. Compare the reactivity of a bromide ion and a bromine atom. Why is there a difference?

3. Why are carbon compounds that contain mercury a greater environmental concern than elemental mercury?

4. Use the activity series of metals to justify why mercury can remain for one or two years in the atmosphere without reacting.

5. High-temperature combustion has been used to safely dispose of some toxic industrial compounds. Why is mercury not “destroyed” when coal containing mercury impurities is burned?

6. Give an example of the role of luck in Bottenheim’s research.

7. Why do you think mercury concentrations may be highest in organisms at the top of Arctic food chains?

8. Scientists predict that the amount of sea ice in the Arctic will decrease in the near future. What impact might this have on how quickly mercury disappears in the spring? Why?

9. Research in remote areas like the Arctic is expensive. Much of the funding comes from government sources. What is your opinion about funding for the research to solve the disappearing mercury problem? Is the funding warranted? Is it a waste of money? Explain.
Double Displacement Reactions

Industrial processes produce unwanted by-products (Figure 1). Dissolved toxic metal ions—copper, mercury, and cadmium—are common leftovers in the wastewater. These toxins must be removed before the water can be released into the environment. In Section 4.4, you learned that some metal ions can be recovered as solid metals by allowing a solution of the ions to react with a more active metal.

Another way to remove unwanted ions is to precipitate them from solution. A precipitate is the solid that forms as a result of the reaction of two solutions. Chemists have found that adding sodium hydroxide to a mixture of waste metal ions causes most of the toxic ions to form metal-hydroxide precipitates (Figure 2). The mixture is then filtered to remove the precipitates. Copper ions can be removed from solutions containing waste copper(II) sulfate by adding sodium hydroxide solution.

\[
\text{CuSO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Cu(OH)}_2(s) + \text{Na}_2\text{SO}_4(aq)
\]

A closer look at the equation shows that sodium and copper ions in the reactants have traded places or displaced each other. A chemical reaction in which two elements from different compounds displace each other is called a double displacement reaction. These reactions follow the pattern

\[
\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}
\]

(Figure 3)

Types of Double Displacement Reactions

Double displacement reactions vary, resulting in different products. Some produce precipitates, some release gases, and others result in solutions that are more neutral than the reactant solutions. We can classify double displacement reactions according to the type of product they generate.

Precipitation Reactions

When two solutions react, positive ions (cations) in one solution attract and bond to negative ions (anions) of the other solution. If the new compound precipitates, it must be only slightly soluble in a given solvent. Before we can predict which compound precipitates, we need to understand some key ideas related to solubility. A more detailed discussion of solubility is given in Unit 4: Solutions and Solubility.
SOLUBILITY
As you know, a solution is a homogeneous mixture of a solute dissolved in a solvent. For example, an aqueous salt solution is a mixture of salt (the solute) dissolved in water (the solvent). Solubility is defined as the quantity of solute that will dissolve in a given quantity of solvent. The solubility of a solute depends on factors such as the type of bonding within the solute, the states of both substances, and the temperature of the solvent.

All substances dissolve in water to some degree. Chemists describe a substance as being very soluble (or highly soluble) if a significant quantity of the substance dissolves. A substance that does not dissolve well is only slightly soluble. In the equation of a precipitation reaction, state symbols indicate which of the products forms a precipitate. The state symbol “(s)” is used to indicate the precipitate; the symbol “(aq)” indicates the highly soluble substances that remain in solution.

PREDICTING PRECIPITATES
Selecting the correct reactant to precipitate metal ions from solution requires background knowledge of which metal compounds are soluble and which are not. This information is summarized in a solubility table (Table 1). This table is a summary of the combinations of cations and anions. It tells us the combinations of ions that produce compounds that are very soluble and compounds that are only slightly soluble. For example, the table states that all hydroxides of Group 1 elements are very soluble. Sodium hydroxide is a compound that includes a Group 1 element. Therefore, we would predict that sodium hydroxide dissolves readily in water. Conversely, most other cations form hydroxide compounds that are only slightly soluble. Therefore, mixing a solution of a soluble Cu2⁺ compound with a sodium hydroxide solution should produce a precipitate of copper(II) hydroxide (Figure 4).

Table 1 Solubility of Ionic Compounds at Room Temperature

<table>
<thead>
<tr>
<th>Anions</th>
<th>Cations</th>
</tr>
</thead>
<tbody>
<tr>
<td>Cl⁻</td>
<td>Group 1, NH₄⁺, most, Ag⁺, Pb²⁺, Hg₂⁺, Cu⁺</td>
</tr>
<tr>
<td>Br⁻</td>
<td>Group 1, NH₄⁺, Group 2</td>
</tr>
<tr>
<td>I⁻</td>
<td>Group 1, NH₄⁺, Sr²⁺, Ba²⁺, TI⁺</td>
</tr>
<tr>
<td>S²⁻</td>
<td>Group 1, NH₄⁺, Group 2</td>
</tr>
<tr>
<td>OH⁻</td>
<td>Group 1, NH₄⁺</td>
</tr>
<tr>
<td>SO₄²⁻</td>
<td>Group 1, NH₄⁺</td>
</tr>
<tr>
<td>CO₃²⁻</td>
<td>Group 1, NH₄⁺</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td></td>
</tr>
<tr>
<td>ClO₃⁻</td>
<td></td>
</tr>
<tr>
<td>ClO₄⁻</td>
<td></td>
</tr>
<tr>
<td>Group 1, NH₄⁺</td>
<td></td>
</tr>
<tr>
<td>H⁺</td>
<td></td>
</tr>
</tbody>
</table>

The reactants do not both have to be in solution when they are combined. As long as they are both very soluble and are both added to water, they will dissolve, react, and form the precipitate.

SELECTING A REACTANT TO PRODUCE A PRECIPITATE
How could we treat water contaminated with toxic metal ions, using the solubility table as a tool? Wastewater treatment usually involves two precipitation reactions to ensure that all the metal ions are removed. After the initial precipitation with a hydroxide solution, the mixture is filtered to remove the precipitate. Then another soluble compound is added to the liquid to precipitate any remaining metals. Which anion should that compound contain?

Figure 4 Combining solutions of sodium hydroxide and copper(II) sulfate produces a jellylike precipitate of copper(II) hydroxide.
To remove a range of metals, chemists select an anion that forms slightly soluble compounds with as many metals as possible. In practice, soluble sulfide compounds are usually used (Figure 5).

Chemists can also selectively precipitate certain cations from a mixture by using differences in solubility. For example, the solubility table shows that silver chloride is slightly soluble, but copper(II) chloride is very soluble. Therefore, adding sodium chloride will precipitate only silver ions, Ag⁺(aq), from a solution containing both silver ions and copper(II) ions.

**Tutorial 1  Predicting Double Displacement Reactions**

The following examples illustrate how to predict whether or not a precipitate forms during a double displacement reaction, as well as how to write a balanced chemical equation for the reaction (Figure 6).

**Sample Problem 1: Reactions Involving Two Ionic Compounds**

Will a precipitate form when solutions of potassium chloride and silver nitrate are combined? If you predict that a precipitate will form, write a chemical equation for the reaction. If you predict that no reaction occurs, write *no reaction*.

**Step 1.** Write chemical formulas of the reactants.

\[
\text{KCl}(aq) \quad \text{AgNO}_3(aq)
\]

**Step 2.** Separate the reactants into their ions (Figure 6).

\[
\text{K}^+ \quad \text{Cl}^- \quad \text{Ag}^+ \quad \text{NO}_3^- 
\]

**Step 3.** Combine cations and anions to form new compounds.

\[
\text{KNO}_3 \quad \text{AgCl}
\]

**Step 4.** Check for low solubility.

AgCl is slightly soluble, while KNO₃ is very soluble. Therefore, AgCl should precipitate.

**Step 5.** Write a balanced chemical equation for the reaction. Label the precipitate with the state symbol (s) and soluble compounds with (aq).

\[
\text{KCl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{AgCl(s)} + \text{KNO}_3(aq)
\]
POLYATOMIC IONS

Compounds of polyatomic ions also vary in their solubility. This difference in solubility causes some polyatomic compounds to form precipitates in double displacement reactions. Sample Problem 2 shows how to approach reactions involving polyatomic ions.

Sample Problem 2: Reactions Involving Polyatomic Compounds

Will a precipitate form when solutions of potassium sulfate and iron(III) chloride are combined? If you predict that a precipitate will form, write a chemical equation for the reaction. If you predict that no reaction occurs, write no reaction.

Step 1. Write chemical formulas of the reactants.

\[ K_2SO_4(aq) \quad FeCl_3(aq) \]

Step 2. Separate the reactants into their ions.

\[ K^+ \quad K^+ \quad SO_4^{2-} \quad Fe^{3+} \quad Cl^- \quad Cl^- \]

Step 3. Combine cations and anions to form new compounds.

\[ KCl \quad Fe_2(SO_4)_3 \]

Step 4. Check for low solubility.

Both compounds are very soluble, so no precipitate forms: no reaction.

Practice

1. Which of the following combinations result in the formation of a precipitate? Write a balanced chemical equation if you predict that the reaction will occur. (a) \( Na_2S(aq) + Pb(NO_3)_2(aq) \rightarrow \) (b) \( NH_4Cl(aq) + K_2SO_4(aq) \rightarrow \) (c) \( FeCl_3(aq) + Na_2CO_3(aq) \rightarrow \)

Mini Investigation

Testing Water for Ions

Skills: Controlling Variables, Performing, Observing, Analyzing, Communicating

Water samples are routinely tested to identify the ions present. Information about the ions in your local drinking water is available from the Ministry of the Environment. The concentration of certain ions in bottled water is often provided on the product label.

A solution of silver nitrate is used to detect the presence of chloride ions in water. A white precipitate of silver chloride is a positive test for chloride ions. Similarly, calcium ions in water can be detected by adding a solution of sodium oxalate, \( Na_2C_2O_4(aq) \). The result is a white calcium oxalate precipitate.

Equipment and Materials: eye protection; lab apron; safety gloves; test-tube rack; 6 test tubes; masking tape; marker; samples of distilled water, tap water, and bottled water; dropper bottles of dilute silver nitrate and dilute sodium oxalate

Silver nitrate and sodium oxalate are both toxic. Silver nitrate stains the skin.

Sodium oxalate is an irritant. Avoid skin and eye contact. If you spill these chemicals on your skin, wash the affected area with a lot of cool water.

(Teacher Note: The concentration of each of the dilute solutions is 0.1 mol/L.)

1. Put on your eye protection, lab apron, and safety gloves.
2. Add distilled water, tap water, and bottled water to three test tubes to a depth of about 4 cm. Label the test tubes and place them in the test-tube rack.
3. Create another set of test tubes identical to those in Step 1.
4. Add 4 drops of silver nitrate to each test tube of the first set. Swirl to mix the contents of each test tube.
5. Add 4 drops of sodium oxalate to each test tube of the second set. Swirl to mix the contents of each test tube.

A. Compare the appearance of silver chloride precipitate in the first three test tubes. Suggest reasons to explain any differences you observed.

B. Suggest a human activity that increases the amount of chloride ions in water.

C. Compare the appearance of calcium oxalate precipitate in the second three test tubes.

D. Suggest a possible natural source of calcium ions in water.
Reactions That Produce a Gas

Double displacement involving acids can also produce gases. This can occur in two ways. First, gases may be produced directly. For example, the addition of an acid to a sulfide compound like potassium sulfide, K₂S, results in a double displacement reaction that produces hydrogen sulfide gas:

\[ K₂S(aq) + 2 \text{HCl}(aq) \rightarrow \text{H}_2\text{S}(g) + 2 \text{KCl}(aq) \]

Hydrogen sulfide is toxic and has a distinctive odour of rotten eggs. Spilling an acid onto a sulfide compound releases toxic hydrogen sulfide fumes. For this reason, the use of sulfide compounds has been banned in many schools.

Gases can also be produced when an unstable product of a double displacement reaction decomposes. For example, the pattern of double displacement reactions suggests that magnesium carbonate and sulfuric acid react to produce magnesium sulfate and carbonic acid:

\[ \text{MgCO}_3(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{MgSO}_4(aq) + \text{H}_2\text{CO}_3(aq) \]

However, carbonic acid is unstable and immediately decomposes into water and bubbles of carbon dioxide (Figure 7):

\[ \text{H}_2\text{CO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) \]

Similar reactions occur when acids are added to solutions of sulfite compounds. For example, a solution of sodium sulfite reacts with hydrochloric acid according to the following chemical equation:

\[ \text{Na}_2\text{SO}_3(aq) + 2 \text{HCl}(aq) \rightarrow \text{H}_2\text{SO}_3(aq) + 2 \text{NaCl}(aq) \]

The sulfite product, H₂SO₃, is sulfurous acid. However, bubbles of toxic sulfur dioxide appear as soon as the reactants are combined. This occurs because sulfurous acid is unstable and quickly decomposes into water and sulfur dioxide:

\[ \text{H}_2\text{SO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{SO}_2(g) \]

The production of gases by double displacement reactions is summarized in Table 2.

Neutralization Reactions

In earlier science courses, you learned that the reaction of an acid with a base is called a neutralization reaction. The end result is a mixture with a pH closer to a neutral pH of 7 than either of its reactants. For example, nitric acid, HNO₃(aq), can be neutralized by the addition of a sodium hydroxide solution, NaOH(aq). The chemical equation for the reaction is

\[ \text{HNO}_3(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaNO}_3(aq) \]

\[ \text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB} \]

If you think of the water molecule as consisting of two parts, H—OH, then a neutralization reaction follows the pattern of double displacement reactions.

Consider another example. Antacids are medications that neutralize excess stomach acid. For example, magnesium hydroxide, Mg(OH)₂, is the active ingredient in some antacid medications (Figure 8). Magnesium hydroxide neutralizes excess hydrochloric acid in the stomach according to the equation

\[ \text{Mg(OH)}_2(s) + 2 \text{HCl}(aq) \rightarrow 2 \text{H}_2\text{O}(l) + \text{MgCl}_2(aq) \]

Note that these reactions generally do not produce precipitates. We can detect that a reaction has occurred only by testing the pH of the resulting solution and comparing it with the pH of the reactant solutions.

Figure 7 The reaction of magnesium carbonate and hydrochloric acid results in the production of a gas.

Table 2 Compounds That React with Acids to Produce Gases

<table>
<thead>
<tr>
<th>Compound</th>
<th>Gas produced</th>
</tr>
</thead>
<tbody>
<tr>
<td>sulfides, e.g., Na₂S</td>
<td>hydrogen sulfide, H₂S</td>
</tr>
<tr>
<td>carbonates, e.g., Na₂CO₃</td>
<td>carbon dioxide, CO₂</td>
</tr>
<tr>
<td>sulfites, e.g., K₂SO₃</td>
<td>sulfur dioxide, SO₂</td>
</tr>
</tbody>
</table>

neutralization reaction the reaction of an acid and a base in which the resulting solution has a pH closer to 7 than either of the reactants

Figure 8 Magnesium hydroxide has a low solubility in water. This explains why a mixture of magnesium hydroxide and water is a thick white paste rather than a clear solution.
4.6 Summary

- In a double displacement reaction, two elements trade places to form two new compounds. The net result is that two compounds react to form two new compounds. The pattern for these reactions is \(AB + CD \rightarrow AD + CB\).
- Double displacement reactions can produce a precipitate, a neutralized solution, or a gas.
- Most double displacement reactions produce precipitates that can be predicted using a solubility table.
- The reaction of an acid with a base produces a neutralized solution.
- The reaction of an acid with certain polyatomic compounds can produce sulfides, carbonates, or sulfites.

4.6 Questions

1. How do single and double displacement reactions differ?

2. Classify the reactions represented by these equations as either single or double displacement.
   - (a) \(HI(aq) + AgNO_3(aq) \rightarrow AgI(s) + HNO_3(aq)\)
   - (b) \(Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)\)
   - (c) \(ZnS(g) + 2 HCl(aq) \rightarrow ZnCl_2(aq) + H_2S(g)\)
   - (d) \(Cl_2(g) + 2 NH_4Br(aq) \rightarrow Br_2(l) + 2 NH_4Cl(aq)\)

3. Write the chemical formula for each of the following compounds. Predict the solubility of these compounds in water.
   - (a) lead sulfate (in car batteries)
   - (b) ammonium phosphate (fertilizer)
   - (c) calcium sulfate (a component of drywall)
   - (d) aluminum sulfate (used in water purification)
   - (e) calcium phosphate (in bones)
   - (f) barium sulfate (used during stomach X-rays)
   - (g) ammonium carbonate (smelling salts)
   - (h) calcium carbonate (in shells)

4. Complete the chemical equations of the following reactions. Include state symbols to indicate the state of each compound.
   - (a) \(ZnCl_2(aq) + KOH(aq) \rightarrow\)
   - (b) \(Ni(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow\)
   - (c) \(Ba(OH)_2(aq) + K_2SO_4(aq) \rightarrow\)
   - (d) \(FeSO_4(aq) + K_2PO_4(aq) \rightarrow\)
   - (e) \(ZnS(s) + 2 HCl(aq) \rightarrow\)
   - (f) \(CaCO_3(s) + 2 HNO_3(aq) \rightarrow\)
   - (g) \(MgSO_4(aq) + HCl(aq) \rightarrow\)

5. Will a precipitate form when the following solutions are combined? If you predict that a precipitate will form, write a chemical equation for the reaction. If you predict that no reaction occurs, write no reaction.
   - (a) \(AgNO_3 + K_2SO_4\)
   - (b) \(NH_4Cl + NaS\)
   - (c) \(Pb(NO_3)_2 + Na_2PO_4\)

6. Write a balanced chemical equation (with state symbols) for each of the following double displacement reactions.

7. Silver compounds are expensive. Metal manufacturers therefore want to recover any waste silver ions from solution. One way to do this is to precipitate the silver as a compound and then separate the compound from the mixture using filtration. Use the solubility table to determine a way of precipitating only silver ions from a mixture of dissolved metal ions.

8. Research why soaking in vinegar is a convenient method of removing limescale from the heating coils of a kettle. Write and classify the chemical reactions involved.

9. “Black smokers” emerge from cracks in the ocean floor (Figure 9). These black plumes are created by superheated water passing through rock. Research the chemical reactions that cause these plumes to be black. Summarize your findings in a paragraph.
Investigation 4.4.1

Hydrogen Blast-Off

Hydrogen is an ideal rocket fuel because it releases more energy per tonne when it burns than any other fuel. The chemical equation for the combustion of hydrogen is

\[2 \text{H}_2(g) + \text{O}_2(g) \rightarrow \text{2 H}_2\text{O}(g) + \text{energy}\]

However, a successful launch requires the proportions of hydrogen to oxygen to be just right. In this investigation, you will test-burn two different mixtures of hydrogen and oxygen to see which ratio of the gases would be better for a rocket launch. You will first produce hydrogen and oxygen using these chemical reactions:

\[\text{Mg(s)} + 2 \text{HCl(aq)} \rightarrow \text{H}_2(g) + \text{MgCl}_2(aq)\]
\[2 \text{H}_2\text{O}_2(aq) \rightarrow 2 \text{H}_2\text{O}(l) + \text{O}_2(g)\]

Testable Question

Which ratio of hydrogen to oxygen produces the most energy (the loudest “pop”) when burned:
- 2 parts hydrogen to 1 part oxygen
- 1 part hydrogen to 2 parts oxygen

Hypothesis

Predict which hydrogen-to-oxygen mixture produces the most energy. Give reasons for your hypothesis.

Variables

Identify all major variables that will be measured and/or controlled in this investigation.

Identify the manipulated (independent) and responding (dependent) variables.

Experimental Design

You will use the reactions given in the introduction to prepare two different mixtures of hydrogen and oxygen. You will test the mixtures by squeezing them into a candle flame to see which produces the loudest “pop.”

Equipment and Materials

- eye protection
- lab apron
- scissors
- marker
- well plate
- one-holed stopper (that fits the wells of the well plate)
- candle or Bunsen burner clamped to a retort stand

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**SKILLS MENU**

- Planning
- Controlling
- Variables
- Performing
- Communicating

- Questioning
- Researching
- Hypothesizing
- Predicting
- Observing
- Analyzing
- Evaluating

- Performing
- Communicating

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**Equipment and Materials**

- spark lighter
- toothpick
- dropper bottles of
  - 6% hydrogen peroxide solution, H$_2$O$_2$(aq)
  - dilute hydrochloric acid, HCl(aq)
- 2 plastic pipettes
- 1 cm strip of magnesium, Mg(s)
- yeast

(Teacher Note: The concentration of the dilute hydrochloric acid is 0.1 mol/L.)

The hydrochloric acid and hydrogen peroxide solutions used in this activity are irritants. Wash any spills on skin or clothing immediately with plenty of cold water. Report any spills to your teacher.

An open flame is used. Tie back long hair and secure loose clothing. Never leave the flame unattended.

Procedure

Part A: Preparing the Gas-Collecting Bulbs

1. Prepare two gas-collecting bulbs and one gas delivery stopper assembly, using Figure 1 as a guide.

2. Label one bulb as 1H$_2$:2O$_2$ and the other as 2H$_2$:1O$_2$.

3. Fill the two bulbs with water. Stand them, open end upward, in the well plate.
Part B: Collecting Hydrogen

4. Put on your eye protection and lab apron.
5. Add 15 drops of hydrochloric acid to a clean, dry well of the well plate.
6. Add a 1 cm strip of magnesium to the same well.
7. Seal the well with the stopper assembly. Wait 5 s for the hydrogen gas being produced to displace the air in the well.
8. Place the “1H₂:2O₂” gas-collecting bulb full of water onto the stopper assembly (Figure 2). The fit should be loose to allow water to be displaced.

9. Allow one-third of the bulb to fill with hydrogen gas. (The remaining two-thirds will be water.)
10. Remove the bulb and stand it in a well plate. Always keep the opening of the bulb at the bottom; otherwise, the gas will escape.
11. Place the “2H₂:1O₂” gas-collecting bulb full of water onto the stopper assembly.
12. Allow two-thirds of this bulb to fill with hydrogen gas. (The remaining one-third will be water.)
13. Remove the bulb and stand it in the well plate, as in Step 10.

Part C: Collecting Oxygen Gas

14. Add 15 drops of hydrogen peroxide to a clean, dry well.
15. Use the wide end of the toothpick to add a few grains of yeast.
16. Seal the well with the gas delivery stopper (without a bulb). Ensure that the bottom of the tube is not blocked.
17. Wait 5 s for the oxygen gas being produced to displace the air in the well.
18. Place the “2H₂:1O₂” bulb that is two-thirds full of hydrogen gas onto the stopper assembly (Figure 3).
19. Allow the oxygen gas to displace all but about two drops of water in the bulb. These drops act as a plug to keep the collected gases in the bulb.

20. Return the bulb to its well.
21. Repeat Steps 18 to 20 with the “1H₂:2O₂” bulb.

Part D: Testing the Gas Mixtures

22. Light a candle or burner.
23. Without tilting it, bring the “1H₂:2O₂” bulb to within 2 cm of the flame (Figure 4(a)).

24. Quickly rotate the bulb and squirt its contents into the flame (Figure 4(b)). Record your observations.
25. Repeat Steps 23 and 24 with the other bulb. Record the results when the two gas mixtures are squirted into the flame.

Analyze and Evaluate

(a) What variables were manipulated in this investigation? What was the responding variable?
(b) Classify each of the three reactions (in Parts B, C, and D) that you observed in this activity.
(c) Answer the Question at the beginning of this investigation. Support your answer with evidence.
(d) How would your results differ if you did not wait a few seconds before collecting each gas?
(e) What role does the yeast play in this experiment?
(f) What was the most difficult challenge you faced during the experiment? How did you overcome that challenge?

Apply and Extend

(g) Predict how the reaction of a 1:1 mixture of hydrogen to oxygen would compare with the reactions you tested in this investigation. Justify your prediction.
An Activity Series of Ions

Copper is above silver on the activity series (Figure 5 in Section 4.4). You would therefore expect that a single displacement reaction occurs when copper wire is placed into a solution of silver nitrate (Figure 1).

\[ \text{Cu(s)} + 2 \text{AgNO}_3(\text{aq}) \rightarrow 2 \text{Ag(s)} + \text{Cu(NO}_3)_2(\text{aq}) \]

Silver nitrate is soluble in water, so the silver and nitrate ions separate as the compound dissolves. Therefore, a silver nitrate solution is actually a mixture of silver ions, \( \text{Ag}^+(\text{aq}) \), and nitrate ions, \( \text{NO}_3^-\text{(aq)} \), in water. So, the reaction is really between copper metal and silver ions. The nitrate ions just remain in solution. However, a similar reaction does not occur if a strip of silver is placed in a solution containing \( \text{Cu}^2+(\text{aq}) \) ions. It seems that metal ions—as well as metal elements—differ in their reactivity.

**Figure 1** The reaction of copper wire in a solution of silver nitrate. The blue colour of the solution is evidence that \( \text{Cu}^2+(\text{aq}) \) ions are produced.

**Testable Question**

What is the trend of reactivity of metal ions?

**Prediction**

Predict the trend in reactivity of the metal ions used in this experiment.

**Variables**

Identify all major controlled, manipulated, and responding variables in this experiment.

**Equipment and Materials**

- eye protection
- lab apron
- well plate
- sand paper
- clean strips of magnesium, Mg, zinc, Zn, iron, Fe, tin, Sn, and copper, Cu
- dropper bottles of dilute solutions of:
  - magnesium sulfate, \( \text{MgSO}_4 \)
  - zinc sulfate, \( \text{ZnSO}_4 \)
  - iron(II) sulfate, \( \text{FeSO}_4 \)
  - tin(II) chloride, \( \text{SnCl}_2 \)
  - copper(II) sulfate, \( \text{CuSO}_4 \)

*(Teacher Note: The concentration of each of the five solutions is 0.1 mol/L.)*

**Experimental Design**

You will place five metals in different metal ion solutions to see which combinations of metals and ions result in single displacement reactions.

**Procedure**

1. Design a procedure to determine which metal/metal ion solution combinations result in a single displacement reaction.
2. Proceed, with your teacher’s approval.
3. Record that a reaction occurred only when a new solid forms. Disregard any gas bubbles that may form.
4. Dispose of the contents of the well plate as directed by your teacher.
5. Clean your workstation and wash your hands.

**Observations**

Organize your observations in a table similar to Table 1.

**Table 1 Observation Table**

<table>
<thead>
<tr>
<th>Metal</th>
<th>Ion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg(s)</td>
<td>( \text{Mg}^{2+} (\text{aq}) )</td>
</tr>
<tr>
<td>Zn(s)</td>
<td>( \text{Zn}^{2+} (\text{aq}) )</td>
</tr>
<tr>
<td>Fe(s)</td>
<td>( \text{Fe}^{2+} (\text{aq}) )</td>
</tr>
<tr>
<td>Sn(s)</td>
<td>( \text{Sn}^{2+} (\text{aq}) )</td>
</tr>
<tr>
<td>Cu(s)</td>
<td>( \text{Cu}^{2+} (\text{aq}) )</td>
</tr>
</tbody>
</table>

**Analyze and Evaluate**

(a) What variables were measured/recording and/or manipulated in this investigation? What type of relationship was being tested?

(b) Answer the Testable Question that was posed at the beginning of this investigation. Support your answer with evidence.

(c) Compare the trends in reactivity of metals with trends in the reactivity of their ions.

(d) Use your answer to question (c) and the activity series of metals to rank these ions in order from least to most reactive: \( \text{Ni}^{2+} (\text{aq}) \), \( \text{Ba}^{2+} (\text{aq}) \), \( \text{Hg}^{2+} (\text{aq}) \). Justify your prediction.
(e) A solution containing highly toxic cadmium ions, Cd²⁺(aq), reacts with iron but does not react with tin. Predict the position of cadmium on the activity series compared to the five metals used in this experiment.

(f) Write a chemical equation for each single displacement reaction recorded in your observation table.

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**Investigation 4.6.1**

**Precipitate Patterns**

Wastewater from metal manufacturing processes often contains toxic metal cations. These ions must be removed before the water can be released into the environment. Many of these ions can be removed using precipitation reactions. For example, recycled car batteries contain toxic lead cations, Pb²⁺(aq). These can be removed by adding a solution of potassium iodide, KI(aq) (Figure 1).

In this activity, you will combine different pairs of solutions to discover which cation/anion combinations result in the formation of a precipitate.

**Figure 1** Lead and iodide ions combine to form a precipitate of lead(II) iodide, PbI₂(s).

**Purpose**

To identify which anion could be used to precipitate the most metal cations from an unknown mixture.

**Equipment and Materials**

- eye protection
- lab apron
- safety gloves
- well plate
- dropper bottles of dilute solutions of:
  - sodium sulfate, Na₂SO₄
  - sodium carbonate, Na₂CO₃
  - sodium chloride, NaCl
  - calcium nitrate, Ca(NO₃)₂
  - copper(II) nitrate, Cu(NO₃)₂
  - iron(III) nitrate, Fe(NO₃)₃
  - potassium nitrate, KNO₃
  - magnesium nitrate, Mg(NO₃)₂
  - silver nitrate, AgNO₃
  - zinc nitrate, Zn(NO₃)₂

**Apply and Extend**

(g) Some nuclear power plants use seawater as a coolant. Choosing the correct type of piping to carry seawater is critical because seawater contains trace amounts of dissolved silver and gold ions. Is copper piping a suitable choice? Why?

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**Procedure**

Your teacher may choose to demonstrate a reaction involving the silver ion.

1. Design a procedure to determine which anion precipitates the most cations. Include an observation table in your design.
2. Proceed, with your teacher’s approval.
3. Dispose of the contents of the well plate as directed.
4. Clean your workstation and wash your hands.

**Analyze and Evaluate**

(a) What variables were manipulated in this investigation?

(b) Evaluate your experimental design. How could it have been improved?

(c) Based on your evidence, which anion could be used to precipitate most of the metal cations?

(d) Which anion could be used to selectively remove silver ions from solution? Why?

(e) What evidence suggests that nitrate compounds are soluble in water?

(f) Write the chemical formula for each precipitate that formed.

**Apply and Extend**

(g) Write a balanced chemical equation for each precipitation reaction that occurred.

(h) Why is it necessary to use distilled water to prepare the solutions used in this experiment?

(i) “Hard” water contains a high concentration of calcium ions. Suggest a way to make hard water “softer.”
Summary Questions

1. Create a question-and-answer game based on the Key Concepts listed in the margin on page 6. Divide the questions into four or five categories. The questions within each category should get increasingly more complex. Try to anticipate questions that you think might appear on a test for this chapter.

2. Look back at the Starting Points questions on page 6. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.

Vocabulary

- chemical reaction (p. 8)
- catalyst (p. 8)
- law of conservation of mass (p. 8)
- synthesis reaction (p. 12)
- decomposition reaction (p. 15)
- single displacement reaction (p. 20)
- activity series (p. 21)
- precipitate (p. 28)
- double displacement reaction (p. 28)
- solute (p. 29)
- solvent (p. 29)
- solubility (p. 29)
- neutralization reaction (p. 32)

Career Pathways

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a BSc degree. Others require specialized or postgraduate degrees. Here are just a few pathways to careers mentioned in this chapter.

1. Select an interesting career that relates to the effects of chemical reactions. Research the educational pathway you would need to follow to pursue this career.

2. What is involved in the college diploma program necessary to become a pharmacy assistant? Research at least two programs, and prepare a brief report of your findings.
For each question, select the best answer from the four alternatives.

1. Which of the following is a balanced chemical equation? (4.1)
   (a) \( \text{Al}(s) + \text{Cl}_2(g) \rightarrow \text{Al}_2\text{Cl}_6(s) \)
   (b) \( \text{P}_4(s) + 5 \text{O}_2(g) \rightarrow \text{P}_4\text{O}_{10}(s) \)
   (c) \( 3 \text{N}_2(g) + 2 \text{H}_2(g) \rightarrow 2 \text{NH}_3(g) \)
   (d) \( \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(l) \)

2. What coefficient should be placed in front of HF to balance the chemical equation below?
   \( \text{UO}_2(s) + \text{HF}(g) \rightarrow \text{UF}_4(s) + 2 \text{H}_2\text{O}(l) \) (4.1)
   (a) 1
   (b) 2
   (c) 3
   (d) 4

3. Which of the following equations shows the general pattern for a synthesis reaction? (4.2)
   (a) \( \text{A} + \text{B} \rightarrow \text{AB} \)
   (b) \( \text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB} \)
   (c) \( \text{AB} \rightarrow \text{A} + \text{B} \)
   (d) \( \text{AB} + \text{C} \rightarrow \text{A} + \text{CB} \)

4. Which of the following equations represents a decomposition reaction? (4.2)
   (a) \( \text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) \)
   (b) \( \text{Fe}(s) + \text{HCl}(aq) \rightarrow \text{FeCl}_2(aq) + \text{H}_2(g) \)
   (c) \( 2 \text{NH}_3(g) \rightarrow \text{N}_2(g) + 3 \text{H}_2(g) \)
   (d) \( \text{S} + \text{O} \rightarrow \text{SO}_2 \)

5. Which of the following equations represents a single displacement reaction? (4.4)
   (a) \( 4 \text{Fe}(s) + 3 \text{O}_2(aq) \rightarrow 2 \text{Fe}_2\text{O}_3(s) \)
   (b) \( 2 \text{H}_2\text{O}(l) \rightarrow 2 \text{H}_2(g) + \text{O}_2(g) \)
   (c) \( \text{Zn}(s) + 2 \text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g) \)
   (d) \( \text{Ca(OH)}_2(s) + 2 \text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + 2 \text{H}_2\text{O}(l) \)

6. Based on the activity series of metals, which of the following reactions occurs? (4.4)
   (a) \( \text{Ni}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \)
   (b) \( \text{Cu}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \)
   (c) \( \text{Hg}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \)
   (d) \( \text{Au}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \)

7. Which halogen is most reactive at room temperature? (4.4)
   (a) \( \text{Cl}_2(g) \)
   (b) \( \text{I}_2(s) \)
   (c) \( \text{Br}_2(l) \)
   (d) \( \text{F}_2(g) \)

8. What type of reaction is shown below?
   \( \text{Fe}(s) + \text{Cu(NO}_3)_2(aq) \rightarrow \text{Fe(NO}_3)_2(aq) + \text{Cu}(s) \)
   (4.2, 4.4, 4.6)
   (a) single displacement
   (b) double displacement
   (c) synthesis
   (d) decomposition

9. What type of reaction occurs when hydrogen fluoride solution reacts with potassium hydroxide solution? (4.2, 4.4, 4.6)
   (a) single displacement
   (b) neutralization
   (c) decomposition
   (d) synthesis

10. If lead metal and hydrogen bromide solution are combined, what type of reaction will occur? (4.2, 4.4, 4.6)
   (a) double displacement
   (b) synthesis
   (c) decomposition
   (d) single displacement

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. A catalyst makes a reaction go slower. (4.1)
12. A colour change is usually evidence that a chemical reaction has occurred. (4.1)
13. In a synthesis reaction, a new element is produced from two simpler elements. (4.2)
14. In a single displacement reaction, one element replaces another in a compound. (4.4)
15. A neutralization reaction results in a solution that has a higher pH than the pH of the reactant solutions. (4.6)
16. A precipitation reaction always has an acid and a base as reactants. (4.6)
17. In a double displacement reaction, a highly soluble product forms a precipitate. (4.6)
Knowledge

For each question, select the best answer from the four alternatives.

1. Which statement correctly describes the law of conservation of mass? (4.1)
   (a) The mass of the products equals the mass of the reactants.
   (b) If a reaction has two products, their masses are equal.
   (c) The mass of the products is less than the mass of the reactants.
   (d) If a reaction has two reactants, their masses are equal.

2. Which of the following chemical equations is balanced? (4.1)
   (a) $CS_2(g) + 3 O_2(g) \rightarrow CO_2(g) + SO_2(g)$
   (b) $CS_2(g) + 3 O_2(g) \rightarrow CO_2(g) + 2 SO_2(g)$
   (c) $CS_2(l) + O_2(g) \rightarrow CO_2(g) + SO_2(g)$
   (d) $CS_2(l) + 3 O_2(g) \rightarrow CO_2(g) + 3 SO_2(g)$

3. Which of the following chemical equations is balanced? (4.1)
   (a) $Fe_2O_3(s) + 2 C(s) \rightarrow Fe(s) + 2 CO_2(g)$
   (b) $Ca(OH)_2(aq) + 2 HCl(aq) \rightarrow CaCl_2(aq) + 2 H_2O(aq)$
   (c) $NO(g) + Br_2(g) \rightarrow NOBr(g)$
   (d) $Ca(OH)_2(s) + 2 HCl(aq) \rightarrow CaCl_2(aq) + 2 H_2O(l)$

4. Which of the following equations represents a single displacement reaction? (4.4)
   (a) $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(aq)$
   (b) $S(s) + O_2(g) \rightarrow SO_2(g)$
   (c) $Li(s) + KCl(aq) \rightarrow K(s) + LiCl(aq)$
   (d) $2 H_2O_2(aq) \rightarrow 2 H_2O(l) + O_2(g)$

5. The activity series of metals shows
   (a) the relative atomic numbers of various metals
   (b) the relative reactivity of various elements
   (c) whether or not a compound is soluble
   (d) the ionic charges of different metallic compounds (4.4)

6. What type of reaction is represented by the equation below?
   $2 KClO_3(s) \rightarrow 2 KCl(s) + 3 O_2(g)$ (4.2, 4.4, 4.6)
   (a) single displacement
   (b) double displacement
   (c) neutralization
   (d) decomposition

7. What type of reaction is represented by the equation below?
   $LiCl(aq) + AgNO_3(aq) \rightarrow LiNO_3(aq) + AgCl(s)$ (4.2, 4.4, 4.6)
   (a) single displacement
   (b) double displacement
   (c) synthesis
   (d) decomposition

8. Which of the following equations represents a precipitation reaction? (4.6)
   (a) $H_2O(g) + CO(g) \rightarrow H_2(g) + CO_2(g)$
   (b) $BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$
   (c) $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(s)$
   (d) $LiBr(aq) + AgNO_3(aq) \rightarrow LiNO_3(aq) + AgBr(s)$

9. Which of the following statements correctly describes a decomposition reaction? (4.2)
   (a) A compound breaks down into smaller compounds or elements.
   (b) Two compounds combine to make a larger compound.
   (c) Atoms of two elements switch places between compounds.
   (d) A compound changes state.

10. Which equation most accurately represents the synthesis reaction between lithium and oxygen? (4.2)
    (a) $Li(s) + O_2(g) \rightarrow Li_2O(s)$
    (b) $2 LiO(s) \rightarrow 2 Li(s) + O_2(g)$
    (c) $6 Li(s) + O_2(g) \rightarrow 2 Li_2O(s)$
    (d) $4 Li(s) + O_2(g) \rightarrow 2 Li_2O(s)$

   Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. If a substance changes colour, a chemical change must have occurred. (4.1)

12. We can use the activity series to determine whether a synthesis reaction will occur. (4.2)

Match each chemical equation from the first list with the correct type of reaction from the second list.

13. (a) $MgO(aq) + 2 HCl(aq) \rightarrow MgCl_2 + H_2O$
    (b) $2 HgO(s) \rightarrow 2 Hg(l) + O_2(g)$
    (c) $P_4(s) + 6 Cl_2(g) \rightarrow 4 PCl_3(l)$
    (d) $KI(aq) + F_2(aq) \rightarrow 2 KF(aq) + I_2(s)$
    (i) decomposition
    (ii) synthesis
    (iii) single displacement
    (iv) double displacement (4.2, 4.4, 4.6)
Write a short answer to each question.

14. What must be true for an equation to be balanced? (4.1) [T]

15. Write the general pattern for a decomposition reaction. (4.2) [T]

16. How does a catalyst affect a reaction? (4.1) [T]

17. What is the state of a precipitate that results from a precipitation reaction? (4.6) [T]

18. What does the symbol (aq) indicate? (4.6) [T]

Understanding

19. Balance the following chemical equations. (4.1) [T]
   (a) NaClO₃(s) → NaCl(s) + O₂(g)
   (b) C₂H₂(g) + O₂(g) → H₂O(g) + CO₂(g)
   (c) Na(s) + H₂O(l) → NaOH(aq) + H₂(g)

20. During cellular respiration, an organism breaks down glucose to release usable energy. Other chemicals are produced in the process. Balance the respiration equation shown below.
   \[ C₆H₁₂O₆(aq) + O₂(aq) \rightarrow CO₂(aq) + H₂O(l) + \text{energy} \]

21. (a) List three clues that a chemical reaction has occurred.
    (b) If you observed one of these clues, would that prove that a chemical reaction has occurred? Explain. (4.1) [T]

22. Water can be forced to decompose into hydrogen and oxygen gases. Write the balanced equation for this decomposition reaction. (4.2) [T]

23. (a) Predict whether zinc will react with lead(II) sulfate.
    (b) What information did you use to make your prediction?
    (c) Name one metal (other than zinc) that, according to your information, would react with lead(II) sulfate.
    (d) Name one metal (other than zinc) that would not react with lead(II) sulfate. (4.4) [T]

24. (a) Predict whether nickel will react with sodium chloride.
    (b) What information did you use to make your prediction? (4.4) [T]

25. Describe the solubility of each of the following compounds. (4.6) [T]
   (a) potassium bromide, KBr(s)
   (b) sodium carbonate, Na₂CO₃(s)
   (c) silver acetate, AgC₂H₃O₂(s)

26. (a) Rewrite the chemical equation
   \[ CdSO₄(aq) + H₂S(aq) \rightarrow CdS(\,\text{s}) + H₂SO₄(\,\text{aq}) \]
   so that it includes the missing state symbols.
   (b) Predict any precipitates formed in this reaction.
   (c) Explain how you determined whether there would be any precipitates. (4.6) [T]

27. A precipitate forms in the reaction between solutions of sodium hydroxide and nickel(II) chloride. (4.6) [T]
   (a) Write the chemical formula for the precipitate.
   (b) Explain how you determined the identity of the precipitate.

28. Limestone consists mostly of calcium carbonate, CaCO₃(s). One way to test if a mineral is limestone is to put a small sample of it in hydrochloric acid. If bubbles appear, the sample may be limestone. (4.1, 4.6) [T]
   (a) The reaction produces calcium chloride, carbon dioxide, and water. Write the balanced chemical equation for the reaction.
   (b) What causes the bubbles?
   (c) How useful is this test for limestone? Explain.

29. Nitrous oxide, N₂O(g), can be produced by heating solid ammonium nitrate. Another product of this reaction is water vapour. (4.2, 4.4, 4.6) [T]
   (a) Write the balanced chemical equation for the reaction that produces nitrous oxide.
   (b) What type of reaction is this?
   (c) This reaction can be dangerously explosive. Another way to produce nitrous oxide and liquid water is by mixing ammonia and oxygen gases. Write the balanced chemical equation for this alternative process.
   (d) What type of reaction is described in part (c)?

30. Hydrogen peroxide, H₂O₂(l), can be produced by reacting oxygen gas and hydrogen gas. (4.2, 4.4, 4.6) [T]
   (a) Write the balanced chemical equation for this process.
   (b) What type of reaction is this?
   (c) The reaction depends on the use of a catalyst. Describe the purpose of a catalyst.
   (d) What would be the economic effect if a hydrogen peroxide processing plant did not have the correct catalyst?

31. Predict which type of reaction is likely to take place between the following pairs of reactants. (4.2, 4.4, 4.6) [T]
   (a) calcium metal and oxygen gas
   (b) hydrocyanic acid and sodium hydroxide solution

32. Classify the reactions represented by the following equations. Explain how you arrived at your answers. (4.2, 4.4, 4.6) [T]
   (a) HF(aq) + KOH(aq) → KF(aq) + H₂O(l)
   (b) Al(s) + Cr₂O₃(s) → Al₂O₃(s) + Cr(s)
33. Complete and balance each of the following reaction equations and classify the reaction, giving reasons for your answer. (4.2, 4.4, 4.6)

(a) \( \text{AlCl}_3(aq) + 3 \text{AgNO}_3(aq) \rightarrow \text{AgCl(s)} + \text{Al(NO}_3)_3(aq) \)
(b) \( \text{Li(s)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{LiSO}_4(aq) + \text{H}_2(g) \)
(c) \( \text{H}_2\text{SO}_4(aq) + \text{NaOH(aq)} \rightarrow 2 \text{H}_2\text{O(l)} + \text{Na}_2\text{SO}_4(aq) \)

34. An aqueous solution of lead(II) nitrate is mixed with aqueous sodium iodide. (4.6)

(a) Predict whether these two compounds will react.
(b) Describe how you made your prediction.
(c) If you predict that a reaction will occur, write the balanced chemical reaction for the reaction. Include the states of all reactants and products.
(d) List the precipitates, if any, that form in this reaction.

Analysis and Application

35. A chemist wants to determine whether a certain metal will act as a catalyst in a reaction between zinc and hydrochloric acid. (4.1)

(a) Write the balanced equation for the reaction between zinc and hydrochloric acid.
(b) Write an experimental design outlining an experiment to determine whether the unknown metal is a catalyst for this reaction.

36. Two aqueous solutions, hydrogen iodide and hydrogen peroxide, are mixed. It is unknown whether they give off a gaseous product when they react. Design an experiment to test for the release of a gas using the law of conservation of mass. (4.1)

37. Solid lead(II) nitrate reacts with solid potassium iodide in a double displacement reaction. (4.1)

(a) Write the chemical formula for each reactant and product in this reaction.
(b) Write the balanced equation for this reaction.
(c) How could you use this reaction to demonstrate the law of conservation of mass? Write an experimental design, including any necessary safety precautions.

38. Mercury(II) oxide decomposes when heated in a Bunsen burner flame. (4.2)

(a) Write the chemical formula for mercury(II) oxide.
(b) Write the chemical formulas of the products of the decomposition reaction.
(c) Write the balanced equation for the decomposition of mercury(II) oxide.

39. Hydrogen and bromine are both diatomic elements. (4.2)

(a) Write the formulas for both of these substances.

(b) Predict the product of the synthesis reaction between them.
(c) Write the balanced equation for the synthesis reaction between bromine and hydrogen.

40. A solution of potassium chloride reacts with sodium in a single displacement reaction. (4.4)

(a) Identify the products of this reaction.
(b) Write the balanced equation for this reaction.

41. A chemist wants to know whether zinc reacts with sulfuric acid. (4.4)

(a) How could the chemist find the answer to this question without performing the experiment?
(b) Predict whether zinc reacts with sulfuric acid.

42. The steel hull of a ship is made mostly of iron. Attaching a zinc plate to a ship's hull protects the hull by preventing iron from reacting with water. The zinc plate, however, reacts with water and must be replaced periodically. (4.4)

(a) Use the activity series to suggest a possible explanation for why zinc "protects" the iron in the hull of the ship.
(b) Would a silver plate attached to the ship's hull have the same effect? Explain.
(c) Name one other metal that would be effective in slowing corrosion and one that would not. Explain your choices.

43. A researcher is studying a reaction in which the reactants are aqueous lithium chloride and silver nitrate solution. (4.6)

(a) Write the balanced equation for the reaction that occurs when solutions of lithium chloride and silver nitrate are combined. Include all state symbols.
(b) Describe a chemical test that could be done to be certain that all of the silver ions present had precipitated. Why does your test work?

44. A student combines the following pairs of aqueous solutions. In each case, predict whether a reaction will occur and, if you predict that a reaction will occur, write the balanced equation for this reaction, including all state symbols. (4.6)

(a) magnesium chloride and sodium hydroxide
(b) barium nitrate and potassium carbonate
(c) BaCl_2(aq) and Na_2SO_4(aq)
(d) iron(II) nitrate and potassium sulfate

45. A chemist performs an experiment in which hydrochloric acid is mixed with magnesium hydroxide. (4.2, 4.4, 4.6)

(a) Write the balanced chemical equation for this reaction.
(b) Which type of reaction is this?
46. Silicon carbide, SiC(s), is made by reacting solid silicon dioxide with solid carbon at very high temperatures. The other product of the reaction is carbon monoxide gas. (4.2, 4.4, 4.6)  
(a) Write the balanced chemical equation for this reaction.  
(b) Which type of reaction is this?
47. Write the balanced equation for the neutralization reaction between nitric acid and potassium hydroxide. (4.6) 
48. Hydrofluoric acid is used to etch glass. Used hydrofluoric acid must be neutralized before disposal. (4.6)  
(a) Name one compound that could be used to neutralize the acid.  
(b) Write the balanced equation for the neutralization reaction.

Evaluation
49. A chemical company has a large volume of waste silver nitrate solution. In order to reduce costs, they want to recover the silver from the solution as silver metal. One proposed recovery method is to add solid aluminum to the solution. (4.2, 4.4, 4.6)  
(a) Using what you have learned about chemical reactions, evaluate this proposal.  
(b) If you think the proposal would work, write the balanced chemical equation for the reaction. If you think it would not work, suggest a different chemical reaction to retrieve the silver metal.
50. A water supply has been contaminated with an unknown mixture of toxic metal solutions leaking from a nearby factory. One way to treat the problem is to precipitate the metals from the water. Environmental chemists are considering using sodium carbonate and sodium sulfate for this purpose. Based on the solubility table, which compound would you recommend? (4.2, 4.4, 4.6) 
51. The atmosphere of the planet Venus contains a great deal of sulfuric acid. Sulfuric acid falls as precipitation in much the same way as water falls on Earth. An engineer is designing a probe to send through the atmosphere of Venus. The engineer plans to use a thick plating of copper as the outer skin of the probe to protect it from the corrosive effects of the sulfuric acid. (4.2, 4.4, 4.6)  
(a) Given what you have learned about reactivity, evaluate this proposal.  
(b) List two alternative plating metals for this probe and explain the effectiveness of each of them.

Reflect on Your Learning
52. What did you learn in this chapter that you found most surprising? Explain.  
53. (a) What questions do you still have about chemical reactions?  
(b) What methods can you use to find answers to your questions?  
54. Has the information you have learned in this chapter changed your opinion about any environmental issues related to chemistry? Explain.

Research
55. News stories often report on advances in hydrogen fuel cell technology for transportation. Research how hydrogen fuel cells are used to power vehicles. In your research, find out what problems remain to be solved for this technology. Create a chart that compares the advantages and disadvantages of using this technology. Then write a paragraph that explains your position on the question: Do you think hydrogen-powered cars are a useful technology for society to pursue? Support your position with evidence from your research.  
56. Acid precipitation damages both the natural environment and human-made structures. Research the causes of acid precipitation and what is being done to mitigate its effects. In your research, determine how human activity has affected acid precipitation. Prepare to present your response to the question: What should be done about acid precipitation? Include reasoned arguments to support your suggestions.  
57. Fritz Haber helped develop the Haber process for producing large quantities of ammonia. Research the Haber process to find out about the chemical reaction it uses. Write a short report that explains the importance of this process to our society and addresses the following questions:  
(a) What products would not be possible without the Haber process?  
(b) Do you think the Haber process has helped or hurt society? Explain your reasoning.