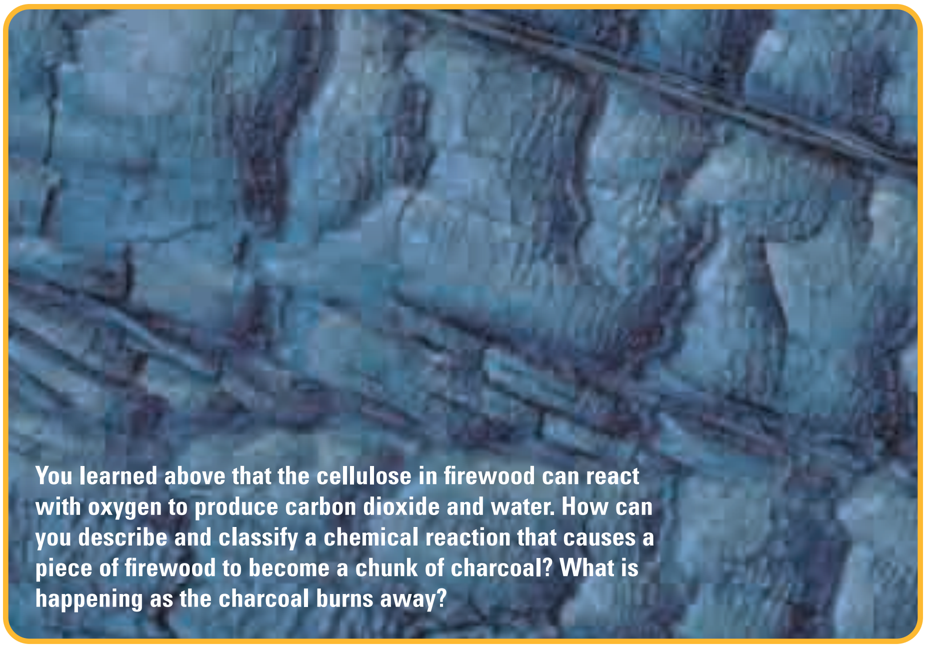
CHAPTER 4

Classifying Reactions: Chemicals in Balance

Picture a starry night and tired hikers sitting around a campfire. Everywhere in this peaceful setting, chemical reactions are taking place. The cellulose in the firewood reacts with the oxygen in the air, producing carbon dioxide and water. The light and heat of the campfire are evidence of the chemical reaction. If someone roasts a marshmallow its sugars react with oxygen. The soft, white marshmallow forms a brown and brittle crust. When someone eats a marshmallow, the chemicals in the stomach react with the sugar molecules to digest them. A person telling a story exhales carbon dioxide with every breath. Carbon dioxide is the product of respiration, another chemical reaction.

In each star in the night sky above, another type of reaction is taking place. This type of reaction is called a **nuclear reaction**, because it involves changes within the nucleus of the atom. Nuclear reactions are responsible for the enormous amounts of heat and light generated by all the stars, including our Sun.

Back on Earth, however, chemical reactions are everywhere in our daily lives. We rely on chemical reactions for everything from powering a car to making toast. In this chapter, you will learn how to write balanced chemical equations for these reactions. You will look for patterns and similarities between the chemical equations, and you will classify the reactions they represent. As well, you will learn how to balance and classify equations for nuclear reactions.



You learned above that the cellulose in firewood can react with oxygen to produce carbon dioxide and water. How can you describe and classify a chemical reaction that causes a piece of firewood to become a chunk of charcoal? What is happening as the charcoal burns away?

Chapter Preview

Concepts and Skills You Will Need

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

- defining and describing the relationships among atomic number, mass number, atomic mass, isotope, and radioisotope (Chapter 2, section 2.1)
- naming chemical compounds (Chapter 3, section 3.4)
- writing chemical formulas (Chapter 3, section 3.4)
- explaining how different elements combine to form covalent and ionic bonds using the octet rule (Chapter 3, sections 3.2 and 3.3)

4.1 Chemical Equations

Section Preview/ Specific Expectations

In this section, you will

- **use** word equations and skeleton equations to describe chemical reactions
- **balance** chemical equations
- **communicate** your understanding of the following terms: *chemical reactions, reactant, product, chemical equations, word equation, skeleton equation, law of conservation of mass, balanced chemical equation*

In Chapter 3, you learned how and why elements combine to form different compounds. In this section, you will learn how to describe what happens when elements and compounds interact with one another to form new substances. These interactions are called **chemical reactions**. A substance that undergoes a chemical reaction is called a **reactant**. A substance that is formed in a chemical reaction is called a **product**.

For example, when the glucose in a marshmallow reacts with oxygen in the air to form water and carbon dioxide, the glucose and oxygen are the reactants. The carbon dioxide and water are the products. Chemists use **chemical equations** to communicate what is occurring in a chemical reaction. Chemical equations come in several forms. All of these forms condense a great deal of chemical information into a short statement.

Word Equations

A **word equation** identifies the reactants and products of a chemical reaction by name. In Chapter 3, you learned that chlorine and sodium combine to form the ionic compound sodium chloride. This reaction can be represented by the following word equation:

$$\text{sodium} + \text{chlorine} \rightarrow \text{sodium chloride}$$

In this equation, "+" means "reacts with" and "→" means "to form." Try writing some word equations in the following Practice Problems.

CHECKPOINT

Write the chemical formulas of the products in the reactions described in Practice Problem 1.

Practice Problems

1. Describe each reaction using a word equation. Label the reactant(s) and product(s).
   (a) Calcium and fluorine react to form calcium fluoride.
   (b) Barium chloride and hydrogen sulfate react to form hydrogen chloride and barium sulfate.
   (c) Calcium carbonate, carbon dioxide, and water react to form calcium hydrogen carbonate.
   (d) Hydrogen peroxide reacts to form water and oxygen.
   (e) Sulfur dioxide and oxygen react to form sulfur trioxide.
2. Yeast can facilitate a reaction in which the sugar in grapes reacts to form ethanol and carbon dioxide. Write a word equation to describe this reaction.

CHEM FACT

Caffeine is an ingredient in coffee, tea, chocolate, and cola drinks. Its chemical name is 1,3,7-trimethylxanthene. You can see how long names like this would become unwieldy in word equations.

Word equations are useful because they identify the products and reactants in a chemical reaction. They do not, however, provide any chemical information about the compounds and elements themselves. If you did not know the formula for sodium chloride, for example, this equation would not help you understand the reaction very well. Another shortcoming of a word equation is that the names for chemicals are often very long and cumbersome. Chemists have therefore devised a more convenient way of representing reactants and products.

Skeleton Equations

Using a chemical formula instead of a chemical name simplifies a chemical equation. It allows you to see at a glance what elements and compounds are involved in the reaction. A **skeleton equation** lists the chemical formula of each reactant on the left, separated by a + sign if more than one reactant is involved, followed by an arrow →. The chemical formula of each product is listed on the right, again separated by a + sign if more than one product is produced. A skeleton equation also shows the state of each reactant by using the appropriate subscript, as shown in Table 4.1.

The reaction of sodium metal with chlorine gas to form sodium chloride can be represented by the following skeleton equation:

$$Na_{(s)} + Cl_{2(g)} \rightarrow NaCl_{(s)}$$

A skeleton equation is more useful to a chemist than a word equation, because it shows the formulas of the compounds involved. It also shows the state of each substance.Try writing some skeleton equations in the following Practice Problems.

Table 4.1 Symbols Used in Chemical Equations

Symbol	Meaning
+	reacts with (reactant side)
+	and (product side)
→	to form
(s)	solid or precipitate
(ℓ)	liquid
(g)	gas
(aq)	in aqueous (water) solution

Practice Problems

3. Write a skeleton equation for each reaction.
- **(a)** Solid zinc reacts with chlorine gas to form solid zinc chloride.
- **(b)** Solid calcium and liquid water react to form solid calcium hydroxide and hydrogen gas.
- **(c)** Solid barium reacts with solid sulfur to produce solid barium sulfide.
- **(d)** Aqueous lead(II) nitrate and solid magnesium react to form aqueous magnesium nitrate and solid lead.

4. In each reaction below, a solid reacts with a gas to form a solid. Write a skeleton equation for each reaction.
- **(a)** carbon dioxide + calcium oxide → calcium carbonate
- **(b)** aluminum + oxygen → aluminum oxide
- **(c)** magnesium + oxygen → magnesium oxide

Why Skeleton Equations Are Incomplete

Although skeleton equations are useful, they do not fully describe chemical reactions. To understand why, consider the skeleton equation showing the formation of sodium chloride (above). According to this equation, one sodium atom reacts with one chlorine molecule containing two chlorine atoms. The product is one formula unit of sodium chloride, containing one atom of sodium and one atom of chlorine. Where has the extra chlorine atom gone?

The Law of Conservation of Mass

All atoms must be accounted for, according to an important law. The **law of conservation of mass** states that *in any chemical reaction, the mass of the products is always equal to the mass of the reactants*. In other words, according to this law, matter can be neither created nor destroyed. Chemical reactions proceed according to the law of conservation of mass, which is based on experimental evidence.

CHEM FACT

Many chemical reactions can go in either direction, so an arrow pointing in the opposite direction is often added to the equation. This can look like $\rightleftarrows$ or $\rightleftharpoons$. To indicate which reaction is more likely to occur, one arrow can be drawn longer than the other: for example: $\overset{\rightarrow}{\longleftarrow}$ or $\overset{\longrightarrow}{\leftarrow}$.

Careers in Chemistry

Food Chemist

How do you make a better tasting sports drink? How can you make a gravy mix that can be ready to serve in 5 min, yet maintain its consistency under a heat lamp for 8 h? Food chemists use their knowledge of chemical reactions to improve food quality and develop new products.

A good example of food chemistry in action is the red pimento stuffing in olives. Chopped pimentos (sweet red peppers) and sodium alginate are mixed. This mixture is then added to a solution of calcium chloride. The sodium alginate reacts with the calcium chloride. Solid calcium alginate forms, which causes the stream of pimento mixture to form a gel instantly. The gelled strip is then sliced thinly and stuffed into the olives.

Food chemists work in universities, government laboratories, and major food companies. To become a food chemist, most undergraduates take a food science degree with courses in chemistry. It is also possible to become a food chemist with an undergraduate chemistry degree plus experience in the food industry. Students can specialize in food chemistry at the graduate level.

Make Career Connections

1. If you are interested in becoming a food chemist, you can look for a summer or part-time job in the food industry.
2. To find out more about food science, search for Agriculture and Agri-Food Canada's web site and the Food Web web site. You can also contact the Food Science department of a university.

History LINK

Jan Baptista van Helmont (1577–1644) was a Flemish physician who left medical practice to devote himself to the study of chemistry. He used the mass balance in an important experiment that laid the foundations for the law of conservation of mass. He showed that a definite quantity of sand could be fused with excess alkali to form a kind of glass. He also showed that when this product was treated with acid, it regenerated the original amount of sand (silica). As well, Van Helmont is famous for demonstrating the existence of gases, which he described as "aerial fluids." Investigate on the Internet or in the library to find out how he did this.

Balanced Chemical Equations

A **balanced chemical equation** reflects the law of conservation of mass. This type of equation shows that there is the same number of each kind of atom on both sides of the equation. Some skeleton equations are, by coincidence, already balanced. For example, examine the reaction of carbon with oxygen to form carbon dioxide, shown in Figure 4.1. In the skeleton equation, one carbon atom and two oxygen atoms are on the left side of the equation, and one carbon atom and two oxygen atoms are on the right side of the equation.

$$C + O_2 \rightarrow CO_2$$

Figure 4.1 This skeleton equation is already balanced.

Most skeleton equations, however, are not balanced, such as the one showing the formation of sodium chloride. Examine Figure 4.2 to see why. There is one sodium atom on each side of the equation, but there are two chlorine atoms on the left side and only one chlorine atom on the right side.

To begin to balance an equation, you can add numbers in front of the appropriate formulas. The numbers that are placed in front of chemical formulas are called **coefficients**. They represent how many of each atom, molecule, or formula unit take part in each reaction. For example, if you add a coefficient of 2 to NaCl in the equation in Figure 4.2, you indicate that two formula units of NaCl are produced in the reaction. Is the equation balanced now? As you can see by examining Figure 4.3, it is not. The chlorine atoms are balanced, but now there is one sodium atom on the left side of the equation and two sodium atoms on the right side.

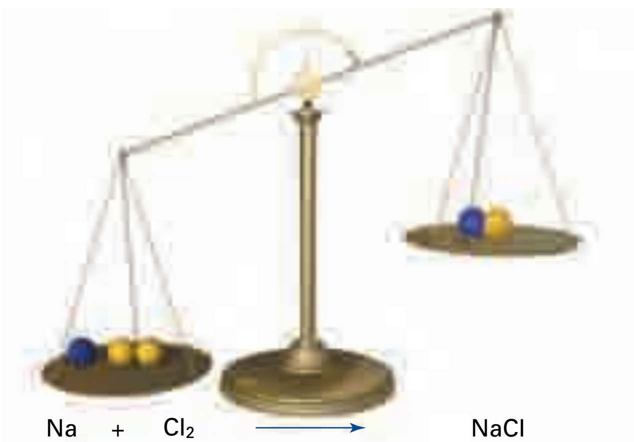


Figure 4.2 This skeleton equation is unbalanced. The mass of the reactants is greater than the mass of the product.

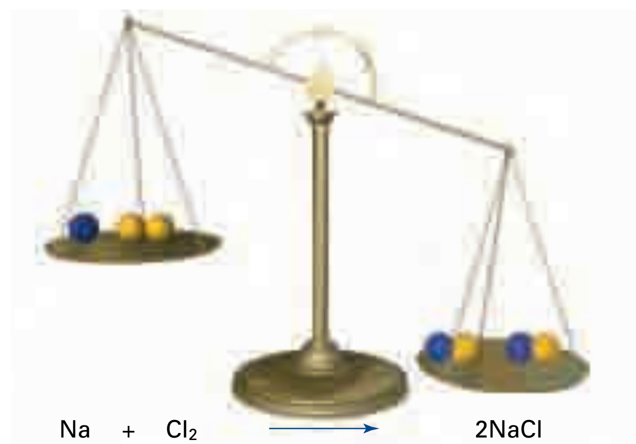


Figure 4.3 The equation is still unbalanced. The mass of the product is now greater than the mass of the reactants.

Add a coefficient of 2 to the sodium on the reactant side. As you can see in Figure 4.4, the equation is now balanced. The mass of the products is equal to the mass of the reactants. This balanced chemical equation satisfies the law of conservation of mass.

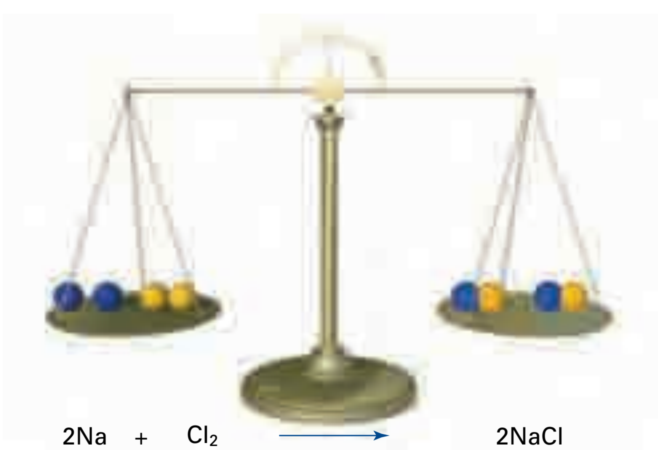


Figure 4.4 The equation is now balanced according to the law of conservation of mass.

mind STRETCH

Why is it *not* acceptable to balance the equation

$$Na + Cl_2 \rightarrow NaCl$$

by changing the formula of NaCl to $NaCl_2$? Would this not satisfy the law of conservation of mass? Write an explanation in your notebook.

You cannot balance an equation by changing any of the chemical formulas. The only way to balance a chemical equation is to put the appropriate numerical coefficient in front of each compound or element in the equation.

Many skeleton equations are simple enough to balance by a back-and-forth process of reasoning, as you just saw with the sodium chloride reaction. Try balancing the equations in the Practice Problems that follow.

Practice Problems

5. Copy each skeleton equation into your notebook, and balance it.

(a) $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$

(b) $P_{4(s)} + O_{2(g)} \rightarrow P_4O_{10(s)}$

(c) $H_{2(g)} + Cl_{2(g)} \rightarrow HCl_{(g)}$

(d) $SO_{2(g)} + H_2O_{(\ell)} \rightarrow H_2SO_{3(aq)}$

6. Indicate whether these equations are balanced. If they are not, balance them.

(a) $4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$

(b) $HgO_{(s)} \rightarrow Hg_{(\ell)} + O_{2(g)}$

(c) $H_2O_{2(aq)} \rightarrow 2H_2O_{(\ell)} + O_{2(g)}$

(d) $2HCl_{(aq)} + Na_2SO_{3(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(\ell)} + SO_{2(g)}$

Steps for Balancing Chemical Equations

More complex chemical equations than the ones you have already tried can be balanced by using a combination of inspection and trial and error. Here, however, are some steps to follow.

Step 1 Write out the skeleton equation. Ensure that you have copied all the chemical formulas correctly.

Step 2 Begin by balancing the atoms that occur in the largest number on either side of the equation. Leave hydrogen, oxygen, and other elements until later.

Step 3 Balance any polyatomic ions, such as sulfate, SO_4^{2-}, that occur on both sides of the chemical equation as an ion unit. That is, do not split a sulfate ion into 1 sulfur atom and 4 oxygen atoms. Balance this ion as one unit.

Step 4 Next, balance any hydrogen or oxygen atoms that occur in a combined and uncombined state. For example, combined oxygen might be in the form of CO_2, while uncombined oxygen occurs as O_2.

Step 5 Finally, balance any other element that occurs in its uncombined state: for example, Na or Cl_2.

Step 6 Check your answer. Count the number of each type of atom on each side of the equation. Make sure that the coefficients used are whole numbers in their lowest terms.

Examine the following Sample Problem to see how these steps work.

Electronic Learning Partner

Go to the Chemistry 11 Electronic Learning Partner for some extra practice balancing chemical equations.

Sample Problem

Balancing Chemical Equations

Problem

Copper(II) nitrate reacts with potassium hydroxide to form potassium nitrate and solid copper(II) hydroxide.
Balance the equation.

$$Cu(NO_3)_{2(aq)} + KOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + KNO_{3(aq)}$$

What Is Required?

The atoms of each element on the left side of the equation should equal the atoms of each element on the right side of the equation.

Plan Your Strategy

Balance the polyatomic ions first (NO_3^-, then OH^-). Check to see whether the equation is balanced. If not, balance the potassium and copper ions. Check your equation again.

Act on Your Strategy

There are two NO_3^- ions on the left, so put a 2 in front of KNO_3.

$$Cu(NO_3)_{2(aq)} + KOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + 2KNO_{3(aq)}$$

To balance the two OH^- ions on the right, put a 2 in front of the KOH.

$$Cu(NO_3)_{2(aq)} + 2KOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + 2KNO_{3(aq)}$$

Check to see that the copper and potassium ions are balanced. They are, so the equation above is balanced.

Check Your Solution

Tally the number of each type of atom on each side of the equation.

$$Cu(NO_3)_{2(aq)} + 2KOH_{(aq)} \rightarrow Cu(OH)_{2(s)} + 2KNO_{3(aq)}$$

	Left Side	Right Side
Cu	1	1
NO_3^-	2	2
K	2	2
OH^-	2	2

Practice Problems

7. Copy each chemical equation into your notebook, and balance it.

(a) $SO_{2(g)} + O_{2(g)} \rightarrow SO_{3(g)}$

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

8. When solid white phosphorus, P_4, is burned in air, it reacts with oxygen to produce solid tetraphosphorus decoxide, P_4O_{10}. When water is added to the P_4O_{10}, it reacts to form aqueous phosphoric acid, H_3PO_4. Write and balance the chemical equations that represent these reactions.

Continued ...

Math LINK

What does it mean when a fraction is expressed in lowest terms? The fraction $\frac{5}{10}$, expressed in lowest terms, is $\frac{1}{2}$. Similarly, the equation

$4H_{2(g)} + 2O_{2(g)} \rightarrow 4H_2O_{(\ell)}$

is balanced, but it can be simplified by dividing all the coefficients by two.

$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)}$

Write the balanced equation $6KClO_{3(s)} \rightarrow 6KCl_{(g)} + 9O_{2(s)}$ so that the coefficients are the lowest possible whole numbers. Check that the equation is still balanced.

Continued ...
FROM PAGE 117

9. Copy each chemical equation into your notebook, and balance it.

(a) $As_4S_{6(s)} + O_{2(g)} \rightarrow As_4O_{6(s)} + SO_{2(g)}$

(b) $Sc_2O_{3(s)} + H_2O_{(\ell)} \rightarrow Sc(OH)_{3(s)}$

(c) $C_2H_5OH_{(\ell)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(\ell)}$

(d) $C_4H_{10(g)} + O_{2(g)} \rightarrow CO_{(g)} + H_2O_{(g)}$

Section Wrap-up

In this section, you learned how to represent chemical reactions using balanced chemical equations. Because there are so many different chemical reactions, chemists have devised different classifications for these reactions. In section 4.2, you will learn about five different types of chemical reactions.

Section Review

1 MC In your own words, explain what a chemical reaction is. Write descriptions of four chemical reactions that you encounter every day.

2 C Write a word equation, skeleton equation, and balanced equation for each reaction.

(a) Sulfur dioxide gas reacts with oxygen gas to produce gaseous sulfur trioxide.

(b) Metallic sodium reacts with liquid water to produce hydrogen gas and aqueous sodium hydroxide.

(c) Copper metal reacts with an aqueous hydrogen nitrate solution to produce aqueous copper(II) nitrate, nitrogen dioxide gas, and liquid water.

3 K/U The equation for the decomposition of hydrogen peroxide, H_2O_2 is $H_2O_{2(aq)} \rightarrow H_2O_{(\ell)} + O_{2(g)}$. Explain why you cannot balance it by writing it as $H_2O_{2(aq)} \rightarrow H_2O_{(\ell)} + O_{(g)}$

4 K/U Balance the following chemical equations.

(a) $Al_{(s)} + O_{2(g)} \rightarrow Al_2O_{3(s)}$

(b) $Na_2S_2O_{3(aq)} + I_{2(aq)} \rightarrow NaI_{(aq)} + Na_2S_4O_{6(aq)}$

(c) $Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + Fe_{(s)}$

(d) $NH_{3(g)} + O_{2(g)} \rightarrow NO_{(g)} + H_2O_{(\ell)}$

(e) $Na_2O_{(s)} + (NH_4)_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(\ell)} + NH_{3(aq)}$

(f) $C_5H_{12(\ell)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

5 I A student places 0.58 g of iron and 1.600 g of copper(II) sulfate in a reaction vessel. The reaction vessel has a mass of 40.32 g, and it contains 100.00 g of water. The aqueous copper sulfate and solid iron react to form solid copper and aqueous iron(II) sulfate. After the reaction, the reaction vessel plus the products have a mass of 142.5 g. Explain the results. Then write a balanced chemical equation to describe the reaction.

Synthesis and Decomposition Reactions

4.2

How are different kinds of compounds formed? In section 4.1, you learned that they are formed by chemical reactions that you can describe using balanced chemical equations. Just as there are different types of compounds, there are different types of chemical reactions. In this section, you will learn about five major classifications for chemical reactions. You will use your understanding of chemical formulas and chemical equations to predict products for each class of reaction.

Section Preview/ Specific Expectations

In this section, you will

- **distinguish** between synthesis, decomposition, and combustion reactions
- **write** balanced chemical equations to represent synthesis, decomposition, and combustion reactions
- **predict** the products of chemical reactions
- **demonstrate** an understanding of the relationship between the type of chemical reaction and the nature of the reactants
- **communicate** your understanding of the following terms: *synthesis reaction, decomposition reaction, combustion reaction, incomplete combustion*

Why Classify?

People use classifications all the time. For example, many types of wild mushrooms are edible, but many others are poisonous—even deadly! How can you tell which is which? Poisonous and deadly mushrooms have characteristics that distinguish them from edible ones, such as odour, colour, habitat, and shape of roots. It is not always easy to distinguish one type of mushroom from another; the only visible difference may be the colour of the mushroom's spores. Therefore, you should never try to eat any wild mushrooms without an expert's advice.

Examine Figure 4.5. Which mushroom looks more appetizing to you? An expert will always be able to distinguish an edible mushroom from a poisonous mushroom based on the characteristics that have been used to classify each type. By classifying, they can predict the effects of eating any wild mushroom.

Figure 4.5 **The mushroom on the left, called a chanterelle, is edible and very tasty. The mushroom on the right is called a death cap. It is extremely poisonous.**

In the same way, you can recognize similarities between chemical reactions and the types of reactants that tend to undergo different types of reactions. With this knowledge, you can predict what will happen when one, two, or more substances react. In this section, you will often see chemical reactions without the subscripts showing the states of matter. They are omitted deliberately because, in most cases, you are not yet in a position to predict the states of the products.

Synthesis Reactions

In a **synthesis reaction**, two or more elements or compounds combine to form a new substance. Synthesis reactions are also known as combination or formation reactions. A general equation for a synthesis reaction is

$$A + B \rightarrow C$$

In a simple synthesis reaction, one element reacts with one or more other elements to form a compound. Two, three, four, or more elements may react to form a single product, although synthesis reactions involving four or more reactants are extremely rare. Why do you think this is so? When two elements react together, the reaction is almost always a synthesis reaction because the product is almost always a single compound. There are several types of synthesis reactions. Recognizing the patterns of the various types of reactions will help you to predict whether substances will take part in a synthesis reaction.

When a metal or a non-metal element reacts with oxygen, the product is an oxide. Figure 4.6 shows a familiar example, in which iron reacts with oxygen according to the following equation:

$$3Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$$

Figure 4.6 **The iron in this car undergoes a synthesis reaction with the oxygen in the air. Iron(III) oxide, also known as rust, is formed.**

Two other examples of this type of reaction are:

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

A second type of synthesis reaction involves the reaction of a metal and a non-metal to form a binary compound. One example is the reaction of potassium with chlorine.

$$2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$$

Synthesis Reactions Involving Compounds

In the previous two types of synthesis reactions, two elements reacted to form one product. There are many synthesis reactions in which one or more compounds are the reactants. For the purpose of this course, however, we will deal only with the two specific types of synthesis reactions involving compounds that you should recognize: oxides and water.

When a non-metallic oxide reacts with water, the product is an acid. You will learn more about acids and the rules for naming them in Chapter 10. The acids that form when non-metallic oxides and water react are composed of hydrogen cations and polyatomic anions containing oxygen and a non-metal. For example, one contributor to acid rain is hydrogen sulfate (sulfuric acid), H_2SO_4, which forms when sulfur trioxide reacts with water. The sulfur trioxide comes from sources such as industrial plants that emit the gas as a byproduct of burning fossil fuels, as shown in Figure 4.7.

$$SO_{3(g)} + H_2O_{(\ell)} \rightarrow H_2SO_{4(aq)}$$

Figure 4.7 **Sulfur trioxide, emitted by this factory, reacts with the water in the air. Sulfuric acid is formed in a synthesis reaction.**

Conversely, when a metallic oxide reacts with water, the product is a metal hydroxide. Metal hydroxides belong to a group of compounds called bases. You will learn more about bases in Chapter 10. For example, when calcium oxide reacts with water, it forms calcium hydroxide, $Ca(OH)_2$. Calcium oxide is also called lime. It can be added to lakes to counteract the effects of acid precipitation.

$$CaO_{(s)} + H_2O_{(\ell)} \rightarrow Ca(OH)_{2(aq)}$$

Sometimes it is difficult to predict the product of a synthesis reaction. The only way to really know the product of a reaction is to carry out the reaction and then isolate and identify the product. For example, carbon can react with oxygen to form either carbon monoxide or carbon dioxide. Therefore, if all you know is that your reactants are carbon and oxygen, you cannot predict with certainty which compound will form. You can only give options.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

$$2C_{(s)} + O_{2(g)} \rightarrow 2CO_{(g)}$$

You would need to analyze the products of the reaction by experiment to determine which compound was formed.

History LINK

Today we have sophisticated lab equipment to help us analyze the products of reactions. In the past, when such equipment was not available, chemists sometimes jeopardized their safety and health to determine the products of the reactions they studied. Sir Humphry Davy (1778–1829), a contributor to many areas of chemistry, thought nothing of inhaling the gaseous products of the chemical reactions that he carried out. He tried to breathe pure CO_2, then known as *fixed air*. He nearly suffocated himself by breathing hydrogen. In 1800, Davy inhaled dinitrogen monoxide, N_2O, otherwise known as nitrous oxide, and discovered its anaesthetic properties. What is nitrous oxide used for today?

CHECKPOINT

As you begin learning about different types of chemical reactions, keep a separate list of each type of reaction. Add to the list as you encounter new reactions.

Try predicting the products of synthesis reactions in the following Practice Problems.

Practice Problems

10. Copy the following synthesis reactions into your notebook. Predict the product of each reaction. Then balance each chemical equation.

(a) $K + Br_2 \rightarrow$

(b) $H_2 + Cl_2 \rightarrow$

(c) $Ca + Cl_2 \rightarrow$

(d) $Li + O_2 \rightarrow$

11. Copy the following synthesis reactions into your notebook. For each set of reactants, write the equations that represent the possible products.

(a) $Fe + O_2 \rightarrow$
(suggest two different synthesis reactions)

(b) $V + O_2 \rightarrow$
(suggest four different synthesis reactions)

(c) $Co + Cl_2 \rightarrow$
(suggest two different synthesis reactions)

(d) $Ti + O_2 \rightarrow$
(suggest three different synthesis reactions)

12. Copy the following equations into your notebook. Write the product of each reaction. Then balance each chemical equation.

(a) $K_2O + H_2O \rightarrow$

(b) $MgO + H_2O \rightarrow$

(c) $SO_2 + H_2O \rightarrow$

13. Ammonia gas and hydrogen chloride gas react to form a solid compound. Predict what the solid compound is. Then write a balanced chemical equation.

Figure 4.8 As electricity passes through the water, it decomposes to hydrogen and oxygen gas.

Decomposition Reactions

In a **decomposition reaction**, a compound breaks down into elements or other compounds. Therefore, *a decomposition reaction is the opposite of a synthesis reaction*. A general formula for a decomposition reaction is:

$$C \rightarrow A + B$$

The substances that are produced in a decomposition reaction can be elements or compounds. In the simplest type of decomposition reaction, a compound breaks down into its component elements. One example is the decomposition of water into hydrogen and oxygen. This reaction occurs when electricity is passed through water. Figure 4.8 shows an apparatus set up for the decomposition of water.

$$2H_2O \rightarrow 2H_2 + O_2$$

More complex decomposition reactions occur when compounds break down into other compounds. An example of this type of reaction is shown in Figure 4.9. The photograph shows the explosive decomposition of ammonium nitrate.

When ammonium nitrate is heated to a high temperature, it forms dinitrogen monoxide and water according to the following balanced equation:

$$NH_4NO_{3(s)} \rightarrow N_2O_{(g)} + 2H_2O_{(g)}$$

Try predicting the products of the decomposition reactions in the following Practice Problems.

Figure 4.9 At high temperatures, ammonium nitrate explodes, decomposing into dinitrogen monoxide and water.

Practice Problems

14. Mercury(II) oxide, or mercuric oxide, is a bright red powder. It decomposes on heating. What are the products of the decomposition of HgO?

15. What are the products of the following decomposition reactions? Predict the products. Then write a balanced equation for each reaction.

(a) $HI \rightarrow$ **(c)** $AlCl_3 \rightarrow$

(b) $Ag_2O \rightarrow$ **(d)** $MgO \rightarrow$

16. Calcium carbonate decomposes into calcium oxide and carbon dioxide when it is heated. Based on this information, predict the products of the following decomposition reactions.

(a) $MgCO_3 \rightarrow$

(b) $CuCO_3 \rightarrow$

Combustion Reactions

Combustion reactions form an important class of chemical reactions. The combustion of fuel—wood, fossil fuel, peat, or dung—has, throughout history, heated and lit our homes and cooked our food. The energy produced by combustion reactions moves our airplanes, trains, trucks, and cars.

A complete **combustion reaction** is the reaction of a compound or element with O_2 to form the most common oxides of the elements that make up the compound. For example, a carbon-containing compound undergoes combustion to form carbon dioxide, CO_2. A sulfur-containing compound reacts with oxygen to form sulfur dioxide, SO_2.

Combustion reactions are usually accompanied by the production of light and heat. In the case of carbon-containing compounds, complete combustion results in the formation of, among other things, carbon dioxide. For example, methane, CH_4, the primary constituent of natural gas, undergoes complete combustion to form carbon dioxide, (the most common oxide of carbon), as well as water. This combustion reaction is represented by the following equation:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

The combustion of methane, shown in Figure 4.10, leads to the formation of carbon dioxide and water.

The complete combustion of any compound that contains carbon, hydrogen, and oxygen (such as ethanol, C_2H_5OH) produces carbon dioxide and water.

Figure 4.10 This photo shows the combustion of methane in a laboratory burner.

Compounds that contain elements other than carbon also undergo complete combustion to form stable oxides. For instance, sulfur-containing compounds undergo combustion to form sulfur dioxide, SO_2, a precursor to acid rain. Complete combustion reactions are often also synthesis reactions. Metals, such as magnesium, undergo combustion to form their most stable oxide, as shown in Figure 4.11.

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

Figure 4.11 **Magnesium metal burns in oxygen. The smoke and ash that are produced in this combustion reaction are magnesium oxide.**

In the absence of sufficient oxygen, carbon-containing compounds undergo **incomplete combustion**, leading to the formation of carbon monoxide, CO, and water. Carbon monoxide is a deadly gas. You should always make sure that sufficient oxygen is present in your indoor environment for your gas furnace, gas stove, or fireplace.

Try the following problems to practise balancing combustion reactions.

CHECKPOINT

Copy the following skeleton equation for the combustion of pentane, C_5H_{12}, into your notebook and balance it.

$C_5H_{12} + O_2 \rightarrow CO_2 + H_2O$

If it took you a long time to balance this equation, chances are that you did not use the quickest method. Try balancing carbon first, hydrogen second, and finally oxygen. What is the advantage of leaving O_2 until the end? Would this method work for incomplete combustion reactions? Would this method work if the "fuel" contained oxygen in addition to C and H? Now try balancing the chemical equation for the combustion of heptane (C_7H_{16}) and the combustion of rubbing alcohol, isopropanol (C_3H_8O).

Practice Problems

17. The alcohol lamps that are used in some science labs are often fuelled with methanol, CH_3OH. Write the balanced chemical equation for the complete combustion of methanol.

18. Gasoline is a mixture of compounds containing hydrogen and carbon, such as octane, C_8H_{18}. Write the balanced chemical equation for the complete combustion of C_8H_{18}.

19. Acetone, C_3H_6O, is often contained in nail polish remover. Write the balanced chemical equation for the complete combustion of acetone.

20. Kerosene consists of a mixture of hydrocarbons. It has many uses including jet fuel and rocket fuel. It is also used as a fuel for hurricane lamps. If we represent kerosene as $C_{16}H_{34}$, write a balanced chemical equation for the complete combustion of kerosene.

Section Wrap-up

In this section, you learned about three major types of reactions: synthesis, decomposition, and combustion reactions. Using your knowledge about these types of reactions, you learned how to predict the products of various reactants. In section 5.3, you will increase your understanding of chemical reactions even further, learning about two major types of chemical reactions. As well, you will observe various chemical reactions in three investigations.

Section Review

1 K/U Write the product for each synthesis reaction. Balance the chemical equation.

(a) $Be + O_2 \rightarrow$

(b) $Li + Cl_2 \rightarrow$

(c) $Mg + N_2 \rightarrow$

(d) $Al + Br_2 \rightarrow$

(e) $K + O_2 \rightarrow$

2 K/U Write the products for each decomposition reaction. Balance the chemical equation.

(a) $K_2O \rightarrow$

(b) $CuO \rightarrow$

(c) $H_2O \rightarrow$

(d) $Ni_2O_3 \rightarrow$

(e) $Ag_2O \rightarrow$

3 C Write a balanced chemical equation for each of the following word equations. Classify each reaction.

(a) With heating, solid tin(IV) hydroxide produces solid tin(IV) oxide and water vapour.

(b) Chlorine gas reacts with crystals of iodine to form iodine trichloride.

4 K/U Write a balanced chemical equation for the combustion of butanol, C_4H_9OH.

5 I A red compound was heated, and the two products were collected. The gaseous product caused a glowing splint to burn brightly. The other product was a shiny pure metal, which was a liquid at room temperature. Write the most likely reaction that would explain these results. Classify the reaction. **Hint:** Remember that the periodic table identifies the most common valences.

6 MC Explain why gaseous nitrogen oxides emitted by automobiles and industries contribute to acid rain. Write balanced chemical reactions to back up your ideas. You may need to look up chemical formulas for your products.

Unit Project Prep

Before you design your Chemistry Newsletter at the end of Unit 1, consider that fuels are composed of compounds containing hydrogen and carbon (hydrocarbons). What kind of reaction have you seen in this section that involves those kinds of compounds? What type of warning would you expect to see on a container of lawnmower fuel? How is the warning related to the types of reaction that involve hydrocarbons?

4.3 Single Displacement and Double Displacement Reactions

Section Preview/ Specific Expectations

In this section, you will

- **distinguish** between synthesis, decomposition, combustion, single displacement, and double displacement reactions
- **write** balanced chemical equations to represent single displacement and double displacement reactions
- **predict** the products of chemical reactions and **test** your predictions through experimentation
- **demonstrate** an understanding of the relationship between the type of chemical reaction and the nature of the reactants
- **investigate**, through experimentation, the reactivity of different metals to produce an activity series
- **communicate** your understanding of the following terms: *single displacement reaction, activity series, double displacement reaction, precipitate, neutralization reactions*

In section 4.2, you learned about three different types of chemical reactions. In section 4.3, you will learn about two more types of reactions. You will learn how performing these reactions can help you make inferences about the properties of the elements and compounds involved.

Single Displacement Reactions

In a **single displacement reaction**, one element in a compound is displaced (or replaced) by another element. Two general reactions represent two different types of single displacement reactions. One type involves a metal replacing a metal cation in a compound, as follows:

$$A + BC \rightarrow AC + B$$

For example, $Zn_{(s)} + Fe(NO_3)_{2(aq)} \rightarrow Zn(NO_3)_{2(aq)} + Fe_{2(s)}$

The second type of single displacement reaction involves a non-metal (usually a halogen) replacing an anion in a compound, as follows:

$$DE + F \rightarrow DF + E$$

For example, $Cl_{2(g)} + CaBr_{2(aq)} \rightarrow CaCl_{2(aq)} + Br_{2(\ell)}$

Single Displacement Reactions and the Metal Activity Series

Most single displacement reactions involve one metal displacing another metal from a compound. In the following equation, magnesium metal replaces the zinc in $ZnCl_2$, thereby liberating zinc as the free metal.

$$Mg_{(s)} + ZnCl_{2(aq)} \rightarrow MgCl_{2(aq)} + Zn_{(s)}$$

The following three reactions illustrate the various types of single displacement reactions involving metals:

1. $Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$
   In this reaction, *one metal replaces another metal in an ionic compound*. That is, copper replaces silver in $AgNO_3$. Because of the +2 charge on the copper ion, it requires two nitrate ions to balance its charge.

2. $Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
   In this reaction, magnesium metal replaces hydrogen from hydrochloric acid, $HCl_{(aq)}$. Since hydrogen is diatomic, it is "liberated" in the form of H_2. This reaction is similar to reaction 1 if
   - you treat hydrochloric acid as an ionic compound (which it technically is not), and if
   - you treat hydrogen as a metal (also, technically, not the case).

3. $2Na_{(s)} + 2H_2O_{(\ell)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$
   Sodium metal displaces hydrogen from water in this reaction. Again, since hydrogen is diatomic, it is produced as H_2. As above, you can understand this reaction better if
   - you treat hydrogen as a metal, and if
   - you treat water as an ionic compound, $H^+(OH^-)$.

All of the reactions just described follow the original general example of a single displacement reaction:

$$A + BC \rightarrow AC + B$$

Figures 4.12 and 4.13 show two examples of single displacement reactions.

When analyzing single displacement reactions, use the following guidelines:

- Treat hydrogen as a metal.
- Treat acids, such as HCl, as ionic compounds of the form H^+Cl^-. (Treat sulfuric acid, H_2SO_4, as $H^+H^+SO_4^{2-}$).
- Treat water as ionic, with the formula $H^+(OH^-)$.

Figure 4.12 **Lithium metal reacts violently with water in a single displacement reaction. Lithium must be stored under kerosene or oil to avoid reaction with atmospheric moisture, or oxygen.**

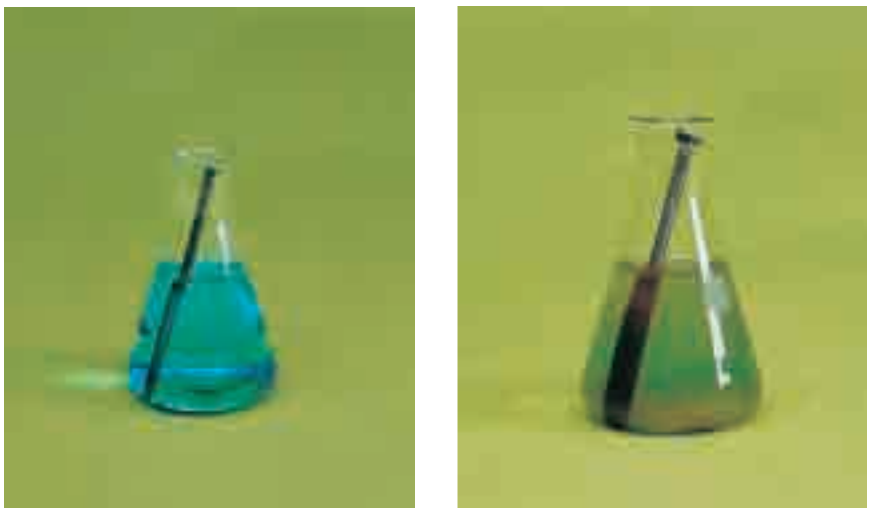


Figure 4.13 **When an iron nail is placed in a solution of copper(II) sulfate, a single displacement reaction takes place.**
$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$
Notice the formation of copper metal on the nail.

Practice Problems

21. Each of the following incomplete equations represents a single displacement reaction. Copy each equation into your notebook, and write the products. Balance each chemical equation. When in doubt, use the most common valence.

(a) $Ca + H_2O \rightarrow$	**(e)** $Pb + H_2SO_4 \rightarrow$
(b) $Zn + Pb(NO_3)_2 \rightarrow$	**(f)** $Mg + Pt(OH)_4 \rightarrow$
(c) $Al + HCl \rightarrow$	**(g)** $Ba + FeCl_2 \rightarrow$
(d) $Cu + AgNO_3 \rightarrow$	**(h)** $Fe + Co(ClO_3)_2 \rightarrow$

Through experimentation, chemists have ranked the relative reactivity of the metals, including hydrogen (in acids and in water), in an **activity series**. The reactive metals, such as potassium, are at the top of the activity series. The unreactive metals, such as gold, are at the bottom. In Investigation 4-A, you will develop an activity series using single displacement reactions.

Investigation 4-A

MICROSCALE

SKILL FOCUS
- Predicting
- Performing and recording
- Analyzing and interpreting

Creating an Activity Series of Metals

Certain metals, such as silver and gold, are extremely unreactive, while sodium is so reactive that it will react with water. Zinc is unreactive with water. It *will*, however, react with acid. Why will magnesium metal react with copper sulfate solution, while copper metal will not react with aqueous magnesium sulfate? In Chapter 3, you learned that an alloy is a solution of two or more metals. Steel is an alloy that contains mostly iron. Is its reactivity different from iron's reactivity?

Question

How can you rank the metals, including hydrogen, in terms of their reactivity? Is the reactivity of an alloy very different from the reactivity of its major component?

Predictions

Based on what you learned in Chapter 3 about periodic trends, make predictions about the relative reactivity of copper, iron, magnesium, zinc, and tin. Explain your reasons for these predictions.

What do you know about alloys such as bronze, brass, and steel? Based on what you know, make a prediction about whether steel will be more or less reactive than iron, its main component.

Materials

well plate(s): at least a 6 × 8 matrix
wash bottle with distilled water
5 test tubes
test tube rack
dilute $HCl_{(aq)}$
6 small pieces each of copper, iron, magnesium, zinc, tin, steel, galvanized steel, stainless steel
dropper bottles of dilute solutions of $CuSO_4$, $FeSO_4$, $MgSO_4$, $ZnSO_4$, $SnCl_2$

Metal \ Cation or solution	HCl	H_2O	Cu^{2+}	Fe^{2+}	Mg^{2+}	Sn^{2+}
Cu						
Mg						
Sn						
Zn						
Fe						
steel						
galvanized steel						
stainless steel						

Safety Precautions

Handle the hydrochloric acid solution with care. It is corrosive. Wipe up any spills with copious amounts of water, and inform your teacher.

Procedure

1. Place your well plate(s) on a white sheet of paper. Label them according to the matrix on the previous page.
2. Place a rice-grain-sized piece of each metal in the appropriate well. Record the appearance of each metal.
3. Put enough drops of the appropriate solution to completely cover the piece of metal.
4. Record any changes in appearance due to a chemical reaction. In reactions of metal with acid, look carefully for the formation of bubbles. If you are unsure about any observation, repeat the experiment in a small test tube. This will allow you to better observe the reaction.
5. If you believe that a reaction has occurred, write "r" on the matrix. If you believe that no reaction has occurred, write "nr" on the matrix.
6. Dispose of the solutions in the waste beaker supplied by your teacher. Do not pour anything down the drain.

Analysis

1. For any reactions that occurred, write the corresponding single displacement reaction.
2. (a) What was the most reactive metal that you tested?
   (b) What was the least reactive metal that you tested?
3. Look at Figure 4.12. Lithium reacts violently with water to form aqueous lithium hydroxide and hydrogen gas. Do you expect lithium to react with hydrochloric acid?
   (a) Write the balanced chemical equation for this reaction.
   (b) Is lithium more or less reactive than magnesium?
4. What evidence do you have that hydrogen in hydrochloric acid is different from hydrogen in water?

Conclusion

5. (a) Write the activity series corresponding to your observations. Include hydrogen in the form of water and also as an ion (H^+). Do not include the alloys.
   (b) How did the reactivity of the iron compare with the reactivity of the various types of steel?
6. How do you think an activity series for metal would help you predict whether or not a single displacement reaction will occur? Use examples to help you explain your answer.
7. You have learned that an alloy is a homogeneous mixture (solution) of two or more metals. Steel consists of mostly iron.
   (a) Which type of steel appeared to be the most reactive? Which type was the least reactive? Did you notice any differences?
   (b) What other components make up steel, galvanized steel, and stainless steel?

Application

8. For what applications are the various types of steel used? Why would you not use iron for these applications?

The Metal Activity Series

As you can see in Table 4.2, the more reactive metals are at the top of the activity series. The less reactive metals are at the bottom. *A reactive metal will displace or replace any metal in a compound that is below it in the activity series.* Metals from lithium to sodium will displace hydrogen as a gas from water. Metals from magnesium to lead will displace hydrogen as a gas only from acids. Copper, mercury, silver, and gold will not displace hydrogen from acids.

Table 4.2 Activity Series of Metals

Metal	Displaces hydrogen from acids	Displaces hydrogen from cold water	
lithium			Most Reactive
potassium			
barium			
calcium			
sodium			
magnesium			
aluminum			
zinc			
chromium			
iron			
cadmium			
cobalt			
nickel			
tin			
lead			
hydrogen			
copper			
mercury			
silver			
platinum			
gold			Least Reactive

You can use the activity series to help you predict the products of the reaction of a metal and a metal-containing compound. For example, consider the following incomplete equation.

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow$$

You can see from the activity series that iron is above copper. This means that iron is more reactive than copper. This reaction will proceed.

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$$

The copper metal produced is less reactive than iron metal. What about the following incomplete reaction between silver and calcium chloride?

$$Ag_{(s)} + CaCl_{2(aq)} \rightarrow$$

Silver is below calcium in the activity series, meaning that it is less reactive. There would be no reaction between these two substances. Predict whether the substances in the following Practice Problem will react.

PROBLEM TIP

A single displacement reaction always favours the production of the less reactive metal. In other words, the "free" metal that is formed from the compound must always be less reactive than the metal that displaced it. For example,

$$2AgNO_{3(aq)} + Cu_{(s)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$$

Silver metal is more stable than copper metal. In other words, silver is below copper in the activity series.

CHEM FACT

Most metals that we use in everyday life are actually alloys. An alloy is a solid solution of one metal (or non-metal) in another metal. For example, steel is an alloy of iron. Steel has many uses, from construction to the automobile industry. If the iron were not alloyed with other elements, it would not have the physical and chemical properties required, such as hardness and corrosion resistance.

Practice Problems

22. Using the activity series, write a balanced chemical equation for each single displacement reaction. If you predict that there will be no reaction, write "NR."

(a) $Cu + MgSO_4 \rightarrow$

(b) $Zn + FeCl_2 \rightarrow$

(c) $K + H_2O \rightarrow$

(d) $Al + H_2SO_4 \rightarrow$

(e) $Fe + Al_2(SO_4)_3 \rightarrow$

(f) $Ni + NiCl_2 \rightarrow$

(g) $Zn + H_2SO_4 \rightarrow$

(h) $Mg + SnCl_2 \rightarrow$

Single Displacement Reactions Involving Halogens

Non-metals, typically halogens, can also take part in single displacement reactions. For example, molecular chlorine can replace bromine from KBr, an ionic compound, producing bromine and potassium chloride.

$$Cl_{2(g)} + 2KBr_{(aq)} \rightarrow 2KCl_{(aq)} + Br_{2(\ell)}$$

The **activity series for halogens** directly mirrors the position of halogens in the periodic table. It can be shown simply in the following way. Fluorine is the most reactive, and iodine is the least reactive.

$$F > Cl > Br > I$$

In the same way as you used the activity series for metals, you can use the activity series for halogens to predict whether substances will undergo a single displacement reaction. For example, fluorine is above chlorine in the activity series. So, given the reactants fluorine and sodium chloride, you can predict that the following reaction will occur:

$$F_{2(g)} + 2NaCl_{(aq)} \rightarrow 2NaF_{(aq)} + Cl_{2(aq)}$$

On the other hand, iodine is below bromine in the activity series. So, given the reactants iodine and calcium bromide, you can predict that no reaction will occur.

$$I_{2(aq)} + CaBr_{2(aq)} \rightarrow NR$$

Try the following problems to practise using the metal and halogen activity series to predict whether reactions will occur.

CHECKPOINT

Based on what you know about the electronegativity and electron affinity for the halogens, explain the organization of the halogen activity series.

Practice Problems

23. Using the activity series for halogens, write a balanced chemical equation for each single displacement reaction. If you predict that there will be no reaction, write "NR".

(a) $Br_2 + KCl \rightarrow$

(b) $Cl_2 + NaI \rightarrow$

24. Using the appropriate activity series, write a balanced chemical equation for each single displacement reaction. If you predict that there will be no reaction, write "NR".

(a) $Pb + HCl \rightarrow$

(b) $KI + Br_2 \rightarrow$

(c) $KF + Cl_2 \rightarrow$

(d) $Ca + H_2O \rightarrow$

(e) $MgSO_4 + Zn \rightarrow$

(f) $Ni + H_2SO_4 \rightarrow$

Figure 4.14 When a few drops of silver nitrate, $AgNO_3$, are added to a sample of salt water, $NaCl_{(aq)}$, a white precipitate of silver chloride, AgCl, is formed.

Double Displacement Reactions

A **double displacement reaction** involves the exchange of cations between two ionic compounds, usually in aqueous (water) solution. A double displacement reaction is also known as a double replacement reaction. A general equation for a double displacement reaction is:

$$AB + CD \rightarrow CB + AD$$

In this equation, A and C are cations and B and D are anions.

Consider the following situation. You have two unlabelled beakers. One contains distilled water, and the other contains salt water. The two samples look virtually identical. How can you quickly determine which is the salt water without tasting them? (You should never taste anything in a chemistry laboratory.)

A common test for the presence of chloride ions in water is the addition of a few drops of silver nitrate solution. The formation of a white solid indicates the presence of chloride ions, as you can see in Figure 4.14.

A double displacement reaction has occurred. That is, the cations in the reactants have essentially changed places. This switch is modelled in Figure 4.15.

$$NaCl_{(aq)} + AgNO_{3(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

Figure 4.15 Sodium chloride and silver nitrate form ions in solution. When silver ions and chloride ions come into contact, they form a solid.

Since silver chloride is virtually insoluble in water, it forms a solid compound, or precipitate.

Double displacement reactions tend to occur in aqueous solution. Not all ionic compounds, however, will react with one another in this way. You can tell that a double displacement reaction has taken place in the following cases:

- a solid (precipitate) forms
- a gas is produced
- some double displacement reactions also form a molecular compound, such as water. It is hard to tell when water is formed, because often the reaction takes place in water.

Double Displacement Reactions that Form a Precipitate

A **precipitate** is a solid that separates from a solution as the result of a chemical reaction. You will learn more about precipitates in Chapter 9. *Many double displacement reactions involve the formation of a precipitate.*

Examine the double displacement reaction that occurs when aqueous solutions of barium chloride and potassium sulfate are mixed. A white precipitate is immediately formed. The equation for the reaction is

$$BaCl_{2(aq)} + K_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2KCl_{(aq)}$$

You should think about two questions when analyzing a double displacement reaction.

1. How do we determine the products?
2. Which of the products—if any—will precipitate out of solution?

Barium chloride solution contains Ba^{2+} ions and Cl^- ions. Potassium sulfate solution contains K^+ and SO_4^{2-} ions. When they are mixed, the Ba^{2+} ions come in contact with SO_4^{2-} ions. Because barium sulfate is insoluble, the product comes out of solution as a solid. The K^+ ions and Cl^- ions also come into contact with each other, but potassium chloride is soluble, so these ions stay in solution.

How do you know that the precipitate is $BaSO_4$ and *not* KCl? More generally, how can you predict whether a precipitate will be formed in a double displacement reaction? In this chapter, you will be given information on solubility as you need it. You will learn more about how to predict whether a compound is soluble or not in Chapter 9. Barium sulfate, $BaSO_4$ is not soluble in water, while potassium chloride, KCl, is. Therefore, a reaction will take place and barium sulfate will be the precipitate.

In summary, to determine the products and their physical states in a double displacement reaction, you must first "deconstruct" the reactants. Then switch the cations, and "reconstruct" the products using proper chemical formulas. You should then balance the chemical equation. You will be given information to determine which of the products, if any, will form a precipitate. Finally, you can write the physical state—(s) or (aq)—of each product and balance the equation.

Given the following reactants, how would you predict the products of the reaction and their state? Note that many hydroxide compounds, including magnesium hydroxide, are insoluble. Potassium cations form soluble substances with all anions.

$$MgCl_{2(aq)} + KOH_{(aq)} \rightarrow$$

Examine figure 4.16 to see how to separate the compounds into ions, Mg^{2+} and Cl^-; K^+ and OH^-. Then switch the anions and write chemical formulas for the new compounds. Check to ensure your equation is balanced.

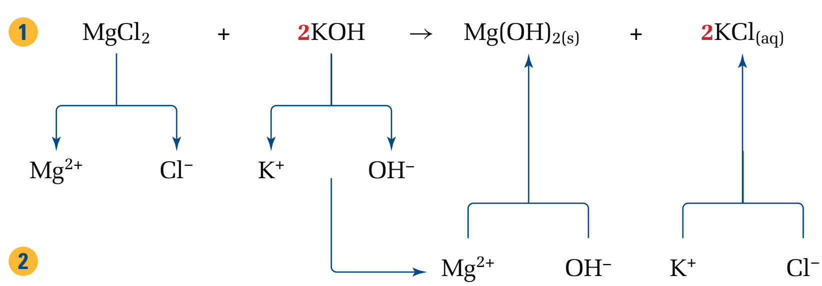


Figure 4.16 **Predicting a double displacement reaction.**

What happens if both products are soluble ionic compounds? Both ionic compounds will be ions dissolved in the water. If neither product precipitates out, no reaction occurs. Try the following problem to practise writing the products of double displacement reactions and predicting their states.

PROBLEM TIP

When aqueous solutions of barium chloride and sodium chromate are mixed, a precipitate is formed. The balanced equation for this double replacement reaction is

$$BaCl_{2(aq)} + Na_2CrO_{4(aq)} \rightarrow BaCrO_{4(?)} + 2NaCl_{(?)}$$

What is the precipitate? From experience, you know that NaCl *is* water soluble. So, *by process of elimination*, the precipitate must be barium chromate, $BaCrO_{4(s)}$.

Practice Problems

25. Write a balanced chemical equation for each double displacement reaction. Write "NR" if you predict that no reaction will occur. Note that K^+, Na^+, and Li^+ ions form soluble compounds with all anions. All nitrate compounds are soluble. Sulfate compounds with Ca^{2+}, Sr^{2+}, Ba^{2+}, Ra^{2+}, and Pb^{2+} are insoluble, but most other sulfate compounds are soluble. Lead(II) iodide is insoluble.

(a) $Pb(NO_3)_{2(aq)} + KI_{(aq)} \rightarrow$

(b) $FeCl_{3(aq)} + Na_2SO_{4(aq)} \rightarrow$

(c) $NaNO_{3(aq)} + MgSO_{4(aq)} \rightarrow$

(d) $Ba(NO_3)_{2(aq)} + MgSO_{4(aq)} \rightarrow$

Double Displacement Reactions That Produce a Gas

In certain cases, you know that a double displacement reaction has occurred because a gas is produced. The gas is formed when one of the products of the double displacement reaction decomposes to give water and a gas.

Consider the reaction of aqueous sodium carbonate (washing soda) and hydrochloric acid, shown in Figure 4.17. Hydrochloric acid is sold at the hardware store under the common name "muriatic acid." If you carry out this reaction, you immediately see the formation of carbon dioxide gas. The first reaction that takes place is a double displacement reaction. Determine the products in the following way. Separate the reactions into ions, and switch the anions. Write chemical formulas for the products and balance the equation.

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2CO_{3(aq)}$$

But this isn't all that is happening! The carbonic acid, H_2CO_3, is unstable and subsequently decomposes to carbon dioxide and water.

$$H_2CO_{3(aq)} \rightarrow H_2O_{(\ell)} + CO_{2(g)}$$

Overall, we can write this two-step reaction as follows:

$$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_2O_{(\ell)} + CO_{2(g)}$$

Figure 4.17 The reaction of hydrochloric acid and sodium carbonate, Na_2CO_3 (washing soda), is a double displacement reaction. This reaction initially forms sodium chloride and carbonic acid, H_2CO_3. The carbonic acid spontaneously decomposes to water and carbon dioxide gas.

Another double displacement reaction results in the formation of gaseous ammonia, NH_3. Ammonia, a pungent-smelling gas, is an important industrial chemical. It is used as a fertilizer and, when dissolved in water, as a household cleaner. Ammonium hydroxide is formed in the reaction below

$$NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow NH_4OH_{(aq)} + NaCl_{(aq)}$$

The ammonium hydroxide, NH_4OH, immediately decomposes to give ammonia and water, according to the equation

$$NH_4OH_{(aq)} \rightarrow NH_{3(g)} + H_2O_{(\ell)}$$

Combining these equations gives

$$NH_4Cl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + NH_{3(g)} + H_2O_{(\ell)}$$

This example illustrates the formation of a gas by an initial double displacement reaction, followed by the decomposition of one of the products to a gas and water.

Practice Problems

26. (a) When sodium sulfite, $Na_2SO_{3(aq)}$, is mixed with hydrogen chloride, $HCl_{(aq)}$ (hydrochloric acid), the odour of sulfur dioxide gas, $SO_{2(g)}$, is detected. Write the balanced chemical equation for this reaction.

(b) Hydrogen sulfide, H_2S, is a poisonous gas that has the odour of rotten eggs. When aqueous calcium sulfide, CaS, is reacted with sulfuric acid, a rotten egg smell is detected. Write the balanced chemical equation for this reaction.

The Formation of Water in a Neutralization Reaction

Neutralization reactions are a special type of double displacement reaction that produces water. Neutralization involves the reaction of an acid with a base to form water and an ionic compound. You will learn more about neutralization reactions in Chapter 10. For example, the neutralization of hydrogen nitrate (nitric acid) with sodium hydroxide (a base) is a double displacement reaction.

$$HNO_{3(aq)} + NaOH_{(aq)} \rightarrow NaNO_{3(aq)} + H_2O_{(\ell)}$$

Often neutralization reactions produce no precipitate or gas. In Chapter 10, you will learn how chemists recognize when neutralization reactions take place.

Practice Problems

27. Write the balanced chemical equation for each neutralization reaction.

(a) $HCl_{(aq)} + LiOH_{(aq)} \rightarrow$

(b) $HClO_{4(aq)} + Ca(OH)_{2(aq)} \rightarrow$

(c) $H_2SO_{4(aq)} + NaOH_{(aq)} \rightarrow$

28. Write the balanced chemical equation for each double replacement reaction. Be sure to indicate the physical state of all products.

(a) $BaCl_{2(aq)} + Na_2CrO_{4(aq)} \rightarrow$ (A precipitate is produced.)

(b) $H_3PO_{4(aq)} + NaOH_{(aq)} \rightarrow$ (Water is produced.)

(c) $K_2CO_{3(aq)} + HNO_{3(aq)} \rightarrow$ (A gas is produced.)

You have learned about a variety of double displacement reactions. In Investigation 4-B, you will make predictions about whether double displacement reactions will take place. Then you will make observations to test your predictions.

In Investigation 4-C, you will perform reactions that involve copper compounds to reinforce what you have learned about the different types of reactions. You will identify the series of reactions that begin by reacting copper and finish by producing copper.

CHEM FACT

To non-chemists, the term "salt" refers solely to sodium chloride. To chemists, "salt" is a generic term that describes an ionic compound with an anion that is not OH^- or O^{2-} and with a cation that is not H^+. Sodium chloride, NaCl, and potassium fluoride, KF, are two examples.

mind STRETCH

Consider the reaction

$$KNO_{3(aq)} + NaI_{(aq)} \rightarrow KI_{(aq)} + NaNO_{3(aq)}$$

The products are both water soluble. Has a chemical reaction occurred? If the water is allowed to evaporate, what compounds will remain?

Investigation 4-B

MICROSCALE

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

Observing Double Displacement Reactions

A double displacement reaction involves the exchange of cations between two ionic compounds, usually in aqueous solution. It can be represented with the general equation

$$AB + CD \rightarrow AD + CB$$

Most often, double displacement reactions result in the formation of a precipitate. However, some double displacement reactions result in the formation of an unstable compound which then decomposes to water and a gas.

The reaction of an acid and a base—a neutralization reaction—is also a type of double displacement reaction. It results in the formation of a salt and water.

Question

How can you tell if a double displacement reaction has occurred? How can you predict the products of a double displacement reaction?

Prediction

For each reaction in Tables A and B, write a balanced chemical equation. Use the following guidelines to predict precipitate formation in Table A.

- Hydrogen, ammonium, and Group I ions form soluble compounds with all negative ions.
- Chloride ions form compounds that are not very soluble when they bond to silver, lead(II), mercury(I), and copper(I) positive ions.
- All compounds that are formed from a nitrate and a positive ion are soluble.
- With the exception of the ions in the first bulleted point, as well as strontium, barium, radium, and thallium positive ions, hydroxide ions form compounds that do not dissolve.
- Iodide ions that are combined with silver, lead(II), mercury(I), and copper(I) are not very soluble.
- Chromate compounds are insoluble, except when they contain ions from the first bulleted point.

Materials

well plate
sheet of white paper
several test tubes
test tube rack
test tube holder
2 beakers (50 mL)
tongs
scoopula
laboratory burner
flint igniter
red litmus paper
wooden splint
wash bottle with distilled water
HCl solution
the following aqueous solutions in dropper bottles: $BaCl_2$, $CaCl_2$, $MgCl_2$, Na_2SO_4, $NaOH$, $AgNO_3$, $Pb(NO_3)_2$, KI, $FeCl_3$, solid Na_2CO_3 and NH_4Cl

Safety Precautions

- Hydrochloric acid is corrosive. Use care when handling it.
- Before lighting the laboratory burner, check that there are no flammable liquids nearby.
- If you accidentally spill a solution on your skin, immediately wash the area with copious amounts of water.
- Wash your hands thoroughly after the experiment.

Procedure

1. Copy Tables A and B into your notebook. Do not write in this textbook.
2. Place the well plate on top of the sheet of white paper.
3. Carry out each of the reactions in Table A by adding several drops of each solution to a well. Record your observations in Table A.

If you are unsure about the formation of a precipitate, repeat the reaction in a small test tube for improved visibility.

4. Place a scoopula tipful of Na_2CO_3 in a 50 mL beaker. Add 5 mL of HCl. Use a burning wooden splint to test the gas produced. Record your observations in Table B.
5. Place a scoopula tipful of NH_4Cl in a test tube. Add 2 mL NaOH. To detect any odour, gently waft your hand over the mouth of the test tube towards your nose. Warm the tube gently (do not boil) over a flame. Record your observations in Table B.
6. Dispose of all chemicals in the waste beaker supplied by your teacher. Do not pour anything down the drain.

Table A Double Displacement Reactions That May Form a Precipitate

Skeleton equation	Observations
$MgCl_2$ + NaOH	
$FeCl_3$ + NaOH	
$BaCl_2$ + Na_2SO_4	
$CaCl_2$ + $AgNO_3$	
$Pb(NO_3)_2$ + KI	

Table B Double Displacement Reactions That May Form a Gas

Reaction	Observations
Na_2CO_3 + HCl	
NH_4Cl + NaOH	

Analysis

1. Write the balanced chemical equation for each reaction in Table A.
2. For each reaction in Table B, write the appropriate balanced chemical equation for the double displacement reaction. Then write a balanced chemical equation for the decomposition reaction that leads to the formation of a gas and water.

Conclusion

3. How did you know when a double displacement reaction had occurred? How did your results compare with your predictions?

Application

4. Suppose that you did not have any information about the solubility of various compounds, but you did have access to a large variety of ionic compounds. What would you need to do before predicting the products of the displacement reactions above? Outline a brief procedure.

Investigation 4-C

MICROSCALE

SKILL FOCUS

Predicting

Performing and Recording

Analyzing and Interpreting

From Copper to Copper

This experiment allows you to carry out the sequential conversion of copper metal to copper(II) nitrate to copper(II) hydroxide to copper(II) oxide to copper(II) sulfate and back to copper metal. This conversion is carried out using synthesis, decomposition, single displacement, and double displacement reactions.

Question

What type of chemical reaction is involved in each step of this investigation?

Prediction

Examine the five reactions outlined in the procedure. Predict what reactions will occur, and write equations to describe them.

Materials

hot plate
glass rod
wash bottle with distilled water
50 mL Erlenmeyer flask
beaker tongs
250 mL beaker containing ice and liquid water
red litmus paper
$Cu(NO_3)_2$ solution
*6 mol/L NaOH solution in dropping bottle
*3 mol/L H_2SO_4 solution in dropping bottle
0.8 g of flaked zinc

Safety Precautions

- Constantly stir, or swirl, any precipitate-containing solution that is being heated to avoid a sudden boiling over, or *bumping*.
- Unplug any hot plate not in use.
- Do not allow electrical cords to hang over the edge of the bench.
- NaOH and H_2SO_4 solutions are corrosive. Handle them with care. If you accidentally spill a solution on your skin, wash the area immediately with copious amounts of cool water. Inform your teacher.

Procedure

Reaction A: Reaction of Copper Metal to form Copper(II) Nitrate

$$Cu_{(s)} + 4HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2NO_{2(g)} + 2H_2O_{(\ell)}$$

CAUTION Your teacher will carry out steps 1 to 4 in the fumehood before class. Concentrated nitric acid is required, and the $NO_{2(g)}$ produced is poisonous. Furthermore, this reaction is quite slow. Your teacher will perform a brief demonstration of this reaction so that you may record observations.

1. Place 0.100 g (100 mg) of Cu in a 50 mL Erlenmeyer flask.
2. Add 2 mL of 6 mol/L $HNO_{3(aq)}$ to the flask in the fumehood.
3. Warm the flask on a hot plate in the fumehood. The heating will continue until all the Cu dissolves and the evolution of brown $NO_{2(g)}$ ceases.
4. Cool the flask in a cool water bath.
5. Add about 2 mL of distilled water to the flask containing the $Cu(NO_3)_2$ solution.

*The unit mol/L refers to concentration. You will learn more about this in Unit 3. For now, you should know that 6 mol/L NaOH and 3 mol/L H_2SO_4 are highly corrosive solutions, and you should treat them with respect.

Reaction B: Preparation of Copper(II) Hydroxide

6. At room temperature, while stirring with a glass rod, add 6 mol/L NaOH, drop by drop, until the solution is basic to red litmus paper. (Red litmus paper turns blue in basic solution.) Do not put the red litmus paper in the solution. Dip the glass rod into the solution and touch it to the red litmus paper. Record your observations.

Reaction C: Preparation of Copper(II) Oxide

7. While constantly stirring the solution with a glass rod, heat the mixture from step 6 on a hot plate until a black precipitate is formed. If necessary, use the wash bottle to wash loose any unreacted light blue precipitate that is adhering to the side of the flask.

8. When all of the light blue precipitate has reacted to form the black precipitate, cool the flask in an ice bath or a cool water bath for several minutes.

Reaction D: Preparation of Copper(II) Sulfate Solution

9. Carefully add about 6 mL of 3 mol/L sulfuric acid to the flask. Stir it until all the black precipitate has dissolved. Record your observations. **CAUTION** The sulfuric acid is highly corrosive. If any comes in contact with your skin, rinse the area thoroughly and immediately with water.

Reaction E: Regeneration of Copper Metal

10. In the fumehood or in a well ventilated area, carefully add about 0.8 g of powdered zinc to the solution of copper(II) sulfate. Stir or swirl the solution until the blue colour disappears. Record your observations. **CAUTION** You should wear a mask for this step to avoid breathing in the powdered zinc.

11. When the reaction is complete, add 5 mL of 3 mol/L sulfuric acid while stirring or swirling the solution. This removes any unreacted zinc but does not affect the copper metal. Carefully decant the liquid into a clean waste container. Wash the copper metal carefully several times with water. Return the copper metal to your teacher. Wash your hands. **CAUTION** This sulfuric acid is highly corrosive. If any comes in contact with your skin, rinse the area thoroughly and immediately with water.

Analysis

1. What type of reaction is occurring in reactions A through E?

2. Write a balanced chemical equation for reactions B through E.

3. Explain why H_2SO_4 reacts with Zn but not with Cu. (See step 11 in the procedure.)

4. Could another metal have been used in place of Zn in step 10? Explain.

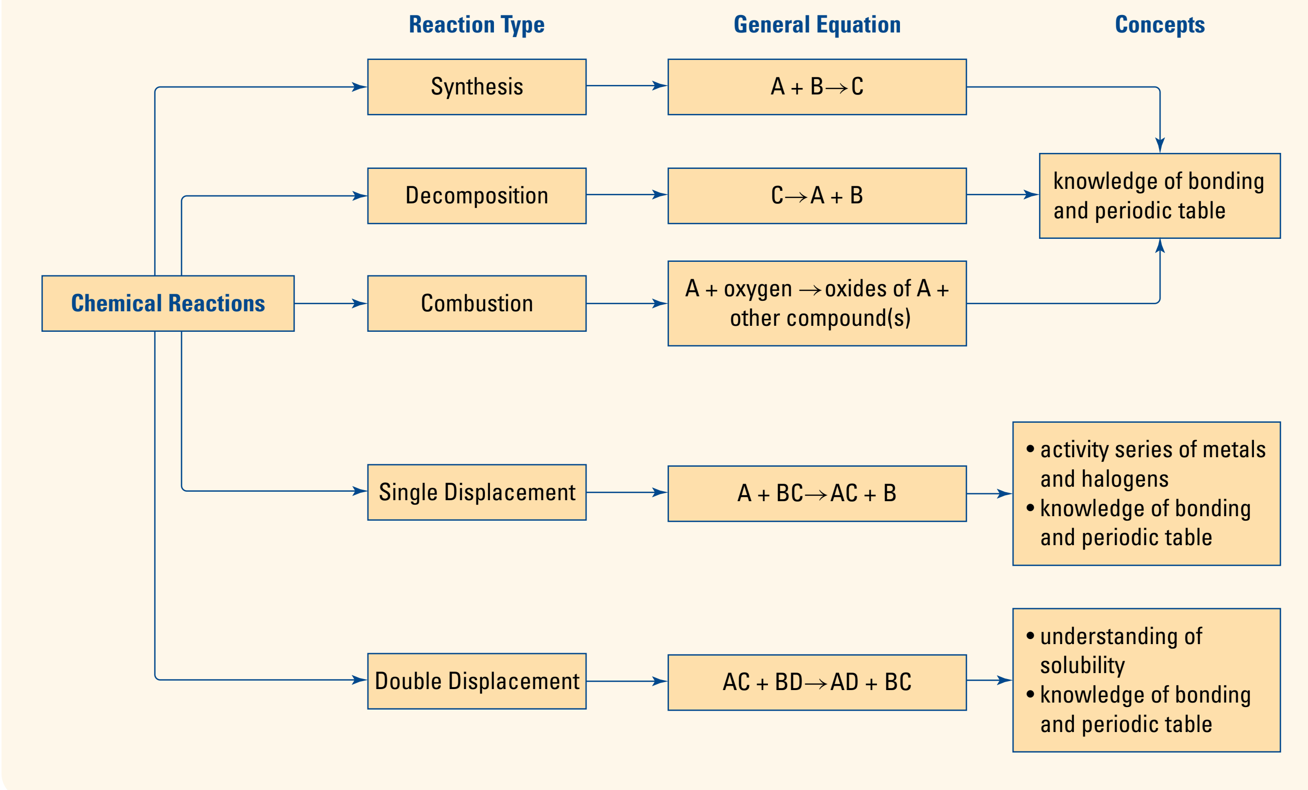
5. Why was powdered Zn used in step 10, rather than a single piece of Zn?

6. You used 0.100 g of Cu metal in reaction A. How much copper should theoretically be recovered at the end of reaction E?

Conclusion

7. Create a flowchart that shows each step of the reaction series. Include the balanced chemical equations.

Concept Organizer   Predicting Chemical Reactions



Section Wrap-up

In sections 4.2 and 4.3, you have examined five different types of chemical reactions: synthesis, decomposition, combustion, single displacement, and double displacement. Equipped with this knowledge, you can examine a set of reactants and predict what type of reaction will occur and what products will be formed. The Concept Organizer above provides a summary of the types of chemical reactions.

Section Review

Unit Project Prep

Before you design your Chemistry Newsletter at the end of Unit 1, take a look at some of the labels of chemical products in your home. Are there any warnings about mixing different products together? Use what you know about chemical reactions to explain why mixing some chemical products might be dangerous.

1 K/U Write the product(s) of each single displacement reaction. If you predict that there will be no reaction, write "NR." Balance each chemical equation.

(a) $Li + H_2O \rightarrow$

(b) $Sn + FeCl_2 \rightarrow$

(c) $F_2 + KI \rightarrow$

(d) $Al + MgSO_4 \rightarrow$

(e) $Zn + CuSO_4 \rightarrow$

(f) $K + H_2O \rightarrow$

2 K/U Complete each double displacement reaction. Be sure to indicate the physical state of each product. Then balance the equation. **Hint:** Compounds containing alkali metal ions are soluble. Calcium chloride is soluble. Iron(III) hydroxide is insoluble.

(a) $NaOH + Fe(NO_3)_3 \rightarrow$

(b) $Ca(OH)_2 + HCl \rightarrow$

(c) $K_2CrO_4 + NaCl \rightarrow$

(d) $K_2CO_3 + H_2SO_4 \rightarrow$

3 **(a)** C Explain why the following chemical equation represents a double displacement reaction followed by a decomposition reaction.

$(NH_4)_2SO_{4(aq)} + KOH_{(aq)} \rightarrow NH_{3(g)} + H_2O_{(\ell)} + K_2SO_{4(aq)}$

(b) Balance the chemical equation in part (a).

4 K/U Identify each reaction as synthesis, combustion (complete or incomplete), decomposition, single displacement, or double displacement. Balance the equations, if necessary.

(a) $C_8H_{18(\ell)} + O_{2(g)} \rightarrow CO_{(g)} + H_2O_{(g)}$

(b) $Pb_{(s)} + H_2SO_{4(aq)} \rightarrow PbSO_{4(s)} + H_{2(g)}$

(c) $Al_2(SO_4)_{3(aq)} + K_2CrO_{4(aq)} \rightarrow K_2SO_{4(aq)} + Al_2(CrO_4)_{3(s)}$

(d) $C_3H_7OH_{(\ell)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

(e) $(NH_4)_2Cr_2O_{7(s)} \rightarrow N_{2(g)} + H_2O_{(g)} + Cr_2O_{3(s)}$

(f) $Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$

(g) $N_2O_{4(g)} \rightarrow 2NO_{2(g)}$

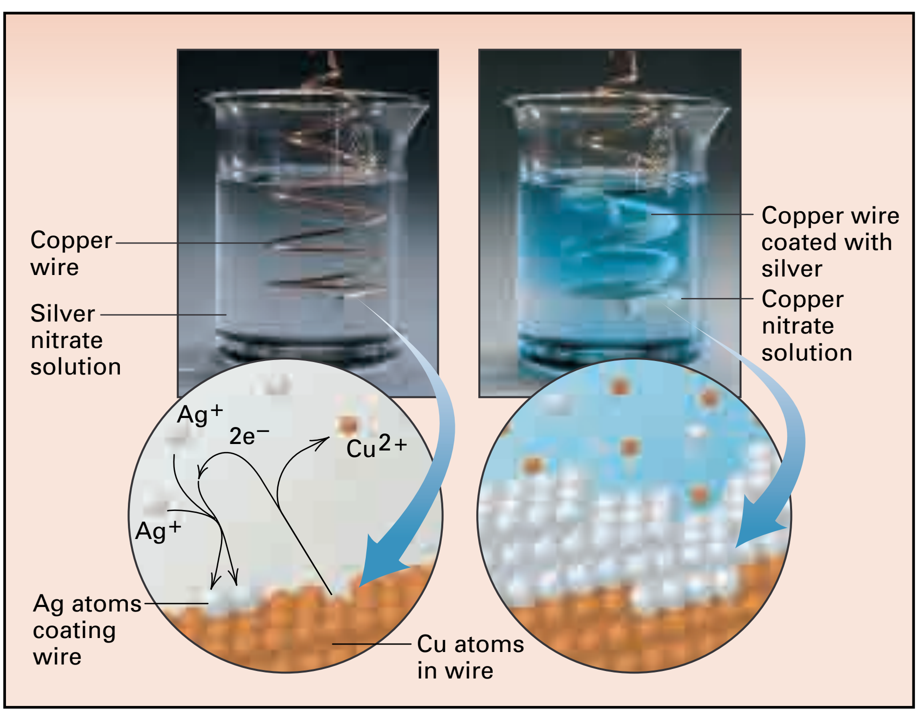
5 MC Biosphere II was created in 1991 to test the idea that scientists could build a sealed, self-sustaining ecosystem. The carbon dioxide levels in Biosphere II were lower than scientists had predicted. Scientists discovered that the carbon dioxide was reacting with calcium hydroxide, a basic compound in the concrete.

(a) Write two balanced equations to show the reactions. Then classify the reactions. Hint: In the first reaction, carbon dioxide reacts with water in the concrete to form hydrogen carbonate. Hydrogen carbonate, an acid, reacts with calcium hydroxide, a base.

(b) Why do you think scientists failed to predict that this would happen?

(c) Suggest ways that scientists could have combatted the problem.

6 K/U What reaction is shown in the figure below? Write a balanced chemical equation to describe the reaction, then classify it.



4.4 Simple Nuclear Reactions

Section Preview/ Specific Expectations

In this section, you will

- **balance** simple nuclear equations
- **communicate** your understanding of the following terms: *nuclear reactions, nuclear equation, alpha* (α) *particle emission, beta* (β) *decay, beta particle, gamma* (γ) *radiation, nuclear fission, nuclear fusion*

You have seen some chemical reactions that involve the formation and decomposition of different compounds. These reactions involve the rearrangement of atoms due to the breaking and formation of chemical bonds. Chemical bonds involve the interactions between the electrons of various atoms. There is another class of reactions, however, that are not *chemical*. These reactions involve changes that occur within the nucleus of atoms. These reactions are called **nuclear reactions**.

We know that nuclear weapons are capable of mass destruction, yet radiation therapy, shown in Figure 4.18, is a proven cancer fighter. Smoke detectors, required by law in all homes, rely on the radioactive decay of americium-241. The human body itself is radioactive, due to the presence of radioactive isotopes including carbon-14, phosphorus-32, and potassium-40. Most people view radioactivity and nuclear reactions with a mixture of fascination, awe, and fear. Since radioactivity is all around us, it is important to understand what it is, how it arises, and how we can deal with it safely.

Figure 4.18 **This patient is about to undergo radiation therapy.**

Types of Radioactive Decay and Balancing Nuclear Equations

There are three main types of radioactive decay: alpha particle emission, beta particle emission, and the emission of gamma radiation. When an unstable isotope undergoes radioactive decay, it produces one or more different isotopes. We represent radioactive decay using a **nuclear equation.** Two rules for balancing nuclear equations are given below.

Rules for Balancing Nuclear Equations

1. The sum of the mass numbers (written as superscripts) on each side of the equation must balance.
2. The sum of the atomic numbers (written as subscripts) on each side of the equation must balance.

Alpha Decay

Alpha (α) particle emission, or alpha decay, involves the loss of one alpha particle. An α particle is a helium nucleus, ${}^{4}_{2}He$, composed of two protons and two neutrons. Since it has no electrons, an alpha particle carries a charge of +2.

One example of alpha particle emission is the decay of radium. This decay is shown in the following equation:

$$ {}^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He $$

Notice that the sum of the mass numbers on the right (222 + 4) equals the mass on the left (226). As well, the atomic numbers balance (88 = 86 + 2). Thus, this nuclear equation is balanced.

Media LINK

The media are full of references to radioactivity. For example, the comic book hero Spiderman gained the abilities of a spider after being bitten by a radioactive spider. Bart Simpson reads *Radioactive Man* comics. Can you think of any other references to radioactivity in popular culture? What kind of reputation do radioactivity and nuclear reactions have? Do you think this reputation is deserved?

In another example of alpha particle emission, Berkelium-248 is formed by the decay of a certain radioisotope according to the balanced nuclear equation:

$$^{a}_{b}X \rightarrow {}^{248}_{97}\mathrm{Bk} + {}^{4}_{2}\mathrm{He}$$

Given this information, what is $^{a}_{b}X$? You can use your knowledge of how to balance a nuclear equation to determine the identity of a radioisotope undergoing alpha particle decay.

The total of the atomic masses on the right side is $(248 + 4) = 252$. The total of the atomic numbers on the right is $(97 + 2) = 99$. Therefore, $a = 252$ and $b = 99$. From the periodic table, you see that element number 99 is Es, einsteinium. The missing atom is $^{252}_{99}\mathrm{Es}$, so the balanced nuclear equation is:

$$^{252}_{99}\mathrm{Es} \rightarrow {}^{248}_{97}\mathrm{Bk} + {}^{4}_{2}\mathrm{He}$$

Try the following problems to practise balancing alpha emission nuclear reactions.

Practice Problems

29. Uranium was the first element shown to be radioactive. Complete the following reaction representing the alpha decay of uranium-238.

$$^{238}_{92}\mathrm{U} \rightarrow \underline{\hspace{3em}} + {}^{4}_{2}\mathrm{He}$$

30. Radon-222, $^{222}_{86}\mathrm{Rn}$, is known to decay by alpha particle emission. Write a balanced nuclear equation and name the element produced in this decay process.

31. Plutonium-242 decays by emitting an alpha particle. Write the balanced nuclear equation for this reaction.

32. Neodymium-144, $^{144}_{60}\mathrm{Nd}$, decays by alpha particle emission. Write the balanced nuclear equation for this nuclear decay.

Beta Decay

Beta (β) decay occurs when an isotope emits an electron, called a **beta particle**. Because of its tiny mass and –1 charge, a beta particle, is represented as $^{0}_{-1}\mathrm{e}$. For example, hydrogen-3, or tritium, emits a beta particle to form helium-3 as illustrated by the equation:

$$^{3}_{1}\mathrm{H} \rightarrow {}^{3}_{2}\mathrm{He} + {}^{0}_{-1}\mathrm{e}$$

Notice that the total of the atomic masses and the total of the atomic numbers on each side of the nuclear equation balance. What is happening, however to the hydrogen-3 nucleus as this change occurs? In effect, the emission of a beta particle is accompanied by the conversion, inside the nucleus, of a neutron into a proton:

$$\underset{\text{neutron}}{^{1}_{0}\mathrm{n}} \rightarrow \underset{\text{proton}}{^{1}_{1}\mathrm{H}} + \underset{\text{electron }(\beta\text{ particle})}{^{0}_{-1}\mathrm{e}}$$

History LINK

Marie Curie discovered the element polonium, Po, in 1898. She named polonium after Poland, her homeland. Curie won two Nobel Prizes, one in Physics (1903) for sharing in the discovery of radioactivity, and one in Chemistry (1911) for the discovery of radium, which has been used to treat cancer. Radium-226 undergoes alpha decay to yield radon-222.

CHEM FACT

One of the most harmful potential sources of radiation in the home is radon gas. Radon-222 is a product of the decay of uranium-containing rocks beneath Earth's surface. Since radon is denser than air, it can build up to dangerous levels in basements when it seeps through cracks in walls and floors. Simple radon detectors can be purchased at hardware stores.

> **CHEM FACT**
>
> The nucleus of the most common isotope of hydrogen consists of one proton. Therefore, a proton can be represented by H^+ or ${}^{1}_{1}H$.

Carbon-14 is a radioactive isotope of carbon. Its nucleus emits a beta, particle to form a nitrogen-14 nucleus, according to the balanced nuclear equation shown below in Figure 4.19.

$${}^{14}_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$$

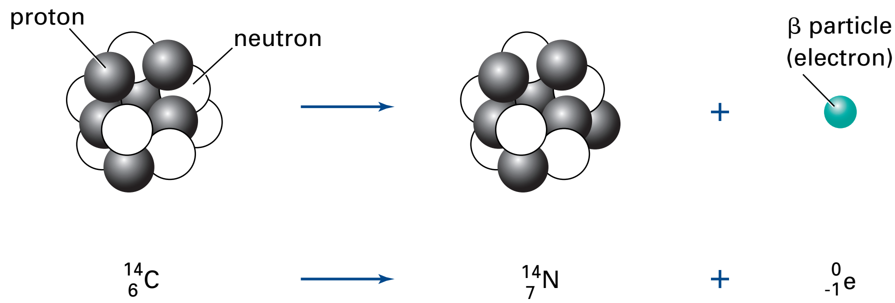


Figure 4.19 **Carbon-14 decays by emitting a beta particle and converting to nitrogen-14. Notice that a neutron in the nucleus of carbon is converted to a protron as the β particle is emitted.**

> **CHEM FACT**
>
> The terms radiation and radioactivity are often confused. Radiation refers to electromagnetic radiation—everything from gamma rays, to X-rays, to visible light, to microwaves, to radio and television signals. Radioactivity, on the other hand, involves the emission of particles or energy from an unstable nucleus.

Radioactive waste from certain nuclear power plants and from weapons testing can lead to health problems. For example, ions of the radioactive isotope strontium-90, an alkali metal, exhibit chemical behaviour similar to calcium ions. This leads to incorporation of the ions in bone tissue, sending ionizing radiation into bone marrow, and possibly causing leukemia. Given the following equation for the decay of strontium-90, how would you complete it?

$${}^{90}_{38}Sr \rightarrow \boxed{} + {}^{0}_{-1}e$$

Since both atomic numbers and mass numbers must balance, you can find the other product.

The mass number of the unknown element is equal to 90 + 0 = 90. The atomic number of the unknown element is equal to 38 – (–1) = 39. From the periodic table, you can see that element 39 is yttrium, Y. The balanced nuclear equation is therefore

$${}^{90}_{38}Sr \rightarrow {}^{90}_{39}Y + {}^{0}_{-1}e$$

You can check your answer by ensuring that the total mass number and the total atomic number on each side of the equation are the same.

Mass numbers balance: 90 = 90 + 0

Atomic numbers balance: 38 = 39 + (–1)

Try the following problems to practise balancing beta emission equations.

Practice Problems

33. Write the balanced nuclear equation for the radioactive decay of potassium-40 by emission of a β particle.

34. What radioisotope decays by β particle emission to form ${}^{47}_{21}Sc$?

35. Complete the following nuclear equation:

$${}^{73}_{31}Ga \rightarrow {}^{0}_{-1}e + \boxed{}$$

Canadians in Chemistry

Harriet Brooks

It was the Nobel that just missed being Canadian. In 1907, Ernest Rutherford left Montréal's McGill University for a position in England. The following year, he received the Nobel Prize for Chemistry for his investigations of the chemistry of radioactive substances. Most of the work, however, had been done in Montréal. Moreover, one young Canadian woman had played an important role in putting it on the right track.

Harriet Brooks is nearly forgotten today, even though she helped to show that elements could be transformed. For over a century, chemists had rejected the dream of the ancient alchemists who thought that they might turn lead into gold with the help of the philosopher's stone. They believed that elements were forever fixed and unchangeable.

Then Harriet Brooks arrived on the scene. When she joined Rutherford's team, she was asked to measure the atomic mass of the isotopes that make up the mysterious vapour given off by radium. She determined that its atomic mass was between 40 and 100, whereas radium was known to have an atomic mass of over 140. Surely this was not just a gaseous form of radium. Somehow radium was turning into another element!

It turned out that Brooks' result was a mistake. Radon—as the mystery gas is now known—has almost the same atomic mass as radium. Brooks' result was a fruitful mistake, however. Her experiment led to a basic understanding of radioactivity and isotopes.

Why did Rutherford win a Nobel Prize for Chemistry? Both he and Brooks worked as physicists. By proving that elements transformed, Rutherford, Brooks, and their co-workers revolutionized traditional chemistry.

Gamma Radiation

Gamma (γ) radiation is high energy electromagnetic radiation. It often accompanies either alpha or beta particle emission. Since gamma radiation has neither mass nor charge, it is represented as ${}^{0}_{0}\gamma$, or simply γ. For example, cesium-137 is a radioactive isotope that is found in nuclear fall-out. It decays with the emission of a beta particle and gamma radiation, according to the equation

$$ {}^{137}_{55}\mathrm{Cs} \rightarrow {}^{137}_{56}\mathrm{Ba} + {}^{0}_{-1}\mathrm{e} + {}^{0}_{0}\gamma $$

How is gamma radiation produced in a radioactive decay? When a radioactive nucleus emits an alpha or beta particle, the nucleus is often left in an unstable, high-energy state. The "relaxation" of the nucleus to a more stable state is accompanied by the emission of gamma radiation.

Nuclear Fission and Fusion

All cases of radioactive decay involve the atom's nucleus. Since these processes do not involve the atom's electrons, they occur regardless of the chemical environment of the nucleus. For example, radioactive hydrogen-3, or tritium, will decay by β particle emission whether it is contained in a water molecule or hydrogen gas, or in a complex protein.

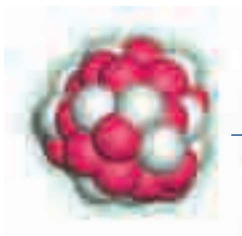
Mode	Emission	Decay					Change in ...		
							Mass numbers	Atomic numbers	Number of neutrons
α Decay	α ($^{4}_{2}He$)	 reactant	→	 product	+	 α expelled	−4	−2	−2
β Decay	$^{0}_{-1}\beta$	$^{1}_{0}n$  in nucleus	→	$^{1}_{1}p$  in nucleus	+	$^{0}_{-1}\beta$  β expelled	0	+1	−1
γ Emission	$^{0}_{0}\gamma$	 excited nucleus	→	 stable nucleus	+	$^{0}_{0}\gamma$ γ photon radiated	0	0	0

Figure 4.20 **A summary of alpha decay, beta decay, and gamma emission**

Many chemical reactions, once begun, can be stopped. For example, a combustion reaction, such as a fire, can be extinguished before it burns itself out. Nuclear decay processes, on the other hand, cannot be stopped.

The principles of balancing nuclear equations apply to all nuclear reactions. **Nuclear fission** occurs when a highly unstable isotope splits into smaller particles. Nuclear fission usually has to be induced in a particle accelerator. Here, an atom can absorb a stream of high-energy particles such as neutrons, $^{1}_{0}n$. This will cause the atom to split into smaller fragments.

For example, when uranium-235 absorbs a high energy neutron, $^{1}_{0}n$, it breaks up, or undergoes fission as follows:

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{87}_{35}Br + \square + 3^{1}_{0}n$$

How would you identify the missing particle? Notice that three neutrons, $^{1}_{0}n$, have a mass number of 3 and a total atomic number of 0. The total atomic mass on the left side is (235 + 1) = 236. On the right we have (87 + 3(1)) = 90, and so (236 − 90) = 146 remains. The missing particle must have a mass number of 146.

The total atomic number on the left is 92. The total atomic number on the right, not including the missing particle, is 35. This means that (92 − 35) = 57 is the atomic number of the missing particle. From the periodic table, atomic number 57 corresponds to La, lanthanum. The balanced nuclear equation is

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{87}_{35}Br + ^{146}_{57}La + 3^{1}_{0}n$$

Check your answer by noting that the total mass number and the total atomic number are the same on both sides.

Nuclear fusion occurs when a target nucleus absorbs an accelerated particle. The reaction that takes place in a hydrogen bomb is a fusion reaction, as are the reactions that take place within the Sun, shown in Figure 4.21. Fusion reactions require very high temperatures to proceed but produce enormous amounts of energy. The fusion reaction that takes place in a hydrogen bomb is represented by the following equation:

$$^{6}_{3}\text{Li} + ^{1}_{0}\text{n} \rightarrow ^{3}_{1}\text{H} + ^{4}_{2}\text{He}$$

Notice that the total mass numbers and the total atomic numbers are the same on both sides.

Figure 4.21 The Sun's interior has a temperature of about 15 000 000 °C, due to the energy provided by nuclear fusion reactions.

Practice Problems

36. Astatine can be produced by the bombardment of a certain atom with alpha particles, as follows:

$\underline{\hspace{2em}} + ^{4}_{2}\text{He} \rightarrow ^{211}_{85}\text{At} + 2^{1}_{0}\text{n}$

Identify the atom.

37. Balance the following equation by adding a coefficient.

$^{252}_{96}\text{Cf} + ^{10}_{5}\text{B} \rightarrow ^{257}_{101}\text{Md} + \underline{\hspace{2em}}\ ^{1}_{0}\text{n}$

38. How many neutrons are produced when U-238 is bombarded with C-12 nuclei in a particle accelerator? Balance the following equation.

$^{238}_{92}\text{U} + ^{12}_{6}\text{C} \rightarrow ^{246}_{98}\text{Cf} + \underline{\hspace{2em}}\ ^{1}_{0}\text{n}$

39. Aluminum-27, when it collides with a certain nucleus, transforms into phosphorus-30 along with a neutron. Write a balanced nuclear equation for this reaction.

Section Wrap-Up

In this chapter, you learned how atoms can interact with each other and how unstable isotopes behave. In the first three sections, you learned about chemical reactions. In section 4.4, you learned about different types of nuclear reactions: reactions in which atoms of one element change into atoms of another element. In Unit 2, you will learn how stable isotopes contribute to an indirect counting method for atoms and molecules.

Section Review

1 K/U Draw a chart in your notebook to show alpha decay, beta decay, gamma decay, nuclear fusion, and nuclear fission. Write a description and give an example of each type of reaction. Illustrate each example with a drawing.

2 K/U Complete each nuclear equation. Then state the type of nuclear reaction that each equation represents.

(a) $^{232}_{90}\text{Th} + \underline{\hspace{2em}} \rightarrow ^{233}_{90}\text{Th}$

(b) $^{233}_{91}\text{Pa} \rightarrow ^{233}_{92}\text{U} + \underline{\hspace{2em}}$

(c) $^{226}_{88}\text{Ra} \rightarrow \underline{\hspace{2em}} + ^{4}_{2}\text{He}$

(d) $^{210}_{83}\text{Bi} \rightarrow ^{206}_{81}\text{Tl} + \underline{\hspace{2em}}$

(e) $^{210}_{83}Bi \rightarrow + ^{206}_{81}Tl +$ ______

(f) ______ $+ ^{1}_{0}n \rightarrow ^{90}_{38}Sr + ^{143}_{54}Xe + 3^{1}_{0}n$

(g) $^{6}_{3}Li + ^{2}_{1}H \rightarrow 2$ ______

3 MC Nuclear reactors have complex cooling systems that absorb the heat given off by the fission reaction. The absorbed heat is used to produce steam to drive a generator, thus producing electrical energy. Cooling the steam for re-use requires a large amount of cool water, which is usually obtained from a nearby river or lake. A large amount of hot water is then released into the river or lake. Do you think this is a form of pollution? What kinds of problems might warm water cause?

4 K/U Alpha or beta particle emission from a radioactive nucleus is often, but not always, accompanied by gamma rays. Why does the presence of gamma rays *not* affect how a nuclear equation of this type is balanced?

5 K/U Write a balanced nuclear equation to describe each of the following statements. Classify the reactions.

(a) Radon-222 undergoes alpha decay, forming polonium-218.

(b) When hydrogen-2 (deuterium) and hydrogen-3 (tritium) react, they form an alpha particle and a subatomic particle.

(c) Bismuth-214 undergoes beta decay, emitting one electron and forming a different nucleus.

(d) When a neutron collides with uranium-235, it forms krypton-92 and one other nucleus.

(e) Polonium-218 decays to lead-214, emitting one other particle.

(f) Strontium-90 emits a subatomic particle, forming yttrium-90.

CHAPTER 4 Review

Reflecting on Chapter 4

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

- Distinguish between chemical reactions and nuclear reactions.
- Summarize guidelines for balancing chemical equations.
- Summarize the different types of chemical reactions.
- Summarize the types of nuclear decay.
- Explain why knowing the solubility of compounds is important to predicting double displacement reactions.
- Summarize guidelines for balancing nuclear equations
- Describe how to use the activity series of metals and the activity series of halogens.

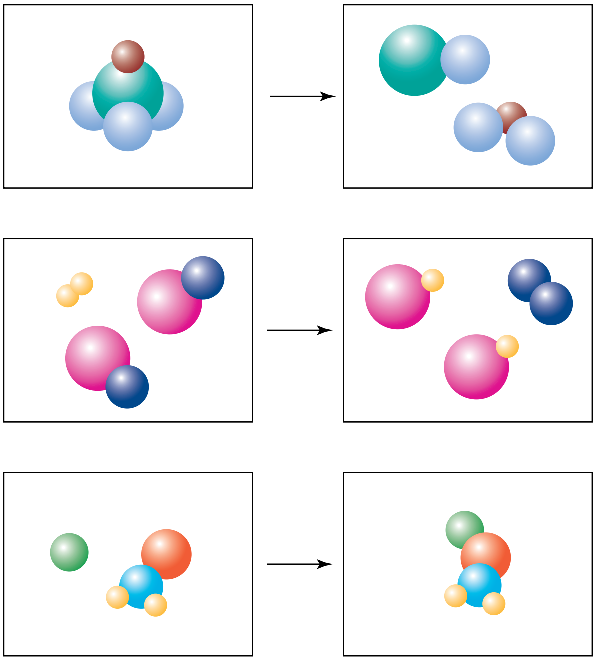
Reviewing Key Terms

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

activity series	neutralization reactions
alpha (α) particle emission	nuclear equation
balanced chemical equation	nuclear fission
beta particle	nuclear fusion
beta (*β*) decay	nuclear reactions
chemical equations	precipitate
chemical reactions	product
combustion reaction	reactant
decomposition reaction	single displacement reactions
double displacement reaction	skeleton equation
gamma (γ) radiation	synthesis reaction
incomplete combustion	word equation
law of conservation of mass	

Knowledge/Understanding

1. How can you tell if a chemic.al reaction has occurred?
2. Explain why the mixing of red paint with white paint does not constitute a chemical reaction, even though the "product" has a different appearance.
3. Explain how balancing a chemical equation satisfies the law of conservation of mass.
4. Copy each chemical equation into your notebook, and balance it.
   (a) $PdCl_{2(aq)} + HNO_{3(aq)} \rightarrow Pd(NO_3)_{2(aq)} + HCl_{(aq)}$
   (b) $Cr_{(s)} + HCl_{(aq)} \rightarrow CrCl_{2(aq)} + H_{2(g)}$
   (c) $FeO_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$
5. What type of chemical reaction is illustrated in each diagram below?



6. Classify each reaction as synthesis, decomposition, single displacement, double displacement, or combustion. Also, balance each chemical equation.
   (a) $H_{2(g)} + CuO_{(s)} \rightarrow Cu_{(s)} + H_2O_{(g)}$
   (b) $Ag_{(s)} + S_{8(s)} \rightarrow Ag_2S_{(s)}$
   (c) $C_4H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$
   (d) $NH_{3(g)} + HCl_{(g)} \rightarrow NH_4Cl_{(s)}$
   (e) $Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$
   (f) $RbCl_{(s)} + O_{2(g)} \rightarrow RbClO_{4(s)}$
   (g) $Cu_2S_{(s)} + O_{2(g)} \rightarrow Cu_2O_{(g)} + SO_{2(g)}$
7. Why is the solubility chart useful for analyzing double displacement reactions?
8. Nitrogen dioxide is a component of smog. It is produced in an automobile engine's combustion chamber. When exposed to sunlight, nitrogen dioxide forms nitrogen monoxide and oxygen. What type of reaction is this?

9. Write a balanced chemical equation corresponding to each word equation.
   (a) The reaction between aqueous sodium hydroxide and iron(III) nitrate produces a precipitate.
   (b) Powdered antimony reacts with chlorine gas to produce antimony trichloride.
   (c) Mercury(II) oxide is prepared from its elements.
   (d) Ammonium nitrite decomposes into nitrogen gas and water.
   (e) Aluminum metal reacts with a solution of zinc sulfate to produce aluminum sulfate and metallic zinc.
10. Consider the unbalanced chemical equation corresponding to the formation of solid lead(II) chromate, $PbCrO_4$:
    $Pb(NO_3)_{2(aq)} + K_2CrO_{4(aq)} \rightarrow PbCrO_{4(s)} + KNO_{3(aq)}$
    (a) What type of chemical reaction is this?
    (b) Balance the equation.
11. In general, what is formed when an oxide of a non-metal reacts with water? Give an example.
12. In general, what is formed when an oxide of a metal reacts with water? Give an example.
13. Complete and balance each nuclear equation. Then classify the reaction.
    (a) ${}^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{4}_{2}He +$ ▭
    (b) ${}^{239}_{92}U \rightarrow$ ▭ $+ {}^{0}_{-1}\beta$
    (c) ${}^{239}_{93}Np \rightarrow {}^{239}_{94}Pu +$ ▭
    (d) ${}^{238}_{92}U \rightarrow {}^{234}_{90}Th +$ ▭ $+ 2{}^{0}_{0}\gamma$
14. Write the product(s) for each reaction. If you predict that there will be no reaction, write "NR." Balance each chemical equation.
    (a) $BaCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow$
    (b) $Fe_{(s)} + CuSO_{4(aq)} \rightarrow$
    (c) $C_2H_{2(g)} + O_{2(g)} \rightarrow$
    (d) $PCl_{5(s)} \rightarrow$ ▭ $+ Cl_{2(g)}$
    (e) $Mg_{(s)} + Fe_2O_3 \rightarrow$
    (f) $Ca_{(s)} + Cl_{2(g)} \rightarrow$
15. Iron often occurs as an oxide, such as Fe_2O_3. In the steel industry, Fe_2O_3 is reacted with carbon monoxide to produce iron metal and carbon dioxide. Write the balanced chemical equation for this reaction, and classify it.
16. Calcium chloride is often used to melt ice on roads and sidewalks, or to prevent it from forming. Calcium chloride can be made by reacting hydrochloric acid with calcium carbonate. Write the balanced chemical equation corresponding to this reaction, and classify it.

Inquiry

17. An American penny is composed of a zinc core clad in copper. Some of the copper is filed away, exposing the zinc, and placed in a solution of hydrochloric acid. Describe what will occur.
18. What will happen to a silver earring that is accidentally dropped into toilet bowl cleaner that contains hydrochloric acid?

Communication

19. Explain why it is advisable to store chemicals in tightly sealed bottles out of direct sunlight.
20. Why is smoking not allowed near an oxygen source? What would happen if a match were struck in an oxygen-rich atmosphere?
21. Even if a smoker is very careful not to let a lighted cigarette come in contact with liquid gasoline, why is it very dangerous to smoke when refuelling an automobile?
22. Solutions that have been used to process film contain silver ions, $Ag^{+}_{(aq)}$.
    (a) Explain how you could recover the silver, in the form of an ionic compound.
    (b) How could you recover the silver as silver metal?

Making Connections

23. Calcium oxide, CaO (lime), is used to make mortar and cement.
    (a) State two reactions that could be used to make lime. Classify each reaction, based on the types of reactions studied in this chapter.
    (b) In construction, cement is prepared by mixing the powdered cement with water. Write the chemical equation that represents the reaction of calcium oxide with water. Why are we cautioned not to expose skin to dry cement mix *and* wet cement? It may help you to know that bases are often corrosive. They can burn exposed skin.

Answers to Practice Problems and Short Answers to Section Review Questions

Practice Problems: 1.(a) calcium + fluorine (reactants) → calcium fluoride (product) **(b)** barium chloride + hydrogen sulfate → hydrogen chloride + barium sulfate **(c)** calcium carbonate + carbon dioxide + water → calcium hydrogen carbonate **(d)** hydrogen peroxide → water + oxygen **(e)** sulfur dioxide + oxygen → sulfur trioxide **2.** Sugar → ethanol + carbon dioxide **3.(a)** $Zn_{(s)} + Cl_{2(g)} \rightarrow ZnCl_{2(s)}$ **(b)** $Ca_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$ **(c)** $Ba_{(s)} + S_{(s)} \rightarrow BaS_{(s)}$ **(d)** $Pb(NO_3)_{2(aq)} + Mg_{(s)} \rightarrow Mg(NO_3)_{2(aq)} + Pb_{(s)}$ **4.(a)** $CO2_{(g)} + CaO_{(s)} \rightarrow CaCO_{3(s)}$ **(b)** $Al_{(s)} + O_{2(g)} \rightarrow Al_2O_{3(s)}$ **(c)** $Mg_{(s)} + O_{2(g)} \rightarrow MgO_{(s)}$ **5.(a)** $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$ **(b)** $P_{4(s)} + 5O_{2(g)} \rightarrow P_4O_{10(s)}$ **(c)** $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$ **(d)** $SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$ **6.(a)** balanced **(b)** $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$ **(c)** $H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$ **(d)** balanced **7.(a)** $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$ **(b)** $BaCl_{2(aq)} + Na_2SO_{4(aq)} \rightarrow NaCl_{(aq)} + BaSO_{4(s)}$ **8.** $P_{4(s)} + 5O_{2(g)} \rightarrow P_4O_{10(s)}; P_4O_{10(s)} + 6H_2O_{(l)} \rightarrow 4H_3PO_{4(aq)}$ **9.(a)** $As_4S_{6(s)} + 9O_{2(g)} \rightarrow As_4O_{6(s)} + 6SO_{2(g)}$ **(b)** $Sc_2O_{3(s)} + 3H_2O_{(l)} \rightarrow 2Sc(OH)_{3(s)}$ **(c)** $C_2H_5OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(l)}$ **(d)** $2C_4H_{10(g)} + 9O_{2(g)} \rightarrow 8CO_{(g)} + 10H_2O_{(g)}$ **10.(a)** $2K + Br_2 \rightarrow 2KBr$ **(b)** $H_2 + Cl_2 \rightarrow 2HCl$ **(c)** $Ca + Cl_2 \rightarrow CaCl_2$ **(d)** $Li + O_2 \rightarrow LiO_2$ **11.(a)** products are Fe_2O_3, FeO **(b)** possible products: V_2O_5, VO, V_2O_3, VO_2 **(c)** possible products: TiO_2, TiO, Ti_2O_3 **12.(a)** $K_2O + H_2O \rightarrow 2KOH$ **(b)** $MgO + H_2O \rightarrow Mg(OH)_2$ **(c)** $SO_2 + H_2O \rightarrow H_2SO_3$ **13.** $NH_{3(g)} + HCl_{(g)} \rightarrow NH_4Cl_{(s)}$ **14.** Hg, O_2 **15.(a)** $2HI \rightarrow H_2 + I_2$ **(b)** $2Ag_2O \rightarrow 4Ag + O_2$ **(c)** $2AlCl_3 \rightarrow 2Al + 3Cl_2$ **(d)** $MgO \rightarrow Mg + O_2$ **16.(a)** $MgCO_3 \rightarrow MgO + CO_2$ **(b)** $CuCO_3 \rightarrow CuO + CO_2$ **17.** $2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$ **18.** $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$ **19.** $C_3H_6O + O_2 \rightarrow CO_2 + H_2O$ **20.** $2C_{16}H_{34} + 49O_2 \rightarrow 32CO_2 + 34H_2O$ **21.(a)** $Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$ **(b)** $Zn + Pb(NO_3)_2 \rightarrow Zn(NO_3)_2 + Pb$ **(c)** $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$ **(d)** $Li + AgNO_3 \rightarrow Ag + LiNO_3$ **(e)** $Pb + H_2SO_4 \rightarrow PbSO_4 + H_2$ **(f)** $2Mg + Pt(OH)_4 \rightarrow 2Mg(OH)_2 + Pt$ **(g)** $Ba + FeCl_2 \rightarrow BaCl_2 + Fe$ **(h)** $Fe + Co(ClO_3)_2 \rightarrow Fe(ClO_3)_3 + Co$ **22.(a)** NR **(b)** $Zn + FeCl_2 \rightarrow ZnCl_2 + Fe$ **(c)** $K + H_2O \rightarrow KOH + H_2$ **(d)** $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$ **(e)** NR **(f)** NR **(g)** $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$ **(h)** $Mg + SnCl_2 \rightarrow MgCl_2 + Sn$ **23.(a)** NR **(b)** $Cl_2 + 2NaI \rightarrow 2NaCl + I_2$ **24.(a)** $2Pb + 2HCl \rightarrow 2PbCl + H_2$ **(b)** $KI + Br_2 \rightarrow KBr + I_2$ **(c)** NR **(d)** $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$ **(e)** NR **(f)** $Ni + H_2SO_4 \rightarrow NiSO_4 + H_2$ **25.(a)** $Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow 2KNO_{3(aq)} + PbI_{2(s)}$ **(b)** NR **(c)** NR **(d)** $Ba(NO_3)_{2(aq)} + Mg(SO_4)_{(aq)} \rightarrow BaSO_{4(s)} + Mg(NO_3)_{2(aq)}$ **26.(a)** $Na_2SO_{3(aq)} + 2HCl_{(aq)} \rightarrow SO_{2(g)} + 2NaCl_{(aq)} + H_2O_{(l)}$ **(b)** $CaS_{(aq)} + H_2SO_{4(aq)} \rightarrow H_2S_{(g)} + CaSO_{4(l)}$ **27.(a)** $HCl_{(aq)} + LiOH_{(aq)} \rightarrow H_2O_{(l)} + LiCl_{(aq)}$ **(b)** $HclO_{4(aq)} + Ca(OH)_{2(aq)} \rightarrow H_2O_{(l)} + Ca(ClO_4)_{2(aq)}$ **(c)** $H_2SO_{4(aq)} + NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$ **28.(a)** $BaCl_{2(aq)} + Na_2CrO_{4(aq)} \rightarrow BaCrO_{4(s)} + 2NaCl_{(aq)}$ **(b)** $HNO_{3(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaNO_{3(aq)}$ **(c)** $K_2CO_{3(aq)} + 2HNO_{3(aq)} \rightarrow H_2O_{(l)} + 2KNO_{3(aq)} + CO_{2(g)}$ **29.** [234/90]Th **30.** [222/86]Rn → [4/2]He + [218/82]Pb **31.** [242/94]Pu → [4/2]He + [238/92]U **32.** [144/60]Nd → [4/2]He + [140/58]Ce **33.** [40/19]K → [0/ – 1]e + [40/20]Ca **34.** [47/20]Ca **35.** [73/31]Ga → [0/ – 1]e + [73/32]Ge **36.** [208/83]Bi **37.** 5 **38.** 4 **39.** [27/13]Al+ [4/2]He → [30/15]P + [1/0]n

Section Review: 4.1: 2.(a) $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$ **(b)** $Na_{(s)} + H_2O_{(l)} \rightarrow H_{2(g)} + NaOH_{(aq)}$ **(c)** $Cu_{(s)} + HNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + NO_{2(g)} + H_2O_{(l)}$ **4.(a)** $4Al_{(s)} + 3O_{2(g)} \rightarrow 42Al_2O_{3(s)}$ **(b)** $2Na_2S_2O_{3(aq)} + I_{2(aq)} \rightarrow 2NaI_{(aq)} + Na_2S_4O_{6(aq)}$ **(c)** $2Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2Fe_{(s)}$ **(d)** $4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(l)}$ **(e)** $Na_2O_{(s)} + (NH_4)_2SO_{4(aq)} + H_2O_{(l)} + NH_{3(aq)}$ **(f)** $C_5H_{12(l)} + 8O_{2(g)} \rightarrow 5CO_{2(g)} + 6H_2O_{(g)}$ **5.** $Fe_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + FeSO_{4(aq)}$ **4.2: 1.(a)** $Be + O_2 \rightarrow BeO$ **(b)** $2Li + Cl_2 \rightarrow 2LiCl$ **(c)** $Mg + N_2 \rightarrow Mg_3N_2$ **(d)** $Ca + Br_2 \rightarrow CaBr_2$ **2.(a)** $2K_2O \rightarrow O_2 + 4K$ **(b)** $2CuO \rightarrow 2Cu + O_2$ **(c)** $2H_2O \rightarrow 2H_2 + O_2$ **(d)** $2Ni_2O_3 \rightarrow 4Ni + 3O_2$ **(e)** $2Ag_2O \rightarrow 4Ag + O_2$ **3.(a)** $Sn(OH)_{4(s)} \rightarrow SnO_{2(s)} + 2H_2O_{(g)}$, decomposition **(b)** $3Cl_2(g) + I_{2(s)} \rightarrow 2ICl_3$ synthesis **(c)** $C_4H_9OH + 6O_2 \rightarrow 5H_2O + 4CO_2$ **5.** $2HgO_{(s)} \rightarrow O_{2(g)} + 2Hg_{(s)}$ decomposition **4.3: 1.(a)** $Li + H_2O \rightarrow Li_2O + H_2$ **(b)** NR **(c)** $F_2 + 2KI \rightarrow 2KF + I_2$ **(d)** NR **(e)** $Zn + CuSO_4 \rightarrow Cu + ZnSO_4$ **(f)** $K + H_2O \rightarrow K_2O + H_2$ **2.(a)** $NaOH_{(aq)} + Fe(NO_3)_{3(aq)} \rightarrow NaNO_{3(aq)} + Fe(OH)_{3(s)}$ **(b)** $Ca(OH)_{2(aq)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O)_{(l)}$ **(c)** NR **(d)** $K_2CO_{3(s)} + H_2SO_{4(aq)} \rightarrow K_2SO_{4(aq)} + CO_{2(g)} + H_2O_{(l)}$ **3.(b)** $(NH_4)_2SO_{4(aq)} + 2KOH_{(aq)} \rightarrow 2NH_{3(g)} + 2H_2O_{(l)} + K_2SO_{4(aq)}$ **4.(a)** incomplete combustion **(b)** single displacement **(c)** double displacement **(d)** complete combustion **(e)** decomposition **(f)** synthesis **(g)** decomposition **5.4: 2.(a)** [1/0]n **(b)** [0/ – 1]e **(c)** [222/2]Rn **(d)** [4/2]He **(e)** [236/92]U **(f)** [4/2]He