SECTION 4.3

Reactions in Industry

At the beginning of Unit 2, you read about a chemical reaction that is used to weld railroad tracks together. This is just one of the many reactions that are involved in the processes required to make the products you use in your daily life. Chemical reactions are also involved in cleaning up the problems that result from these processes.

Key Terms

matte
leaching

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Thermite Reactions

Now that you have learned more about chemical reactions, you can re-examine thermite reactions. When this type of reaction is used to weld railroad tracks, the reactants, aluminum and iron oxide are powdered and thoroughly mixed. The reaction is a single displacement reaction in which aluminum displaces iron according to the following balanced chemical equation:

$$\begin{array}{ccccccc} \text{aluminum} & + & \text{iron oxide} & \rightarrow & \text{iron} & + & \text{aluminum oxide} \\ 8\text{Al(s)} & + & 3\text{Fe}_3\text{O}_4\text{(s)} & \rightarrow & 9\text{Fe}(\ell) & + & 4\text{Al}_2\text{O}_3\text{(s)} \end{array}$$

Once the reaction begins, it releases enough energy to melt the iron.

In **Figure 4.13**, an aluminum wrench is hitting rusty iron, causing a small-scale thermite reaction that produces sparks. In industrial settings or workshops, the grinding or cutting of iron or steel can produce powdered iron oxides. Care must be taken in such areas to avoid an unexpected thermite reaction when using aluminum objects.

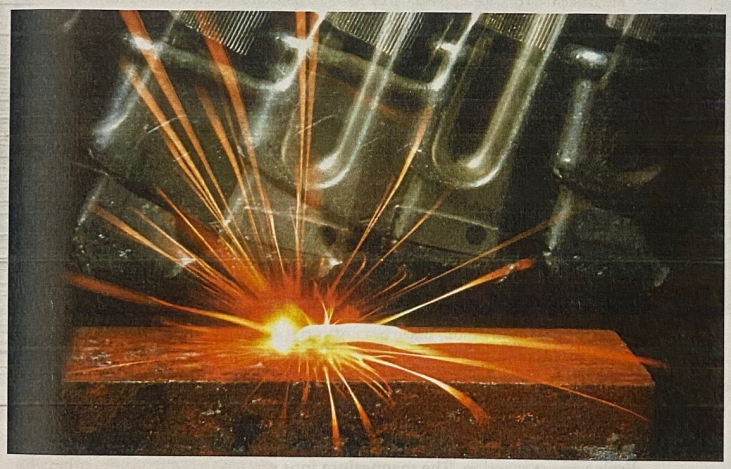

Figure 4.13 The friction from the impact of an aluminum wrench with rusty iron generated enough heat to initiate a small thermite reaction, as shown by the bright sparks flying away from the point of impact.

Analyze *The thermite reaction shown here occurs in the solid state. How is this different from the other single displacement reactions you have studied?*

In addition to iron oxide, other metals can be used in thermite reactions. For example, a thermite reaction involving copper(II) oxide can be used to produce pure copper according to the following chemical equation:

$$\begin{array}{ccccccc} \text{aluminum} & + & \text{copper(II) oxide} & \rightarrow & \text{copper} & + & \text{aluminum oxide} \\ 2\text{Al(s)} & + & 3\text{CuO(s)} & \rightarrow & 3\text{Cu}(\ell) & + & \text{Al}_2\text{O}_3\text{(s)} \end{array}$$

This reaction is often used to produce pure copper for welding electrical conductors. The copper(II) oxide is held in a heat-resistant reaction chamber. The thermite reaction heats the copper enough to melt it. The liquid copper is then allowed to flow into a mold that surrounds the ends of the conductors. When the copper cools into a solid, it forms a weld that allows an electric current to flow between the conductors.

Magnesium Mining from Seawater

Many metals, such as copper, zinc, and gold, are extracted from solid ores. Although magnesium is abundant in Earth's crust, it is extracted from seawater, not rocks. As a result, plants that produce magnesium, such as the one shown in **Figure 4.14**, are usually located on the coast. Magnesium ions are the second most abundant cations found in seawater, with only sodium ions in greater abundance. The process of producing metallic magnesium requires several chemical reactions.

Figure 4.14 This plant processes seawater to extract magnesium.

Steps in the Process of Magnesium Mining

Figure 4.15 shows the main steps in extracting magnesium from seawater. The steps in the flowchart match the reactions described below.

Calcium carbonate is decomposed to produce calcium oxide.

↓

Calcium hydroxide is produced by a synthesis reaction between calcium oxide and water.

↓

Magnesium hydroxide precipitates when calcium hydroxide is mixed with seawater.

↓

A neutralization reaction between magnesium hydroxide and hydrochloric acid produces magnesium chloride.

↓

Magnesium and chlorine are produced by the decomposition of magnesium chloride.

Figure 4.15 Use this flowchart to help you understand the sequence of chemical reactions used in the extraction of magnesium from seawater.

Decomposition of Calcium Carbonate

First, calcium carbonate from seashells is decomposed to produce calcium oxide. Recall, from Chapter 3, that carbon dioxide is a product when a metal carbonate decomposes. The chemical equation is

$$\underset{\text{calcium carbonate}}{CaCO_3(s)} \rightarrow \underset{\text{calcium oxide}}{CaO(s)} + \underset{\text{carbon dioxide}}{CO_2(g)}$$

Synthesis of Calcium Hydroxide

The calcium oxide undergoes a synthesis reaction to form calcium hydroxide:

$$\underset{\text{calcium oxide}}{CaO(s)} + \underset{\text{water}}{H_2O(\ell)} \rightarrow \underset{\text{calcium hydroxide}}{Ca(OH)_2(aq)}$$

Double Displacement between Calcium Hydroxide and Magnesium Compounds

The calcium hydroxide reacts with the magnesium ions in seawater in a double displacement reaction. The magnesium ions are separated from the other ions in seawater, such as sodium, chloride, and bromide ions, as the precipitate magnesium hydroxide:

$$\underset{\text{calcium hydroxide}}{Ca(OH)_2(aq)} + \underset{\text{magnesium ions}}{Mg^{2+}(aq)} \rightarrow \underset{\text{magnesium hydroxide}}{Mg(OH)_2(s)} + \underset{\text{calcium ions}}{Ca^{2+}(aq)}$$

Neutralization of Magnesium Hydroxide

The solid magnesium hydroxide is filtered out and undergoes neutralization with hydrochloric acid:

$$\underset{\text{magnesium hydroxide}}{Mg(OH)_2(s)} + \underset{\text{hydrochloric acid}}{2HCl(aq)} \rightarrow \underset{\text{magnesium chloride}}{MgCl_2(aq)} + \underset{\text{water}}{2H_2O(\ell)}$$

As in the earlier furnaces, the sulfur dioxide gas is collected, so it can be sent to a plant that produces sulfuric acid. The small furnaces make it easier to recover the gases and provide a steady stream of sulfur dioxide to the acid plant. Meanwhile, the copper undergoes a final purification step in the anode furnace.

Because copper is an excellent conductor of electricity, a major use of copper is in electrical wiring. Copper is also melted with other metals to make alloys. For example, bronze is mainly an alloy of copper and tin, and brass is mainly an alloy of copper and zinc.

Gold and Cyanide Leaching

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Gold is a relatively non-reactive metal. As a result, it can be found in nature in its uncombined form. As shown in **Figure 4.18**, however, most gold is mixed into the rock that surrounds it and must be separated. The most cost-effective method for removing the gold involves treating the crushed rock with a sodium cyanide solution to dissolve the gold.

Figure 4.18 Much of the rock mined as gold ore contains particles of gold that are too tiny to be seen without a microscope. Some rocks, however, such as the one shown here, contain visible particles of gold.

Leaching is the process of converting a metal to a soluble form to extract the metal. A commonly used process for the extraction of gold from ore is to react it with a solution of sodium cyanide, NaCN(aq), to form sodium dicyanoaurate(I), $Na[Au(CN)_2](aq)$, and sodium hydroxide. The overall reaction for leaching gold is

leaching a process that is used to extract a metal by dissolving the metal in an aqueous solution

$$4Au(s) + 8NaCN(aq) + O_2(g) + 2H_2O(\ell) \rightarrow 4Na[Au(CN)_2](aq) + 4NaOH(aq)$$

The gold is recovered from the solution through displacement by zinc, according to the following equation:

$$2Na[Au(CN)_2](aq) + Zn(s) \rightarrow 2NaCN(aq) + Zn(CN)_2(aq) + 2Au(s)$$

The cyanide solution can be recycled and used again to convert gold to a soluble form in an aqueous solution.

In industry, the most important use of gold is in the manufacture of electronics components. Gold is an excellent conductor of electricity and is resistant to corrosion, so it is used in small amounts in electronic devices, including cell phones and computers. Gold is also an ingredient in some medications. In addition, it is used to fill cavities in teeth and to make crowns to cover and protect teeth. Gold is highly suited to such uses because it does not corrode, it does not trigger allergic reactions, and it is easy to shape into the desired form. Gold coins have been used for many centuries as currency. Other uses of gold include the production of jewellery, watches, and art objects.

Waste and Spill Treatment

Because of the potentially harmful effects of the chemicals that are used and formed during metal production, steps are taken to reduce emissions and to respond to spills.

Sulfur Dioxide Waste

Some industries, such as coal-burning power plants, use scrubbers to remove sulfur dioxide, $SO_2(g)$, from exhaust gases to prevent its release into the atmosphere. If sulfur dioxide is released into the atmosphere, it can eventually become sulfuric acid in rain and snow.

Because many metals, such as copper, are found as sulfide ores, large amounts of sulfur dioxide are commonly formed during metal extraction and purification. Sulfur dioxide can be converted into sulfuric acid, which is either used in some of the purification steps or sold. You may recognize the synthesis reactions that are involved in forming sulfuric acid from Chapter 3:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$
$$SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(aq)$$

Cyanide Spills

Cyanide leaching allows gold to be extracted from ores that would have too low a gold content to be profitable. Unfortunately, cyanide is deadly in very small amounts if ingested. A cyanide spill may occur if the wall of a holding pond breaks or if a storm produces a large amount of rain and causes the holding pond to overflow. Two methods are commonly used to treat a cyanide spill.

Use of Sodium Hypochlorite

The first method is a two-reaction process. In the first reaction, sodium hypochlorite, the active ingredient in many chlorine bleaches, is added to the cyanide solution:

$$NaCN(aq) + NaOCl(aq) \rightarrow NaCNO(aq) + NaCl(aq)$$

Sodium cyanate, NaCNO(aq), much less toxic than sodium cyanide. However, the second reaction entirely eliminates any toxicity. Additional sodium hypochlorite is used to convert the sodium cyanate into non-toxic compounds:

$$2NaCNO(aq) + 3NaOCl(aq) + H_2O(\ell) \rightarrow 3NaCl(aq) + N_2(g) + 2NaHCO_3(aq)$$

Use of Iron(II) Sulfate

The second method involves adding iron(II) sulfate, which binds the toxic and soluble cyanide ions into non-hazardous, complex iron(II) cyanide ions. The iron(II) cyanide ions form precipitates with many metal ions, such as zinc and iron, as shown in **Figure 4.19**.

First, a double displacement reaction occurs:

$$FeSO_4(aq) + 2NaCN(aq) \rightarrow Fe(CN)_2(aq) + Na_2SO_4(aq)$$

Next, a synthesis reaction forms complex ion, iron(II) cyanide ions:

$$Fe(CN)_2(aq) + 4NaCN(aq) \rightarrow Na_4Fe(CN)_6(aq)$$

Finally, double displacement reactions, such as the reaction below, cause the iron(II) cyanide ions to form precipitates with other metal ions that are present.

$$3Na_4Fe(CN)_6(aq) + 4FeCl_3(aq) \rightarrow Fe_4[Fe(CN)_6]_3(s) + 12NaCl(aq)$$

Figure 4.19 The precipitate that is formed in the reaction between sodium iron(II) cyanide and iron(III) chloride is a pigment known as Prussian blue. This pigment is used in blueprints and paints.

Infer ***Why is forming a precipitate helpful when cleaning up a cyanide spill?***

Extracting Metals from Ores

Gold and copper are important metals for Canada's economy. Several different methods are used for extracting these metals. Each method combines physical processes, such as grinding and filtering, and chemical processes to obtain the desired product.

Copper Smelting

A smelter is a facility that uses heat to extract metal from ore. **Figure 4.17** shows a smelter that is used to refine copper at the Kidd Creek copper and zinc mine in Timmins, Ontario. An important part of the design of this type of smelter is the different heights of the furnaces. Molten products flow from one furnace continuously into the next furnace because of the height difference, reducing the need for large buckets and transfer equipment.

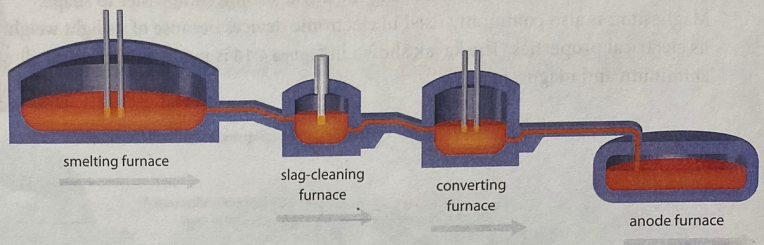

Figure 4.17 These furnaces produce increasingly pure copper through a series of reactions.
***Identify** how the construction of the smelter allows gravity to move the material from one furnace to the next.*

matte an impure copper(I) sulfide mixture that is formed by smelting the sulfide ore

Smelting Furnace

The smelting furnace is the first stage of copper production. An impure copper sulfide **matte**, $Cu_2S(\ell)$, forms. Oxygen is used to separate the iron from the copper in the mineral chalcopyrite, $CuFeS_2(s)$, ore according to this overall reaction:

$$2CuFeS_2(s) + 4O_2(g) \rightarrow Cu_2S(\ell) + 2FeO(\ell) + 3SO_2(g)$$

The iron(II) oxide is reacted with sand and limestone to convert it into a low density compound called *slag* which floats to the top of the molten mixture. The copper matte is more dense and sinks to the bottom.

Suggested Investigation
Inquiry Investigation 4-E, From Copper to Copper

The sulfur dioxide gas that is produced in these reactions and in later steps of the process is collected and sent to a plant that produces sulfuric acid. This prevents the release of large amounts of pollutants that contribute to the formation of acid precipitation.

Slag-Cleaning Furnace

The products of the smelting furnace move to the slag-cleaning furnace, where they separate due to differences in density. The matte flows into the converting furnace, while the undesired material, called *slag*, is sent to storage.

Converting Furnace and Anode Furnace

In the next step, matte is purified further, to about 99 percent, in the converting furnace through further displacement reactions. Air is blown through the molten mixture and oxygen in the air reacts with the copper matte in a two step reaction. First, some copper(I) sulfide is converted to copper(I) oxide. The copper(I) oxide then reacts with more copper(I) sulfide to form metallic copper and sulfur dioxide.

$$2Cu_2S(\ell) + 3O_2(g) \rightarrow 2Cu_2O(\ell) + 2SO_2(g)$$
$$Cu_2S(\ell) + 2Cu_2O(\ell) \rightarrow 6Cu(\ell) + SO_2(g)$$

Decomposition of Magnesium Chloride

The magnesium chloride is dried, melted, and then decomposed through electrolysis to form magnesium metal:

$$\text{magnesium chloride} \rightarrow \text{magnesium} + \text{chlorine}$$
$$MgCl_2(\ell) \rightarrow Mg(\ell) + Cl_2(g)$$

Industrial Uses of Magnesium

The main industrial use for magnesium is in the manufacturing of aluminum-magnesium alloys. An *alloy* is a mixture of two or more metals. Because magnesium is a less dense metal than aluminum, their alloys are lighter in weight than pure aluminum. In addition, the alloys are stronger and more resistant to corrosion than pure aluminum. Beverage cans are often made of aluminum to which a small amount of magnesium has been added, making the metal stronger and easier to shape. Magnesium is also commonly used in electronic devices because of its light weight and its electrical properties. The kayak shown in **Figure 4.16** is made, in part, of an alloy of aluminum and magnesium.

Figure 4.16 The frame of this kayak is constructed using tubing that is made from an alloy of aluminum and magnesium.

Infer *What properties of aluminum-magnesium tubing make it suitable for constructing a kayak?*

Learning Check

19. Describe the displacement that occurs in a thermite reaction involving iron oxide.

20. What product makes a thermite reaction useful for welding?

21. Explain how a thermite reaction can be used to produce pure copper.

22. How are seashells used to help extract magnesium from seawater?

23. During the extraction of magnesium from seawater, why is a precipitate of magnesium formed, if it is converted into soluble magnesium chloride in the next step?

24. Refer to **Figure 4.15**. What is an advantage of the production of chlorine in the final step of extracting magnesium from seawater?