

BIG Idea Chemical energy can be converted to electric energy and electric energy to chemical energy.

20.1 Voltaic Cells

MAIN Idea In voltaic cells, oxidation takes place at the anode, yielding electrons that flow to the cathode, where reduction occurs.

20.2 Batteries

MAIN Idea Batteries are voltaic cells that use spontaneous reactions to provide energy for a variety of purposes.

20.3 Electrolysis

MAIN Idea In electrolysis, a power source causes nonspontaneous reactions to occur in electrochemical cells.

ChemFacts

- A camera trap captured this image of a mountain lion. Camera traps are a noninvasive way to study animals.
- In a camera trap, a camera is attached to a sensor that triggers the camera's shutter when an animal approaches.
- The sensor is usually an infrared sensor, but pressure pads are also used.
- Batteries power both the camera and the sensor. Heat, humidity, and cold challenge battery performance.
- To conserve batteries, traps can be set to operate only at certain times.



Lithium-ion camera battery



Eight alkaline C-cells



Camera trap



Start-Up Activities

LAUNCH Lab

How can you make a battery from a lemon?

You can purchase a handy package of portable power at any convenience store—a battery. You can also light a bulb with a lemon. How are these power sources alike?



Procedure

1. Read and complete the lab safety form.
2. Insert **zinc** and **copper strips** into a **lemon**, about 2 cm apart from each other.
3. Attach the black lead from a **voltmeter** to the zinc and the red lead to the copper. Read the potential difference (voltage) shown on the voltmeter, and record your reading.
4. Remove one of the metals from the lemon, and observe what happens to the potential difference on the voltmeter.

Analysis

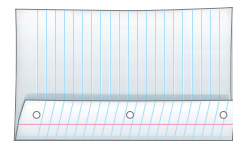
1. **Explain** the purpose of the zinc and copper metals.
2. **Infer** What is the function of the lemon?

Inquiry Do you think that you can make a battery from foods other than lemons? Form a hypothesis about other foods—or types of foods—that can be made into batteries. Then, design a battery using another food. If your teacher approves your design, test your hypothesis.

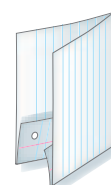
FOLDABLES™ Study Organizer

Electrochemical Cells Make this Foldable to help you compare voltaic cells to electrolytic cells.

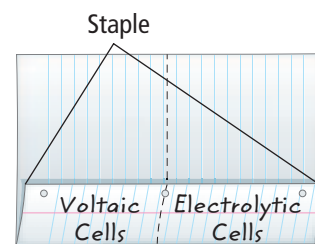
- ▶ **STEP 1** Fold up the bottom of a horizontal sheet of paper about 5 cm.



- ▶ **STEP 2** Fold the sheet in half.



- ▶ **STEP 3** Open the paper and staple the bottom flap to make two compartments. Label as shown.



FOLDABLES Use this Foldable with Sections 20.1 and 20.3. As you read about electrochemical cells, summarize information on index cards or quarter-sheets of paper and store them in the appropriate compartments.

Chemistry online

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- ▶ find the Try at Home Lab, Old Pennies

Section 20.1

Objectives

- Describe a way to obtain electrical energy from a redox reaction.
- Identify the parts of a voltaic cell, and explain how each part operates.
- Calculate cell potentials, and determine the spontaneity of redox reactions.

Review Vocabulary

oxidation: the loss of electrons from the atoms of a substance; increases an atom's oxidation number

reduction: the gain of electrons by the atoms of a substance; decreases the atom's oxidation number

New Vocabulary

salt bridge
electrochemical cell
voltaic cell
half-cell
anode
cathode
reduction potential
standard hydrogen electrode

Voltaic Cells

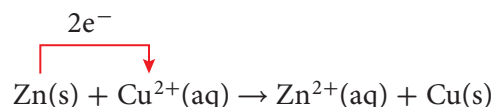
MAIN Idea In voltaic cells, oxidation takes place at the anode, yielding electrons that flow to the cathode, where reduction occurs.

Real-World Reading Link What could you do with half of a dollar bill? Without the other half, you cannot spend it. Voltaic cells have two half-cells, and both are required to produce energy.

Redox in Electrochemistry

Electrochemistry is the study of the redox processes by which chemical energy is converted to electrical energy and vice versa. Electrochemical processes are useful in industry and critically important for biological functioning.

In Chapter 19, you read that all redox reactions involve a transfer of electrons from the species that is oxidized to the species that is reduced. **Figure 20.1** and **Figure 20.2** illustrate the simple redox reaction in which zinc atoms are oxidized to form zinc (Zn^{2+}) ions. The two electrons donated from each zinc atom are accepted by a copper (Cu^{2+}) ion, which becomes an atom of copper metal. The following net ionic equation illustrates the electron transfer that occurs.

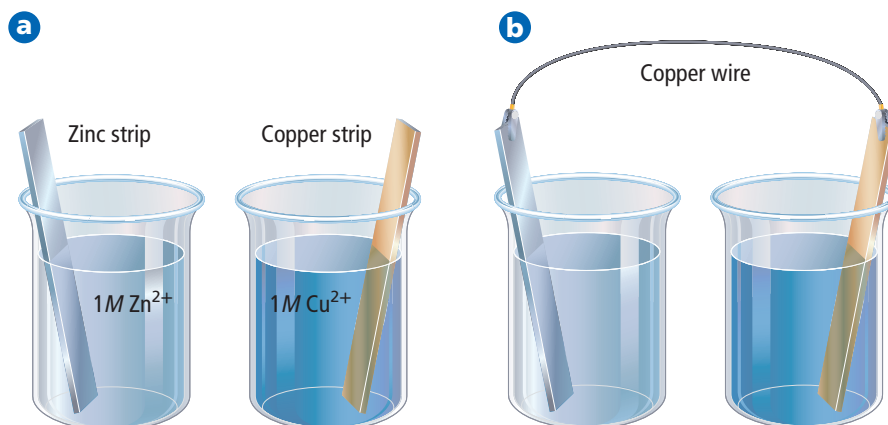


Half-reactions Two half-reactions make up this redox process:

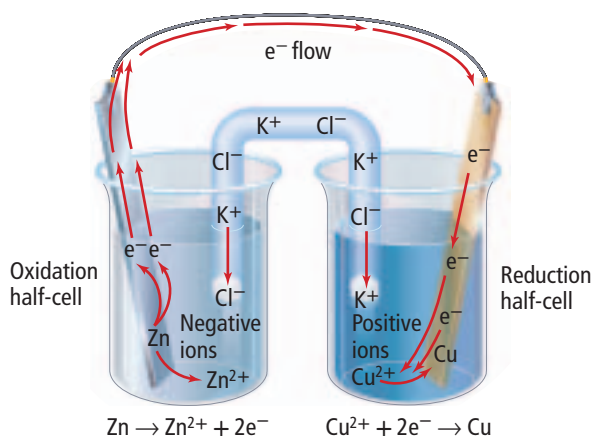


What do you think would happen if you separated the oxidation half-reaction from the reduction half-reaction? Can a redox reaction occur? Consider **Figure 20.1a**, in which a zinc strip is immersed in a solution of zinc sulfate and a copper strip is immersed in a solution of copper(II) sulfate.

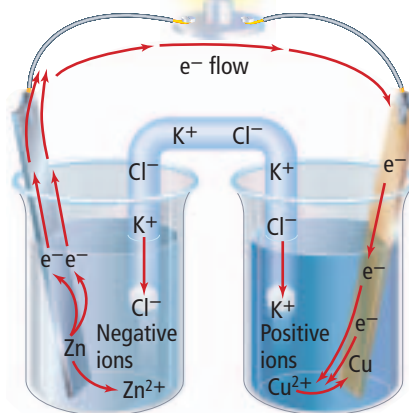
■ **Figure 20.1** These containers are constructed and arranged so that zinc will be oxidized on one side, while copper ions will be reduced on the other. In **a**, zinc metal is immersed in 1M zinc sulfate solution, and copper metal in 1M copper sulfate. In **b**, a wire joining the zinc and copper strips provides a pathway for the flow of electrons, but the pathway is not complete. Electron transfer is still not possible.



The salt bridge and the wire provide an unbroken pathway for electrical charge to flow.



The redox reaction provides energy to light the bulb.



■ **Figure 20.2** The addition of a salt bridge completes the pathway. Negative ions move through the salt bridge to the zinc side. Positive ions move through the bridge to the copper side.

Concepts in Motion

Interactive Figure To see an animation of a voltaic cell, visit glencoe.com.

Two problems prevent a redox reaction from occurring. First, there is no way for zinc atoms to transfer electrons to copper(II) ions. This problem can be solved by connecting the zinc and copper strips with a metal wire, as shown in **Figure 20.1b**. The wire can serve as a pathway for electrons to flow from the zinc strip to the copper strip.

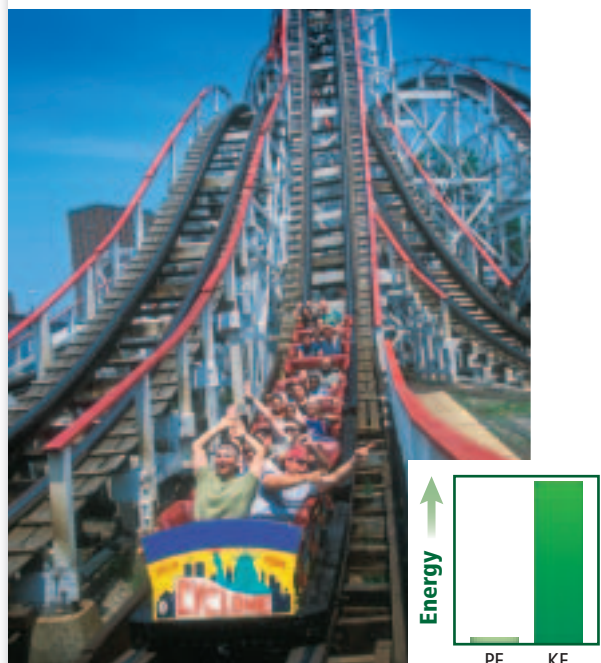
The second problem is that when the metal strips are placed in their solutions, oxidation begins at the zinc strip and reduction begins at the copper strip—but these reactions cannot continue. The reason is that as zinc is oxidized, positive zinc ions build up around the zinc electrode. As copper in the copper sulfate solution is reduced, negative sulfate ions build up around the copper electrode. The buildup of charges stops any further reaction. To solve this problem, a salt bridge must be added to the system. A **salt bridge** is a pathway to allow the passage of ions from one side to another, as shown in **Figure 20.2**. A salt bridge consists of a tube containing a conducting solution of a soluble salt, such as KCl, held in place by an agar gel or other permeable plug. Ions can move through the plug, but the solutions in the two beakers cannot mix.

When the connecting metal wire and the salt bridge are in place, the spontaneous redox reaction begins. Electrons flow through the wire from the oxidation half-reaction to the reduction half-reaction, while positive and negative ions move through the salt bridge. A flow of charged particles is called an electric current. In **Figure 20.2**, the flow of electrons through the wire and the flow of ions through the salt bridge make up the electric current. The energy of the flowing electrons can be used to light a bulb, as shown in **Figure 20.2**.

Electrochemical cells The device shown in **Figure 20.2** is a type of electrochemical cell called a voltaic cell. An **electrochemical cell** is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction. A **voltaic cell** is a type of electrochemical cell that converts chemical energy to electrical energy by a spontaneous redox reaction. The voltaic cell, also shown in **Figure 20.3**, is named for Alessandro Volta (1745–1827), the Italian physicist who is credited with its invention in 1800.

■ **Figure 20.3** This replica of one of Alessandro Volta's first cells consists of discs of zinc and copper arranged in alternating layers and separated by cloth or cardboard soaked in an acidic solution. Electric current increased with the number of metal discs used.





■ **Figure 20.4** The roller coaster at the top of the track has high potential energy relative to track below because of the difference in height. Similarly, an electrochemical cell has potential energy to produce a current because there is a difference in the ability of the electrodes to move electrons from the anode to the cathode.

Chemistry of Voltaic Cells

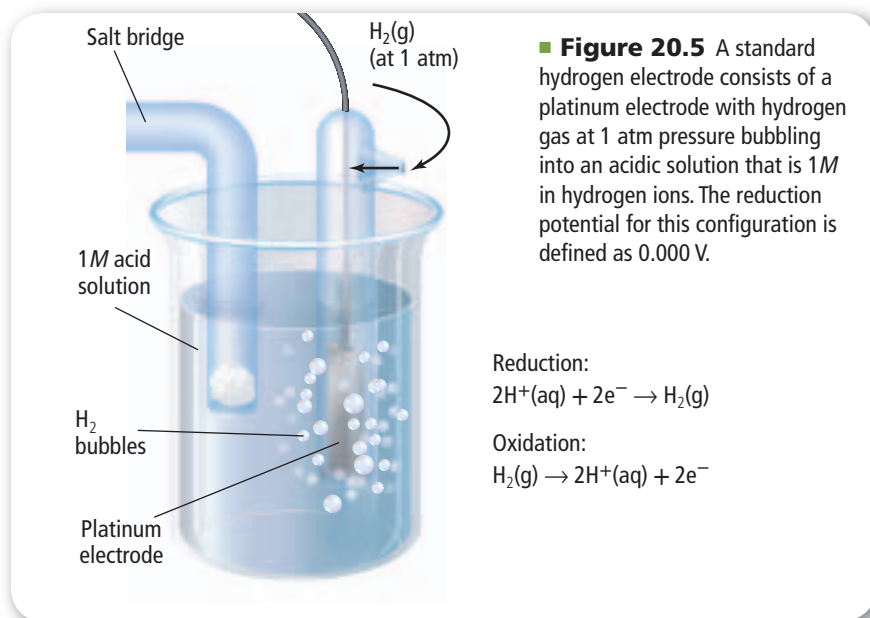
An electrochemical cell consists of two parts called **half-cells**, in which the separate oxidation and reduction reactions take place. Each half-cell contains an electrode and a solution containing ions. An electrode is an electrically conductive material, usually a metallic strip or graphite, that conducts electrons into and out of the solution in the half cell. In **Figure 20.2**, the beaker with the zinc electrode is where the oxidation half of the redox reaction takes place. The beaker with the copper electrode is where the reduction half of the reaction takes place. The reaction that takes place in each half-cell is called a half-cell reaction. The electrode where oxidation takes place is called the **anode**. The electrode where reduction takes place is called the **cathode**.



Reading Check Identify which of the beakers in **Figure 20.2** contains the anode.

Voltaic cells and energy Recall from Chapter 15 that an object's potential energy is due to its position or composition. In electrochemistry, electric potential energy is a measure of the amount of current that can be generated from a voltaic cell to do work. Electric charge can flow between two points only when a difference in electric potential energy exists between the two points. In an electrochemical cell, these two points are the two electrodes. Electrons generated at the anode, the site of oxidation, are thought to be pushed or driven toward the cathode by the electromotive force (EMF). This force is due to the difference in electric potential energy between the two electrodes and is referred to as the cell potential. A volt is a unit used to measure cell potential. The electric potential difference of a voltaic cell is an indication of the energy that is available to move electrons from the anode to the cathode.

Consider the analogy illustrated in **Figure 20.4**. The roller coaster is momentarily stationary at the top of the track. Then, it plummets from its high position because of the difference in gravitational potential energy (PE) between the top and bottom of the track. The kinetic energy (KE) attained by the roller coaster is determined by the difference in height (potential energy) between the top and bottom parts of the track. Similarly, the energy of the electrons flowing from the anode to the cathode in a voltaic cell is determined by the difference in electric potential energy between the two electrodes. In redox terms, the voltage of a cell is determined by comparing the difference in the tendency of the two electrode materials to accept electrons. The greater the difference, the greater the potential energy difference between the two electrodes and the larger the voltage of the cell will be.

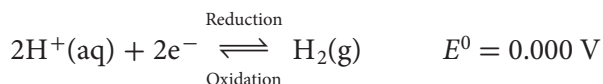


The force of gravity always causes a diver to fall downward to a lower energy state, never upward to a higher energy state. When a diver steps off a diving board, his or her spontaneous motion is always downward. Similarly, in the zinc-copper cell, under standard conditions, copper(II) ions at the cathode accept electrons more readily than the zinc ions at the anode. Thus, the redox reaction occurs spontaneously only when the electrons flow from the zinc to the copper.

Calculating Electrochemical Cell Potentials

Recall that gaining electrons is called reduction. Building on this fact, the tendency of a substance to gain electrons is its **reduction potential**. The reduction potential of an electrode cannot be determined directly because the reduction half-reaction must be coupled with an oxidation half-reaction. When two half-reactions are coupled, the voltage generated corresponds to the difference in potential between the half-reactions. The electrical potential difference between two points is expressed in volts (V).

The standard hydrogen electrode Long ago, chemists decided to measure the reduction potential of all electrodes against one electrode. They chose to use the standard hydrogen electrode. The **standard hydrogen electrode** consists of a small sheet of platinum immersed in a hydrochloric acid (HCl) solution that has a hydrogen-ion concentration of 1M. Hydrogen gas (H₂), at a pressure of 1 atm is bubbled in and the temperature is maintained at 25°C, as shown in **Figure 20.5**. The potential, also called the standard reduction potential (E^0), of this standard hydrogen electrode is defined as 0.000 V. This electrode can act as an oxidation half-reaction or a reduction half-reaction, depending on the half-cell to which it is connected. The two possible reactions at the hydrogen electrode are the following.



VOCABULARY

ACADEMIC VOCABULARY

Correspond

to be in agreement or to match
Her directions correspond with the map.

FOLDABLES

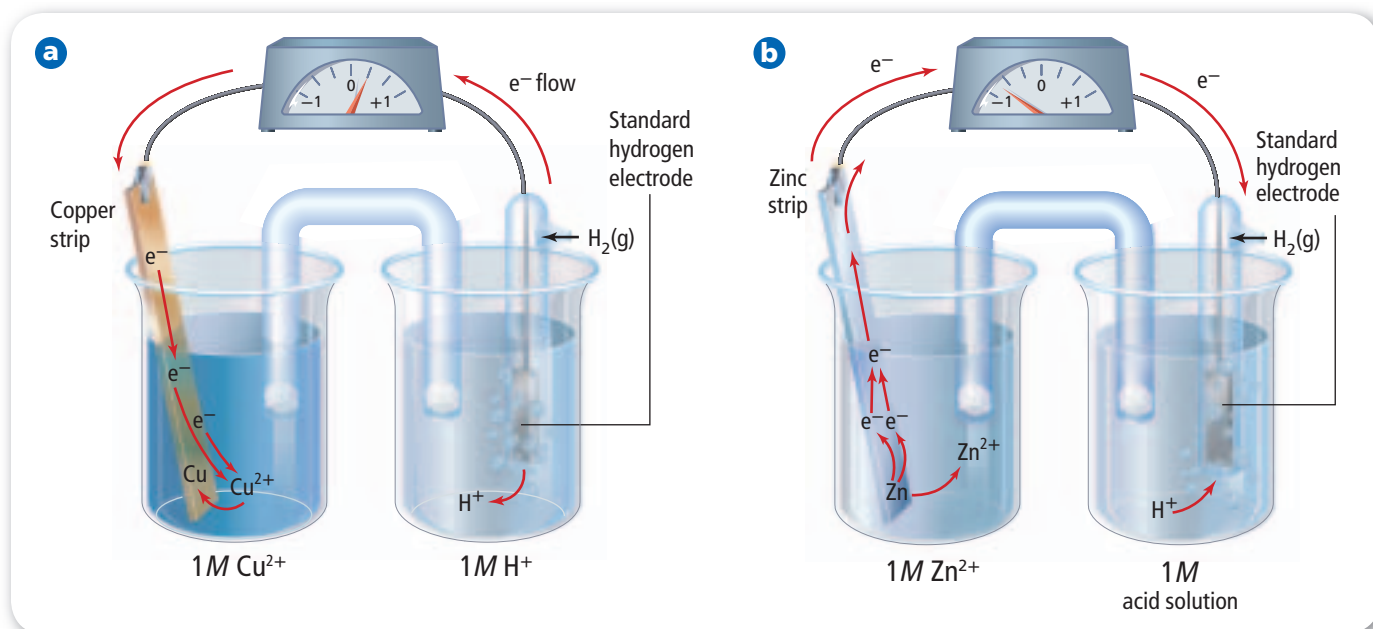
Incorporate information from this section into your Foldable.

Half-cell potentials Over the years, chemists have measured and recorded the standard reduction potentials of many different half-cells. **Table 20.1** lists some common half-cell reactions in order of increasing reduction potential. The values in the table were obtained by measuring the potential when each half-cell was connected to a standard hydrogen half-cell. All of the half-reactions in **Table 20.1** are written as reductions. However, in any voltaic cell, which always contains two half-reactions, the half-reaction with the lower reduction potential will proceed in the opposite direction and will be an oxidation reaction. In other words, the half-reaction that is more positive will proceed as a reduction and the half-reaction that is more negative will proceed as an oxidation.

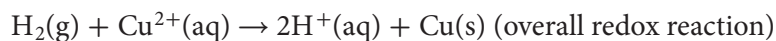
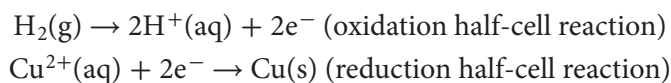
The electrode being measured must also be under standard conditions, that is, immersed in a 1M solution of its ions at 25°C and 1 atm. The superscript zero in the notation E^0 is a shorthand way of indicating “measured under standard conditions.”

Table 20.1 Standard Reduction Potentials

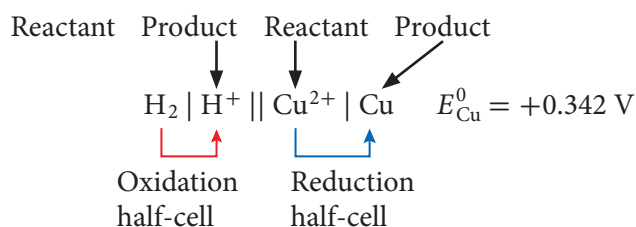
Half-Reaction	E^0 (V)	Half-Reaction	E^0 (V)
$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.0401	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	+0.153
$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.868	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.3419
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	+0.401
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.372	$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.5355
$\text{Be}^{2+} + 2\text{e}^- \rightarrow \text{Be}$	-1.847	$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.771
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.662	$\text{NO}_3^- + 2\text{H}^+ + \text{e}^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}$	+0.775
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.185	$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	+0.7973
$\text{Cr}^{2+} + 2\text{e}^- \rightarrow \text{Cr}$	-0.913	$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.7996
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.8277	$\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	+0.851
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.7618	$2\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	+0.920
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.744	$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	+0.957
$\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$	-0.47627	$\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-$	+1.066
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.447	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$	+1.18
$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.4030	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.229
$\text{PbI}_2 + 2\text{e}^- \rightarrow \text{Pb} + 2\text{I}^-$	-0.365	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	+1.35827
$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.3588	$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	+1.498
$\text{Co}^{2+} + 2\text{e}^- \rightarrow \text{Co}$	-0.28	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.507
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.257	$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	+1.692
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.1375	$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.776
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.1262	$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	+1.92
$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.037	$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	+2.010
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.0000	$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	+2.866



Determining electrochemical cell potentials You can use **Table 20.1** to calculate the electric potential of a voltaic cell consisting of a copper electrode and a zinc electrode under standard conditions. The first step is to determine the standard reduction potential for the copper half-cell (E_{Cu}^0). When the copper electrode is attached to a standard hydrogen electrode, as in **Figure 20.6a**, electrons flow from the hydrogen electrode to the copper electrode, and copper ions are reduced to copper metal. The E^0 , measured by a voltmeter, is $+0.342\text{ V}$. The positive voltage indicates that Cu^{2+} ions at the copper electrode accept electrons more readily than do H^+ ions at the standard hydrogen electrode. Therefore, oxidation takes place at the hydrogen electrode and reduction takes place at the copper electrode. The oxidation and reduction half-cell reactions and the overall reaction are



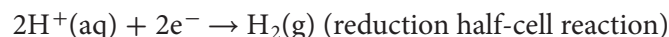
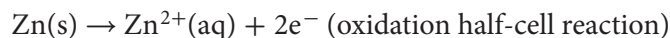
This reaction can be written in a form called cell notation.



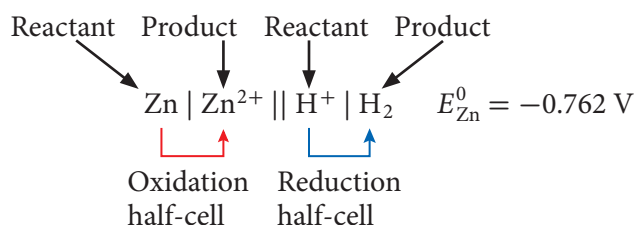
The two participants in the oxidation reaction are written first and in the order they appear in the oxidation half-reaction—reactant | product. They are followed by a double vertical line (||) representing the wire and salt bridge connecting the half-cells. Then, the two participants in the reduction reaction are written in the same reactant | product order. Note that for positive values of E^0 , it is customary to place a plus sign before the voltage.

Figure 20.6 a. When a $\text{Cu} | \text{Cu}^{2+}$ electrode is connected to the hydrogen electrode, electrons flow toward the copper strip and reduce Cu^{2+} ions to Cu atoms. The voltage of this reaction is $+0.342\text{ V}$. **b.** When a $\text{Zn} | \text{Zn}^{2+}$ electrode is connected to the hydrogen electrode, electrons flow away from the zinc strip and zinc atoms are oxidized to Zn^{2+} ions. The voltage of this reaction is -0.762 V .

The next step is to determine the standard reduction potential for the zinc half-cell (E_{Zn}^0). When the zinc electrode is measured against the standard hydrogen electrode under standard conditions, as in **Figure 20.6b**, electrons flow from the zinc electrode to the hydrogen electrode. The E^0 of the zinc half-cell, measured by a voltmeter, is -0.762 V. This means that the H^+ ions at the hydrogen electrode accept electrons more readily than do the zinc ions. Thus, the hydrogen ions have a higher reduction potential than the zinc ions. Recall that the hydrogen electrode is assigned a zero potential, so the reduction potential of the zinc electrode must have a negative value. The two half-cell reactions and the overall reaction are written as follows.



This reaction can be written in the following cell notation.



The final step in calculating electrochemical cell potential is to combine the copper and zinc half-cells as a voltaic cell. This means calculating the voltaic cell's standard potential using the following formula.

Formula for Cell Potential

$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

E_{cell}^0 represents the overall standard cell potential.

$E_{\text{reduction}}^0$ represents the standard half-cell potential for the reduction.

$E_{\text{oxidation}}^0$ represents the standard half-cell potential for the oxidation.

The standard potential of a cell is the standard potential of the reduction half-cell minus the standard potential of the oxidation half-cell.

Because reduction occurs at the copper electrode and oxidation occurs at the zinc electrode, the E^0 values are substituted as follows.

$$\begin{aligned} E_{\text{cell}}^0 &= E_{\text{Cu}^{2+}|\text{Cu}}^0 - E_{\text{Zn}^{2+}|\text{Zn}}^0 \\ &= +0.342 \text{ V} - (-0.762 \text{ V}) \\ &= +1.104 \text{ V} \end{aligned}$$

Notice that the negative sign in the formula automatically changes the sign of the oxidation half-reaction, so you do not reverse the sign of the standard reduction potentials listed in **Table 20.1** when they are used for the oxidation half-reaction.

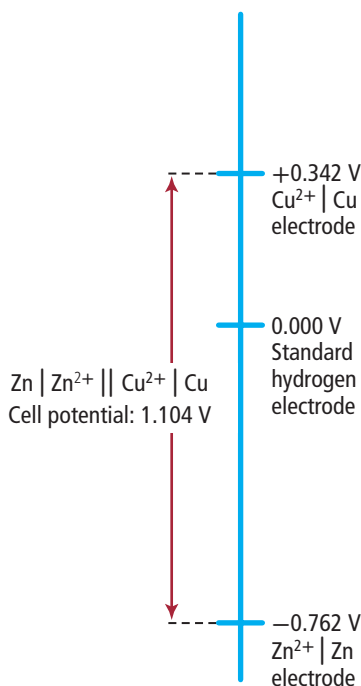
The graph in **Figure 20.7** shows how the zinc half-cell with the lower reduction potential and the copper half-cell with the higher reduction potential are related.

■ **Figure 20.7** This simple graph illustrates how the overall cell potential is derived from the difference in reduction potential of two electrodes. Compare it to a number line. The potential difference between the zinc and copper electrodes is $+1.104$ V.



Graph Check

Identify the metal, copper or zinc, that is easier to oxidize than hydrogen.

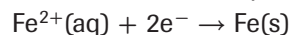
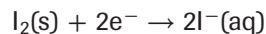


EXAMPLE Problem 20.1

Math Handbook

Solving Algebraic
Equations
pages 954–955

Calculate a Cell Potential The following reduction half-reactions represent the half-cells of a voltaic cell.



Determine the overall cell reaction and the standard cell potential. Describe the cell using cell notation.

1 Analyze the Problem

You are given the half-cell equations and can find standard reduction potentials in **Table 20.1**. The half-reaction with the lower reduction potential will be an oxidation. With this information, you can write the overall cell reaction, calculate the standard cell potential, and describe the cell in cell notation.

Known

Standard reduction potentials for the half-cells

$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

Unknown

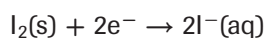
overall cell reaction = ?

$$E_{\text{cell}}^0 = ?$$

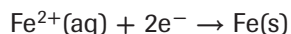
cell notation = ?

2 Solve for the Unknown

Find the standard reduction potentials of each half-reaction in **Table 20.1**.

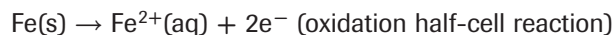


$$E_{\text{I}_2|\text{I}^-}^0 = +0.536 \text{ V}$$



$$E_{\text{Fe}|\text{Fe}^{2+}}^0 = -0.447 \text{ V}$$

The reduction of iodine has the higher reduction potential, so this half-reaction proceeds in the forward direction as a reduction. The iron half-reaction proceeds in the reverse direction as an oxidation.



Rewrite the iron half-reaction in the correct direction.



Add the two equations.

The overall cell reaction is $\text{I}_2(\text{s}) + \text{Fe}(\text{s}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq})$.

Calculate the standard cell potential.

$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

State the formula for cell potential.

$$E_{\text{cell}}^0 = E_{\text{I}_2|\text{I}^-}^0 - E_{\text{Fe}|\text{Fe}^{2+}}^0$$

Substitute $E_{\text{I}_2|\text{I}^-}^0$ and $E_{\text{Fe}|\text{Fe}^{2+}}^0$ in the generic equation.

$$E_{\text{cell}}^0 = +0.536 \text{ V} - (-0.447 \text{ V})$$

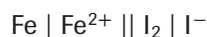
Substitute $E_{\text{I}_2|\text{I}^-}^0 = +0.536 \text{ V}$ and $E_{\text{Fe}|\text{Fe}^{2+}}^0 = -0.447 \text{ V}$.

$$E_{\text{cell}}^0 = +0.983 \text{ V}$$

Describe the cell using cell notation.



First, write the oxidation half-reaction using cell notation: reactant then product.



Next, write the reduction half-reaction to the right. Separate the half-cells by a double vertical line.

Cell notation: $\text{Fe} | \text{Fe}^{2+} || \text{I}_2 | \text{I}^-$

3 Evaluate the Answer

The calculated potential is reasonable given the potentials of the half-cells. E^0 is reported to the correct number of significant figures.



Personal Tutor For an online tutorial on calculating potentials, visit glencoe.com.

For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation. Refer to Chapter 19 to review writing and balancing redox equations.

- $\text{Pt}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Pt}(\text{s})$ and $\text{Sn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Sn}(\text{s})$
- $\text{Co}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Co}(\text{s})$ and $\text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-} \rightarrow \text{Cr}(\text{s})$
- $\text{Hg}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Hg}(\text{l})$ and $\text{Cr}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cr}(\text{s})$
- Challenge** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.
 $\text{NO}_3^{-} + 4\text{H}^{+} + 3\text{e}^{-} \rightarrow \text{NO} + 2\text{H}_2\text{O}$
 $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^{-} \rightarrow 4\text{OH}^{-}$

Using Standard Reduction Potentials

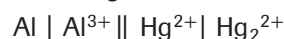
The Example Problems showed you how to use the data from **Table 20.1** to calculate the standard potential (voltage) of voltaic cells. Another important use of standard reduction potentials is to determine if a proposed reaction under standard conditions will be spontaneous. How can standard reduction potentials indicate spontaneity? Electrons in a voltaic cell always flow from the half-cell with the lower standard reduction potential to the half-cell with the higher reduction potential, giving a positive cell voltage. To predict whether any proposed redox reaction will occur spontaneously, simply write the process in the form of half-reactions and look up the reduction potential of each. Use the values to calculate the potential of a voltaic cell operating with these two half-cell reactions. If the calculated potential is positive, the reaction is spontaneous. If the value is negative, the reaction is not spontaneous. However, the reverse of a nonspontaneous reaction will occur because it will have a positive cell voltage, which means that the reverse reaction is spontaneous.



Reading Check Identify the sign of the potential of a redox reaction that occurs spontaneously.

Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use **Table 20.1** to help you determine the correct half-reactions.

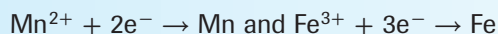
- $\text{Sn}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + \text{Cu}(\text{s})$
- $\text{Mg}(\text{s}) + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Pb}(\text{s}) + \text{Mg}^{2+}(\text{aq})$
- $2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) + 10\text{Hg}^{2+}(\text{aq}) \rightarrow 2\text{MnO}_4^{-}(\text{aq}) + 16\text{H}^{+}(\text{aq}) + 5\text{Hg}_2^{2+}(\text{aq})$
- $2\text{SO}_4^{2-}(\text{aq}) + \text{Co}^{2+}(\text{aq}) \rightarrow \text{Co}(\text{s}) + \text{S}_2\text{O}_8^{2-}(\text{aq})$
- Challenge** Using **Table 20.1**, write the equation and determine the cell voltage (E°) for the following cell. Is the reaction spontaneous?



Problem-Solving Strategy

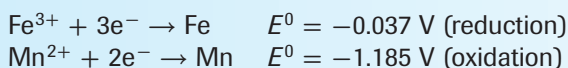
Determining Cell Potentials

The five steps that follow summarize the procedure for calculating the potential of a voltaic cell in which a spontaneous redox reaction occurs. Suppose you must write the equation for and calculate the potential of a cell made up of these half-reactions:

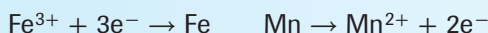


A table of reduction potentials, such as **Table 20.1**, is all that is required.

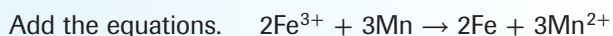
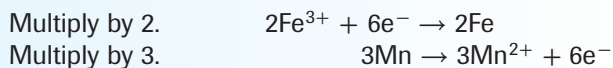
1. Find the two half-reactions on **Table 20.1**.
2. Compare the two half-cell potentials. The half-cell with the higher reduction potential is the cell in which reduction will occur. Oxidation will occur in the half-cell with the lower reduction potential.



3. Write the equation for the reduction as it is in **Table 20.1**. Write the equation for the oxidation in the opposite direction.



4. Balance the electrons in the two half-cell equations by multiplying each by a factor. Add the equations.



5. Equalizing the electrons lost and gained does not affect the E° for the overall reaction. Use the formula: $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{reduction}} - E^{\circ}_{\text{oxidation}}$ to obtain the cell potential.

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{Fe}^{3+}|\text{Fe}} - E^{\circ}_{\text{Mn}^{2+}|\text{Mn}} = -0.037 \text{ V} - (-1.185 \text{ V}) \\ &= +1.148 \text{ V} \end{aligned}$$

Apply the Strategy

Determine E°_{cell} for the spontaneous redox reaction that occurs between magnesium and nickel.

Section 20.1 Assessment

Section Summary

- In a voltaic cell, oxidation and reduction take place at electrodes separated from each other.
- The standard potential of a half-cell reaction is its voltage when paired with a standard hydrogen electrode under standard conditions.
- The reduction potential of a half-cell is negative if it undergoes oxidation when connected to a standard hydrogen electrode. The reduction potential of a half-cell is positive if it undergoes reduction when connected to a standard hydrogen electrode.
- The standard potential of a voltaic cell is the difference between the standard reduction potentials of the half-cell reactions.

10. **MAIN Idea** Describe the conditions under which a redox reaction causes an electric current to flow through a wire.
11. **Identify** the components of a voltaic cell. Explain the role of each component in the operation of the cell.
12. **Write** the balanced equation for the spontaneous cell reaction that occurs in a cell with these reduction half-reactions.
 - a. $\text{Ag}^{+}(\text{aq}) + \text{e}^{-} \rightarrow \text{Ag}(\text{s})$ and $\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Ni}(\text{s})$
 - b. $\text{Mg}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Mg}(\text{s})$ and $2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{H}_2(\text{g})$
 - c. $\text{Sn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Sn}(\text{s})$ and $\text{Fe}^{3+}(\text{aq}) + 3\text{e}^{-} \rightarrow \text{Fe}(\text{s})$
 - d. $\text{PbI}_2(\text{s}) + 2\text{e}^{-} \rightarrow \text{Pb}(\text{s}) + 2\text{I}^{-}(\text{aq})$ and $\text{Pt}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Pt}(\text{s})$
13. **Determine** the standard potential for electrochemical cells in which each equation represents the overall cell reaction. Identify the reactions as spontaneous or nonspontaneous as written.
 - a. $2\text{Al}^{3+}(\text{aq}) + 3\text{Cu}(\text{s}) \rightarrow 3\text{Cu}^{2+}(\text{aq}) + 2\text{Al}(\text{s})$
 - b. $\text{Hg}^{2+}(\text{aq}) + 2\text{Cu}^{+}(\text{aq}) \rightarrow 2\text{Cu}^{2+}(\text{aq}) + \text{Hg}(\text{l})$
 - c. $\text{Cd}(\text{s}) + 2\text{NO}_3^{-}(\text{aq}) + 4\text{H}^{+}(\text{aq}) \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{NO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$
14. **Design** a concept map for Section 20.1, starting with the term *electrochemical cell*. Incorporate all the new vocabulary terms in your map.

Section 20.2

Objectives

- Describe the structure, composition, and operation of the typical carbon-zinc dry-cell battery.
- Distinguish between primary and secondary batteries, and give two examples of each type.
- Explain the structure and operation of the hydrogen-oxygen fuel cell.
- Describe the process of corrosion of iron and methods to prevent corrosion.

Review Vocabulary

reversible reaction: a reaction that can take place in both the forward and reverse directions

New Vocabulary

battery
dry cell
primary battery
secondary battery
fuel cell
corrosion
galvanization

Batteries

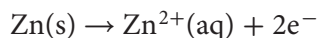
MAIN Idea Batteries are voltaic cells that use spontaneous reactions to provide energy for a variety of purposes.

Real-World Reading Link Take a moment to list some of the places where you know batteries are used. Your list might include flashlights, cars, cell phones, radios, calculators, watches, and toys, among many others. Are the batteries in these devices all the same?

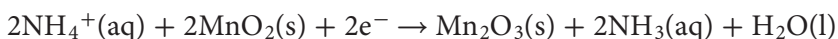
Dry Cells

Some of the spontaneous cell reactions you have been reading about provide the energy of the batteries that you use every day. A **battery** is one or more voltaic cells in a single package that generates electric current. From the time of its invention in the 1860s until recently, the most commonly used voltaic cell was the zinc-carbon dry cell, shown in **Figure 20.8**.

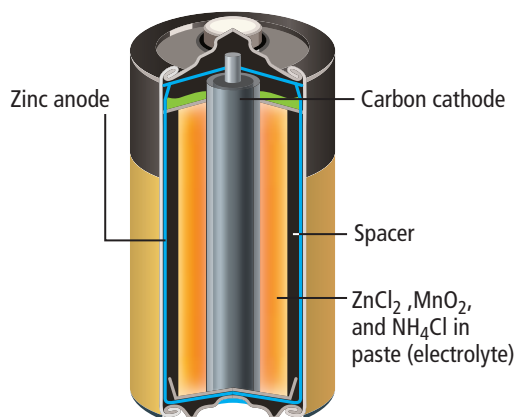
Zinc-carbon dry cells A **dry cell** is an electrochemical cell in which the electrolyte is a moist paste. The paste in a zinc-carbon dry cell consists of zinc chloride, manganese(IV) oxide, ammonium chloride, and a small amount of water inside a zinc case. The zinc shell is the cell's anode, where the oxidation of zinc metal occurs according to the following equation.



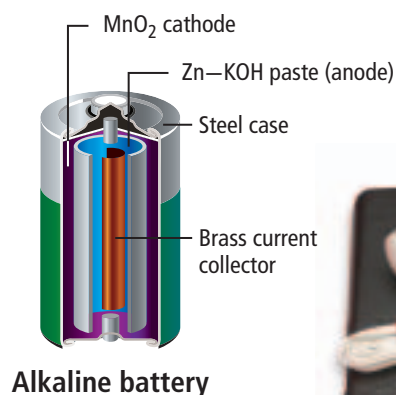
A carbon (graphite) rod in the center of the dry cell serves as the cathode, but the reduction half-cell reaction takes place in the paste. The carbon rod in this type of dry cell is called an inactive cathode because it is made of a material that does not participate in the redox reaction. However, the inactive electrode has the important purpose of conducting electrons. The reduction half-cell reaction for this dry cell is as follows.



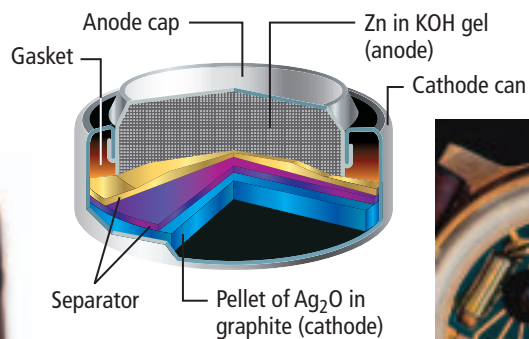
■ **Figure 20.8** The so-called dry cell contains a moist paste in which the cathode half-reaction takes place. In the zinc-carbon dry cell, the zinc case acts as the anode.



Zinc-carbon dry cell



The alkaline battery uses powdered zinc in an alkaline paste as the anode.



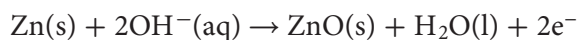
Although it looks different, the silver battery is much like the alkaline battery. The silver battery uses silver oxide (Ag_2O) in the cathode instead of manganese(IV) oxide (MnO_2).



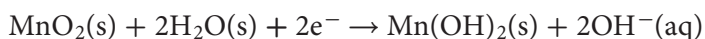
■ **Figure 20.9** Alkaline batteries are more efficient than zinc-carbon dry cells and are useful when smaller batteries are needed. Silver button batteries are even smaller, making them well suited to devices such as watches.

In the zinc-carbon dry cell, a spacer made of a porous material and damp from the liquid in the paste separates the paste from the zinc anode. The spacer acts as a salt bridge to allow the transfer of ions, much like the model voltaic cell you studied in Section 20.1. The zinc-carbon dry cell produces a voltage of 1.5 V until the reduction product, ammonia, comes out of its aqueous solution as a gas. At that point, the voltage drops to a level that makes the battery useless.

Alkaline batteries A more efficient alkaline dry cell, shown in **Figure 20.9**, is replacing the standard zinc-carbon dry cell in many applications. In the alkaline cell, the zinc is in a powdered form, which provides more surface area for reaction. The zinc is mixed in a paste with potassium hydroxide, a strong base, and the paste is contained in a steel case. The cathode mixture is manganese(IV) oxide, also mixed with potassium hydroxide. The anode half-cell reaction is as follows.

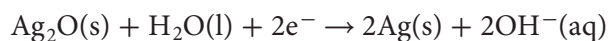



The cathode half-cell reaction is as follows.

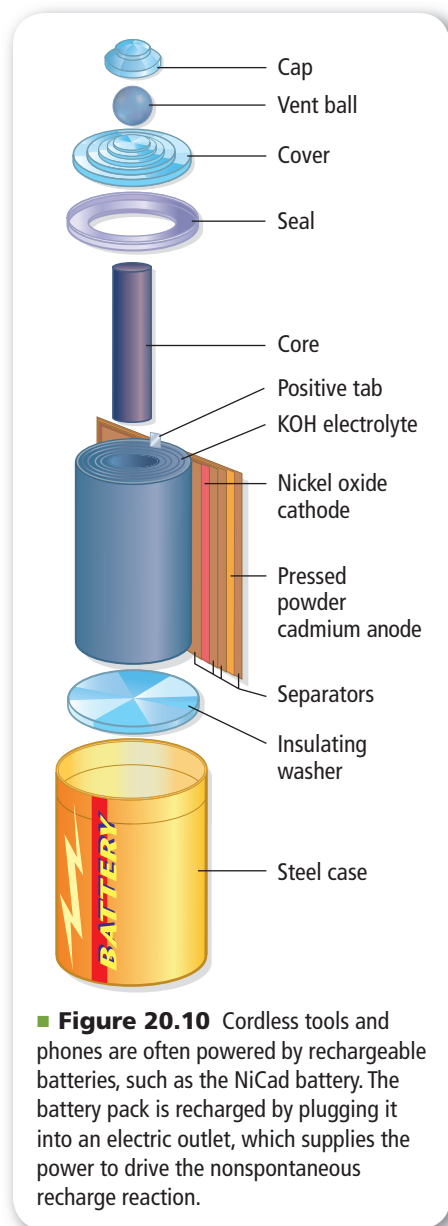


Alkaline batteries do not need the carbon rod cathode, so they can be made smaller and are more useful in small devices.

Silver batteries The silver battery shown in **Figure 20.9** is even smaller and is used to power devices such as hearing aids, watches, and cameras. The silver battery uses the same anode half-reaction as the alkaline battery, with the following cathode half-reaction.



 **Reading Check** Identify the half-reaction that occurs in both alkaline and silver batteries.

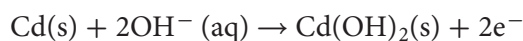


■ **Figure 20.10** Cordless tools and phones are often powered by rechargeable batteries, such as the NiCad battery. The battery pack is recharged by plugging it into an electric outlet, which supplies the power to drive the nonspontaneous recharge reaction.

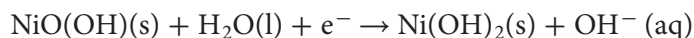
Primary and secondary batteries Batteries are divided into two types, depending on their chemical processes. The zinc-carbon, alkaline-zinc, and silver cells are classified as primary batteries.

Primary batteries produce electric energy by means of redox reactions that are not easily reversed. These cells deliver current until the reactants are gone, and then the battery must be discarded. Other batteries, called **secondary batteries**, depend on reversible redox reactions, so they are rechargeable. A car battery and the battery in a laptop computer are examples of secondary batteries, which are sometimes called storage batteries.

The storage batteries that power devices such as cordless drills and screwdrivers, shavers, and camcorders are usually nickel-cadmium rechargeable batteries, sometimes called NiCad batteries, as shown in **Figure 20.10**. For maximum efficiency, the anode and cathode are long, thin ribbons of material separated by a layer through which ions can pass. The ribbons are wound into a tight coil and packaged in a steel case. The anode reaction that occurs when the battery is used to generate electric current is the oxidation of cadmium in the presence of a base.



The cathode reaction is the reduction of nickel from the +3 to the +2 oxidation state.

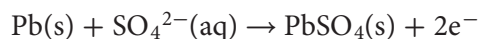


When the battery is recharged, these reactions are reversed.

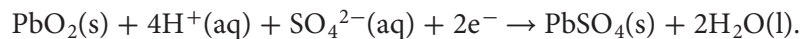
Lead-Acid Storage Battery

Another common storage battery is the lead-acid battery used in automobiles. Most auto batteries contain six cells that generate about 2 V each for a total output of 12 V. The anode of each cell consists of two or more grids of porous lead, and the cathode consists of lead grids filled with lead(IV) oxide. This type of battery should probably be called a lead-lead(IV) oxide battery, but the term *lead-acid* is commonly used because the battery's electrolyte is a solution of sulfuric acid. The lead-acid battery is not a dry cell.

The following equation represents the oxidation half-cell reaction at the anode where lead is oxidized from the zero oxidation state to the +2 oxidation state in PbSO_4 .



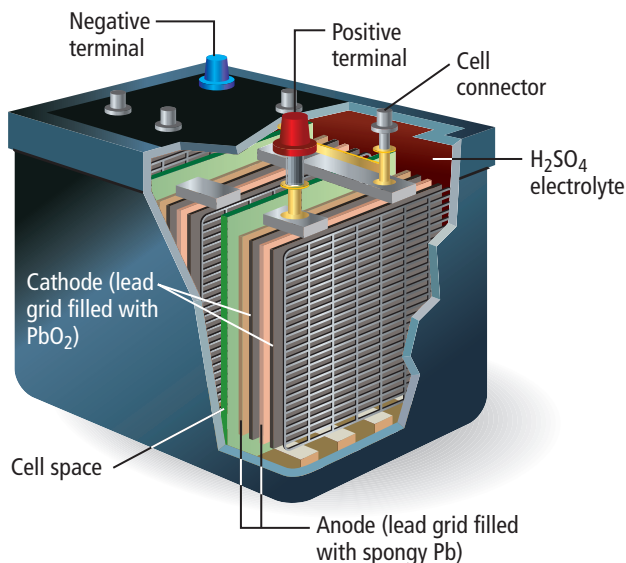
The reduction of lead from the +4 to the +2 oxidation state takes place at the cathode. The half-cell reaction for the cathode is



The overall reaction is



By looking at the half-cell reactions, you can see that lead(II) sulfate (PbSO_4) is the reaction product in both oxidation and reduction. Also, Pb , PbO_2 , and PbSO_4 are solid substances, so they stay in place where they are formed. Thus, whether the battery is discharging or charging, the reactants are available where they are needed.




Lead-acid batteries contain lead plates and lead(IV) oxide plates. The electrolyte is a solution of sulfuric acid. When the battery is in use, the sulfuric acid is depleted and the electrolyte becomes less dense.



Low electrolyte levels can result in a dead battery. Jumper cables conduct current from a car with a good battery to start a car with a dead one.

Sulfuric acid serves as the electrolyte in the battery, but, as the overall cell equation shows, it is depleted as the battery generates electric current. What happens when the battery is recharging? In this case, the reactions reverse, forming lead and lead(IV) oxide and releasing sulfuric acid, shown as $4\text{H}^+(\text{aq}) + 2\text{SO}_4^{2-}(\text{aq})$ in the equation.

The lead-storage battery shown in **Figure 20.11** is a good choice for motor vehicles because it provides a large initial supply of energy to start the engine, has a long shelf life, and is reliable at low temperatures.

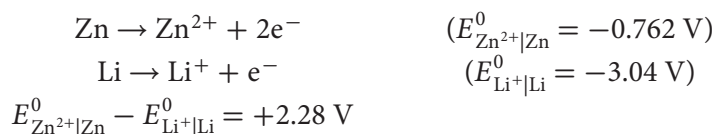
 **Reading Check Identify** the species that is oxidized and the species that is reduced when the lead-acid battery is charging.

Lithium Batteries

Although lead-acid batteries are reliable and suitable for many applications, engineers have been working to develop batteries with less mass and higher capacity to power devices from wristwatches to electric cars. For applications in which a battery is the key component and must provide a significant amount of power, such as for the operation of an electric car, lead-acid batteries are too heavy to be feasible.

The solution is to develop lightweight batteries that store a large amount of energy for their size. Engineers have focused their attention on the element lithium for two reasons: lithium is the lightest known metal and has the lowest standard reduction potential of the metallic elements, -3.04 V, as shown in **Table 20.1**. A battery that oxidizes lithium at the anode can generate almost 2.3 V more than a similar battery in which zinc is oxidized.

Compare the zinc and lithium oxidation half-reactions and their standard reduction potentials.



■ **Figure 20.11** The lead-acid battery used in automobiles discharges when it starts the car and charges when the engine is running.

VOCABULARY

WORD ORIGIN

Capacity

capac-, capax, from Latin, meaning *containing or capable of holding a great deal*



Lithium batteries often deliver either 3 V or 9 V and come in many sizes to fit different devices.



Lithium battery packs power this experimental car to a maximum speed of 113 km/h. The car has a range of over 320 km.

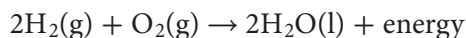
■ **Figure 20.12** The light weight, long life, and high potential of a lithium battery make it an excellent choice for a variety of purposes.

Lithium batteries can be either primary or secondary batteries, depending on which reduction reactions are coupled to the oxidation of lithium. For example, some lithium batteries use the same cathode reaction as zinc-carbon dry cells, the reduction of manganese(IV) oxide (MnO_2) to manganese(III) oxide (Mn_2O_3). These batteries produce an electric current of about 3 V compared to 1.5 V for zinc-carbon cells. Lithium batteries last much longer than other kinds of batteries. As a result, they are often used in watches, computers, and cameras to maintain time, date, memory, and personal settings—even when the device is turned off. **Figure 20.12** shows a range of available lithium batteries and a developing application.

✓ **Reading Check** List three advantages of lithium batteries.

Fuel Cells

When hydrogen burns in air, it does so explosively, with the evolution of light and heat.



Can this reaction occur under controlled conditions inside a cell?

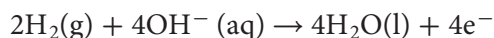
Connection to Physics A **fuel cell** is a voltaic cell in which the oxidation of a fuel is used to produce electric energy. Fuel cells differ from other batteries because they are provided with a continual supply of fuel from an external source. Many people think the fuel cell is a modern invention, but the first one was demonstrated in 1839 by William Grove (1811–1896), a British electrochemist. He called his cell a “gas battery.” It was not until the 1950s, when scientists began working in earnest on the space program, that efficient, practical fuel cells were developed. If astronauts were to fly a space shuttle, supplies of water were needed to support their lives on board and a reliable source of electricity was needed to power the shuttle’s many systems. Both of these primary needs were met with the development of the hydrogen fuel cell that controls the oxidation of hydrogen and provides both electricity and water. The cell produces no by-products to require disposal or storage on a space journey.

Real-World Chemistry Fuel Cells

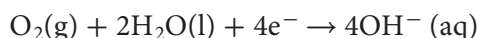


Reducing pollution One of the largest sources of air pollution in many cities is vehicles. In some European cities, experimental buses powered by hydrogen fuel cells are making a difference. Exhaust from these buses contain no carbon dioxide and no oxides of nitrogen or sulfur. Pure water is their only product.

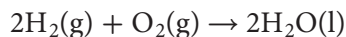
How a fuel cell works As in other voltaic cells, a fuel cell has an anode and a cathode and requires an electrolyte so that ions can migrate between electrodes. A common electrolyte in a fuel cell is an alkaline solution of potassium hydroxide. Each electrode is a hollow chamber of porous carbon walls that allows contact between the inner chamber and the electrolyte surrounding it. The following oxidation half-reaction takes place at the anode.



The reaction uses the hydroxide ions that are abundant in the alkaline electrolyte and releases electrons to the anode. Electrons from the oxidation of hydrogen flow through the external circuit to the cathode where the following reduction half-reaction takes place.




The electrons reduce oxygen in the presence of water to form four hydroxide ions, which replenish the hydroxide ions used up at the anode. When the two half-reactions are combined, the equation is the same as the equation for the burning of hydrogen in oxygen.



Because the fuel for the cell is provided from an outside source, fuel cells never run down as batteries do. They keep producing electricity as long as fuel is available.

Some fuel cells use fuels other than hydrogen. For example, methane replaces hydrogen in some cells, but has the disadvantage of producing carbon dioxide as an exhaust gas. Fuel cells such as the one shown in **Figure 20.13** use a plastic sheet called a proton-exchange membrane (PEM), which eliminates the need for a liquid electrolyte.

 **Reading Check Compare** fuel cells with other voltaic cells to find an important way in which they are different.

CAREERS IN CHEMISTRY

Alternative Energy Specialist

If you like to invent new things and make them work, you might be interested in helping to develop new sources of energy for this increasingly energy-dependent world. Technologies include wind and solar power, geothermal energy, harnessing the energy of tides, utilizing the temperature gradients in bodies of water, and many more. For more information on chemistry careers, visit glencoe.com.

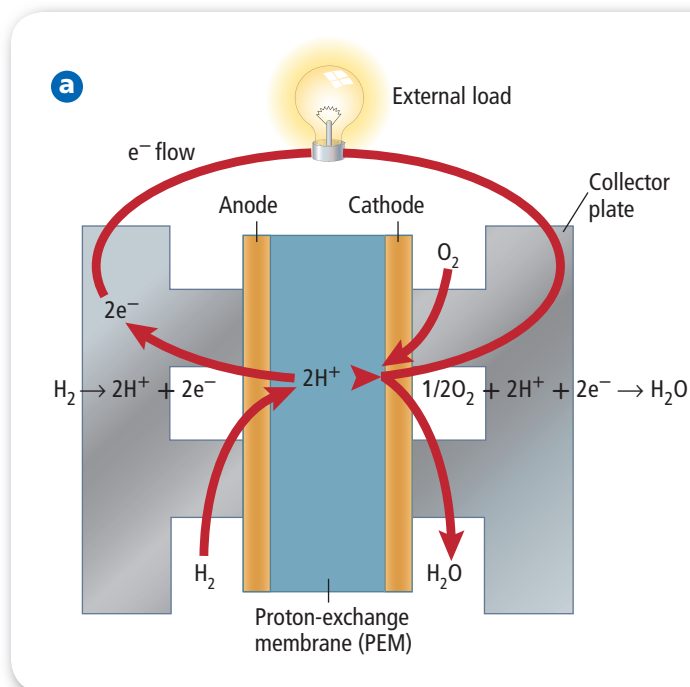


Figure 20.13 a. In this fuel cell, hydrogen is the fuel. The half-reactions are separated by a proton-exchange membrane so that the electrons lost in oxidation flow through an external circuit to reach the site of reduction. As electrons travel through the external circuit, they can do useful work, such as running electric motors. The by-product of this redox reaction is water. **b.** A "stack" of PEM-type cells can generate enough energy to power an electric car.

DATA ANALYSIS LAB

Based on Real Data*

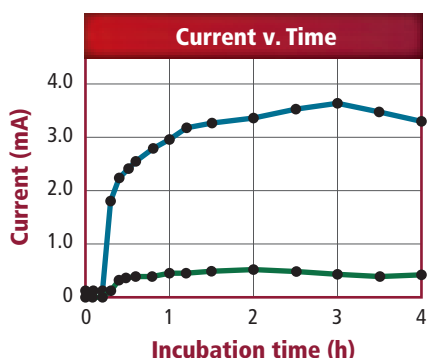
Interpret Graphs

How can you get electric current from microbes?

Scientists have studied the use of microbes as biofuel cells. A biofuel cell directly converts microbial metabolic energy into electric current. An electron mediator facilitates transfer of electrons to an electrode. An electron mediator is a compound that taps into the electron transport chain of cells and steals the electrons that are produced.

Data and Observations

The graph shows the current produced in a biofuel cell with (blue line) and without (green line) the use of an electron mediator.



*Data obtained from: Hyun Park, Doo and J. Gregory Zeikus. April, 2000. Electricity Generation in Microbial Fuel Cells Using Neutral Red as an Electronophore. *Applied and Environmental Microbiology* 66, No. 4:1292–1297.

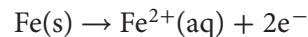
Think Critically

- Infer** the approximate time when the electron mediator was introduced.
- Determine** Did the introduction of the electron mediator make a difference in the current production? Explain your answer.
- Analyze** What is the highest current obtained by the cell?

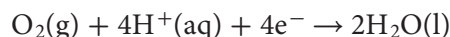
Corrosion

In this chapter, you have examined the spontaneous redox reactions in voltaic cells. Spontaneous redox reactions also occur in nature. An example is the corrosion of iron, usually called rusting. **Corrosion** is the loss of metal resulting from an oxidation-reduction reaction of the metal with substances in the environment. Although rusting is usually thought of as a reaction between iron and oxygen, it is more complex. Both water and oxygen must be present for rusting to occur. For this reason, an iron object, such as the one shown in **Figure 20.14**, that has been left exposed to air and moisture is especially susceptible to rust. The portion that is in contact with the moist ground rusted first.

Rusting usually begins where there is a pit or a small break in the surface of the iron. This region becomes the anode of the cell as iron atoms begin to lose electrons, as illustrated in **Figure 20.15**.



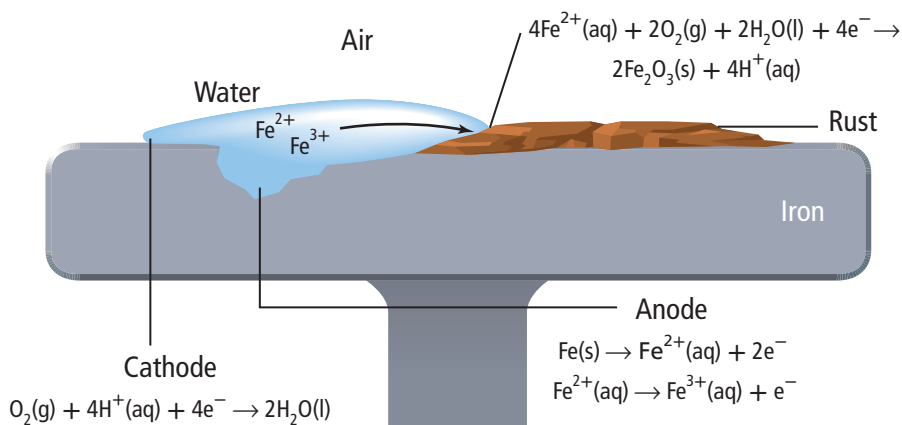
The iron(II) ions become part of the water solution, while the electrons move through the iron to the cathode region. In effect, the piece of iron becomes the external circuit as well as the anode. The cathode is usually located at the edge of the water drop where water, iron, and air come in contact. Here, the electrons reduce oxygen from the air in the following half-reaction.



The supply of H^+ ions is probably furnished by carbonic acid formed when CO_2 from air dissolves in water.

■ **Figure 20.14** Left unattended in the presence of air and moisture, this iron barrel is slowly being oxidized to rust (Fe_2O_3).

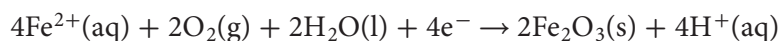




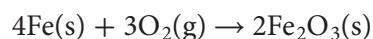
■ **Figure 20.15** Corrosion occurs when air, water, and iron set up a voltaic cell similar to the conditions shown at the surface of this iron I-beam. An I-beam is a large piece of iron shaped like the capital letter I and used in the construction of large buildings.

Name the two species that are oxidized at the anode.

Next, the Fe^{2+} ions in solution are oxidized to Fe^{3+} ions by reacting with oxygen dissolved in the water. The Fe^{3+} ions combine with oxygen to form insoluble Fe_2O_3 , rust.



Combining the three equations yields the overall cell reaction for the corrosion of iron.



Rusting is a slow process because water droplets have few ions and therefore, are not good electrolytes. However, if the water contains abundant ions, as in seawater or in regions where roads are salted in the winter, corrosion occurs much faster because the solutions are excellent electrolytes.

Preventing corrosion Corrosion of cars, bridges, ships, the structures of buildings, and other metallic objects causes more than \$100 billion in damage a year in the United States. For this reason, several means to minimize corrosion have been devised. One example is to apply a coat of paint to seal out both air and moisture, but because paint deteriorates over time, objects such as the bridge shown in **Figure 20.16** must be repainted often.



■ **Figure 20.16** Because corrosion can cause considerable damage, it is important to find ways to prevent rust and deterioration. Paint or another protective coating is one way to protect steel structures from corrosion.

Observe Corrosion

Which metal will corrode?

Procedure



1. Read and complete the lab safety form.
2. Use **sandpaper** to buff the surfaces of four **iron nails**. Wrap two nails with **magnesium ribbon** and two nails with **copper**. Wrap the metals tightly so that the nails do not slip out.
3. Place each of the nails in a separate **beaker**. Add **distilled water** to one of the beakers containing a copper-wrapped nail and one of the beakers containing a magnesium-wrapped nail. Add enough distilled water to just cover the wrapped nails. Add **salt water** to **two additional beakers**. Record your observations of the nails in each beaker.



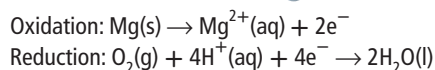
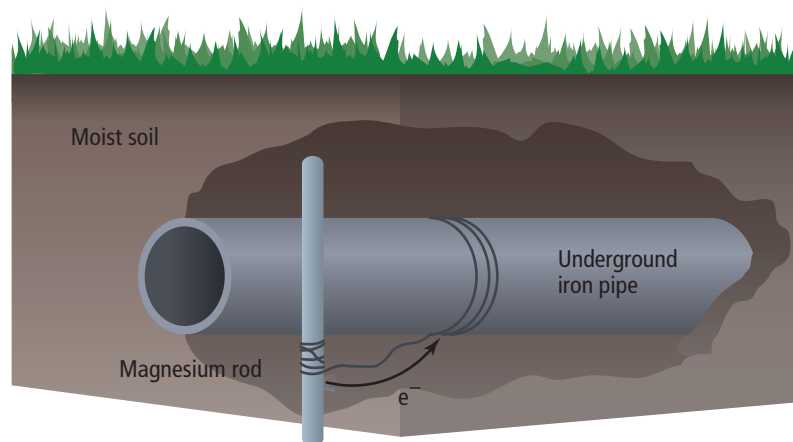
4. Let the beakers stand overnight in the warmest place available. Examine the nails and solutions the next day, and record your observations.

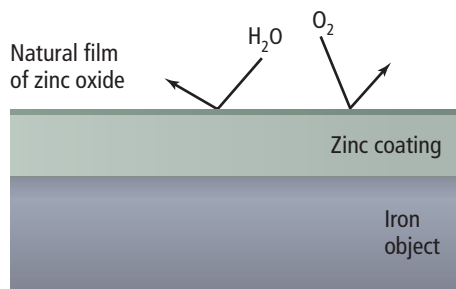
Analysis

1. **Describe** the difference between copper-wrapped nails in the distilled water and the salt water after they have been standing overnight.
2. **Describe** the difference between the magnesium-wrapped nails in the distilled water and in the salt water.
3. **Explain** the difference between a copper-wrapped nail and a magnesium-wrapped nail.

The steel hulls of ships are constantly in contact with salt water, so the prevention of corrosion is vital. Although the hull can be painted, another method is used to minimize corrosion. Blocks of metals, such as magnesium, aluminum, or titanium, are placed in contact with the steel hull. These blocks oxidize more easily than iron and become the anode of the corrosion cell. They are called sacrificial anodes because they are corroded, while the iron in the hull is spared. A similar technique is used to protect underground iron pipes. Magnesium bars are attached to the pipe by wires, and these bars corrode instead of the pipe, as shown in **Figure 20.17**.

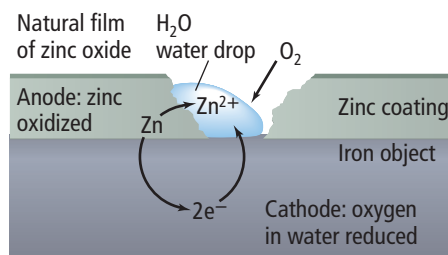
■ **Figure 20.17** Sacrificial anodes of magnesium or other active metals are used to prevent corrosion. A magnesium rod attached to an underground iron pipe helps prevent corrosion by being oxidized itself.





Galvanized object with zinc coating intact

The zinc coating seals the iron from air and water by forming a barrier of zinc oxide that repels water and oxygen.



Galvanized object with zinc coating broken

If the zinc coating breaks, the zinc acts as a sacrificial anode. The zinc coating, rather than the iron object, is oxidized.

■ **Figure 20.18** Galvanization helps prevent corrosion in two ways.

Another approach to preventing corrosion is to coat iron with another metal that is more resistant to corrosion. In the **galvanization** process, iron is coated with a layer of zinc by either dipping the object into molten zinc or by electroplating the zinc onto it. Although zinc is more readily oxidized than iron, it is one of the self-protecting metals, a group that also includes aluminum and chromium. When exposed to air, these metals oxidize at the surface, creating a thin metal-oxide coating that seals the metal from further oxidation.

Galvanization protects iron in two ways. As long as the zinc layer is intact, water and oxygen cannot reach the iron's surface. Inevitably, the zinc coating cracks. When this happens, zinc protects iron from rapid corrosion by becoming the anode of the voltaic cell set up when water and oxygen contact iron and zinc at the same time. **Figure 20.18** illustrates how these two forms of corrosion protection work.

Section 20.2 Assessment

Section Summary

- Primary batteries can be used only once; secondary batteries can be recharged.
- When a battery is recharged, electric energy supplied to the battery reverses the direction of the battery's spontaneous reaction.
- Fuel cells are batteries in which the substance oxidized is a fuel from an external source.
- Methods of preventing corrosion are painting, coating with another metal, or using a sacrificial anode.

- 15. **MAIN Idea Identify** what is reduced and what is oxidized in the zinc-carbon dry-cell battery. What features make the alkaline dry cell an improvement over the earlier type of dry-cell battery?
- 16. **Explain** what happens when a battery is recharged.
- 17. **Describe** the half-reactions that occur in a hydrogen fuel cell, and write the equation for the overall reaction.
- 18. **Describe** the function of a sacrificial anode. How is the function of a sacrificial anode similar to galvanization?
- 19. **Explain** why lithium is a good choice for the anode of a battery.
- 20. **Calculate** Use data from **Table 20.1** to calculate the cell potential of the hydrogen-oxygen fuel cell described on page 723.
- 21. **Design an Experiment** Use your knowledge of acids from Chapter 18 to devise a method for determining whether a lead-acid battery can deliver full charge or is beginning to run down.

Section 20.3

Objectives

- Describe how it is possible to reverse a spontaneous redox reaction in an electrochemical cell.
- Compare the reactions involved in the electrolysis of molten sodium chloride with those in the electrolysis of brine.
- Discuss the importance of electrolysis in the smelting and purification of metals.

Review Vocabulary

redox reaction: an oxidation-reduction reaction

New Vocabulary

electrolysis
electrolytic cell

■ **Figure 20.19** The zinc-copper electrochemical cell can be a voltaic cell or an electrolytic cell.

Infer In each electrochemical cell, which metal is oxidized? Which is reduced?

Electrolysis

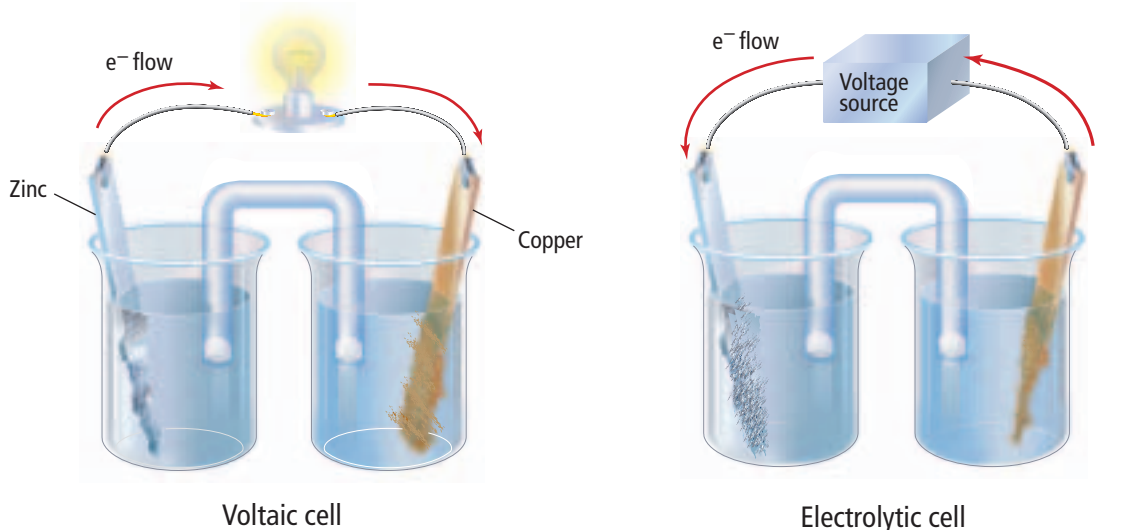
MAIN Idea In electrolysis, a power source causes nonspontaneous reactions to occur in electrochemical cells.

Real-World Reading Link When you ride a bicycle downhill, you don't have to do any work—you just coast. What is different when you ride uphill? You have to provide a lot of energy by pedaling.

Reversing Redox Reactions

When a battery generates electric current, electrons given up at the anode flow through an external circuit to the cathode, where they are used in a reduction reaction. A secondary battery is one that can be recharged by passing a current through it in the opposite direction. To help you understand the process, study the electrochemical cells in **Figure 20.19**. The beakers on the left contain zinc strips in solutions of zinc ions. The beakers on the right contain copper strips in solutions of copper ions. One electrochemical cell is supplying power to a lightbulb by means of a spontaneous redox reaction. Electrons flow spontaneously from the zinc side to the copper side, creating an electric current. The reaction continues until the zinc strip is used up, and then the reaction stops. However, the cell can be regenerated if current is applied in the reverse direction using an external voltage source. The voltage source is required because the reverse reaction is nonspontaneous. If the voltage source is applied long enough, the cell will return to nearly its original strength.

The use of electrical energy to bring about a chemical reaction is called **electrolysis**. An electrochemical cell in which electrolysis occurs is called an **electrolytic cell**. For example, when a secondary battery is recharged, it is acting as an electrolytic cell.



In this voltaic cell, the oxidation of zinc supplies the electrons to light the bulb and reduce copper ions. The spontaneous reaction continues until the zinc is used up.

When an outside voltage is applied, the flow of electrons is reversed and the nonspontaneous reaction occurs, which restores the conditions of the cell.

Applications of Electrolysis

Recall that voltaic cells convert chemical energy to electrical energy as a result of a spontaneous redox reaction. Electrolytic cells do the opposite; they use electrical energy to drive a nonspontaneous reaction. A common example is the electrolysis of water. This reaction is the opposite of burning hydrogen in a fuel cell.



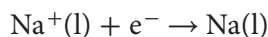
The electrolysis of water is one method by which hydrogen gas can be generated for commercial use.

Electrolysis of molten NaCl Just as electrolysis can decompose water into its elements, it can also separate molten sodium chloride into sodium metal and chlorine gas. This process is carried out in a chamber called a Down's cell, as illustrated in **Figure 20.20**. The electrolyte in the cell is the molten sodium chloride itself. Remember that ionic compounds can conduct electricity only when their ions are free to move, such as when they are dissolved in water or are in the molten state.

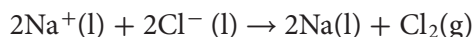
At the anode, chloride ions are oxidized to chlorine (Cl_2) gas.



At the cathode, sodium ions are reduced to sodium metal.



The net cell reaction is the following.



The importance of the Down's cell can best be appreciated in terms of the important roles that both sodium and chlorine play in your life. Chlorine is used throughout the world to purify water for drinking and swimming. Many cleaning products you might use, including household bleach, contain chlorine compounds. You depend on a host of other products, such as paper, plastics, insecticides, textiles, dyes, and paints, that either contain chlorine, or chlorine was used in their production.

In its pure form, sodium is used as a coolant in nuclear reactors and in sodium vapor lamps used for outdoor lighting. In its combined form in ionic compounds, you need only look on the contents list of consumer products to find a variety of sodium salts in the products you use and the foods you eat.

 **Reading Check** Explain why the sodium chloride must be molten in the Down's cell.

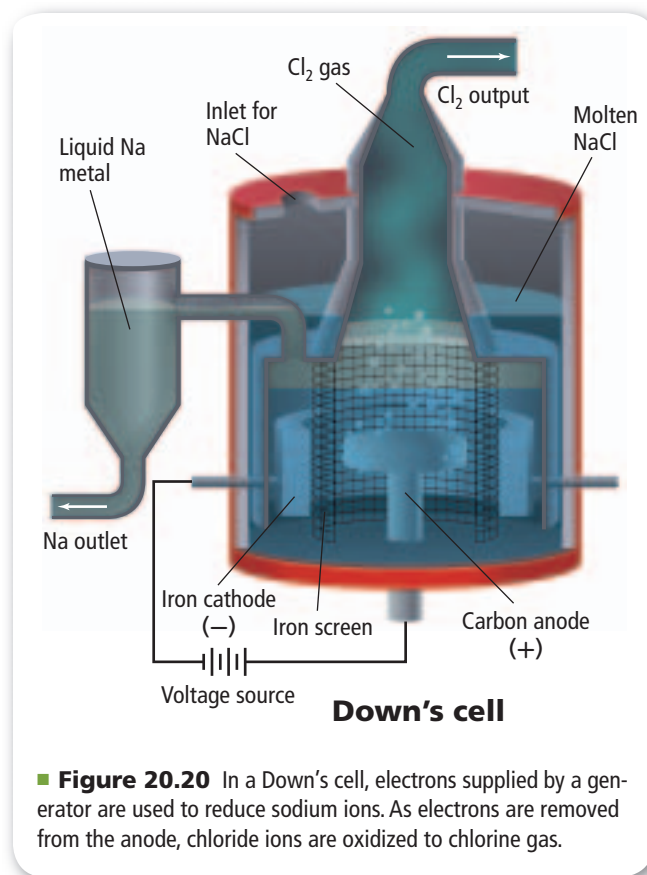
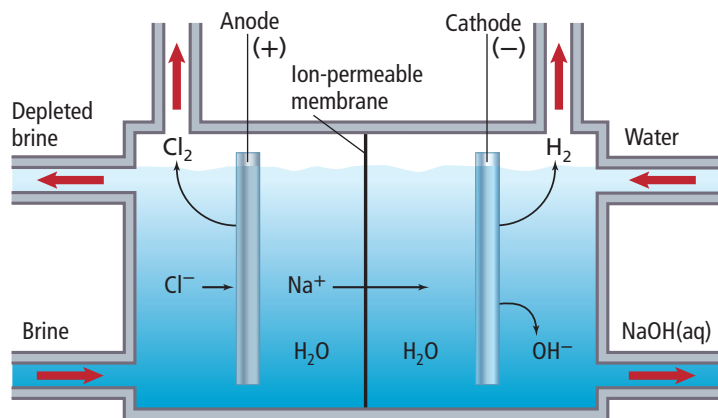


Figure 20.20 In a Down's cell, electrons supplied by a generator are used to reduce sodium ions. As electrons are removed from the anode, chloride ions are oxidized to chlorine gas.

FOLDABLES

Incorporate information from this section into your Foldable.



Commercial facilities use an electrolytic process to obtain hydrogen gas, chlorine gas, and sodium hydroxide from brine.

■ **Figure 20.21** In the electrolysis of brine (aqueous NaCl), sodium is not a product because water is easier to reduce.



Chlorine gas is used to manufacture polyvinyl chloride products, such as these pipes for water distribution.

VOCABULARY

SCIENCE USAGE V. COMMON USAGE

Reduce

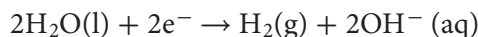
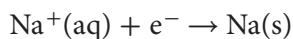
Science usage: to decrease an atom's oxidation number by the addition of electrons

Zinc reduces copper(II) ions to copper atoms by releasing two electrons.

Common usage: to diminish in size, amount, extent, or number

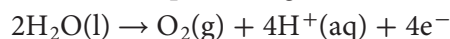
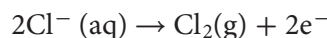
The number of dancers had to be reduced because the stage was too small.

Electrolysis of brine The decomposition of brine, an aqueous solution of sodium chloride, is also carried out by means of electrolysis. **Figure 20.21** shows a typical electrolytic cell and the products of the electrolysis. Two reactions are possible at the cathode: the reduction of sodium ions and the reduction of hydrogen in water molecules.



However, the reduction of sodium (Na^+) does not occur because water is easier to reduce, and thus is reduced preferentially.

Two reactions are also possible at the anode: the oxidation of chloride ions and the oxidation of oxygen in water molecules.



Because the desired product is chlorine (Cl_2), the concentration of chloride ions is kept high in order to favor this half-reaction. The overall cell reaction is as follows.



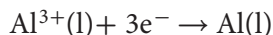
All three products are commercially important substances.



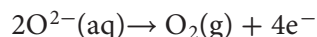
Reading Check Name the species that is oxidized and the species that is reduced in the electrolysis of brine.

Aluminum production Until the late nineteenth century, aluminum metal was more precious than gold because no one knew how to purify it in large quantities. In 1886, 22-year-old Charles Martin Hall (1863–1914) developed a process to produce aluminum by electrolysis. He used heat from a blacksmith forge, electricity from homemade batteries, and his mother's iron skillets as electrodes. At almost the same time, one of Le Châtelier's students, Paul L. T. Héroult (1863–1914), also 22 years old, discovered the same process. Today, it is called the Hall-Héroult process and is illustrated in **Figure 20.22**.

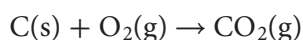
In the modern version of the Hall–Héroult process, aluminum metal is obtained by electrolysis of aluminum oxide, which is refined from bauxite ore ($\text{Al}_2\text{O}_3 \cdot 2\text{H}_2\text{O}$). The aluminum oxide is dissolved at 1000°C in molten synthetic cryolite (Na_3AlF_6), another aluminum compound. The cell is lined with graphite, which forms the cathode for the reaction, as shown in **Figure 20.22**. Another set of graphite rods is immersed in the molten solution as an anode. The following half-reaction occurs at the cathode.



The molten aluminum settles to the bottom of the cell and is drawn off periodically. Oxide ions are oxidized at the cathode in this half-reaction.

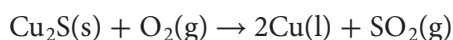


Because temperatures are high, the liberated oxygen reacts with the carbon of the anode to form carbon dioxide.

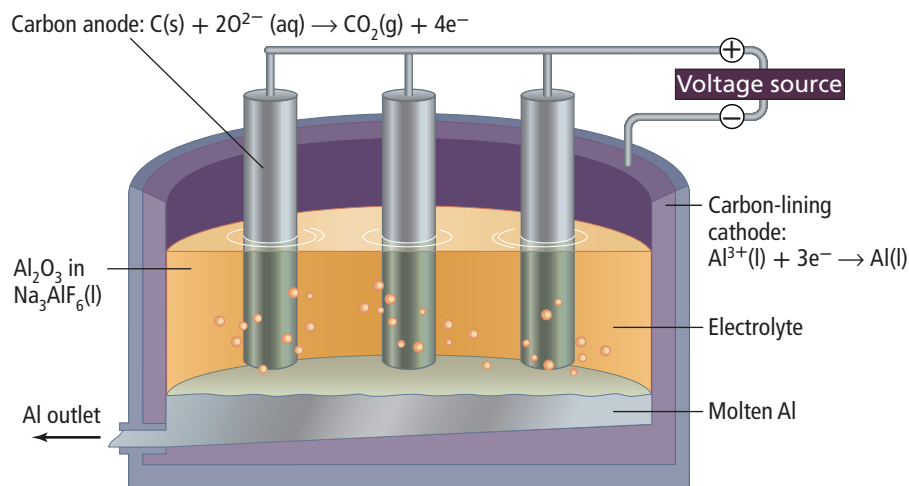


The Hall–Héroult process uses huge amounts of electrical energy. For this reason, aluminum is often produced in plants built close to large hydroelectric power stations, where electrical energy is less expensive. The vast amount of electricity needed to produce aluminum from ore is the primary reason for recycling aluminum. Recycled aluminum has already undergone electrolysis, so the only energy required to make it usable again is the heat needed to melt it in a furnace.

Purification of ores Electrolysis is also used in the purification of metals such as copper. Most copper is mined in the form of the ores chalcopyrite (CuFeS_2), chalcocite (Cu_2S), and malachite ($\text{Cu}_2\text{CO}_3(\text{OH})_2$). The sulfides are most abundant and yield copper metal when heated strongly in the presence of oxygen.



■ **Figure 20.22** The Hall–Héroult process operates at 900°C in smelters similar to this one. Note that carbon (graphite) serves as both the anode and the cathode. Recycled aluminum is often fed into the cell with the new aluminum.



Every ton of aluminum that is recycled saves huge quantities of electrical energy that would be used to produce new aluminum from ore.



■ **Figure 20.23** Power is needed to oxidize silver at the anode and reduce silver at the cathode. In an electrolytic cell used for silver plating, the object to be plated is the cathode where silver ions in the electrolyte solution are reduced to silver metal and deposited on the object.

The copper obtained from this process contains many impurities and must be refined, so the molten copper is cast into large, thick plates. These plates are then used as an anode in an electrolytic cell containing a solution of copper(II) sulfate. The cathode of the cell is a thin sheet of pure copper. As current is passed through the cell, copper atoms in the impure anode are oxidized to copper(II) ions. The copper ions migrate through the solution to the cathode, where they are reduced to copper atoms. These atoms become part of the cathode, while impurities fall to the bottom of the cell.

Electroplating Objects can be electroplated with a metal such as silver in a method similar to that used to refine copper. The object to be silver-plated is the cathode of an electrolytic cell. The anode is a silver bar or sheet, as shown in **Figure 20.23**. At the anode, silver is oxidized to silver ions as electrons are removed by the power source. At the cathode, the silver ions are reduced to silver metal by electrons from the external power source. The silver forms a thin coating over the object being plated. Current passing through the cell must be carefully controlled in order to get a smooth, even metal coating.

Other metals are also used for electroplating. You might have costume jewelry that is electroplated with gold. Or you might admire an automobile whose steel parts such as the bumper have been made more corrosion-resistant by being electroplated first with nickel and then with chromium.

Section 20.3 Assessment

Section Summary

- ▶ In an electrolytic cell, an outside source of power causes a nonspontaneous redox reaction to occur.
- ▶ The electrolysis of molten sodium chloride yields sodium metal and chlorine gas. The electrolysis of brine yields hydrogen gas, sodium hydroxide, and chlorine gas.
- ▶ Metals such as copper are purified in an electrolytic cell.
- ▶ Electrolysis is used to electroplate objects and to produce pure aluminum from its ore.

- 22. **MAIN Idea** Define *electrolysis* and relate the definition to the spontaneity of redox reactions.
- 23. **Explain** why the products of the electrolysis of brine and the electrolysis of molten sodium chloride are different.
- 24. **Describe** how impure copper obtained from the smelting of ore is purified by electrolysis.
- 25. **Explain**, by referring to the Hall-Héroult process, why recycling aluminum is very important.
- 26. **Describe** the anode and cathode of an electrolytic cell in which gold is to be plated on an object.
- 27. **Explain** why producing a kilogram of silver from its ions by electrolysis requires much less electric energy than producing a kilogram of aluminum from its ions.
- 28. **Calculate** Use **Table 20.1** to calculate the voltage of the Down's cell. Should the potential be positive or negative?
- 29. **Summarize** Write a short paragraph summarizing each of the three objectives for Section 20.3 in your own words.

HOW IT WORKS

The Pacemaker: Helping a Broken Heart

Your heart is made of cardiac muscle tissue that contracts and relaxes continuously. This beating results from electric impulses moving along pathways throughout your heart. A group of specialized cells in the upper wall of the heart's right atrium—upper chamber—generates electric impulses. If these cells fail to function or the electric impulse pathways are interrupted, the heart does not beat normally. A pacemaker is an electrical device that can monitor and correct an irregular heartbeat. How does it work?

1 Leads Insulated wires called leads carry electric signals between the heart and the pacemaker. A lead is implanted into a blood vessel and then into a chamber of the heart. Pacemakers might use one, two, or three leads, each in a different chamber.

2 Pacemaker Each lead is connected to the pacemaker, which contains a battery, an antenna, and computer circuits. The pacemaker is then implanted under the skin below the collarbone. Surgery is also required to replace a pacemaker when its battery is low or the circuits fail.



Lead tip

4 Function The tip of the lead attaches to heart tissue, transmits electric impulses, and senses the heartbeat. Pacemakers can stimulate a sluggish heartbeat, smooth an irregular beat, or activate the beating of an arrested heart.

3 Battery Most pacemakers use lithium-iodine batteries. Lithium is the anode and iodine is the cathode. These batteries are ideal for pacemakers because they are long-lived (5 to 8 years), fail gradually, and do not release gases.

WRITING in Chemistry

Summarize The pacemaker is an example of an advancement in medical technology. Write a paragraph about the ways your life or that of someone you know has changed because of medical technology and how it might change in the future. Visit glencoe.com to learn about pacemakers.

MEASURE VOLTAIC CELL POTENTIALS

Background: When two different half-cells are joined, a potential difference is produced. A voltmeter measures the potential difference of combinations of half-cells.

Question: How do measured potentials of voltaic cells compare to calculated potentials?

Materials

metal strips of copper, aluminum, zinc, and magnesium (about 0.6 cm × 1.3 cm)

1M copper(II) nitrate

1M aluminum nitrate

1M zinc nitrate

1M magnesium nitrate

24-well microplate

Beral-type pipette (5)

voltmeter

filter paper (6 pieces 0.6 cm × 2.5 cm)

1M potassium nitrate

forceps

steel wool or sandpaper

table of standard reduction potentials

Safety Precautions



WARNING: Steel wool might have sharp points that can pierce skin. Handle with care.

Procedure

1. Read and complete the lab safety form.
2. Plan how you will arrange voltaic cells using the four metal combinations in the 24-well microplate. Have your instructor approve your plan.
3. Soak the strips of filter paper in 2 mL of potassium nitrate solution. The strips are the salt bridges for the cells. Use forceps to handle the salt bridges.
4. Construct voltaic cells using the four metals and 1 mL of each of the solutions. Put the metals into the wells that contain the appropriate solution; for example, put the zinc metal in the solution with zinc nitrate. Use a different salt bridge for each voltaic cell. Connect the leads from the voltmeter to the metals. If you get a negative value for potential difference, switch the leads on the metals.

Voltaic Cell Potential Data

Anode metal			
Cathode metal			
Measured cell potential (V)			
Anode half-reaction and standard potential			
Cathode half-reaction and standard potential			
Theoretical cell potential			
% Error			

5. Record in the data table which metals are the anode and the cathode in each cell. The black lead of the voltmeter will be attached to the anode. The red lead will be attached to the cathode.
6. Record the cell potential of each cell.
7. **Cleanup and Disposal** Use forceps to remove the metal strips from the microplate. Rinse them with water, then clean them with steel wool or sandpaper. Rinse the microplate.

Analyze and Conclude

1. **Apply** In the data table, write the equations for half-reactions occurring at the anode and cathode in each of the voltaic cells. Find the half-reaction potentials in **Table 20.1**, and record them in the table.
2. **Calculate** and record the theoretical potential for each voltaic cell.
3. **Predict** Using your data, rank the metals in order of most active to least active.
4. **Error Analysis** Calculate the percent error of the voltaic cell potential. Why is the percent error large for some voltaic cells and small for others?

INQUIRY EXTENSION

Design an Experiment that would reduce the percent error discussed in Question 4.



BIG Idea Chemical energy can be converted to electric energy and electric energy to chemical energy.

Section 20.1 Voltaic Cells

MAIN Idea In voltaic cells, oxidation takes place at the anode, yielding electrons that flow to the cathode, where reduction occurs.

Vocabulary

- anode (p. 710)
- cathode (p. 710)
- electrochemical cell (p. 709)
- half-cell (p. 710)
- reduction potential (p. 711)
- salt bridge (p. 709)
- standard hydrogen electrode (p. 711)
- voltaic cell (p. 709)

Key Concepts

- In a voltaic cell, oxidation and reduction take place at electrodes separated from each other.
- The standard potential of a half-cell reaction is its voltage when paired with a standard hydrogen electrode under standard conditions.
- The reduction potential of a half-cell is negative if it undergoes oxidation when connected to a standard hydrogen electrode. The reduction potential of a half-cell is positive if it undergoes reduction when connected to a standard hydrogen electrode.
- The standard potential of a voltaic cell is the difference between the standard reduction potentials of the half-cell reactions.

$$E_{\text{cell}}^0 = E_{\text{reduction}}^0 - E_{\text{oxidation}}^0$$

Section 20.2 Batteries

MAIN Idea Batteries are voltaic cells that use spontaneous reactions to provide energy for a variety of purposes.

Vocabulary

- battery (p. 718)
- corrosion (p. 724)
- dry cell (p. 718)
- fuel cell (p. 722)
- galvanization (p. 727)
- primary battery (p. 720)
- secondary battery (p. 720)

Key Concepts

- Primary batteries can be used only once; secondary batteries can be recharged.
- When a battery is recharged, electric energy supplied to the battery reverses the direction of the battery's spontaneous reaction.
- Fuel cells are batteries in which the substance oxidized is a fuel from an external source.
- Methods of preventing corrosion are painting, coating with another metal, or using a sacrificial anode.

Section 20.3 Electrolysis

MAIN Idea In electrolysis, a power source causes nonspontaneous reactions to occur in electrochemical cells.

Vocabulary

- electrolysis (p. 728)
- electrolytic cell (p. 728)

Key Concepts

- In an electrolytic cell, an outside source of power causes a nonspontaneous redox reaction to occur.
- The electrolysis of molten sodium chloride yields sodium metal and chlorine gas. The electrolysis of brine yields hydrogen gas, sodium hydroxide, and chlorine gas.
- Metals such as copper are purified in an electrolytic cell.
- Electrolysis is used to electroplate objects and to produce pure aluminum from its ore.

Section 20.1

Mastering Concepts

30. What feature of an oxidation-reduction reaction allows it to be used to generate an electric current?
31. Describe the process that releases electrons in a zinc-copper voltaic cell.
32. What is the function of a salt bridge in a voltaic cell?
33. What information do you need in order to determine the standard voltage of a voltaic cell?
34. In a voltaic cell represented by $\text{Al}|\text{Al}^{3+}||\text{Cu}^{2+}|\text{Cu}$, what is oxidized and what is reduced as the cell delivers current?
35. Under what conditions are standard reduction potentials measured?

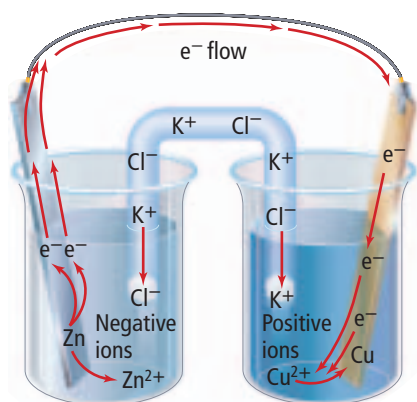


Figure 20.24

36. In Figure 20.24, identify the metal that is being oxidized. Identify the cathode.
37. A salt bridge is filled with KNO_3 . Explain why it is necessary that the potassium ions move through the salt bridge to the cathode.
38. Recall that a reducing agent is the substance being oxidized and an oxidizing agent is the substance being reduced. Use Table 20.1 to select an oxidizing agent that will convert Au to Au^{3+} but will not convert Co^{2+} to Co^{3+} .

Mastering Problems

39. Using Table 20.1, write the standard cell notation for each cell in which each of the following half-cells is connected to the standard hydrogen electrode.
 - a. $\text{Zn}|\text{Zn}^{2+}$
 - b. $\text{Hg}|\text{Hg}^{2+}$
 - c. $\text{Cu}|\text{Cu}^{2+}$
 - d. $\text{Al}|\text{Al}^{3+}$
40. Write the balanced chemical equation for the standard cell notations listed below.
 - a. $\text{I}^-|\text{I}_2||\text{Fe}^{3+}|\text{Fe}^{2+}$
 - b. $\text{Sn}|\text{Sn}^{2+}||\text{Ag}^+|\text{Ag}$
 - c. $\text{Zn}|\text{Zn}^{2+}||\text{Cd}^{2+}|\text{Cd}$

41. Calculate the cell potential for the following voltaic cells.

- a. $2\text{Ag}^+(\text{aq}) + \text{Pb}(\text{s}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$
- b. $\text{Mn}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{Ni}$
- c. $\text{I}_2(\text{aq}) + \text{Sn}(\text{s}) \rightarrow 2\text{I}^-(\text{aq}) + \text{Sn}^{2+}(\text{aq})$

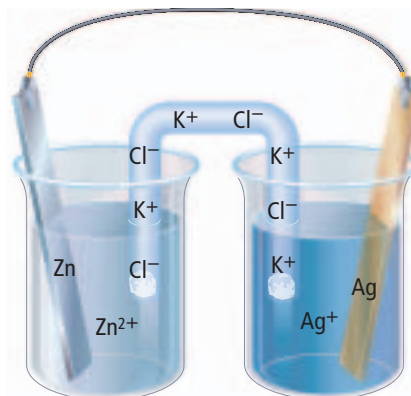


Figure 20.25

42. Figure 20.25 illustrates a voltaic cell consisting of a strip of zinc in a 1.0M solution of zinc nitrate and a strip of silver in a 1.0M solution of silver nitrate. Use the diagram and Table 20.1 to answer these questions.
 - a. Identify the anode.
 - b. Identify the cathode.
 - c. Where does oxidation occur?
 - d. Where does reduction occur?
 - e. In which direction is the current flowing through the connecting wire?
 - f. In which direction are positive ions flowing through the salt bridge?
 - g. What is the cell potential at 25°C and 1 atm ?

Section 20.2

Mastering Concepts

43. What part of a zinc-carbon dry cell is the anode? Describe the reaction that takes place there.
44. How do primary and secondary batteries differ?
45. **Lead-Acid Battery** What substance is reduced in a lead-acid storage battery? What substance is oxidized? What substances are produced in each reaction?
46. **Biofuel Cell** At the cathode of a biofuel cell, Fe^{3+} in potassium hexacyanoiron(III) ($\text{K}_3[\text{Fe}(\text{CN})_6]$) is reduced to Fe^{2+} in potassium hexacyanoiron(II) ($\text{K}_4[\text{Fe}(\text{CN})_6]$). At the anode, reduced nicotinamide-adenine-dinucleotide (NADH) is oxidized to NAD^+ . Use the following standard reduction potential to determine the potential of the cell.

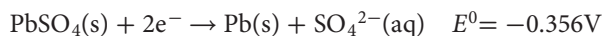
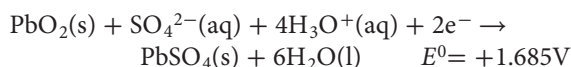
$$\text{NAD}^+ + \text{H}^+ + 2\text{e}^- \rightarrow \text{NADH} \quad E^0 = -0.320\text{ V}$$

$$[\text{Fe}(\text{CN})_6]^{3-} + 1\text{e}^- \rightarrow [\text{Fe}(\text{CN})_6]^{4-} \quad E^0 = +0.36\text{ V}$$

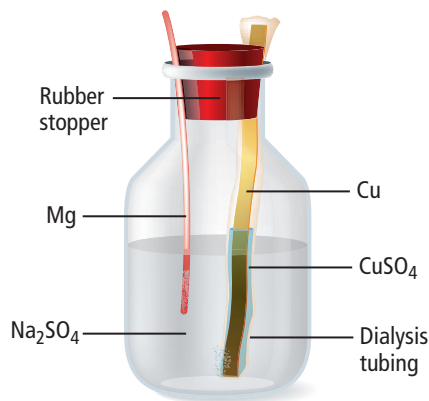
- 47. Fuel Cells** List two ways in which a fuel cell differs from an ordinary battery.
- 48. Galvanization** What is galvanization? How does galvanizing iron protect it from corrosion?
- 49. Batteries** Explain why a lead storage battery does not produce a current when the level of H_2SO_4 is low.
- 50. Steel Wool** is a bundle of filaments made of steel, an alloy of iron and carbon. Which would be the best way to store steel wool?
- Store it in water.
 - Store it in open air.
 - Store it with a desiccant.
- 51. Corrosion Protection** List three ways metals can be protected from corrosion.

Mastering Problems

- 52.** Half-reactions for a lead-acid storage battery are below.



What is the standard cell potential for one cell in a car battery?



■ **Figure 20.26**

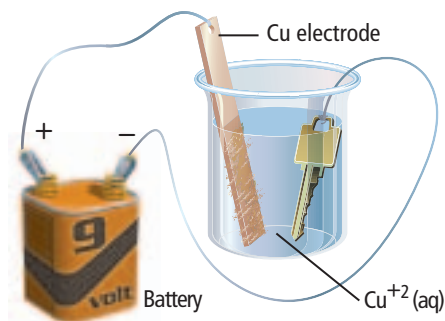
- 53.** The setup in **Figure 20.26** acts as a battery.
- Determine the reaction that takes place at the copper strip.
 - Determine the reaction that takes place at the magnesium wire.
 - Identify the anode.
 - Identify the cathode.
 - Calculate the standard cell potential for this battery.
- 54.** You design a battery that uses a half-cell containing Sn and Sn^{2+} and another half-cell containing Cu and Cu^{2+} . The copper electrode is the cathode, and the tin electrode is the anode. Draw the battery and write the half-reactions that occur in each half-cell. What is the maximum voltage this battery can produce?

Section 20.3

Mastering Concepts

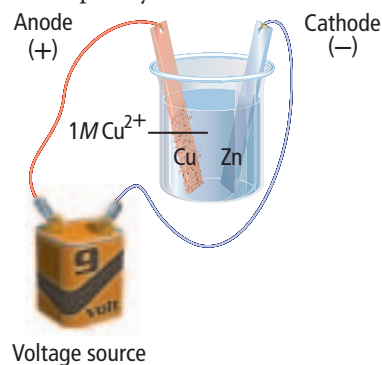
- 55.** How can the spontaneous redox reaction of a voltaic cell be reversed?
- 56.** Where does oxidation take place in an electrolytic cell?
- 57. Down's Cell** What reaction takes place at the cathode when molten sodium chloride is electrolyzed?
- 58. Industry** Explain why the electrolysis of brine is done on a large scale at many sites around the world.
- 59. Recycling** Explain how recycling aluminum conserves energy.
- 60.** Describe what happens at the anode and the cathode in the electrolysis of $\text{KI}(\text{aq})$.

Mastering Problems



■ **Figure 20.27**

- 61. Electroplating** **Figure 20.27** shows a key being electroplated with copper in an electrolytic cell. Where does oxidation occur? Explain your answer.

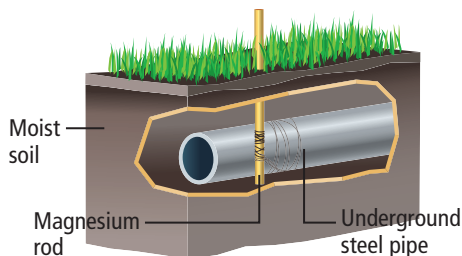


■ **Figure 20.28**

- 62.** Answer the following questions based on **Figure 20.28**.
- Which electrode grows? Write the reaction that occurs at this electrode.
 - Which electrode disappears? Write the reaction that occurs at this electrode.
- 63.** Using **Figure 20.28**, explain what happens to the copper ions in solution.

Mixed Review

64. Why do electrons flow from one electrode to the other in a voltaic cell?
65. **Aluminum Production** What substance is electrolyzed in the industrial process to produce aluminum metal?
66. Write the oxidation and reduction half-reactions for a silver-chromium voltaic cell. Identify the anode, cathode, and the direction of electron flow.
67. Determine whether each redox reaction is spontaneous or nonspontaneous.
- $\text{Mn}^{2+}(\text{aq}) + 2\text{Br}^{-}(\text{aq}) \rightarrow \text{Br}_2(\text{l}) + \text{Mn}(\text{s})$
 - $2\text{Fe}^{2+}(\text{aq}) + \text{Sn}^{2+}(\text{aq}) \rightarrow 2\text{Fe}^{3+}(\text{aq}) + \text{Sn}(\text{s})$
 - $\text{Ni}^{2+}(\text{aq}) + \text{Mg}(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Ni}(\text{s})$
 - $\text{Pb}^{2+}(\text{aq}) + 2\text{Cu}^{+}(\text{aq}) \rightarrow \text{Pb}(\text{s}) + 2\text{Cu}^{2+}(\text{aq})$
68. Determine the voltage of the cell in which each half-cell is connected to a $\text{Ag} | \text{Ag}^{+}$ half-cell.
- $\text{Be}^{2+} | \text{Be}$
 - $\text{S} | \text{S}^{2-}$
 - $\text{Au}^{+} | \text{Au}$
 - $\text{I}_2 | \text{I}^{-}$
69. **Corrosion** Explain why water is necessary for the corrosion of iron.
70. **Space Travel** The space shuttle uses a H_2/O_2 fuel cell to produce electricity.
- What is the reaction at the anode? At the cathode?
 - What is the standard cell potential for the fuel cell?
71. **Fuel Cells** Explain how the oxidation of hydrogen in a fuel cell differs from the oxidation of hydrogen when it burns in air.
72. **Copper Refining** In the electrolytic refining of copper, what factor determines which piece of copper is the anode and which is the cathode?
73. **Storage Batteries** Lead-acid batteries and other rechargeable batteries are sometimes called storage batteries. What is being stored in these batteries?

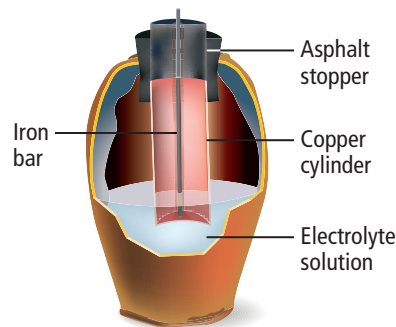


■ Figure 20.29

74. **Corrosion Prevention** Figure 20.29 shows how buried steel pipes can be protected against corrosion. The steel pipe is connected to a more active metal that corrodes instead of the steel.
- What is the cathode? What is the anode?
 - Describe how the magnesium metal protects the steel.

Think Critically

75. **Predict** Suppose that scientists had chosen the $\text{Cu}^{2+} | \text{Cu}$ half-cell as a standard instead of the $\text{H}^{+} | \text{H}_2$ half-cell? What would the potential of the hydrogen electrode be if the copper electrode were the standard? How would the relationships among the standard reduction potentials change?
76. **Apply** Suppose that you have a voltaic cell in which one half-cell is made up of a strip of tin immersed in a solution of tin(II) ions.
- How could you tell by measuring voltage whether the tin strip was acting as a cathode or an anode in the cell?
 - How could you tell by simple observation whether the tin strip was acting as a cathode or an anode?
77. **Hypothesize** The potential of a half-cell varies with concentration of reactants and products. For this reason, standard potentials are measured at 1M concentration. Maintaining a pressure of 1 atm is especially important in half-cells that involve gases as reactants or products. Suggest a reason why gas pressure is critical in these cells.



■ Figure 20.30

78. **Analyze** An earthen vessel was discovered in 1938 near Baghdad. This ancient vessel contained an iron bar surrounded by a copper cylinder, as shown in Figure 20.30. When filled with an electrolyte such as vinegar, this vessel might have acted as a battery.
- Identify the cathode.
 - Identify the anode.
 - Calculate the standard cell potential of this battery.
79. **Apply** During electrolysis, an electrolytic cell releases bromine vapor and hydrogen gas. After electrolysis, the cell is found to contain a concentrated solution of potassium hydroxide. What was the composition of the cell before electrolysis began?
80. **Hypothesize** Suppose in galvanization, copper was plated on iron instead of zinc. Would copper continue to protect the iron from corrosion, as zinc does, if the copper coating became broken or cracked? Explain.

Challenge Problem

81. A battery is assembled using tin and mercury, which have the following reduction half-reactions:
- $$\text{Sn}^{2+} + 2e^{-} \rightarrow \text{Sn}$$
- $$\text{Hg}^{2+} + 2e^{-} \rightarrow \text{Hg}$$
- Write a balanced equation for the cell's reaction.
 - What is oxidized and what is reduced? Identify the oxidizing agent and the reducing agent.
 - Which reaction occurs at the anode? At the cathode?
 - What is the cell potential? Use **Table 20.1**.
 - If sodium sulfate solution is in the salt bridge, in which direction do the sulfate ions move?

Cumulative Review

82. If the volume of a sample of chlorine gas is 8.2 L at 1.2 atm and 306 K, what volume will the gas occupy at STP? (*Chapter 13*)
83. What is meant by solvation? Explain how this process is important for the dissolving of ionic salts in water. (*Chapter 14*)
84. Explain how the molarity of a solution is different from its molality. (*Chapter 14*)
85. Define the calorie. State how the calorie is related to the Calorie and the joule. (*Chapter 15*)
86. Explain why you would find an aluminum chair to be hotter to sit on than a wooden bench after each had been in the sunlight for the same amount of time. (*Chapter 15*)
87. What does a negative sign for the free energy of a reaction tell you about the reaction? ($\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$) (*Chapter 15*)
88. According to the collision model of chemical reactions, how is it possible that two molecules can collide but not react? (*Chapter 16*)
89. List five factors that can affect the rate of a reaction. (*Chapter 16*)
90. The decomposition reaction $A_2B \rightarrow 2A + B$ proceeds to equilibrium at 499°C. Analysis of the equilibrium mixture shows $[A_2B] = 0.855 \text{ mol/L}$, $[A] = 2.045 \text{ mol/L}$, and $[B] = 1.026 \text{ mol/L}$. What is K_{eq} ? (*Chapter 17*)
91. What is the solubility in mol/L of silver iodide, AgI. K_{sp} for AgI is 3.5×10^{-17} . (*Chapter 17*)
92. If you have a solution of a strong acid, is that the same as having a concentrated solution of the acid? Explain your answer. (*Chapter 18*)
93. What are the oxidation numbers for the elements in the ion PO_4^{3-} ? (*Chapter 19*)

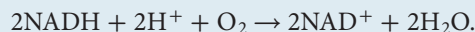
Additional Assessment

WRITING in Chemistry

94. **Sunken Ships** Study of the sunken ocean liner *Titanic* has opened the possibility that deterioration of the steel hull might be partly due to the presence of rusticle communities. Research how the biological activity of rusticle communities results in the oxidation of iron. Write an essay that describes the role of rusticle communities in the destruction of the *Titanic*.
95. **Statue of Liberty** Several years ago, the supporting structure of the Statue of Liberty became so corroded that it had to be replaced entirely. Find out what the structure was made of and why it corroded so badly. Write a report that explains the chemical processes involved and include a time line of the statue, starting in France before 1886.

DBQ Document-Based Questions

Electrochemical Biological Reactions *Standard reduction potentials for some important biological reactions are given in Table 20.2. The strongest oxidizing agent generally available in biological systems is molecular oxygen. Consider the oxidation of reduced nicotinamide-adenine-dinucleotide (NADH) by molecular oxygen. The reaction is the following.*



Data obtained from: Fromm, James Richard. 1997. Biochemical Electrochemistry.

Table 20.2 Aqueous Standard Reduction Potentials at 25°C and pH 7.00

Electrode Couple	E° (V)
$2\text{H}^{+} + 2e^{-} \rightarrow \text{H}_2(\text{g})$	-0.4141
$\text{NAD}^{+} + \text{H}^{+} + 2e^{-} \rightarrow \text{NADH}$	-0.320
$\text{HOCCCOCH}_3^{*} + 2\text{H}^{+} + 2e^{-} \rightarrow \text{HOCCCHOHCH}_3^{**}$	+0.19
$\text{Fe}^{3+} + e^{-} \rightarrow \text{Fe}^{2+}$	+0.769
$\text{O}_2(\text{g}) + 4\text{H}^{+} + 4e^{-} \rightarrow 2\text{H}_2\text{O}$	+0.8147

* HOCCCOCH_3 is pyruvic acid

** HOCCCHOHCH_3 is l-lactic acid.

96. Write the two half-reactions that take place in this reaction.
97. Calculate the cell potential of this reaction using **Table 20.1** on page 712 and **Table 20.2**.
98. Will NAD^{+} oxidize Fe^{2+} to Fe^{3+} ? Explain your answer.

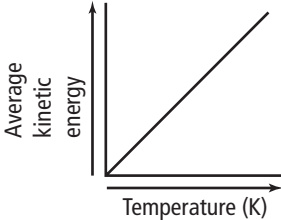
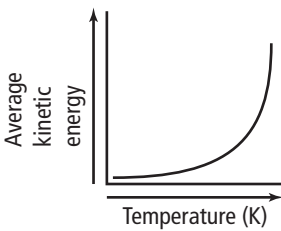
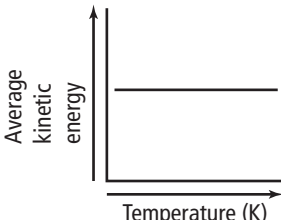
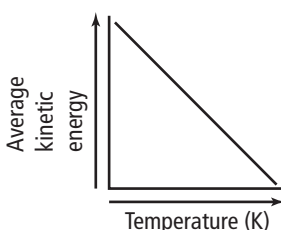
Cumulative Standardized Test Practice

Multiple Choice

Use the table below to answer Questions 1 to 4.

Selected Standard Reduction Potentials at 25°C, 1 atm, and 1M Ion Concentration	
Half-Reaction	E° (V)
$\text{Mg}^{2+} + 2\text{e}^{-} \rightarrow \text{Mg}$	-2.372
$\text{Al}^{3+} + 3\text{e}^{-} \rightarrow \text{Al}$	-1.662
$\text{Pb}^{2+} + 2\text{e}^{-} \rightarrow \text{Pb}$	-0.1262
$\text{Ag}^{+} + \text{e}^{-} \rightarrow \text{Ag}$	0.7996
$\text{Hg}^{2+} + 2\text{e}^{-} \rightarrow \text{Hg}$	0.851

- Which metal ion is most easily reduced?
 - Mg^{2+}
 - Hg^{2+}
 - Ag^{+}
 - Al^{3+}
- On the basis of the standard reduction potentials shown above, which standard cell notation correctly represents its voltaic cell?
 - $\text{Ag} | \text{Ag}^{+} || \text{Al}^{3+} | \text{Al}$
 - $\text{Mg} | \text{Mg}^{2+} || \text{H}^{+} | \text{H}_2$
 - $\text{H}_2 | \text{H}^{+} || \text{Pb}^{2+} | \text{Pb}$
 - $\text{Pb} | \text{Pb}^{2+} || \text{Al}^{3+} | \text{Al}$
- A voltaic cell consists of a magnesium bar dipping into a 1M Mg^{2+} solution and a silver bar dipping into a 1M Ag^{+} solution. What is the standard potential of this cell?
 - 1.572 V
 - 3.172 V
 - 0.773 V
 - 3.971 V
- Assuming standard conditions, which cell will produce a potential of 2.513 V?
 - $\text{Al} | \text{Al}^{3+} || \text{Hg}^{2+} | \text{Hg}$
 - $\text{H}_2 | \text{H}^{+} || \text{Hg}^{2+} | \text{Hg}$
 - $\text{Mg} | \text{Mg}^{2+} || \text{Al}^{3+} | \text{Al}$
 - $\text{Pb} | \text{Pb}^{2+} || \text{Ag}^{+} | \text{Ag}$

- Which statement is NOT true of batteries?
 - Batteries are compact forms of voltaic cells.
 - Secondary batteries are storage batteries.
 - Batteries can consist of a single cell.
 - The redox reaction in a rechargeable battery is reversible.
- Which is NOT a characteristic of a base?
 - tastes bitter
 - conducts electricity
 - reacts with some metals
 - feels slippery
- A carbonated soft drink has a pH of 2.5. What is the concentration of H^{+} ions in the soft drink?
 - $3 \times 10^{-12}\text{M}$
 - $3 \times 10^{-3}\text{M}$
 - $4 \times 10^{-1}\text{M}$
 - $1 \times 10^1\text{M}$
- Which graph correctly shows the relationship between average kinetic energy of particles and the temperature of a sample?
 - 
 - 
 - 
 - 

Short Answer

Use the description below to answer Questions 9 to 11.

In an experimental setup, chlorine gas and nitrogen gas are in separate containers separated by a closed stopcock. One hour after the stopcock is opened, the gases have completely mixed.

- Five minutes after the stopcock is opened, which gas will have traveled farther, the nitrogen or the chlorine?
- Give the ratio of the speed of nitrogen gas to the speed of chlorine gas.
- Evaluate this statement: After one hour, the gas particles stop moving because they have completely mixed.

Extended Response

Use the table below to answer Question 12.

Standard Reduction Potentials at 25°C, 1 atm, and 1M Solution	
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.7996
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.744

- Based on the standard reduction potentials given above, if a silver electrode and a chromium electrode are connected in a voltaic cell, which electrode will undergo oxidation and which will undergo reduction? Explain how you can tell.
- Use Le Châtelier's principle to explain why the instructions for a chemical experiment sometimes instruct the chemist to cool the reaction in an ice bath.

NEED EXTRA HELP?

If You Missed Question . . .	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Review Section . . .	20.1	20.1	20.1	20.1	20.2	18.1	18.3	15.1	12.1	12.1	12.1	8.5	17.2	17.3	15.2

SAT Subject Test: Chemistry

- The hydrogen sulfide produced as a by-product of petroleum refinement can be used to produce elemental sulfur: $2\text{H}_2\text{S}(\text{g}) + \text{SO}_2(\text{g}) \rightarrow 3\text{S}(\text{l}) + 2\text{H}_2\text{O}(\text{g})$.

What is the equilibrium constant expression for this reaction?

- A. $K_{\text{eq}} = \frac{[\text{H}_2\text{O}]}{[\text{H}_2\text{S}][\text{SO}_2]}$ D. $K_{\text{eq}} = \frac{[\text{S}]^3[\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2[\text{SO}_2]}$
 B. $K_{\text{eq}} = \frac{[\text{H}_2\text{S}]^2[\text{SO}_2]}{[\text{H}_2\text{O}]^2}$ E. $K_{\text{eq}} = \frac{[2\text{H}_2\text{O}]^2}{[2\text{H}_2\text{S}]^2[\text{SO}_2]}$
 C. $K_{\text{eq}} = \frac{[\text{H}_2\text{O}]^2}{[\text{H}_2\text{S}]^2[\text{SO}_2]}$

- Which shows the correct graph of the activation energy needed for an endothermic reaction?

