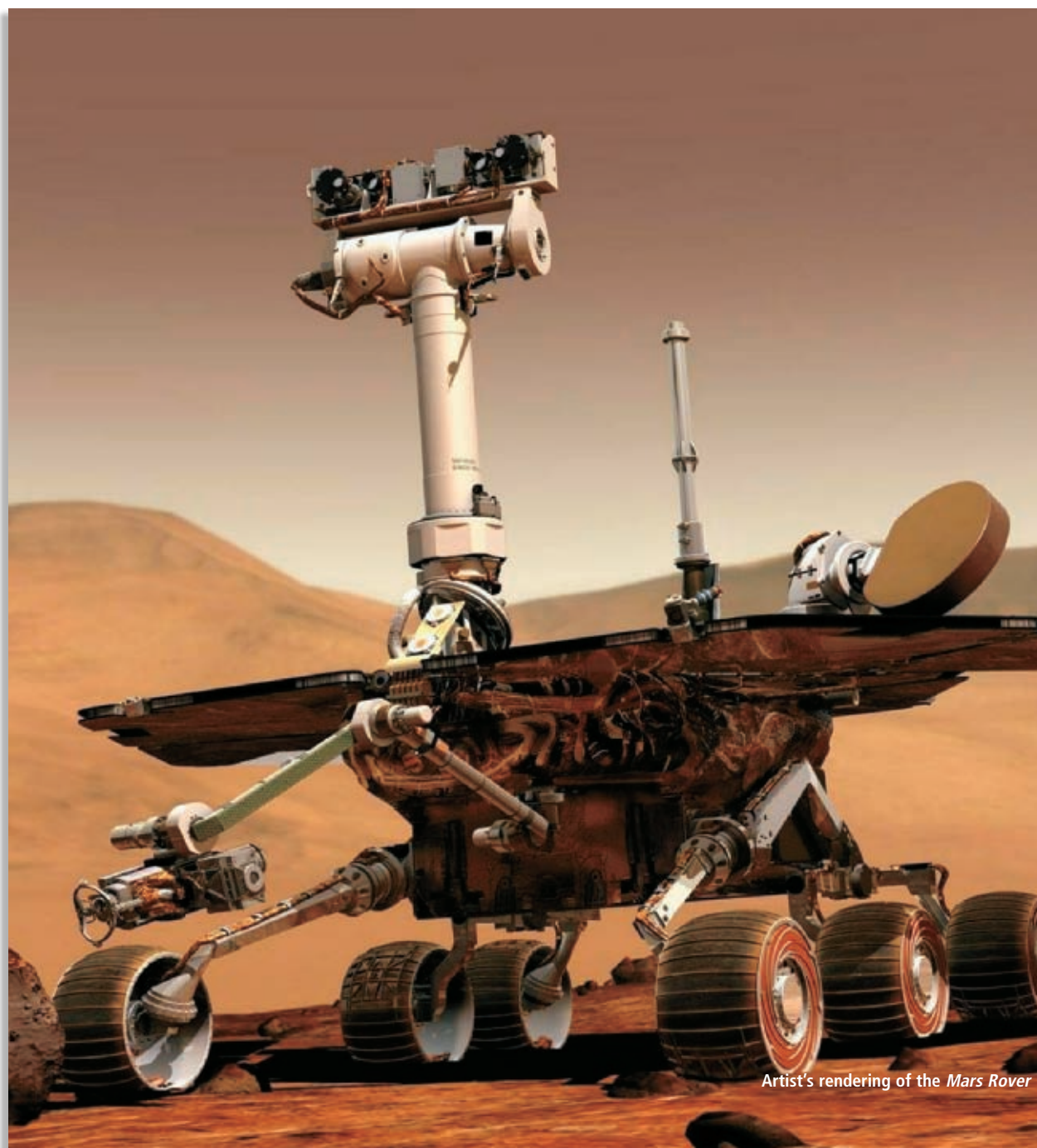


# *Electrochemistry*

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*Electrochemical reactions provide energy in all kinds of applications.*



Artist's rendering of the Mars Rover

# Introduction to Electrochemistry

## SECTION 1

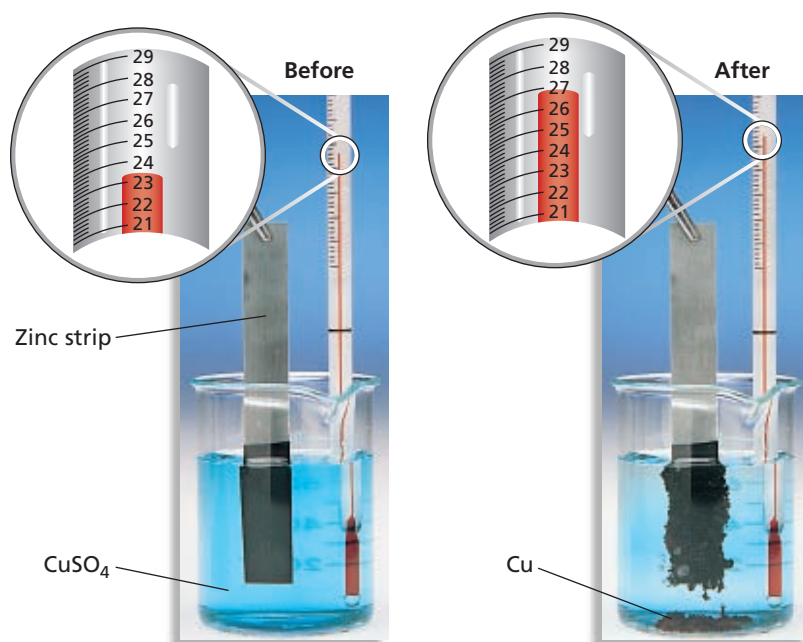
### OBJECTIVES

- Identify parts of an electrochemical cell and their functions.
- Write electrode half reactions for cathodes and anodes.

**O**xidation-reduction reactions involve energy changes. Because these reactions involve electron transfer, the net *release* or net *absorption* of energy can occur in the form of electrical energy rather than as heat. This property allows for a great many practical applications of redox reactions. *The branch of chemistry that deals with electricity-related applications of oxidation-reduction reactions is called electrochemistry.*

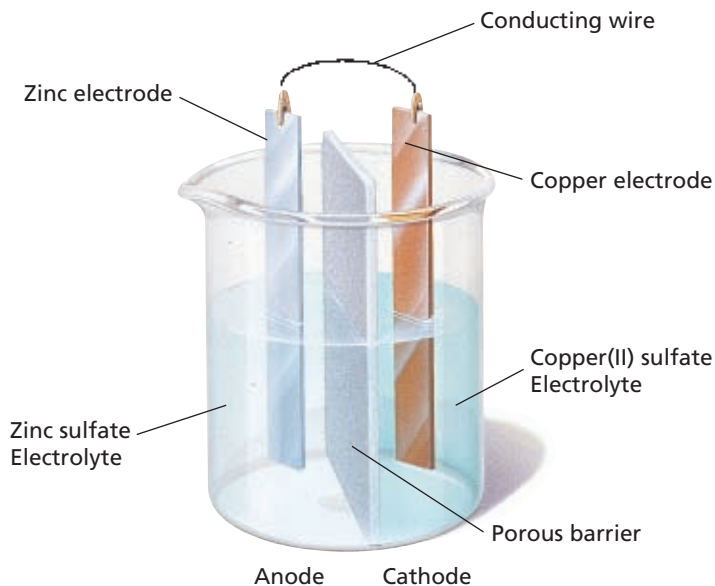
## Electrochemical Cells

Oxidation-reduction reactions involve a transfer of electrons. If the two substances are in contact with one another, a transfer of energy as heat accompanies the electron transfer. In **Figure 1** a zinc strip is in contact with a copper(II) sulfate solution. The zinc strip loses electrons to the copper(II) ions in solution. Copper(II) ions accept the electrons and fall out of solution as copper atoms. As electrons are transferred between zinc atoms and copper(II) ions, energy is released as heat, as indicated by the rise in temperature.



**FIGURE 1** Energy as heat given off when electrons are transferred directly from Zn atoms to Cu<sup>2+</sup> ions causes the temperature of the aqueous CuSO<sub>4</sub> solution to rise.

**FIGURE 2** An electrochemical cell consists of two electrodes. Each electrode is in contact with an electrolyte; the electrode and the electrolyte make up a half-cell. The two electrodes are connected by a wire, and a porous barrier separates the two electrolytes.



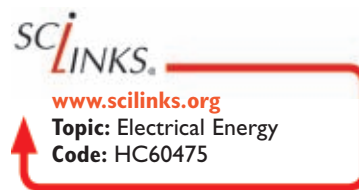
If, however, we separate the substance that is oxidized from the substance that is reduced, the electron transfer is accompanied by a transfer of electrical energy instead of energy as heat. One means of separating oxidation and reduction half-reactions is with a *porous barrier*, or *salt bridge*. This barrier prevents the metal atoms of one half-reaction from mixing with the ions of the other half-reaction. Ions in the two solutions can move through the porous barrier, which keeps a charge from building up on the electrodes. Electrons can be transferred from one side to the other through an external connecting wire. Electric current moves in a closed loop path, or *circuit*, so this movement of electrons through the wire is balanced by the movement of ions in solution.

Altering the system in **Figure 1** as just described would simply involve separating the copper and zinc, as shown in **Figure 2**. The Zn strip is in an aqueous solution of  $\text{ZnSO}_4$ . The Cu strip is in an aqueous solution of  $\text{CuSO}_4$ . Both solutions conduct electricity, so, as you learned in Chapter 12, they are classified as electrolytes. An **electrode** is a conductor used to establish electrical contact with a nonmetallic part of a circuit, such as an electrolyte. In **Figure 2**, the Zn and Cu strips are electrodes. A single electrode immersed in a solution of its ions is a **half-cell**.

### The Half-Cells

In the half-cell that contains the Zn electrode in aqueous  $\text{ZnSO}_4$  solution, the half-reaction is  $\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^-$ . The Zn metal loses two electrons to form  $\text{Zn}^{2+}$  ions in solution, and therefore oxidation is taking place in this half-cell. *The electrode where oxidation occurs is called the anode.* In the half-cell that contains the Cu electrode in aqueous  $\text{CuSO}_4$  solution, the half-reaction is  $\text{Cu}^{2+}(aq) + 2e^- \longrightarrow \text{Cu}(s)$ . In this half-reaction, the  $\text{Cu}^{2+}$  ions in solution gain electrons to become Cu solid; that is, reduction is taking place. *The electrode where reduction occurs is called the cathode.*

Recall from Chapter 19 that these two half-reactions cannot occur separately. Both oxidation and reduction must occur in an electrochemical reaction. The two half-cells taken together make an electrochemical cell. In the Zn/Cu electrochemical cell, the electrons move from the Zn electrode through the wire and down the Cu electrode to the  $\text{Cu}^{2+}$  ions at the electrode-solution interface. The  $\text{Cu}^{2+}$  ions are reduced to solid Cu, and the resulting Cu atoms attach themselves to the surface of the Cu electrode. For this reaction, charge is carried through the barrier by a combination of  $\text{Zn}^{2+}(\text{aq})$  ions moving from the anode to the cathode and the  $\text{SO}_4^{2-}(\text{aq})$  ions moving from the cathode to the anode.



## The Complete Cell

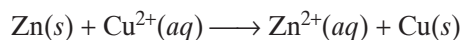
An electrochemical cell may be represented by the following notation:

anode electrode | anode solution || cathode solution | cathode electrode

The double line represents the salt bridge, or porous barrier, between the two half-cells. For the present cell, the cell notation is



The electrochemical reaction can be found by adding the anode half-reaction to the cathode half-reaction. This overall (or net) reaction is the following redox reaction:



Although the two half-reactions occur at the same time, they occur at different places in the cell. Thus, for the reaction to proceed, electrons must pass through the wire that connects the two half-cells.

An electrochemical cell that consists of the Zn and Cu reaction described above is called the *Daniell Cell*, named for the English chemist John Frederick Daniell. The Daniell Cell can generate enough electricity to light up the light bulb shown in **Figure 3**. In electrochemical cells, either a chemical reaction produces electrical energy or an electric current produces a chemical change.



**FIGURE 3** The light bulb is powered by the reaction in this cell.

## SECTION REVIEW

- Why is the use of a salt bridge or porous barrier necessary in an electrochemical cell?
- Given the  $\text{Cu}^{2+}(\text{aq}) | \text{Cu}(s)$  and  $\text{Mg}^{2+}(\text{aq}) | \text{Mg}(s)$  half-reactions, where  $\text{Cu}^{2+}(\text{aq}) | \text{Cu}(s)$  is the cathode reaction,
  - write the overall reaction.
  - write the cell notation.

- Write the half-reaction in which  $\text{I}^{-}(\text{aq})$  changes to  $\text{I}_2(s)$ . Would this reaction occur at the anode or cathode?

### Critical Thinking

- RELATING IDEAS** Is the net chemical result of an electrochemical cell a redox reaction? Explain your answer.

## SECTION 2

### OBJECTIVES

- Describe the operation of voltaic cells, including dry cells, lead-acid batteries, and fuel cells.
- Identify conditions that lead to corrosion and ways to prevent it.
- Describe the relationship between voltage and the movement of electrons.
- Calculate cell voltage/potentials from a table of standard electrode potentials.

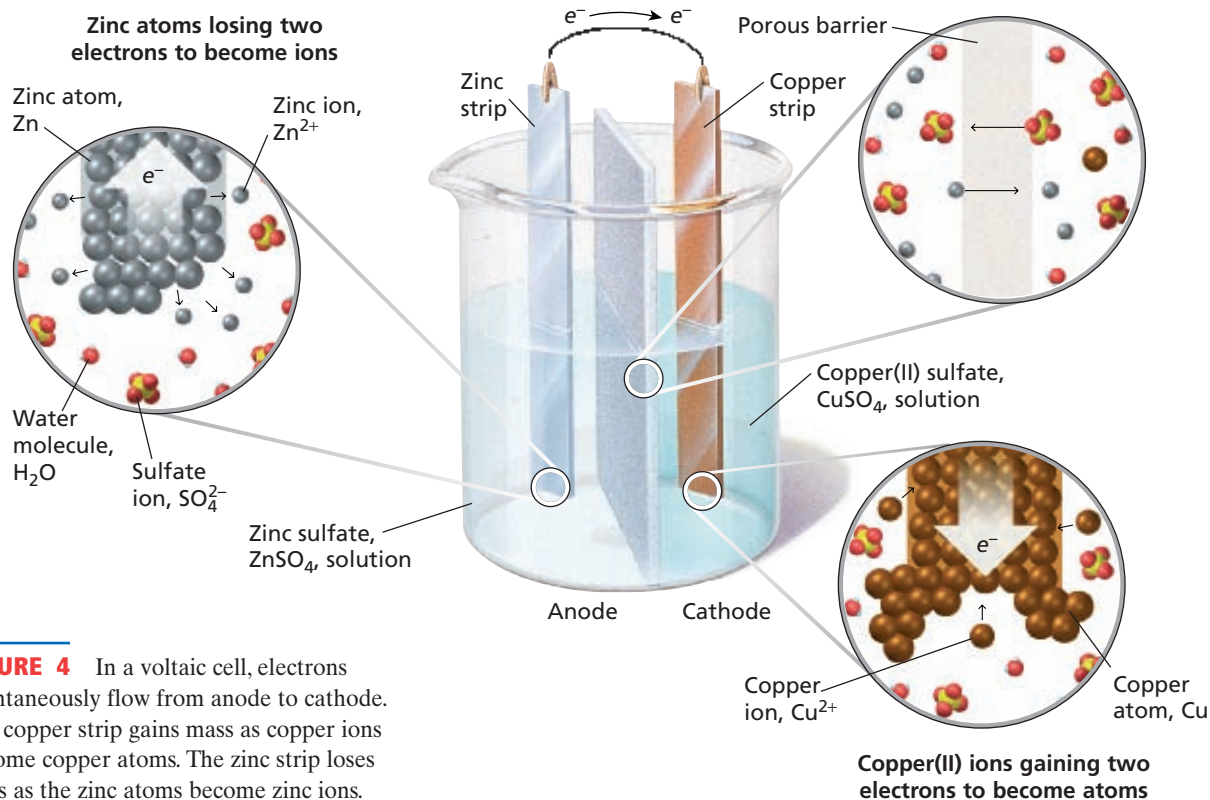
# Voltaic Cells

**V**oltaic cells use spontaneous oxidation-reduction reactions to convert chemical energy into electrical energy. Voltaic cells are also called *galvanic cells*. The most common application of voltaic cells is in batteries.

## How Voltaic Cells Work

**Figure 4** shows an example of a voltaic cell: the  $\text{Zn} \parallel \text{Cu}$  electrochemical cell discussed in the previous section.

Electrons given up at the anode pass along the external connecting wire to the cathode. The movement of electrons through the wire must be balanced by the movement of ions in the solution. Thus, in **Figure 4**, sulfate ions in the  $\text{CuSO}_4$  solution can move through the barrier into the  $\text{ZnSO}_4$  solution.

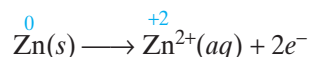


**FIGURE 4** In a voltaic cell, electrons spontaneously flow from anode to cathode. The copper strip gains mass as copper ions become copper atoms. The zinc strip loses mass as the zinc atoms become zinc ions.

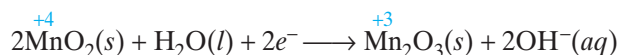
The dry cells pictured in **Figure 5** are common sources of electrical energy. Like the wet cell previously described, dry cells are voltaic cells. The three most common types of dry cells are the zinc-carbon battery, the alkaline battery, and the mercury battery. They differ in the substances being oxidized and reduced.

## Zinc-Carbon Dry Cells

Batteries such as those used in flashlights are zinc-carbon dry cells. These cells consist of a zinc container, which serves as the anode, filled with a moist paste of  $\text{MnO}_2$ , carbon black,  $\text{NH}_4\text{Cl}$ , and  $\text{ZnCl}_2$ , as illustrated in **Figure 6a**. When the external circuit is closed, zinc atoms are oxidized at the negative electrode, or anode.



Electrons move across the circuit and reenter the cell through the carbon rod. The carbon rod is the cathode or positive electrode. Here  $\text{MnO}_2$  is reduced in the presence of  $\text{H}_2\text{O}$  according to the following half-reaction.

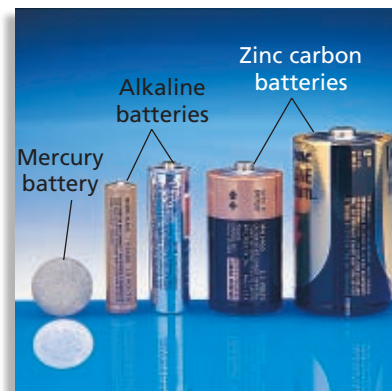
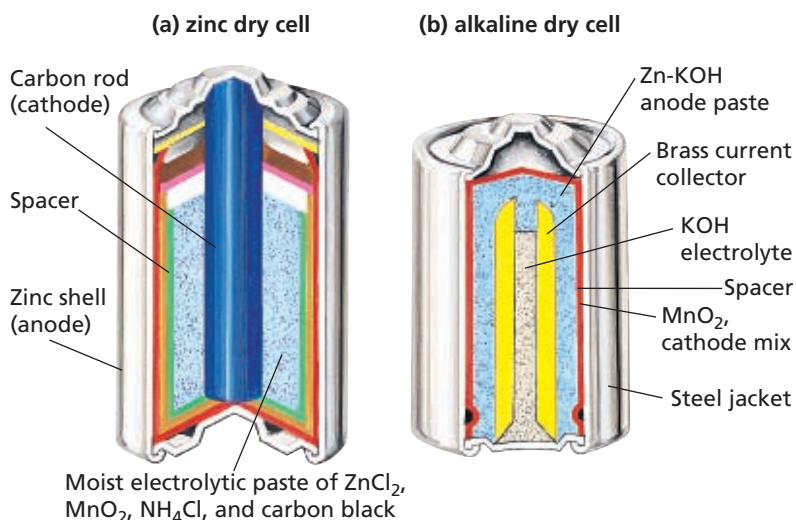


## Alkaline Batteries

The batteries found in a portable compact disc player or other small electronic device are frequently alkaline dry cells. These cells do not have a carbon rod cathode, as in the zinc-carbon cell. The absence of the carbon rod allows them to be smaller. **Figure 6b** shows a model of an alkaline battery. This cell uses a paste of Zn metal and potassium hydroxide instead of a solid metal anode. The half-reaction at the anode is as follows.



The reduction half-reaction, the reaction at the cathode, is exactly the same as that for the zinc-carbon dry cell.

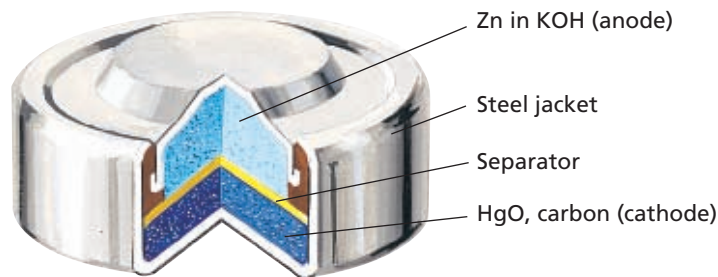


**FIGURE 5** Many common batteries are simple voltaic dry cells.



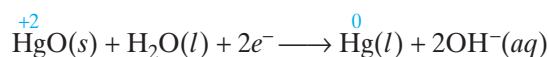
**FIGURE 6** (a) In a zinc dry cell, zinc is oxidized to  $\text{Zn}^{2+}$  at the anode, and manganese(IV) is reduced to manganese(III) at the cathode. (b) KOH makes the electrolyte paste in this battery basic. Thus, it is called an alkaline dry cell.

**FIGURE 7** It is important that mercury batteries be recycled and not just discarded because mercury is a poisonous substance.



## Mercury Batteries

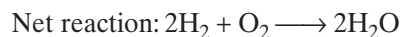
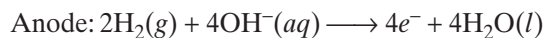
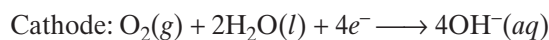
The tiny batteries found in hearing aids, calculators, and camera flashes are mercury batteries, as shown in **Figure 7**. The anode half-reaction is identical to that found in the alkaline dry cell. However, the cathode, or reduction, half-reaction is different. The cathode half-reaction is described by the following equation.



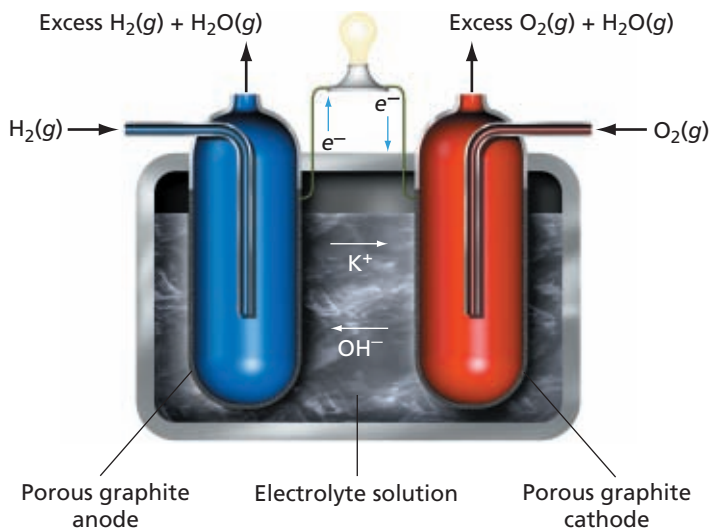
## Fuel Cells

A fuel cell is a voltaic cell in which the reactants are being continuously supplied and the products are being continuously removed. Therefore, unlike a battery, a fuel cell could, in principle, work forever, changing chemical energy into electrical energy.

Fuel cells based on the reactions listed below and shown in **Figure 8** are used in the United States space program.



Fuel cells are very efficient and have very low emissions.



**FIGURE 8** The reactions in this fuel cell take place at carbon electrodes that contain metal catalysts. The water formed is removed as a gas.

SCILINKS

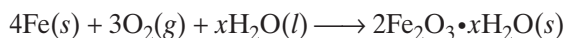
[www.scilinks.org](http://www.scilinks.org)

Topic: Fuel Cells

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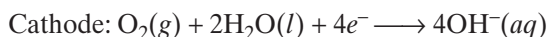
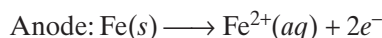
## Corrosion and Its Prevention

Corrosion is an electrochemical process that has a large economic impact. Approximately 20% of all the iron and steel produced is used to repair or replace corroded structures. One of the metals most commonly affected by corrosion is iron. Rust, hydrated iron(III) oxide, forms by the following overall reaction.



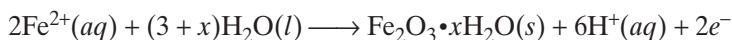
The amount of hydration of the iron oxide, reflected by the coefficient  $x$  in the equation above, varies, and affects the color of the rust formed.

The mechanism for the corrosion of iron contains the following electrochemical reactions.

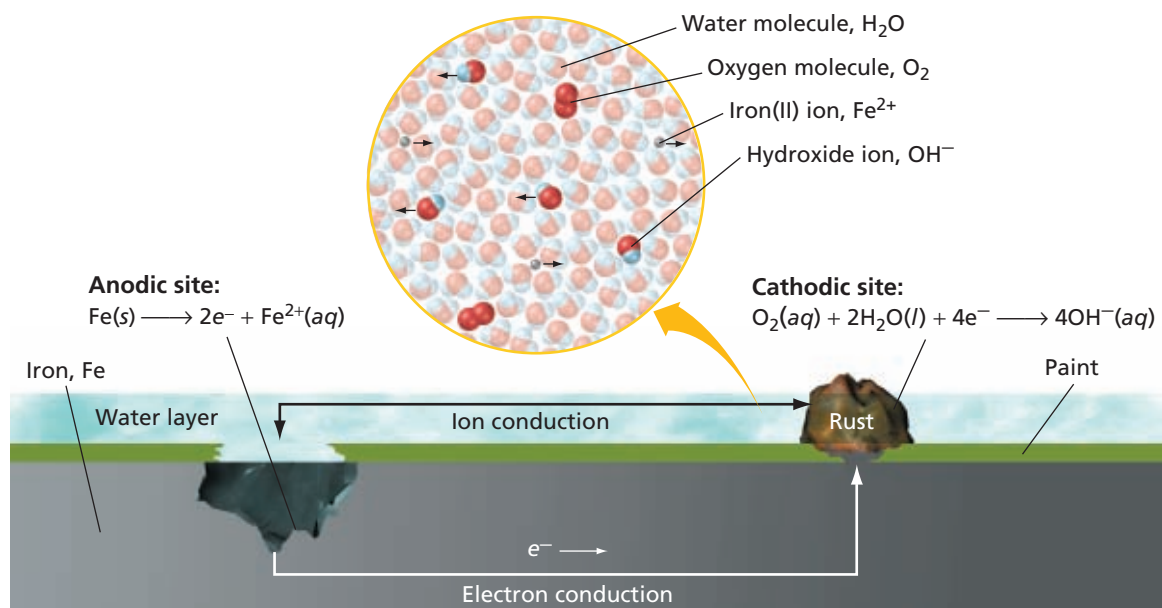


The anode and cathode reactions occur at different regions of the metal surface. The electric circuit is completed by electron flow through the metal itself, which acts like the wire in an electrochemical cell. The water on the surface of the metal serves as the salt bridge. Thus, for corrosion to occur, water and oxygen must be present with the iron.

As shown in **Figure 9**, when the iron is exposed to water and oxygen, the iron metal at the anodic site is oxidized to  $\text{Fe}^{2+}$  ions. The electrons released at this site travel along the metal (like the wire of a cell) to the cathodic region, where oxygen is reduced. The  $\text{Fe}^{2+}$  ions travel along the moisture toward the cathodic regions. At the cathode, the  $\text{Fe}^{2+}$  ions are further oxidized to  $\text{Fe}^{3+}$  ions and form rust, which is hydrated iron oxide,  $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$ .



**FIGURE 9** The cathodic reaction happens where the  $\text{O}_2$  concentration is high. The anodic reaction happens in a region where the  $\text{O}_2$  concentration is low, such as a pit in the metal.







**FIGURE 10** The Alaskan pipeline is cathodically protected by a parallel zinc cable.

### extension

#### Careers in Chemistry

Go to [go.hrw.com](http://go.hrw.com) for a full-length article on careers in chemical engineering.



Keyword: HC6ELEX

The presence of salt or high acidity speed the corrosion process, because the movement of electrons is facilitated by the extra ions present.

How can corrosion be prevented? One way, which makes use of electrochemical properties of metals, is to coat steel with zinc in a process called *galvanizing*. Zinc is more easily oxidized than iron; therefore, zinc will react before the iron is oxidized. This is called *cathodic protection*, and the more easily oxidized metal used is called a *sacrificial anode*.

The Alaskan oil pipeline, shown in **Figure 10**, is an example of steel that is cathodically protected. However, instead of coating, zinc is connected to the pipe by a wire. The zinc will oxidize before the iron in the steel does. As the zinc anode corrodes, it gives electrons to the cathode, the steel, and as the zinc anode does so, it prevents the steel from corroding. As the zinc dissolves, it needs to be replaced. As long as zinc metal that is able to corrode is present, the steel will be protected from corrosion.

## Electrical Potential

In a voltaic cell, the oxidizing agent at the cathode pulls the electrons through the wire away from the reducing agent at the anode. The “pull,” or driving force on the electrons, is called the *electric potential*. Electric potential, or voltage, is expressed in units of volts (V), which is the potential energy per unit charge. Current is the movement of the electrons and is expressed in units of amperes, or amps (A).

Electrical potential is analogous to gravitational potential. Just as water flows from a position of higher gravitational potential to a position of lower gravitational potential, electrons flow from higher electric potential to lower electrical potential.

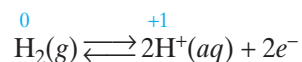
### Electrode Potentials

Reconsider the voltaic cell shown in **Figure 4**. There are two electrodes, Zn and Cu. These two metals each have different tendencies for accepting electrons. *This tendency for the half-reaction of either copper or zinc to occur as a reduction half-reaction in an electrochemical cell can be quantified as a reduction potential.* There are two half-cells in **Figure 4**: a strip of zinc placed in a solution of  $\text{ZnSO}_4$  and a strip of copper placed in a solution of  $\text{CuSO}_4$ . *The difference in potential between an electrode and its solution is known as electrode potential.* When these two half-cells are connected and the reaction begins, a difference in potential is observed between the electrodes. This potential difference, or voltage, is proportional to the energy required to move a certain electric charge between the electrodes. A voltmeter connected across the  $\text{Zn}||\text{Cu}$  voltaic cell measures a potential difference of about 1.10 V when the solution concentrations of  $\text{Zn}^{2+}$  and  $\text{Cu}^{2+}$  ions are each 1 M.

The potential difference measured across the complete voltaic cell is easily measured and equals the sum of the electrode potentials for the two half-reactions. But an individual electrode potential cannot be meas-

ured directly. This is because there can be no transfer of electrons unless both the anode and the cathode are connected to form a complete circuit. A relative value for the potential of a half-reaction can be determined by connecting it to a standard half-cell as a reference. This standard half-cell, shown in **Figure 11**, is called a standard hydrogen electrode, or SHE. It consists of a platinum electrode dipped into a 1.00 M acid solution surrounded by hydrogen gas at 1 atm pressure and 25°C. Other electrodes are ranked according to their ability to reduce hydrogen under these conditions.

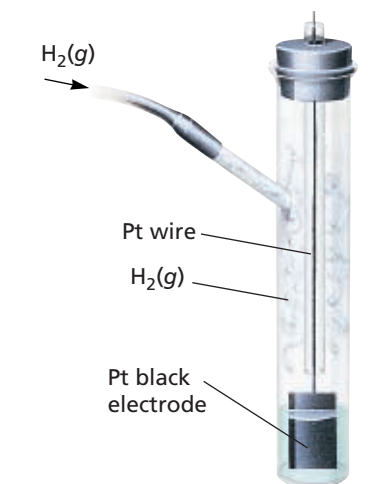
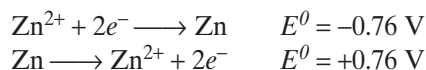
The anodic reaction for the standard hydrogen electrode is described by the forward half-reaction in the following equilibrium equation.



The cathodic half-reaction is the reverse. An arbitrary potential of 0.00 V is assigned to both of these half-reactions. *The potential of a half-cell under standard conditions measured relative to the standard hydrogen electrode is a **standard electrode potential**,  $E^0$ .* Electrode potentials are expressed as *potentials for reduction*. These reduction potentials provide a reliable indication of the tendency of a substance to be reduced. **Figure 12** shows how the SHE is used to find the electrode potentials of the zinc and copper half-cells. Half-reactions for some common electrodes and their standard electrode potentials are listed in **Table 1** on the next page.

Effective oxidizing agents, such as  $\text{Cu}^{2+}$  and  $\text{F}_2$ , have positive  $E^0$  values. Half-reactions with negative reduction potentials prefer oxidation over reduction. Negative  $E^0$  values indicate that the metal or other electrode is more willing to give up electrons than hydrogen. Effective reducing agents, such as Li and Zn, have negative  $E^0$  values.

When a half-reaction is written as an oxidation reaction, the sign of its electrode potential is reversed, as shown for the oxidation and reduction half-reactions for zinc.



**FIGURE 11** A hydrogen electrode is the standard reference electrode for measuring electrode potentials. The electrode surface in contact with the solution is actually a layer of hydrogen adsorbed onto the surface of the platinum.

**FIGURE 12** The electrode potentials of zinc and copper half-cells are measured by coupling them with a standard hydrogen electrode.

**TABLE 1** Standard Reduction Potentials

Half-cell reaction	Standard electrode potential, $E^0$ (in volts)	Half-cell reaction	Standard electrode potential, $E^0$ (in volts)
$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87	$Fe^{3+} + 3e^- \rightleftharpoons Fe$	-0.04
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.49	$Pb^{2+} + 2e^- \rightleftharpoons Pb$	-0.13
$Au^{3+} + 3e^- \rightleftharpoons Au$	+1.42	$Sn^{2+} + 2e^- \rightleftharpoons Sn$	-0.14
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36	$Ni^{2+} + 2e^- \rightleftharpoons Ni$	-0.23
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33	$Co^{2+} + 2e^- \rightleftharpoons Co$	-0.28
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.21	$Cd^{2+} + 2e^- \rightleftharpoons Cd$	-0.40
$Br_2 + 2e^- \rightleftharpoons 2Br^-$	+1.07	$Fe^{2+} + 2e^- \rightleftharpoons Fe$	-0.41
$Hg^{2+} + 2e^- \rightleftharpoons Hg$	+0.85	$S + 2e^- \rightleftharpoons S^{2-}$	-0.51
$Ag^+ + e^- \rightleftharpoons Ag$	+0.80	$Cr^{3+} + 3e^- \rightleftharpoons Cr$	-0.74
$Hg_2^{2+} + 2e^- \rightleftharpoons 2Hg$	+0.80	$Zn^{2+} + 2e^- \rightleftharpoons Zn$	-0.76
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77	$Al^{3+} + 3e^- \rightleftharpoons Al$	-1.66
$MnO_4^- + e^- \rightleftharpoons MnO_4^{2-}$	+0.56	$Mg^{2+} + 2e^- \rightleftharpoons Mg$	-2.37
$I_2 + 2e^- \rightleftharpoons 2I^-$	+0.54	$Na^+ + e^- \rightleftharpoons Na$	-2.71
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34	$Ca^{2+} + 2e^- \rightleftharpoons Ca$	-2.76
$Cu^{2+} + e^- \rightleftharpoons Cu^+$	+0.16	$Ba^{2+} + 2e^- \rightleftharpoons Ba$	-2.90
$S + 2H^+(aq) + 2e^- \rightleftharpoons H_2S(aq)$	+0.14	$K^+ + e^- \rightleftharpoons K$	-2.93
$2H^+(aq) + 2e^- \rightleftharpoons H_2$	0.00	$Li^+ + e^- \rightleftharpoons Li$	-3.04

The potential difference across the zinc/hydrogen cell is  $-0.76$  V, so zinc is considered to have an electrode potential of  $-0.76$ . The negative number indicates that electrons flow from the zinc electrode, where zinc is oxidized, to the hydrogen electrode, where aqueous hydrogen ions are reduced.

A copper half-cell coupled with the standard hydrogen electrode gives a potential difference measurement of  $+0.34$  V. This positive number indicates that  $Cu^{2+}(aq)$  ions are more readily reduced than  $H^+(aq)$  ions.

Standard electrode potentials can be used to predict if a redox reaction will occur spontaneously. A spontaneous reaction will have a positive value for  $E_{cell}^0$ , which is calculated using the following equation.

$$E_{cell}^0 = E_{cathode}^0 - E_{anode}^0$$

The half-reaction that has the more negative standard reduction potential will be the anode. Oxidation occurs at the anode, so the anode half-cell reaction will be the reverse of the reduction reaction found in **Table 1**. For this reason, the total potential of a cell is calculated by *subtracting* the standard reduction potential for the reaction at the anode ( $E_{anode}^0$ ) from the standard reduction potential for the reaction at the cathode ( $E_{cathode}^0$ ).

**SAMPLE PROBLEM A**

For more help, go to the *Math Tutor* at the end of this chapter.

Write the overall cell reaction, and calculate the cell potential for a voltaic cell consisting of the following half-cells: an iron (Fe) electrode in a solution of  $\text{Fe}(\text{NO}_3)_3$  and a silver (Ag) electrode in a solution of  $\text{AgNO}_3$ .

**SOLUTION**

- 1 ANALYZE** Given: A half-cell consists of  $\text{Fe}(s)$  with  $\text{Fe}(\text{NO}_3)_3(aq)$  and a second half-cell consists of  $\text{Ag}(s)$  with  $\text{AgNO}_3(aq)$ .  
Unknown:  $E^0_{\text{cell}}$
- 2 PLAN** 1. Look up  $E^0$  for each half-reaction (written as reductions) in Table 1.  

$$\text{Fe}^{3+}(aq) + 3e^- \longrightarrow \text{Fe}(s) \quad E^0 = -0.04 \text{ V} \quad \text{Ag}^+(aq) + e^- \longrightarrow \text{Ag}(s) \quad E^0 = +0.80 \text{ V}$$
 2. Determine the cathode and anode.  
 Fe in  $\text{Fe}(\text{NO}_3)_3$  is the anode because it has a lower reduction potential than Ag. Ag in  $\text{Ag}(\text{NO}_3)$  is therefore the cathode.
- 3 COMPUTE** 1. Determine the overall cell reaction. Multiply the Ag half-reaction by 3 so that the number of electrons lost in that half-reaction equals the number of electrons gained in the oxidation of iron. Reverse the iron half-reaction to be an oxidation half-reaction.  

$$3\text{Ag}^+(aq) + \text{Fe}(s) \longrightarrow 3\text{Ag}(s) + \text{Fe}^{3+}(aq)$$
 2. Calculate the cell potential by  $E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}}$ . Note that when a half-reaction is multiplied by a constant, the  $E^0$  value is not multiplied by that constant but remains the same.  

$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}} = +0.80 \text{ V} - (-0.04 \text{ V}) = +0.84 \text{ V}$$
- 4 EVALUATE** The calculated value for  $E^0_{\text{cell}}$  is positive, which confirms that it is a voltaic cell, as the problem states.

**PRACTICE**

Answers in Appendix E

- For each pair of half-cells, determine the overall electrochemical reaction that proceeds spontaneously and the  $E^0$  value.
  - $\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}$  and  $\text{Ni}^{2+}/\text{Ni}$
  - SHE and  $\text{Fe}^{2+}/\text{Fe}^{3+}$

**extension**

Go to [go.hrw.com](http://go.hrw.com) for more practice problems that ask you to calculate cell potential.



Keyword: HC6ELEX

**SECTION REVIEW**

- What is a voltaic cell?
- What is electrode potential, and how is it used to calculate information about an electrochemical cell?
- Given the  $\text{Na}^+/\text{Na}$  and  $\text{K}^+/\text{K}$  half-cells, determine the overall electrochemical reaction that proceeds spontaneously and the  $E^0$  value.
- Given the  $\text{MnO}_2/\text{Mn}^{2+}$  and  $\text{Cr}^{3+}/\text{Cr}$  half-cells, determine the overall electrochemical reaction that occurs spontaneously and the  $E^0$  value.

**Critical Thinking**

- EVALUATING IDEAS** A sacrificial anode is allowed to corrode. Why is use of a sacrificial anode considered to be a way to prevent corrosion?

# Chemistry in Action



## Fuel-Cell Cars

The car of the future is quiet, has low emissions, and requires less maintenance than cars of today do. These are the promises of fuel-cell cars. The development of fuel-cell cars has been pushed by state and federal programs, legislation, and funding. There are a number of issues that need to be resolved before the large-scale manufacture of fuel-cell cars becomes feasible. These include the fuel-storage capacity of the car, and the fueling infrastructure.

The technology currently favored for automobiles is the proton-exchange membrane (PEM) fuel cell. The system works by combining hydrogen with oxygen taken from the outside air. The movement of the electrons (produced by the redox reaction) along an external route generates electricity, which is used eventually to turn the car wheels. The process is two to three times more efficient than that of the internal-combustion engine, and the only products are water and a moderate amount of energy as heat. In most cases, a traditional car battery is used to operate the car's electrical system.

In principle, fuel cells can run on any hydrogen-rich fuel. Several candidate fuels that can power fuel cells are being investigated: methanol, ethanol, natural gas, gasoline, and liquid or compressed hydrogen. If the fuel is not hydrogen gas, the engine needs to include a reformer that can convert the other fuel (such as methanol) to hydrogen.



▲ Eventually, fuel-cell cars such as this one may replace gasoline-powered cars.

The range of fuel-cell powered cars is limited by the amount of fuel that can be stored in the storage tank. Current storage now limits the cars to around 100 miles before refueling. Researchers in this area are looking into carbon-adsorption systems, which are refrigerated and pressurized tanks that can store large amounts of hydrogen. Hydrogen is very combustible, but safe hydrogen tanks have been developed. So today, the experts do not have any more concern about the safety of hydrogen storage than they do about gasoline storage.

However, the landscape is not yet dotted with hydrogen refueling stations. Current experimental users such as governmental institutions must have their own hydrogen refueling stations. Auto companies say there needs to be a consumer infrastructure for the hydrogen fuel-cell cars before consumers will buy. On the other hand, possible developers of such hydrogen stations say that there needs to be a demand before the infrastructure can form.

The use of fuels such as methanol solves the storage and fuel infrastructure problem, because these fuels can be stored similar to the way gasoline is stored. Therefore, these fuels can be sold at current gasoline stations.

One should also consider the emissions and energy use involved in making and delivering the hydrogen fuel in addition to the emissions from the hydrogen fuel cell. Producing hydrogen fuel itself could involve substantial emissions. Even so, fuel-cell cars are still more efficient and have greater environmental benefits than the internal-combustion cars of today.

Currently, all major car companies are developing and researching fuel-cell cars. Most companies have prototype cars and some cars are in use by governmental institutions. Which type will ultimately succeed in the mass market is yet unknown. Experts agree, though, that for zero-emission cars, hydrogen is the only major fuel option that has been identified to date.

### Questions

1. Fuel cells that use hydrogen as fuel produce water and moderate amounts of energy in the form of heat as the only emissions. What kinds of negative effects might cars that use hydrogen-rich fuels have on the environment?
2. In the PEM fuel cell, the hydrogen ions are produced at which electrode: the anode or cathode?

# Electrolytic Cells

## SECTION 3

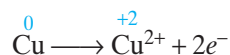
### OBJECTIVES

- Describe the nature of electrolytic cells.
- Describe the process of electrolysis in the decomposition of water and in production of metals.
- Explain the process of electroplating.
- Describe the chemistry of a rechargeable cell.

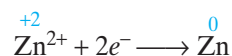
Some oxidation-reduction reactions do not occur spontaneously but can be driven by electrical energy. *If electrical energy is required to produce a redox reaction and bring about a chemical change in an electrochemical cell, it is an **electrolytic cell**.* Most commercial uses of redox reactions make use of electrolytic cells.

## How Electrolytic Cells Work

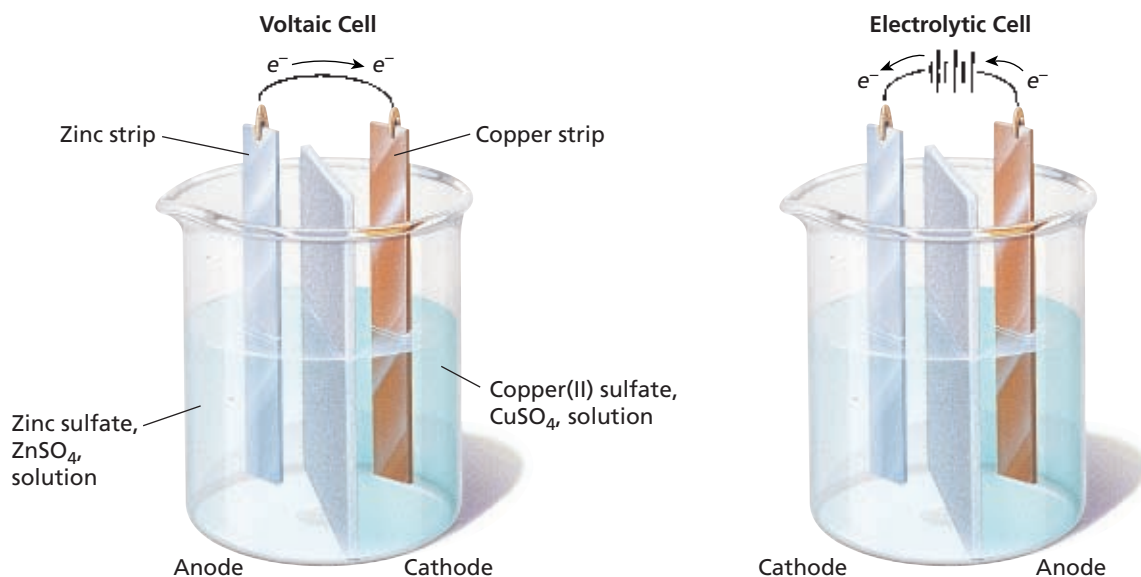
A comparison of electrolytic and voltaic cells can be seen in **Figure 13**. The voltaic cell shown in **Figure 13** has a copper cathode and a zinc anode. If a battery is connected so that the positive terminal contacts the copper electrode and the negative terminal contacts the zinc electrode, the electrons move in the opposite direction. The battery forces the cell to reverse its reaction; the zinc electrode becomes the cathode, and the copper electrode becomes the anode. The half-reaction at the anode, in which copper metal is oxidized, can be written as follows:



The reduction half-reaction of zinc at the cathode is written as follows:



**FIGURE 13** The direction in which the electrons move reverses if a voltaic cell is connected to a direct current source to become an electrolytic cell.



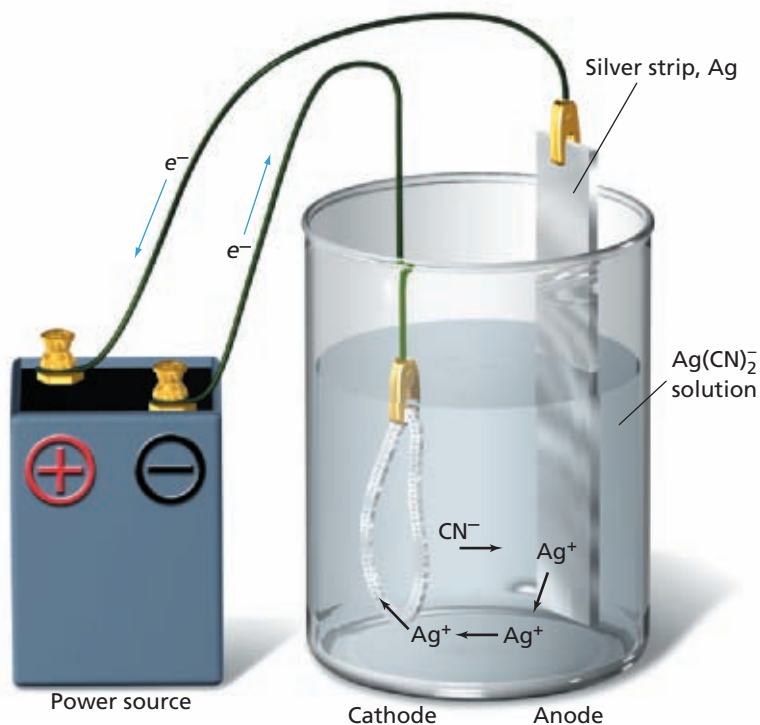
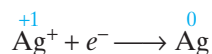
There are two important differences between the voltaic cell and the electrolytic cell.

1. The anode and cathode of an electrolytic cell are connected to a battery or other direct-current source, whereas a voltaic cell serves as a source of electrical energy.
2. Electrolytic cells are those in which electrical energy from an external source causes *nonspontaneous* redox reactions to occur. Voltaic cells are those in which *spontaneous* redox reactions produce electricity. In an electrolytic cell, electrical energy is converted to chemical energy; in a voltaic cell, chemical energy is converted to electrical energy.

## Electroplating

An electrolytic process in which a metal ion is reduced and a solid metal is deposited on a surface is called **electroplating**. An electroplating cell contains a solution of a salt of the plating metal, an object to be plated (the cathode), and a piece of the plating metal (the anode). A silver-plating cell contains a solution of a soluble silver salt and a silver anode. The cathode is the object to be plated. The silver anode is connected to the positive electrode of a battery or to some other source of direct current. The object to be plated is connected to the negative electrode.

A cell in which silver is being electroplated onto a bracelet can be seen in **Figure 14**. Silver ions are reduced at the cathode according to the following half-reaction and deposited as metallic silver when electrons flow through the circuit.



**FIGURE 14** The bracelet in this cell is being coated with a thin layer of silver. Silver ions are replaced in the solution as the pure silver anode dissolves.

SCILINKS

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Topic: Electroplating

Code: HC60492

Meanwhile, metallic silver is oxidized at the anode according to the following half-reaction.

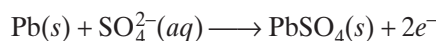


In effect, silver is transferred from the anode to the cathode of the cell.

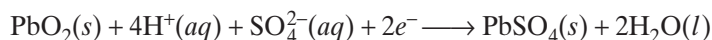
## Rechargeable Cells

A rechargeable cell combines the oxidation-reduction chemistry of both voltaic cells and electrolytic cells. When a rechargeable cell converts chemical energy to electrical energy, it operates as a voltaic cell. But when the cell is recharged, it operates as an electrolytic cell, converting electrical energy to chemical energy.

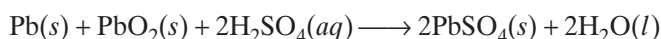
The standard 12 V automobile battery, shown in **Figure 15**, is a set of six rechargeable cells. The anode in each cell is lead submerged in a solution of  $\text{H}_2\text{SO}_4$ . The anode half-reaction is described below.



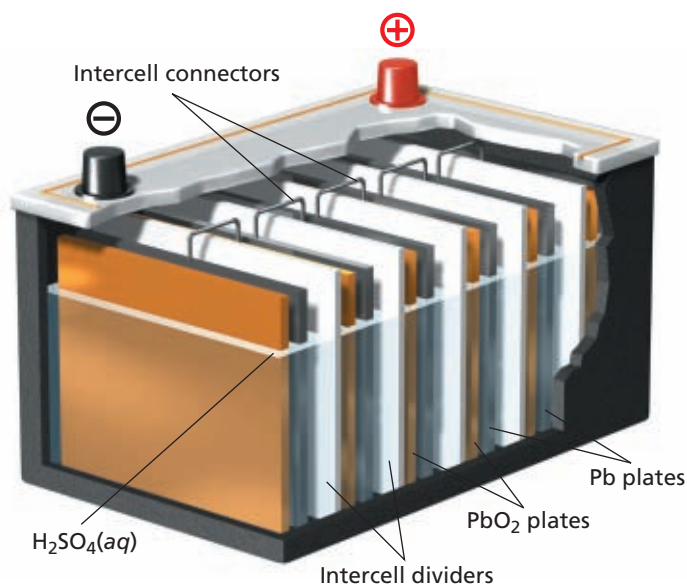
At the cathode,  $\text{PbO}_2$  is reduced according to the following equation.



The net oxidation-reduction reaction for the discharge cycle is:



A car's battery produces the electric energy needed to start its engine. Sulfuric acid, present as its ions, is consumed, and lead(II) sulfate accumulates as a white powder on the electrodes. Once the car is running, the half-reactions are reversed by a voltage produced by the alternator. The  $\text{Pb}$ ,  $\text{PbO}_2$ , and  $\text{H}_2\text{SO}_4$  are regenerated. A battery can be recharged as long as all reactants necessary for the electrolytic reaction are present, and all reactions are reversible.



**FIGURE 15** The rechargeable cells of a car battery produce electricity from reactions between lead(IV) oxide, lead, and sulfuric acid.



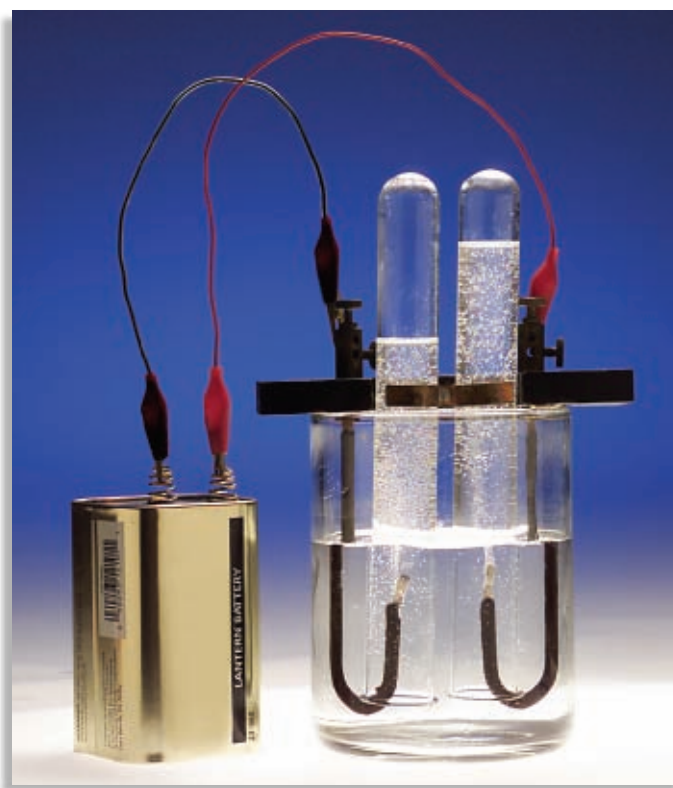
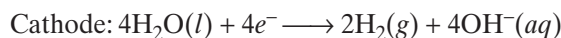
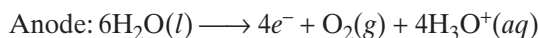
## Electrolysis

Electroplating and recharging a battery are examples of electrolysis. **Electrolysis** is the process of passing a current through a cell for which the cell potential is negative and causing an oxidation-reduction reaction to occur. That is, electrical energy is used to force a nonspontaneous chemical reaction to occur. For this cell reaction to occur, the external voltage must be greater than the potential that would be produced by the spontaneous reverse cell reaction.

Electrolysis is of great industrial importance. It is used to purify many metals from the ores in which they are found chemically combined in Earth's crust.

### Electrolysis of Water

The electrolysis of water, shown in **Figure 16**, leads to the cell reaction in which water is broken down into its elements,  $\text{H}_2$  and  $\text{O}_2$ . Recall that hydrogen gas and oxygen gas combine spontaneously to form water and are used to power fuel cells, which produce electricity. Therefore, the reverse process (electrolysis of water) is nonspontaneous and requires electrical energy. The following half-reactions occur at the anode and cathode.



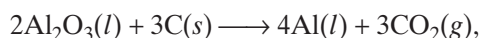
**FIGURE 16** Electrical energy from the battery is used to break down water. Hydrogen forms at the cathode (left tube), and oxygen forms at the anode (right tube).

## Aluminum Production by Electrolysis

Aluminum is the most abundant metal in Earth's crust. It is a relatively reactive metal; therefore, in nature, it is found as its oxide in an ore called *bauxite*. Aluminum is now very useful commercially, but it was not until 1886 that a process to obtain pure aluminum metal was discovered. Charles M. Hall (from the United States) and Paul Héroult (from France) simultaneously, but independently, determined a practical method for producing aluminum; it is an electrolytic process called the Hall-Héroult process.

Bauxite ore contains not only aluminum oxide ( $\text{Al}_2\text{O}_3$ ), but oxides of iron, silicon, and titanium. Therefore, the aluminum oxide (called alumina) must first be separated from the other compounds in the ore. The pure hydrated alumina ( $\text{Al}_2\text{O}_3 \cdot n\text{H}_2\text{O}$ ) is obtained by treating bauxite with sodium hydroxide, which dissolves the alumina but does not dissolve the other compounds in the ore. The alumina solution is then separated from the remaining solid compounds and reprecipitated to obtain pure alumina. The purified alumina is dissolved in molten cryolite,  $\text{Na}_3\text{AlF}_6$ , at  $970^\circ\text{C}$  in an electrochemical cell, and the aluminum ions are reduced to aluminum metal. The liquid aluminum is denser than the molten cryolite and alumina; therefore, the molten aluminum metal settles to the bottom of the cell and is drained off periodically.

The electrolytic solution contains a large number of aluminum-containing ions, and the chemistry of the electrochemical reaction is not completely understood. Scientists still debate the exact species that participate in the half-reactions; but the overall cell reaction is



where carbon is the anode and steel is the cathode in the cell.

The aluminum metal produced in this process is 99.5% pure. The Hall-Héroult process made the production of aluminum economically feasible. However, this process is the largest single user of electrical energy in the United States—nearly 5% of the national total. Recycling aluminum saves almost 95% of the cost of production. Aluminum recycling is one of the most economically worthwhile recycling programs that has been developed.

### Chemistry in Action

#### Sodium Production by Electrolysis

Sodium is such a reactive metal that preparing it through a chemical process can be dangerous. Today, sodium is produced mainly by the electrolysis of molten sodium chloride. The melting point of sodium chloride is about  $800^\circ\text{C}$ ; but mixing it with solid calcium chloride lowers the melting point to around  $600^\circ\text{C}$ . This mixture is introduced into an electrolytic cell called the Downs Cell to retrieve the sodium metal, and the lower melting point allows the cell to use less electrical energy to run.

## SECTION REVIEW

1. Describe an electrolytic cell.
2. Explain the process of electroplating.
3. What is a rechargeable cell?
4. Give an example of how electrolytic cells are used in industry.

### Critical Thinking

5. **APPLYING CONCEPTS** Copper ore contains zinc metal, which is oxidized along with Cu during the electrolytic purification process. However, the  $\text{Zn}^{2+}$  ions are not then reduced when the  $\text{Cu}^{2+}$  ions are reduced to Cu at the cathode to obtain purified copper metal. Explain how Zn can be oxidized with Cu, but their ions not be reduced together.

## Introduction to Electrochemistry

### Vocabulary

electrochemistry  
electrode  
half-cell  
anode  
cathode

- Electrochemistry is the branch of chemistry that deals with electricity-related applications of redox reactions.
- The electrode immersed in an electrolyte solution is a half-cell.
- The anode is the electrode where oxidation takes place. The cathode is the electrode where reduction occurs.
- The cell consists of electrodes connected by a wire along which the electron travels and a salt bridge (or porous barrier) through which ions transfer to balance the charge.
- An electrochemical cell is a system of electrodes and electrolytes in which either chemical reactions produce electrical energy or electric current produces chemical change.

## Voltaic Cells

### Vocabulary

voltaic cell  
reduction potential  
electrode potential  
standard electrode potential

- A voltaic cell, sometimes called a galvanic cell, uses a spontaneous redox reaction to produce electrical energy. Examples of voltaic cells are batteries and fuel cells.
- Fuel cells are voltaic cells in which the reactants are continuously supplied and the products are continuously removed.
- The potential difference must be measured across a complete cell because no transfer of electrons can occur unless both the anode and cathode are connected to form a complete circuit. Thus, the standard electrode potential for a half-cell is measured against the standard hydrogen electrode (SHE).
- Standard reduction potentials,  $E^0$ , are stated as reduction half-reactions. Effective oxidizing agents have positive  $E^0$  values, while effective reducing agents have negative  $E^0$  values.
- A voltaic cell has an  $E_{cell}^0$  value that is positive.
- Corrosion occurs when iron is exposed to oxygen and water. One of the best methods to prevent corrosion is by the use of sacrificial anodes.

## Electrolytic Cells

### Vocabulary

electrolytic cell  
electroplating  
electrolysis

- Electrolytic cells are cells in which electrical energy from an external source causes a nonspontaneous reaction to occur.
- An electrolytic cell has an  $E_{cell}^0$  value that is negative.
- Electrolysis has great economic impact. Applications of electrolytic cells are electroplating of metallic surfaces, rechargeable batteries, aluminum production, and purification of metals.

# CHAPTER REVIEW

For more practice, go to the Problem Bank in Appendix D.

## Introduction to Electrochemistry

### SECTION 1 REVIEW

1. In the half-cell  $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \longrightarrow \text{Zn}(\text{s})$ , what is the electrode and is this half-reaction an anodic reaction or a cathodic reaction?
2. What role does the porous barrier play?
3. For each of the following pairs of half-cells, write the overall reaction and the cell notation. Assume the first half-cell given in each pair is the cathodic half-cell.
  - a.  $\text{Ag}^{+}/\text{Ag}$ ,  $\text{Co}^{2+}/\text{Co}$
  - b.  $\text{Au}^{3+}/\text{Au}$ ,  $\text{Zn}^{2+}/\text{Zn}$
  - c.  $\text{Hg}^{2+}/\text{Hg}$ ,  $\text{K}^{+}/\text{K}$
4. Describe the components of an electrochemical cell and how the electrical charge travels through these components.

## Voltaic Cells

### SECTION 2 REVIEW

5. Describe a voltaic cell, and give two examples of a voltaic cell.
6. What is the essential advantage of a fuel cell over batteries in the generation of electrical energy?
7. Explain why corrosion is a voltaic cell.
8. Discuss the advantages and disadvantages of corrosion-prevention methods.
9. Which half reaction would more likely be an oxidation reaction: one with a standard reduction potential of  $-0.42\text{ V}$ , or one with a standard reduction potential of  $+0.42\text{ V}$ ?
10. Why are dry-cell batteries called *dry cells*, even though their chemistry involves water?
11.
  - a. Explain what is meant by the potential difference between the two electrodes in an electrochemical cell.
  - b. How is this potential difference measured? What units are used?
12. The standard hydrogen electrode is assigned an electrode potential of  $0.00\text{ V}$ . Explain why this voltage is assigned.

13.
  - a. What information is provided by the standard reduction potential of a given half-cell?
  - b. What does the relative value of the reduction potential of a given half-reaction indicate about its oxidation-reduction tendency?
14. When the cell  $\text{Ba}(\text{s}) \mid \text{Ba}^{2+}(\text{aq}) \parallel \text{Sn}^{2+}(\text{aq}) \mid \text{Sn}(\text{s})$  is running, what observations can be made?

### PRACTICE PROBLEMS

15. For each of the following pairs of half-cells, determine the overall electrochemical reaction that proceeds spontaneously:
  - a.  $\text{Na}^{+}/\text{Na}$ ,  $\text{Ni}^{2+}/\text{Ni}$
  - b.  $\text{F}_2/\text{F}^{-}$ ,  $\text{S}/\text{H}_2\text{S}$
  - c.  $\text{Br}_2/\text{Br}^{-}$ ,  $\text{Cr}^{3+}/\text{Cr}$
  - d.  $\text{MnO}_4^{-}/\text{Mn}^{2+}$ ,  $\text{Co}^{2+}/\text{Co}$
16. Determine the values of  $E^{\circ}$  for the cells in the previous problem.
17. Suppose chemists had chosen to make the  $\text{I}_2 + 2\text{e}^{-} \longrightarrow 2\text{I}^{-}$  half-cell the standard electrode and had assigned it a potential of zero volts.
  - a. What would be the  $E^{\circ}$  value for the  $\text{Br}_2 + 2\text{e}^{-} \longrightarrow 2\text{Br}^{-}$  half-cell?
  - b. What would be the  $E^{\circ}$  value for the  $\text{Al}^{3+} + 3\text{e}^{-} \longrightarrow \text{Al}$  half-cell?
  - c. How much change would be observed in the  $E^{\circ}$  value for the reaction involving  $\text{Br}_2 + \text{I}^{-}$  if the  $\text{I}_2$  half-cell is the standard?
18. If a strip of Ni were dipped into a solution of  $\text{AgNO}_3$ , what would be expected to occur? Explain, using  $E^{\circ}$  values and equations.

## Electrolytic Cells

### SECTION 3 REVIEW

19. What reaction happens at the cathode in an electrolysis process?
20. Explain why water cannot be used in the electrochemical cell during the production of aluminum.
21. Calculate the voltage of a cell in which the overall reaction is the electrolysis of aqueous cadmium chloride into its elements.
22. According to electrochemical data, can Ni be plated onto a zinc metal object using a nickel nitrate solution? Explain.

23. Distinguish between a voltaic cell and an electrolytic cell in terms of the nature of the reaction involved.
24. a. What is electroplating?  
b. Identify the anode and cathode in such a process.

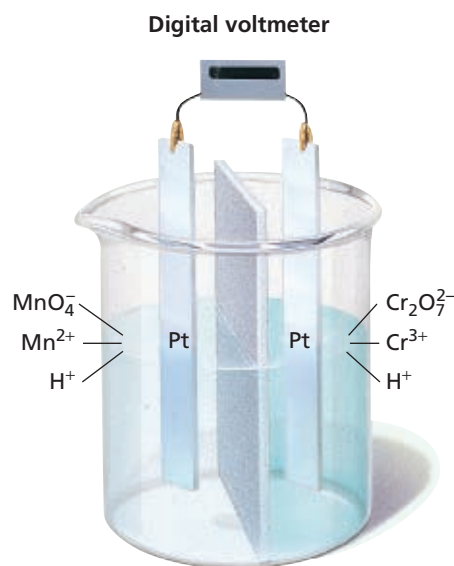
## MIXED REVIEW

25. Predict whether each of the following reactions will occur spontaneously as written by determining the  $E^0$  value for potential reaction. Write and balance the overall equation for each reaction that does occur spontaneously.
- $\text{Mg} + \text{Sn}^{2+} \longrightarrow$
  - $\text{K} + \text{Al}^{3+} \longrightarrow$
  - $\text{Li}^+ + \text{Zn} \longrightarrow$
  - $\text{Cu} + \text{Cl}_2 \longrightarrow$
26. Why is it possible for alkaline batteries to be smaller than zinc-carbon dry cells?
27. Draw a diagram of a voltaic cell whose two half-reactions consist of Ag in  $\text{AgNO}_3$  and Ni in  $\text{NiSO}_4$ . Identify the anode and cathode, and indicate the directions in which the electrons and ions are moving.
28. Can a solution of  $\text{Sn}(\text{NO}_3)_2$  be stored in an aluminum container? Explain, using  $E^0$  values.
29. A voltaic cell is made up of a cadmium electrode in a solution of  $\text{CdSO}_4$  and a zinc electrode in a solution of  $\text{ZnSO}_4$ . The two half-cells are separated by a porous barrier.
- Which is the cathode, and which is the anode?
  - In which direction are the electrons flowing?
  - Write balanced equations for the two half-reactions, and write a net equation for the combined reaction.
30. Would the following pair of electrodes make a good battery? Explain.
- $$\text{Cd} \longrightarrow \text{Cd}^{2+} + 2e^-$$
- $$\text{Fe} \longrightarrow \text{Fe}^{2+} + 2e^-$$
31. a. What would happen if an aluminum spoon were used to stir a solution of  $\text{Zn}(\text{NO}_3)_2$ ?  
b. Could a strip of Zn be used to stir a solution of  $\text{Al}(\text{NO}_3)_3$ ? Explain, using  $E^0$  values.

32. How do the redox reactions for each of the following types of batteries differ?
- zinc-carbon
  - alkaline
  - mercury
33. a. Why are some standard reduction potentials positive and some negative?  
b. Compare the  $E^0$  value for a metal with the reactivity of that metal.

## CRITICAL THINKING

34. **Applying Models** Explain how the oxidation-reduction chemistry of both the voltaic cell and the electrolytic cell are combined in the chemistry of rechargeable cells.
35. **Applying Ideas** In lead-acid batteries, such as your car battery, the degree of discharge of the battery can be determined by measuring the density of the battery fluid. Explain how this is possible.
36. **Applying Ideas** In lead-acid batteries, the battery cannot be recharged indefinitely. Explain why not.
37. **Interpreting Graphics** A voltaic cell is pictured below. Identify the species that is oxidized if current is allowed to flow.



**USING THE HANDBOOK**

38. Aluminum is described in Group 13 of the *Elements Handbook* as a self-protecting metal, and can be used in preventing corrosion of iron structures. Using electrochemical data, explain how aluminum protects iron structures.

**RESEARCH & WRITING**

39. Go to the library, and find out about the electroplating industry in the United States. What are the top three metals used for plating, and how many metric tons of each are used for electroplating each year in the United States?
40. Investigate the types of batteries being considered for electric cars. Write a report on the advantages and disadvantages of these types of batteries.

**ALTERNATIVE ASSESSMENT**

41. **Performance** Take an inventory of the types of batteries used in your home. Find out the voltage supplied by each battery and what electrochemical reaction each uses. Suggest why that electrochemical reaction is used in each case.

42. In our portable society, batteries have become a necessary power supply. As consumers, we want to purchase batteries that will last as long as possible. Advertisements tell us that some batteries last longer than others, but do they really? Design an investigation to answer the question. Is there a difference in longevity between the major brands of AA batteries? Add a cost-effectiveness component to your design.
43. When someone who has a silver filling in a tooth bites down on an aluminum gum wrapper, saliva acts as an electrolyte. The system is an electrochemical cell that produces a small jolt of pain. Explain what occurs, using half-cell reactions and  $E^{\circ}$  values.

**extension****Graphing Calculator Equilibrium Constant for an Electrochemical Cell**

Go to [go.hrw.com](http://go.hrw.com) for a graphing calculator exercise that asks you to calculate an equilibrium constant for an electrochemical cell.



**Keyword:** HC6ELEX

# Math Tutor

## CALCULATING CELL POTENTIALS

You have learned that electrons are transferred in all oxidation-reduction reactions. Electrons from a substance being oxidized are transferred to another substance being reduced. A voltaic cell is a simple device that physically separates the oxidation reaction from the reduction reaction, thus forcing electrons released during oxidation to travel through a wire to reach the site of reduction. If a device, such as a light bulb or motor, is placed in the circuit, the moving electrons can perform useful work.

The potential, or voltage, of a given voltaic cell depends on how strongly the oxidation process tends to give up electrons in addition to how strongly the reduction process tends to take them. The greater these two combined tendencies, the higher the potential of the cell. A potential has been measured for each half-reaction, as you can see in **Table 1** on p. 664. In a voltaic cell, the process with the more-negative reduction potential will proceed as the oxidation reaction at the anode of the cell. The more positive reaction will proceed as reduction at the cathode. The following example shows you how to determine the potentials of different kinds of voltaic cells.

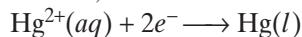
### Problem-Solving TIPS

- In a voltaic cell, the process that has the more negative reduction potential will proceed as the oxidation reaction at the anode of the cell.

### SAMPLE

**Calculate the potential of a voltaic cell in which  $\text{Hg}^{2+}$  ions are reduced to Hg metal while Zn metal is oxidized to  $\text{Zn}^{2+}$  ions.**

Reduction takes place at the cathode, so the cathode half-reaction is

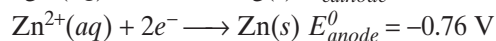


Oxidation takes place at the anode, so the anode half-reaction is



To use the equation for cell potential, rewrite the anode half-reaction as a reduction reaction,  $\text{Zn}^{2+}(\text{aq}) + 2e^{-} \longrightarrow \text{Zn}(\text{s})$ .

Use **Table 1** on page 664 to find the standard reduction potential for each half-reaction. Then, calculate the cell potential.



$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} = +0.85 \text{ V} - (-0.76 \text{ V}) = 1.61 \text{ V}$$

### PRACTICE PROBLEMS

1. Calculate the potential of a voltaic cell in which aluminum metal is oxidized to  $\text{Al}^{3+}$  ions while  $\text{Cu}^{2+}$  ions are reduced to  $\text{Cu}^{+}$  ions.
2. Calculate the potential of a cell in which the reaction is  $\text{Pb}(\text{s}) + \text{Br}_2(\text{l}) \longrightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{Br}^{-}(\text{aq})$ .



# Standardized Test Prep

Answer the following items on a separate piece of paper.

## MULTIPLE CHOICE

- The electrode at which reduction occurs is
  - always the anode.
  - always the cathode.
  - either the anode or the cathode.
  - always the half-cell.
- Refer to the table of standard reduction potentials below. A voltaic cell contains a strip of zinc metal in a solution containing zinc ions in one half-cell. The second is a strip of tin metal in a solution containing tin ions. When this cell operates,
  - Sn is oxidized and  $Zn^{2+}$  is reduced.
  - Sn is reduced and  $Zn^{2+}$  is oxidized.
  - $Sn^{2+}$  is oxidized and Zn is reduced.
  - $Sn^{2+}$  is reduced and Zn is oxidized.
- When a rechargeable cell is being recharged, the cell acts as a(n)
  - fuel cell.
  - electrolytic cell.
  - voltaic cell.
  - Leclanche cell.
- Refer to the table of standard reduction potentials below. The standard potential of the cell  $Sn | Sn^{2+} || Cr^{3+} | Cr$  is  $-0.60$  V. What is the standard reduction potential of the  $Cr^{3+}/Cr$  electrode?
  - $+0.88$  V
  - $+0.74$  V
  - $-0.88$  V
  - $-0.74$  V

Half-cell reaction	Standard reduction potential, $E^0$ (in volts)
$Au^{3+} + 3e^- \rightleftharpoons Au$	+1.50
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	-0.41
$Sn^{2+} + 2e^- \rightleftharpoons Sn$	-0.14
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	-0.76
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	-2.37

- Refer to the table of reduction potentials below. Which metal, Zn or Au, can reduce  $Sn^{2+}$  ions to Sn metal when placed in an aqueous solution of  $Sn^{2+}$  ions?
  - Zn
  - Au
  - Both Zn and Au can reduce  $Sn^{2+}$  ions
  - Neither Zn or Au can reduce  $Sn^{2+}$  ions
- When silver is electroplated onto another metal,  $Ag^+$  is
  - oxidized at the anode.
  - reduced at the anode.
  - oxidized at the cathode.
  - reduced at the cathode.
- Which metal would best provide cathodic protection from corrosion for an iron bridge?
  - Au
  - Sn
  - Cu
  - Mg

## SHORT ANSWER

- In the production of aluminum by the Hall-Héroult process, is the aluminum metal produced at the cathode or the anode?
- The reduction potential for nitric acid is  $0.96$  V (reduction half-reaction is  $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ ). If gold (Au) is placed in a beaker of  $1$  M nitric acid, will the gold dissolve?

## EXTENDED RESPONSE

- Given a voltaic cell consisting of a gold electrode in a gold nitrate solution in one half-cell and a zinc electrode in a zinc nitrate solution in the other half-cell, what occurs once the wire is connected to both electrodes?
- If pure water is placed in a beaker with wires connected to a cathode and anode and electrical voltage is applied, will electrolysis of water occur? Explain your answer.

**Test TIP** Remember that if you can eliminate two of the four answer choices, your chances of choosing the correct answer choice will double.



# Voltaic Cells

## OBJECTIVES

- *Construct a Cu-Zn voltaic cell.*
- *Design and construct two other voltaic cells.*
- *Measure the potential of the voltaic cells.*
- *Evaluate cells by comparing the measured cell voltages with the voltages calculated from standard reduction potentials.*

## MATERIALS

- 0.5 M  $\text{Al}_2(\text{SO}_4)_3$ , 75 mL
- 0.5 M  $\text{CuSO}_4$ , 75 mL
- 0.5 M  $\text{ZnSO}_4$ , 75 mL
- Aluminum strip, 1 cm  $\times$  8 cm
- Copper strip, 1 cm  $\times$  8 cm
- Zinc strip, 1 cm  $\times$  8 cm
- Distilled water
- 100 mL graduated cylinder
- Emery cloth
- 150 mL beakers, 3
- Salt bridge
- Voltmeter
- Wires with alligator clips, 2

## BACKGROUND

In voltaic cells, oxidation and reduction half-reactions take place in separate half-cells, which can consist of a metal electrode immersed in a solution of its metal ions. The electrical potential, or voltage, that develops between the electrodes is a measure of the combined reducing strength of one reactant and oxidizing strength of the other reactant.

## SAFETY



For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

## PREPARATION

1. Follow your teacher's instructions to create the data table that you will use to record your data for three voltaic cells.
2. Remove any oxide coating from strips of aluminum, copper, and zinc by rubbing them with an emery cloth. Keep the metal strips dry until you are ready to use them.
3. Label three 150 mL beakers " $\text{Al}_2(\text{SO}_4)_3$ ," " $\text{CuSO}_4$ ," and " $\text{ZnSO}_4$ ."

## PROCEDURE

1. Pour 75 mL of 0.5 M  $\text{ZnSO}_4$  into the  $\text{ZnSO}_4$  beaker and 75 mL of 0.5 M  $\text{CuSO}_4$  into the  $\text{CuSO}_4$  beaker.
2. Place one end of the salt bridge into the  $\text{CuSO}_4$  solution and the other end into the  $\text{ZnSO}_4$  solution.
3. Place a zinc strip into the zinc solution and a copper strip into the copper solution.

- Using the alligator clips, connect one wire to one end of the zinc strip and the second wire to the copper strip. Take the free end of the wire attached to the zinc strip, and connect it to one terminal on the voltmeter. Take the free end of the wire attached to the copper strip, and connect it to the other terminal on the voltmeter. The needle on the voltmeter should move to the right. If your voltmeter's needle points to the left, reverse the way the wires are connected to the terminals of the voltmeter. Immediately record the voltage reading in your data table, and disconnect the circuit.
- Record the concentration of the solutions and sketch a diagram of your electrochemical cell.
- Rinse the copper and zinc strips with a *very small* amount of distilled water. Collect the rinse from the copper strip in the  $\text{CuSO}_4$  beaker and the rinse from the zinc strip in the  $\text{ZnSO}_4$  beaker. Rinse each end of the salt bridge into the corresponding beaker.
- Use the table of standard reduction potentials in the textbook to calculate the standard voltages for the other cells you can build using copper, zinc, or aluminum. Build these cells and measure their potentials following steps 1–6.

## CLEANUP AND DISPOSAL

- Clean all apparatus and your lab station. Wash your hands. Place the pieces of metal in the containers designated by your teacher. Each solution should be poured in its own separate disposal container. Do not mix the contents of the beakers.



## ANALYSIS AND INTERPRETATION

- Organizing Ideas:** For each cell that you constructed, write the equations for the two half-cell reactions. Obtain the standard half-cell potentials for the half-reactions from **Table 1**, and write these  $E^0$  values after the equations.
- Organizing Ideas:** For each cell you tested, combine the two half-reactions to obtain the equation for the net reaction.
- Organizing Ideas:** Use the  $E^0$  values for the half-reactions to determine the  $E^0$  for each cell.
- Resolving Discrepancies:** Compare the actual cell voltages you measured with the standard cell voltages in item 3. Explain why you would expect a difference.

## CONCLUSIONS

- Inferring Conclusions:** Based on the voltages that you measured, which cell produces the most energy?
- Applying Ideas:** On the basis of your data, which metal is the strongest reducing agent? Which metal ion is the strongest oxidizing agent?
- Applying Ideas:** Indicate the direction of electron flow in each of your cell diagrams.

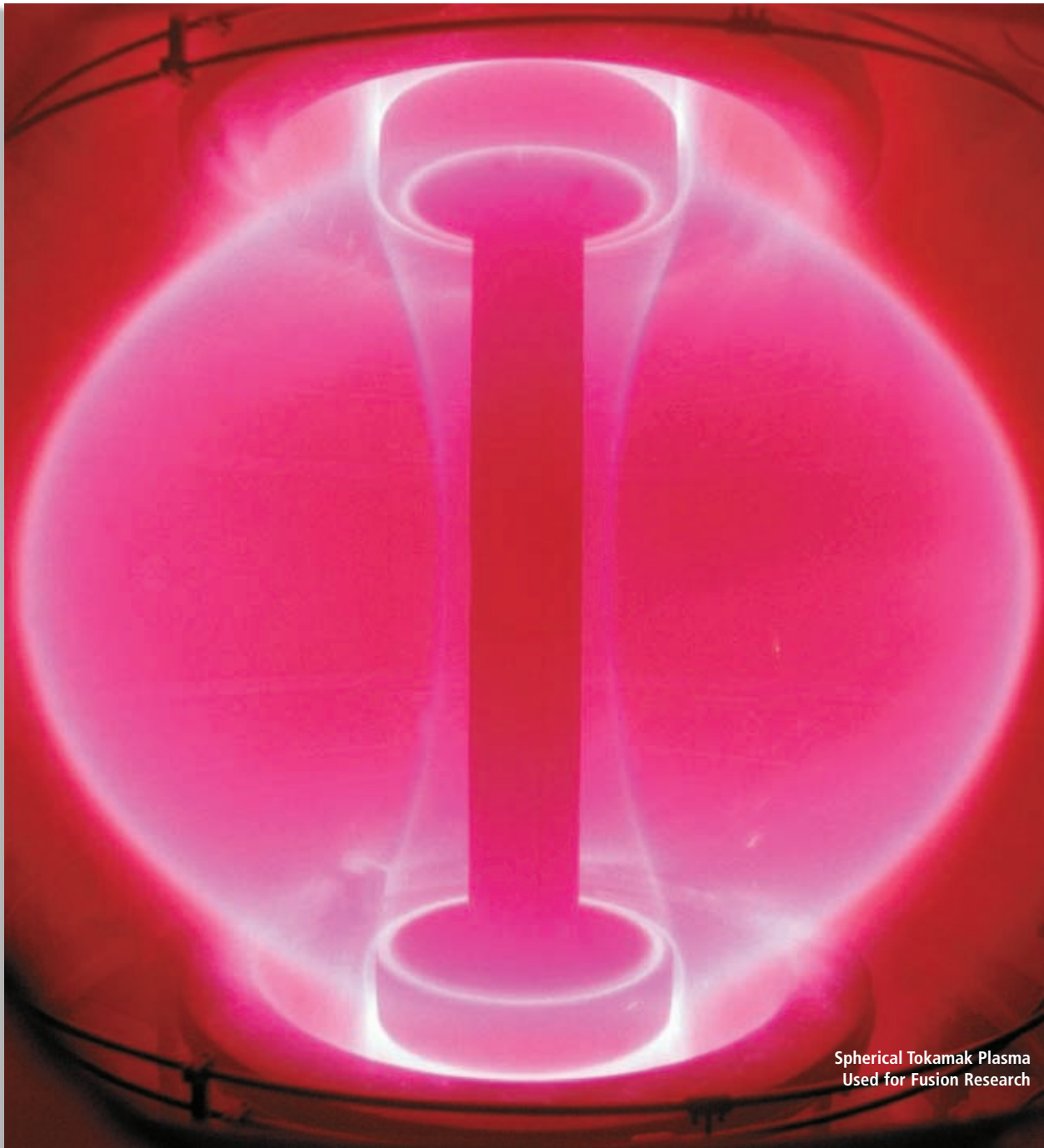
## EXTENSIONS

- Predicting Outcomes:** Describe how and why the reactions would stop if the cells had been left connected.
- Designing Experiments:** Design a method that could use several of the electrochemical cells you constructed to generate more voltage than any individual cell provided. (Hint: consider what would happen if you linked an Al-Zn cell and a Zn-Cu cell. If your teacher approves your plan, test your idea.)

# *Nuclear Chemistry*

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*High temperature plasmas are being studied as potential sources of fusion energy.*



Spherical Tokamak Plasma  
Used for Fusion Research

# The Nucleus

## SECTION 1

### OBJECTIVES

- Explain what a nuclide is, and describe the different ways nuclides can be represented.
- Define and relate the terms *mass defect* and *nuclear binding energy*.
- Explain the relationship between number of nucleons and stability of nuclei.
- Explain why nuclear reactions occur, and know how to balance a nuclear equation.

Atomic nuclei are made of *protons and neutrons*, which are collectively called **nucleons**. In nuclear chemistry, an atom is referred to as a **nuclide** and is identified by the number of protons and neutrons in its nucleus. Nuclides can be represented in two ways. When a symbol such as  ${}^{228}_{88}\text{Ra}$  is used, the superscript is the mass number and the subscript is the atomic number. The same nuclide can also be written as radium-228, where the mass number is written following the element name.

### Mass Defect and Nuclear Stability

Because an atom is made of protons, neutrons, and electrons, you might expect the mass of an atom to be the same as the mass of an equal number of isolated protons, neutrons, and electrons. However, this is not the case. Let's consider a  ${}^4_2\text{He}$  atom as an example. The combined mass of two protons, two neutrons, and two electrons is calculated below.

$$\begin{aligned} 2 \text{ protons: } & (2 \times 1.007\,276 \text{ amu}) = 2.014\,552 \text{ amu} \\ 2 \text{ neutrons: } & (2 \times 1.008\,665 \text{ amu}) = 2.017\,330 \text{ amu} \\ 2 \text{ electrons: } & (2 \times 0.000\,548\,6 \text{ amu}) = 0.001\,097 \text{ amu} \\ \text{total combined mass: } & 4.032\,979 \text{ amu} \end{aligned}$$

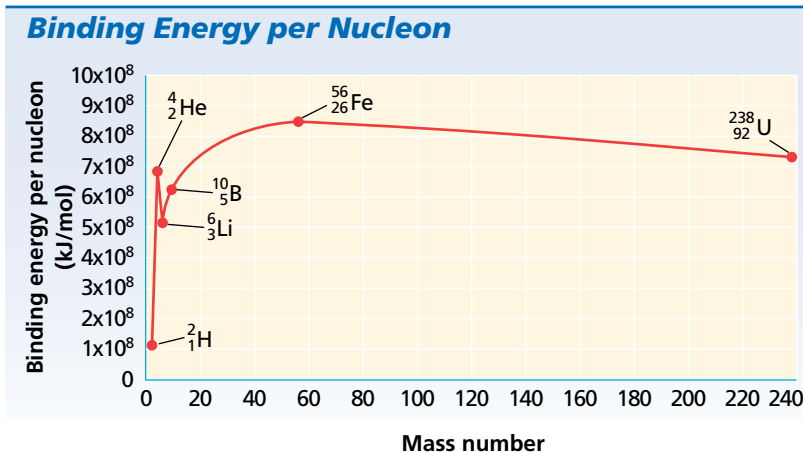
However, the atomic mass of a  ${}^4_2\text{He}$  atom has been measured to be 4.002 602 amu. The measured mass, 4.002 602 amu, is 0.030 377 amu less than the combined mass, 4.032 979 amu, calculated above. *This difference between the mass of an atom and the sum of the masses of its protons, neutrons, and electrons is called the **mass defect**.*

### Nuclear Binding Energy

What causes the loss in mass? According to Albert Einstein's equation  $E = mc^2$ , mass can be converted to energy, and energy to mass. The mass defect is caused by the conversion of mass to energy upon formation of the nucleus. The mass units of the mass defect can be converted to energy units by using Einstein's equation. First, convert 0.030 377 amu to kilograms to match the mass units for energy,  $\text{kg}\cdot\text{m}^2/\text{s}^2$ .

$$0.030\,377 \text{ amu} \times \frac{1.6605 \times 10^{-27} \text{ kg}}{1 \text{ amu}} = 5.0441 \times 10^{-29} \text{ kg}$$

**FIGURE 1** This graph shows the relationship between binding energy per nucleon and mass number. The binding energy per nucleon is a measure of the stability of a nucleus.



The energy equivalent can now be calculated.

$$E = mc^2$$

$$E = (5.0441 \times 10^{-29} \text{ kg})(3.00 \times 10^8 \text{ m/s})^2$$

$$= 4.54 \times 10^{-12} \text{ kg} \cdot \text{m}^2/\text{s}^2 = 4.54 \times 10^{-12} \text{ J}$$

This is the **nuclear binding energy**, the energy released when a nucleus is formed from nucleons. This energy can also be thought of as the energy required to hold the nucleus together. Therefore, the nuclear binding energy is also a measure of the stability of a nucleus.

### Binding Energy per Nucleon

The binding energy per nucleon is used to compare the stability of different nuclides, as shown in **Figure 1**. The *binding energy per nucleon* is the binding energy of the nucleus divided by the number of nucleons it contains. The higher the binding energy per nucleon, the more tightly the nucleons are held together. Elements with intermediate atomic masses have the greatest binding energies per nucleon and are therefore the most stable.

## Nucleons and Nuclear Stability

Stable nuclides have certain characteristics. When the number of protons in stable nuclei is plotted against the number of neutrons, as shown in **Figure 2**, a pattern is observed. The neutron-proton ratios of stable nuclides cluster in a narrow band referred to as the *band of stability*. Among atoms having low atomic numbers, the most stable nuclei are those with a neutron-proton ratio of approximately 1:1. For example,  $^4_2\text{He}$ , a stable isotope of helium with two neutrons and two protons, has a neutron-proton ratio of 1:1. As the atomic number increases, the stable neutron-proton ratio increases to about 1.5:1. For example,  $^{206}_{82}\text{Pb}$ , with 124 neutrons and 82 protons, has a neutron-proton ratio of 1.51:1.

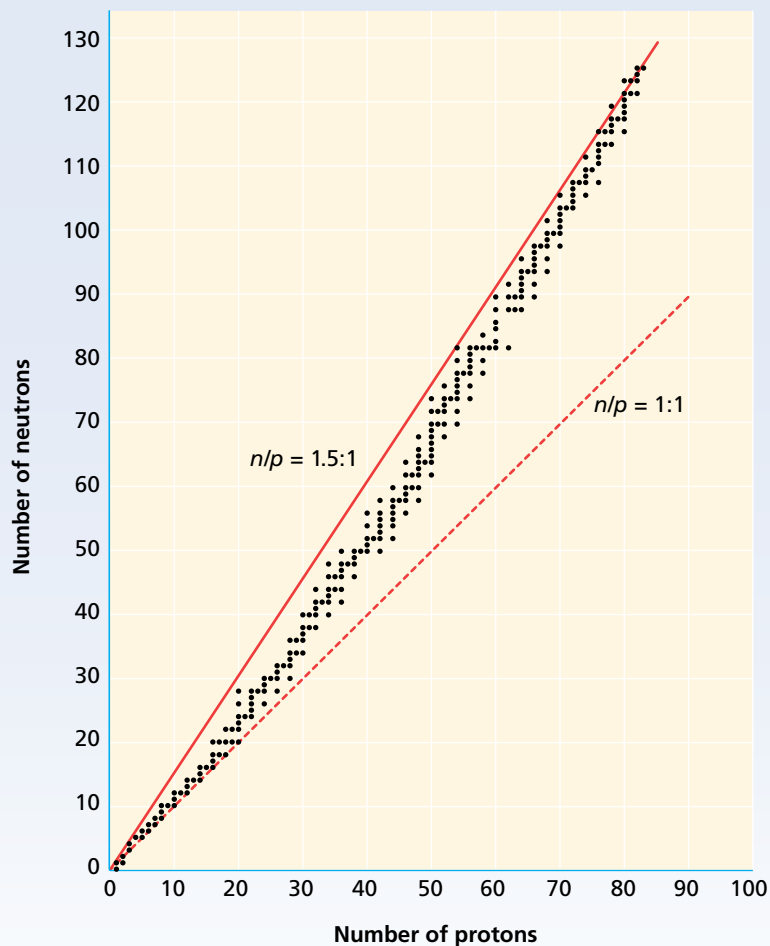
This trend can be explained by the relationship between the nuclear force and the electrostatic forces between protons. A proton in a nucleus repels all other protons through electrostatic repulsion, but the short

### CROSS-DISCIPLINARY

#### Quarks

Many subatomic particles have been identified. Leptons and quarks are the elementary particles of matter. The electron is a lepton. Protons and neutrons are made of quarks. There are six types of quarks that differ in mass and charge. They are named *up*, *down*, *strange*, *charm*, *bottom*, and *top*. Protons consist of two *up* quarks and one *down* quark, and neutrons consist of two *down* quarks and one *up* quark. Although individual quarks have not been isolated, their existence explains the patterns of nuclear binding and decay.

## The Band of Stability

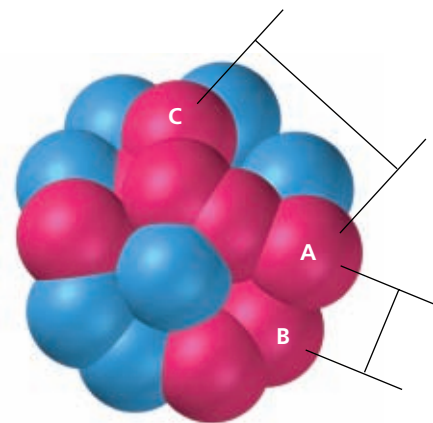


**FIGURE 2** The neutron-proton ratios of stable nuclides cluster together in a region known as the band of stability. As the number of protons increases, the ratio increases from 1:1 to about 1.5:1.

range of the nuclear force allows them to attract only protons very close to them, as shown in **Figure 3**. As the number of protons in a nucleus increases, the repulsive electrostatic force between protons increases faster than the nuclear force. More neutrons are required to increase the nuclear force and stabilize the nucleus. Beyond the atomic number 83, bismuth, the repulsive force of the protons is so great that no stable nuclides exist.

Stable nuclei tend to have even numbers of nucleons. Of the stable nuclides, more than half have even numbers of both protons and neutrons. Only five nuclides have odd numbers of both. This indicates that stability of a nucleus is greatest when the nucleons—like electrons—are paired.

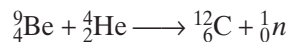
The most stable nuclides are those having 2, 8, 20, 28, 50, 82, or 126 protons, neutrons, or total nucleons. This extra stability at certain numbers supports a theory that nucleons—like electrons—exist at certain energy levels. According to the **nuclear shell model**, nucleons exist in different energy levels, or shells, in the nucleus. The numbers of nucleons that represent completed nuclear energy levels—2, 8, 20, 28, 50, 82, and 126—are called **magic numbers**.



**FIGURE 3** Proton A attracts proton B through the nuclear force but repels it through the electrostatic force. Proton A mainly repels proton C through the electrostatic force because the nuclear force reaches only a few nucleon diameters.

## Nuclear Reactions

Unstable nuclei undergo spontaneous changes that alter the numbers of protons and neutrons. In this process, they give off large amounts of energy and increase their stability. These changes are a type of nuclear reaction. A **nuclear reaction** is a reaction that changes the nucleus of an atom. In equations representing nuclear reactions, the total of the atomic numbers and the total of the mass numbers must be equal on both sides of the equation. An example is shown below.



Notice that when the atomic number changes, the identity of the element changes. A **transmutation** is a change in the identity of a nucleus as a result of a change in the number of its protons.

### SAMPLE PROBLEM A

Identify the product that balances the following nuclear reaction:  ${}^{212}_{84}\text{Po} \longrightarrow {}^4_2\text{He} + \underline{\quad ?}$

**SOLUTION** 1. The total mass number and atomic number must be equal on both sides of the equation.

$$\begin{array}{l} {}^{212}_{84}\text{Po} \longrightarrow {}^4_2\text{He} + \underline{\quad ?} \\ \text{mass number:} \quad 212 - 4 = 208 \qquad \text{atomic number:} \quad 84 - 2 = 82 \end{array}$$

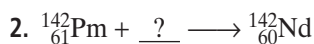
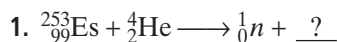
2. The nuclide has a mass number of 208 and an atomic number of 82,  ${}^{208}_{82}\text{Pb}$ .

3. The balanced nuclear equation is  ${}^{212}_{84}\text{Po} \longrightarrow {}^4_2\text{He} + {}^{208}_{82}\text{Pb}$

### PRACTICE

Answers in Appendix E

Using  ${}^1_0n$  to represent a neutron and  ${}^0_{-1}e$  to represent an electron, complete the following nuclear equations:



### extension

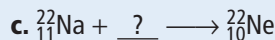
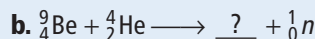
Go to [go.hrw.com](http://go.hrw.com) for more practice problems that ask you to complete nuclear equations.



Keyword: HC6NUCX

## SECTION REVIEW

1. Define mass defect.
2. How is nuclear stability related to the neutron-proton ratio?
3. Complete and balance the following nuclear equations:
  - a.  ${}^{187}_{75}\text{Re} + \underline{\quad ?} \longrightarrow {}^{188}_{75}\text{Re} + {}^1_1\text{H}$



### Critical Thinking

4. **INTERPRETING GRAPHICS** Examine **Figure 2**, and predict if  ${}^7_3\text{Li}$  is a stable isotope of lithium. Explain your answer.

# Radioactive Decay

## SECTION 2

### OBJECTIVES

- Define and relate the terms *radioactive decay* and *nuclear radiation*.
- Describe the different types of radioactive decay and their effects on the nucleus.
- Define the term *half-life*, and explain how it relates to the stability of a nucleus.
- Define and relate the terms *decay series*, *parent nuclide*, and *daughter nuclide*.
- Explain how artificial radioactive nuclides are made, and discuss their significance.

In 1896, Henri Becquerel was studying the possible connection between light emission of some uranium compounds after exposure to sunlight and X-ray emission. He wrapped a photographic plate in a lightproof covering and placed a uranium compound on top of it. He then placed them in sunlight. The photographic plate was exposed even though it was protected from visible light, suggesting exposure by X rays. When he tried to repeat his experiment, cloudy weather prevented him from placing the experiment in sunlight. To his surprise, the plate was still exposed. This meant that sunlight was not needed to produce the rays that exposed the plate. The rays were produced by radioactive decay. **Radioactive decay** is the spontaneous disintegration of a nucleus into a slightly lighter nucleus, accompanied by emission of particles, electromagnetic radiation, or both. The radiation that exposed the plate was **nuclear radiation**, particles or electromagnetic radiation emitted from the nucleus during radioactive decay.

Uranium is a **radioactive nuclide**, an unstable nucleus that undergoes radioactive decay. Studies by Marie Curie and Pierre Curie found that of the elements known in 1896, only uranium and thorium were radioactive. In 1898, the Curies discovered two new radioactive metallic elements, polonium and radium. Since that time, many other radioactive nuclides have been identified. In fact, all of the nuclides beyond atomic number 83 are unstable and thus radioactive.

## Types of Radioactive Decay

A nuclide's type and rate of decay depend on the nucleon content and energy level of the nucleus. Some common types of radioactive nuclide emissions are summarized in **Table 1**.

**TABLE 1** Radioactive Nuclide Emissions

Type	Symbol	Charge	Mass (amu)
Alpha particle	${}^4_2\text{He}$	2+	4.001 5062
Beta particle	${}^0_{-1}\beta$	1-	0.000 5486
Positron	${}^0_{+1}\beta$	1+	0.000 5486
Gamma ray	$\gamma$	0	0

SCILINKS.

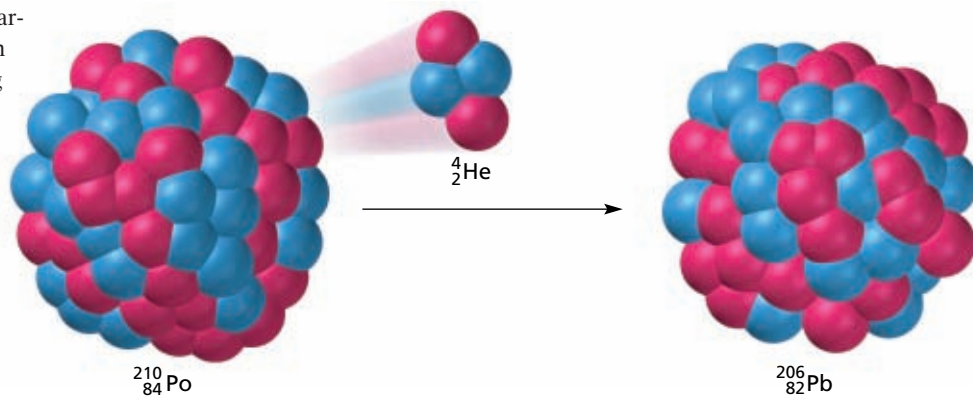
[www.scilinks.org](http://www.scilinks.org)

Topic: Radioactive Decay

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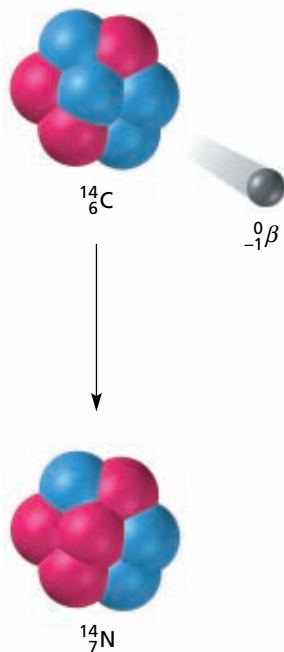
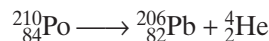


**FIGURE 4** An alpha particle, identical to a helium nucleus, is emitted during the radioactive decay of some very heavy nuclei.



### Alpha Emission

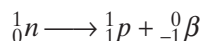
An **alpha particle** ( $\alpha$ ) is two protons and two neutrons bound together and is emitted from the nucleus during some kinds of radioactive decay. Alpha particles are helium nuclei and have a charge of  $2+$ . They are often represented with the symbol  ${}^4_2\text{He}$ . Alpha emission is restricted almost entirely to very heavy nuclei. In these nuclei, both the number of neutrons and the number of protons are reduced in order to increase the stability of the nucleus. An example of alpha emission is the decay of  ${}^{210}_{84}\text{Po}$  into  ${}^{206}_{82}\text{Pb}$ , shown in **Figure 4**. The atomic number decreases by two, and the mass number decreases by four.



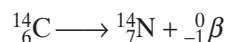
**FIGURE 5** Beta emission causes the transmutation of  ${}^{14}_6\text{C}$  into  ${}^{14}_7\text{N}$ . Beta emission is a type of radioactive decay in which a neutron is converted to a proton with the emission of a beta particle.

### Beta Emission

Nuclides above the band of stability are unstable because their neutron/proton ratio is too large. To decrease the number of neutrons, a neutron can be converted into a proton and an electron. The electron is emitted from the nucleus as a beta particle. A **beta particle** ( $\beta$ ) is an electron emitted from the nucleus during some kinds of radioactive decay.

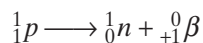


An example of beta emission, shown in **Figure 5**, is the decay of  ${}^{14}_6\text{C}$  into  ${}^{14}_7\text{N}$ . Notice that the atomic number increases by one and the mass number stays the same.

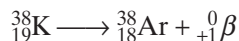


### Positron Emission

Nuclides below the band of stability are unstable because their neutron/proton ratio is too small. To decrease the number of protons, a proton can be converted into a neutron by emitting a positron. A **positron** is a particle that has the same mass as an electron, but has a positive charge, and is emitted from the nucleus during some kinds of radioactive decay.

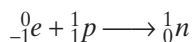


An example of positron emission is the decay of  ${}^{38}_{19}\text{K}$  into  ${}^{38}_{18}\text{Ar}$ . Notice that the atomic number decreases by one but the mass number stays the same.

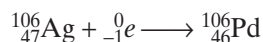


## Electron Capture

Another type of decay for nuclides that have a neutron/proton ratio that is too small is electron capture. *In electron capture, an inner orbital electron is captured by the nucleus of its own atom.* The inner orbital electron combines with a proton, and a neutron is formed.



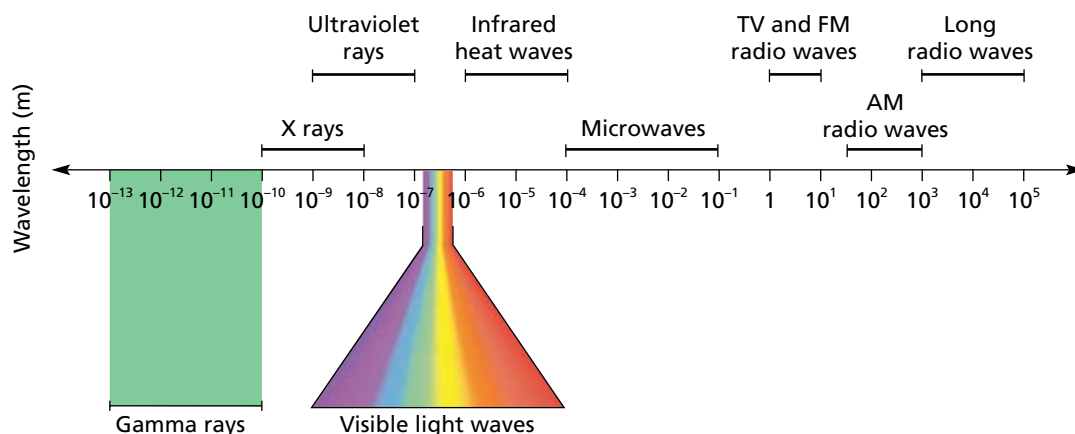
An example of electron capture is the radioactive decay of  ${}^{106}_{47}\text{Ag}$  into  ${}^{106}_{46}\text{Pd}$ . Just as in positron emission, the atomic number decreases by one but the mass number stays the same.



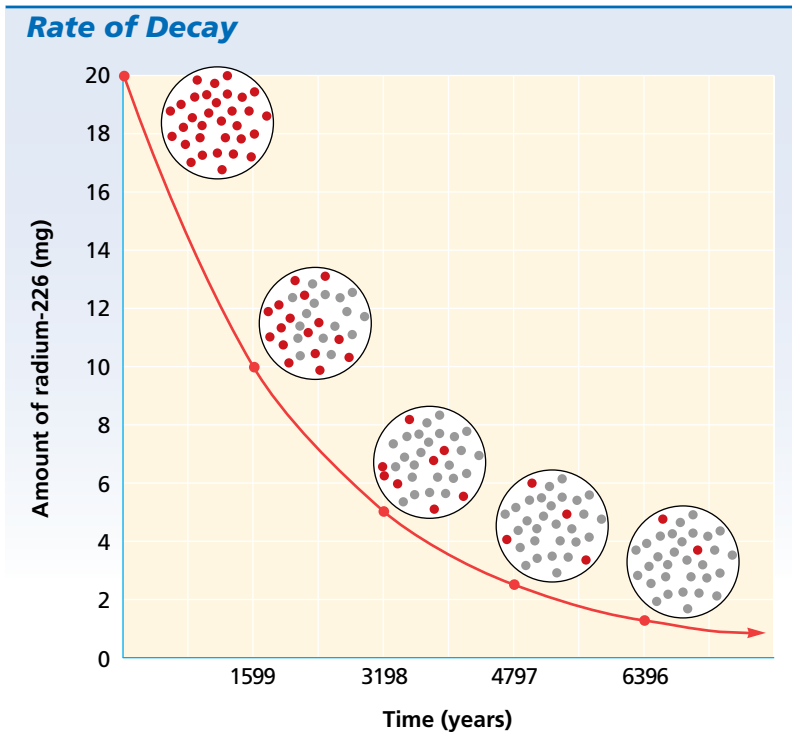
## Gamma Emission

**Gamma rays** ( $\gamma$ ) are high-energy electromagnetic waves emitted from a nucleus as it changes from an excited state to a ground energy state. The position of gamma rays in the electromagnetic spectrum is shown in **Figure 6**. The emission of gamma rays is another piece of evidence supporting the nuclear shell model. According to the nuclear shell model, gamma rays are produced when nuclear particles undergo transitions in nuclear-energy levels. This is similar to the emission of photons (light or X rays) when an electron drops to a lower energy level, which was covered in Chapter 4. Gamma emission usually occurs immediately following other types of decay, when other types of decay leave the nucleus in an excited state.

**FIGURE 6** Gamma rays, like visible light, are a form of electromagnetic radiation, but they have a much shorter wavelength and are much higher in energy than visible light.



**FIGURE 7** The half-life of radium-226 is 1599 years. Half of the remaining radium-226 decays by the end of each additional half-life.



**SCILINKS**  
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**Topic:** Half-Life  
**Code:** HC60709

## Half-Life

No two radioactive isotopes decay at the same rate. **Half-life**,  $t_{1/2}$ , is the time required for half the atoms of a radioactive nuclide to decay. Look at the graph of the decay of radium-226 in **Figure 7**. Radium-226 has a half-life of 1599 years. Half of a given amount of radium-226 decays in 1599 years. In another 1599 years, half of the remaining radium-226 decays. This process continues until there is a negligible amount of radium-226. Each radioactive nuclide has its own half-life. More-stable nuclides decay slowly and have longer half-lives. Less-stable nuclides decay very quickly and have shorter half-lives, sometimes just a fraction of a second. Some representative radioactive nuclides, along with their half-lives, are given in **Table 2**.

**TABLE 2** Representative Radioactive Nuclides and Their Half-Lives

Nuclide	Half-life	Nuclide	Half-life
${}^3_1\text{H}$	12.32 years	${}^{214}_{84}\text{Po}$	163.7 $\mu\text{s}$
${}^{14}_6\text{C}$	5715 years	${}^{218}_{84}\text{Po}$	3.0 min
${}^{32}_{15}\text{P}$	14.28 days	${}^{218}_{85}\text{At}$	1.6 s
${}^{40}_{19}\text{K}$	$1.3 \times 10^9$ years	${}^{238}_{92}\text{U}$	$4.46 \times 10^9$ years
${}^{60}_{27}\text{Co}$	5.27 years	${}^{239}_{94}\text{Pu}$	$2.41 \times 10^4$ years

**SAMPLE PROBLEM B**

For more help, go to the *Math Tutor* at the end of the chapter.

Phosphorus-32 has a half-life of 14.3 days. How many milligrams of phosphorus-32 remain after 57.2 days if you start with 4.0 mg of the isotope?

**SOLUTION**

- 1 ANALYZE** Given: original mass of phosphorus-32 = 4.0 mg  
half-life of phosphorus-32 = 14.3 days  
time elapsed = 57.2 days  
Unknown: mass of phosphorus-32 remaining after 57.2 days
- 2 PLAN** To determine the number of milligrams of phosphorus-32 remaining, we must first find the number of half-lives that have passed in the time elapsed. Then the amount of phosphorus-32 is determined by reducing the original amount by half for every half-life that has passed.
- $$\text{number of half-lives} = \text{time elapsed (days)} \times \frac{1 \text{ half-life}}{14.3 \text{ days}}$$
- amount of phosphorus-32 remaining =  
original amount of phosphorus-32  $\times \frac{1}{2}$  for each half-life
- 3 COMPUTE** number of half-lives =  $57.2 \text{ days} \times \frac{1 \text{ half-life}}{14.3 \text{ days}} = 4 \text{ half-lives}$
- $$\text{amount of phosphorus-32 remaining} = 4.0 \text{ mg} \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} = 0.25 \text{ mg}$$
- 4 EVALUATE** A period of 57.2 days is four half-lives for phosphorus-32. At the end of one half-life, 2.0 mg of phosphorus-32 remains; 1.0 mg remains at the end of two half-lives; 0.50 mg remains at the end of three half-lives; and 0.25 mg remains at the end of four half-lives.

**PRACTICE**

Answers in Appendix E

1. The half-life of polonium-210 is 138.4 days. How many milligrams of polonium-210 remain after 415.2 days if you start with 2.0 mg of the isotope?
2. Assuming a half-life of 1599 years, how many years will be needed for the decay of  $\frac{15}{16}$  of a given amount of radium-226?
3. The half-life of radon-222 is 3.824 days. After what time will one-fourth of a given amount of radon remain?
4. The half-life of cobalt-60 is 5.27 years. How many milligrams of cobalt-60 remain after 52.7 years if you start with 10.0 mg?
5. A sample contains 4.0 mg of uranium-238. After  $4.46 \times 10^9$  years, the sample will contain 2.0 mg of uranium-238. What is the half-life of uranium-238?

**extension**

Go to [go.hrw.com](http://go.hrw.com) for more practice problems that ask you to calculate the half-life or amount of sample remaining.

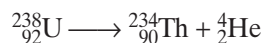


Keyword: HC6NUCX

## Decay Series

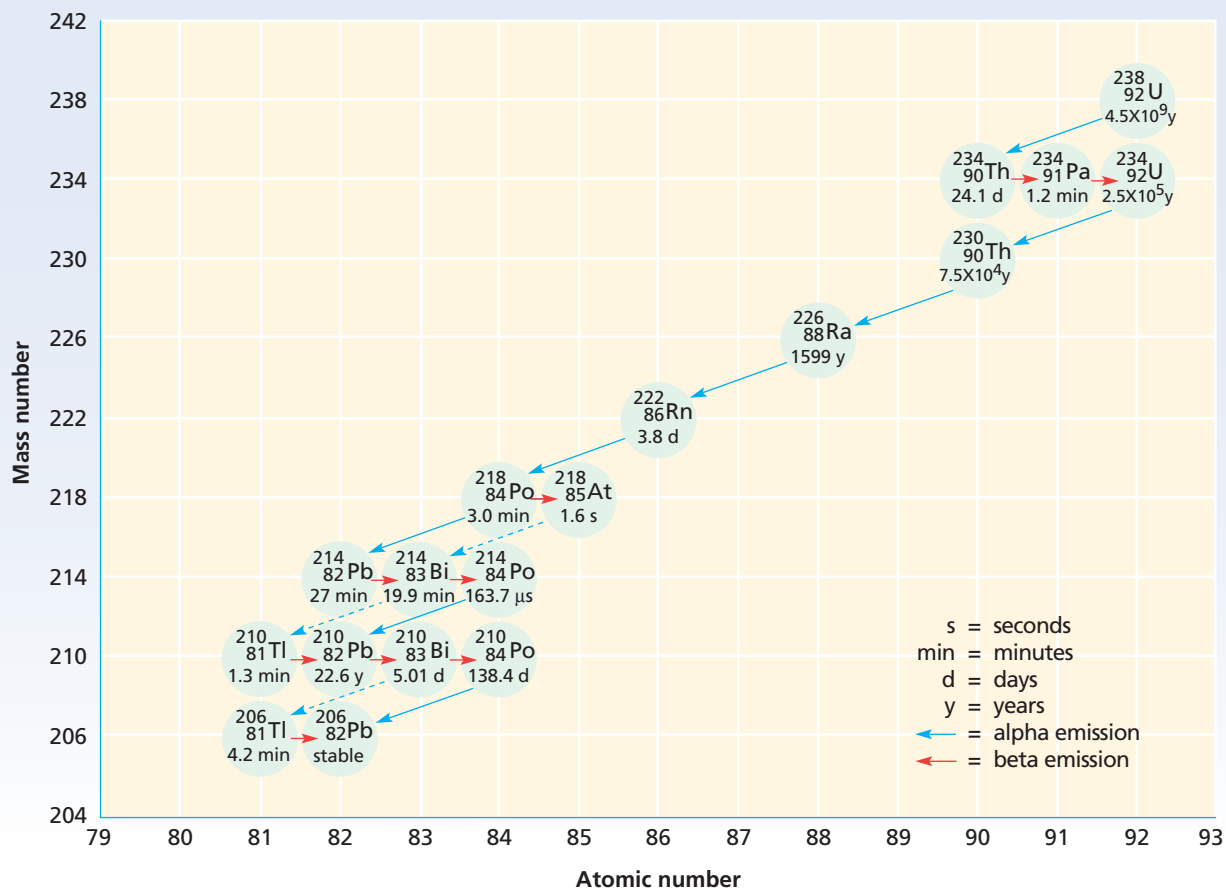
One nuclear reaction is not always enough to produce a stable nuclide. A **decay series** is a series of radioactive nuclides produced by successive radioactive decay until a stable nuclide is reached. The heaviest nuclide of each decay series is called the **parent nuclide**. The nuclides produced by the decay of the parent nuclides are called **daughter nuclides**. All naturally occurring nuclides with atomic numbers greater than 83 are radioactive and belong to one of three natural decay series. The parent nuclides are uranium-238, uranium-235, and thorium-232. The transmutations of the uranium-238 decay series are charted in **Figure 8**.

Locate the parent nuclide, uranium-238, on the chart. As the nucleus of uranium-238 decays, it emits an alpha particle. The mass number of the nuclide, and thus the vertical position on the graph, decreases by four. The atomic number, and thus the horizontal position, decreases by two. The daughter nuclide is an isotope of thorium.

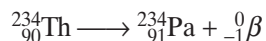


**FIGURE 8** This chart shows the transmutations that occur as  ${}_{92}^{238}\text{U}$  decays to the final, stable nuclide,  ${}_{82}^{206}\text{Pb}$ . Decay usually follows the solid arrows. The dotted arrows represent alternative routes of decay.

### Uranium-238 Decay Series



The half-life of  ${}^{234}_{90}\text{Th}$ , about 24 days, is indicated on the chart. It decays by giving off beta particles. This increases its atomic number, and thus its horizontal position, by one. The mass number, and thus its vertical position, remains the same.



The remaining atomic number and mass number changes shown on the decay chart are also explained in terms of the particles given off. In the final step,  ${}^{210}_{84}\text{Po}$  loses an alpha particle to form  ${}^{206}_{82}\text{Pb}$ . This is a stable, nonradioactive isotope of lead. Notice that  ${}^{206}_{82}\text{Pb}$  contains 82 protons, a magic number. It contains the extra-stable nuclear configuration of a completed nuclear shell.

### extension

#### Historical Chemistry

Go to [go.hrw.com](http://go.hrw.com) for a full-length article on Glenn Seaborg.



Keyword: HC6NUCX

## Artificial Transmutations

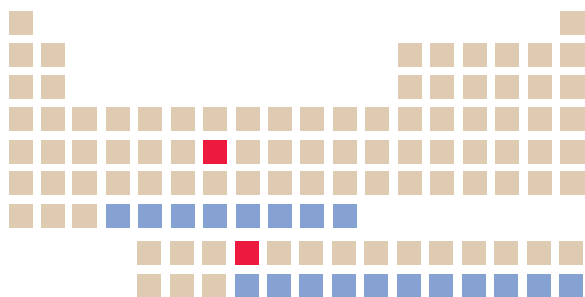
*Artificial radioactive nuclides* are radioactive nuclides not found naturally on Earth. They are made by **artificial transmutations**, *bombardment of nuclei with charged and uncharged particles*. Because neutrons have no charge, they can easily penetrate the nucleus of an atom. However, positively charged alpha particles, protons, and other ions are repelled by the nucleus. Because of this repulsion, great quantities of energy are required to bombard nuclei with these particles. The necessary energy may be supplied by accelerating these particles in the magnetic or electrical field of a particle accelerator. An example of an accelerator is shown in **Figure 9**.

**FIGURE 9** This is an aerial view of the Fermi International Accelerator Laboratory (Fermilab), in Illinois. The particle accelerators are underground. The Tevatron ring, the larger particle accelerator, has a circumference of 4 mi. The smaller ring (top left) is a new accelerator, the Main Injector.



**TABLE 3** Reactions for the First Preparation of Several Transuranium Elements

Atomic number	Name	Symbol	Nuclear reaction
93	neptunium	Np	${}^{238}_{92}\text{U} + {}^1_0n \longrightarrow {}^{239}_{92}\text{U}$ ${}^{239}_{92}\text{U} \longrightarrow {}^{239}_{93}\text{Np} + {}^0_{-1}\beta$
94	plutonium	Pu	${}^{238}_{93}\text{Np} \longrightarrow {}^{238}_{94}\text{Pu} + {}^0_{-1}\beta$
95	americium	Am	${}^{239}_{94}\text{Pu} + 2{}^1_0n \longrightarrow {}^{241}_{95}\text{Am} + {}^0_{-1}\beta$
96	curium	Cm	${}^{239}_{94}\text{Pu} + {}^4_2\text{He} \longrightarrow {}^{242}_{96}\text{Cm} + {}^1_0n$
97	berkelium	Bk	${}^{241}_{95}\text{Am} + {}^4_2\text{He} \longrightarrow {}^{243}_{97}\text{Bk} + 2{}^1_0n$
98	californium	Cf	${}^{242}_{96}\text{Cm} + {}^4_2\text{He} \longrightarrow {}^{245}_{98}\text{Cf} + {}^1_0n$
99	einsteinium	Es	${}^{238}_{92}\text{U} + 15{}^1_0n \longrightarrow {}^{253}_{99}\text{Es} + 7{}^0_{-1}\beta$
100	fermium	Fm	${}^{238}_{92}\text{U} + 17{}^1_0n \longrightarrow {}^{255}_{100}\text{Fm} + 8{}^0_{-1}\beta$
101	mendelevium	Md	${}^{253}_{99}\text{Es} + {}^4_2\text{He} \longrightarrow {}^{256}_{101}\text{Md} + {}^1_0n$
102	nobelium	No	${}^{246}_{96}\text{Cm} + {}^{12}_6\text{C} \longrightarrow {}^{254}_{102}\text{No} + 4{}^1_0n$
103	lawrencium	Lr	${}^{252}_{98}\text{Cf} + {}^{10}_5\text{B} \longrightarrow {}^{258}_{103}\text{Lr} + 4{}^1_0n$

**FIGURE 10** Artificial transmutations filled the gaps in the periodic table, shown in red, and extended the periodic table with the transuranium elements, shown in blue.

### Artificial Radioactive Nuclides

Radioactive isotopes of all the natural elements have been produced by artificial transmutation. In addition, production of technetium and promethium by artificial transmutation has filled gaps in the periodic table. Their positions are shown in **Figure 10**.

Artificial transmutations are also used to produce the transuranium elements. **Transuranium elements** are elements with more than 92 protons in their nuclei. All of these elements are radioactive. The nuclear reactions for the synthesis of several transuranium elements are shown in **Table 3**. Currently, 19 artificially prepared transuranium elements have been named. Others have been reported, but not confirmed. The positions of the named transuranium elements in the periodic table are shown in **Figure 10**.

## SECTION REVIEW

1. Define *radioactive decay*.
2. a. What are the different types of common radioactive decay?  
b. List the types of radioactive decay that convert one nuclide into another.

3. What fraction of a given sample of a radioactive nuclide remains after four half-lives?
4. When does a decay series end?

### Critical Thinking

5. **INTERPRETING CONCEPTS** Distinguish between natural and artificial radioactive nuclides.

# Nuclear Radiation

## SECTION 3

### OBJECTIVES

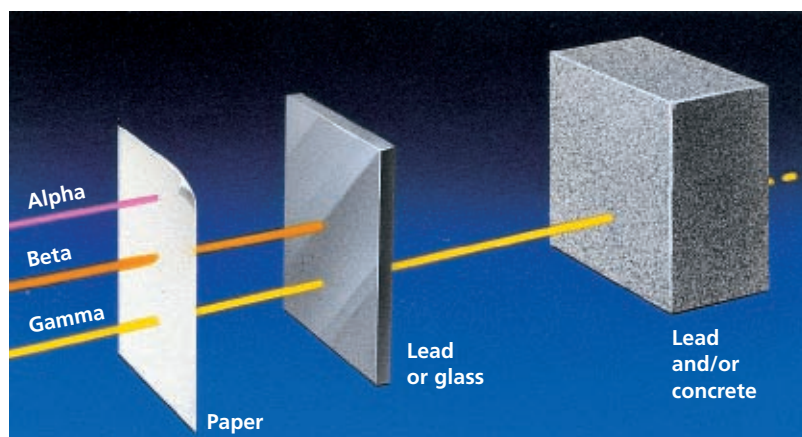
- Compare the penetrating ability and shielding requirements of alpha particles, beta particles, and gamma rays.
- Define the terms *roentgen* and *rem*, and distinguish between them.
- Describe three devices used in radiation detection.
- Discuss applications of radioactive nuclides.

In Becquerel's experiment, nuclear radiation from the uranium compound penetrated the lightproof covering and exposed the film. Different types of nuclear radiation have different penetrating abilities. Nuclear radiation includes alpha particles, beta particles, and gamma rays.

Alpha particles can travel only a few centimeters in air and have a low penetrating ability due to their large mass and charge. They cannot penetrate skin. However, they can cause damage inside the body if a substance that emits alpha particles is ingested or inhaled. Beta particles, which are emitted electrons, travel at speeds close to the speed of light and have a penetrating ability about 100 times greater than that of alpha particles. Beta particles can travel a few meters in air. Gamma rays have the greatest penetrating ability. The penetrating abilities and shielding requirements of different types of nuclear radiation are shown in **Figure 11**.

## Radiation Exposure

Nuclear radiation can transfer the energy from nuclear decay to the electrons of atoms or molecules and cause ionization. The **roentgen (R)** is a unit used to measure nuclear radiation exposure; it is equal to the amount of gamma and X ray radiation that produces  $2 \times 10^9$  ion pairs when it passes through  $1 \text{ cm}^3$  of dry air. Ionization can damage living tissue. Radiation damage to human tissue is measured in rems (roentgen equivalent, man). A **rem** is a unit used to measure the dose of any type of ionizing radiation that factors in the effect that the radiation has on human tissue. Long-term exposure to radiation can cause DNA mutations that result in cancer and



**FIGURE 11** The different penetrating abilities of alpha particles, beta particles, and gamma rays require different levels of shielding. Alpha particles can be shielded with just a sheet of paper. Lead or glass is often used to shield beta particles. Gamma rays are the most penetrating and require shielding with thick layers of lead or concrete, or both.



other genetic defects. DNA can be mutated directly by interaction with radiation or indirectly by interaction with previously ionized molecules.

Everyone is exposed to environmental background radiation. Average exposure for people living in the United States is estimated to be about 0.1 rem per year. However, actual exposure varies. The maximum permissible dose of radiation exposure for a person in the general population is 0.5 rem per year. Airline crews and people who live at high altitudes have increased exposure levels because of increased cosmic ray levels at high altitudes. Radon-222 trapped inside homes also causes increased exposure. Because it is a gas, radon released from certain rocks can move up through the soil into homes through holes in the foundation. Radon trapped in homes increases the risk of lung cancer, especially among smokers.

## Radiation Detection

Film badges, Geiger-Müller counters, and scintillation counters are three devices commonly used to detect and measure nuclear radiation. A film badge and a Geiger-Müller counter are shown in **Figure 12**. As previously mentioned, nuclear radiation exposes film just as visible light does. This property is used in film badges. **Film badges** use exposure of film to measure the approximate radiation exposure of people working with radiation. **Geiger-Müller counters** are instruments that detect radiation by counting electric pulses carried by gas ionized by radiation. Geiger-Müller counters are typically used to detect beta-particles, X rays, and gamma radiation. Radiation can also be detected when it transfers its energy to substances that *scintillate*, or absorb ionizing radiation and emit visible light. **Scintillation counters** are instruments that convert scintillating light to an electric signal for detecting radiation.

**FIGURE 12** Film badges (a) and Geiger-Müller counters (b) are both used to detect nuclear radiation.



(a)



(b)

## Applications of Nuclear Radiation

Many applications are based on the fact that the physical and chemical properties of stable isotopes are essentially the same as those of radioactive isotopes of the same element. A few uses of radioactive nuclides are discussed below.

### Radioactive Dating

**Radioactive dating** is the process by which the approximate age of an object is determined based on the amount of certain radioactive nuclides present. Such an estimate is based on the fact that radioactive substances decay with known half-lives. Age is estimated by measuring either the accumulation of a daughter nuclide or the disappearance of the parent nuclide.

Carbon-14 is radioactive and has a half-life of approximately 5715 years. It can be used to estimate the age of organic material up to about 50 000 years old. Nuclides with longer half-lives are used to estimate the age of older objects; methods using nuclides with long half-lives have been used to date minerals and lunar rocks more than 4 billion years old.

### Radioactive Nuclides in Medicine

In medicine, radioactive nuclides, such as the artificial radioactive nuclide cobalt-60, are used to destroy certain types of cancer cells. Many radioactive nuclides are also used as **radioactive tracers**, which are radioactive atoms that are incorporated into substances so that movement of the substances can be followed by radiation detectors. Detection of radiation from radioactive tracers can be used to diagnose cancer and other diseases. See **Figure 13**.

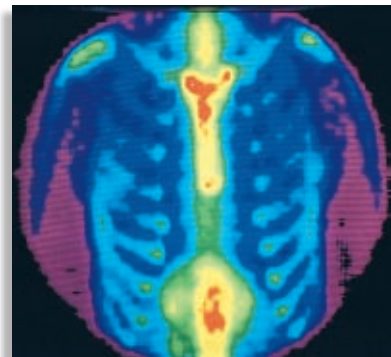
### Radioactive Nuclides in Agriculture

In agriculture, radioactive tracers in fertilizers are used to determine the effectiveness of the fertilizer. The amount of radioactive tracer absorbed by a plant indicates the amount of fertilizer absorbed. Nuclear radiation is also used to prolong the shelf life of food. For example, gamma rays from cobalt-60 can be used to kill bacteria and insects that spoil and infest food.

## Nuclear Waste

### Nuclear Fission and Nuclear Fusion

In nuclear fission, the nucleus of a very heavy atom, such as uranium, is split into two or more lighter nuclei. The products of the fission include the nuclei as well as the nucleons formed from the fragments' radioactive decay. Fission is the primary process powering nuclear reactors, which include those on nuclear-powered submarines and aircraft carriers. Fusion is the opposite process of fission. In fusion, very high temperatures and pressures are used to combine light atoms, such as hydrogen, to make heavier atoms, such as helium. Fusion is the primary



**FIGURE 13** Radioactive nuclides, such as technetium-99, can be used to detect bone cancer. In this procedure, technetium-99 accumulates in areas of abnormal bone metabolism. Detection of the nuclear radiation then shows the location of bone cancer.

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Topic: Radioisotopes

Code: HC61260

process that fuels our sun and the stars. Creating and maintaining a fusion reaction is more complex and expensive than performing fission. Both fission and fusion release enormous amounts of energy that can be converted into energy as heat and electrical energy, and both produce **nuclear waste**. Fission produces more waste than fusion. As new processes are developed to use energy from fission and fusion, a more vexing question arises: how to contain, store, and dispose of nuclear waste.

### Containment of Nuclear Waste

Every radioactive substance has a half-life, which is the amount of time needed for half of a given material to decay. Radioactive waste from medical research, for example, usually has a half-life that is a few months or less. Some of the waste that is produced in a nuclear reactor will take hundreds of thousands of years to decay, and it needs to be contained so that living organisms can be shielded from radioactivity. There are two main types of containment: on-site storage and off-site disposal.

### Storage of Nuclear Waste

The most common form of nuclear waste is spent fuel rods from nuclear power plants. These fuel rods can be contained above the ground by placing them in water pools or in dry casks. Each nuclear reactor in the United States has large pools of water where spent rods can be stored, and some of the radioactive materials will decay. When these pools are full, the rods are moved to dry casks, which are usually made of concrete and steel. Both storage pools and casks are meant for only temporary storage before the waste is moved to permanent underground storage facilities.

### Disposal of Nuclear Waste

Disposal of nuclear waste is done with the intention of never retrieving the materials. Because of this, building disposal sites takes careful planning. Currently, there are 131 disposal sites in 39 states around the United States. The U. S. Department of Energy is developing a new site near Las Vegas, Nevada, called Yucca Mountain, for the permanent disposal of much of this waste. Nuclear waste could be transported there by truck and train beginning in 2017. This plan is controversial—some organizations oppose the idea of the disposal site, and others have proposed alternate plans.

## SECTION REVIEW

1. What is required to shield alpha particles? Why are these materials effective?
2.
  - a. What is the average exposure of people living in the United States to environmental background radiation?
  - b. How does this relate to the maximum permissible dose?

3. What device is used to measure the radiation exposure of people working with radiation?
4. Explain why nuclear radiation can be used to preserve food.

### Critical Thinking

5. **INFERRING CONCLUSIONS** Explain how nuclear waste is contained, stored, and disposed of, and how each method affects the environment.

# Nuclear Fission and Nuclear Fusion

## SECTION 4

### OBJECTIVES

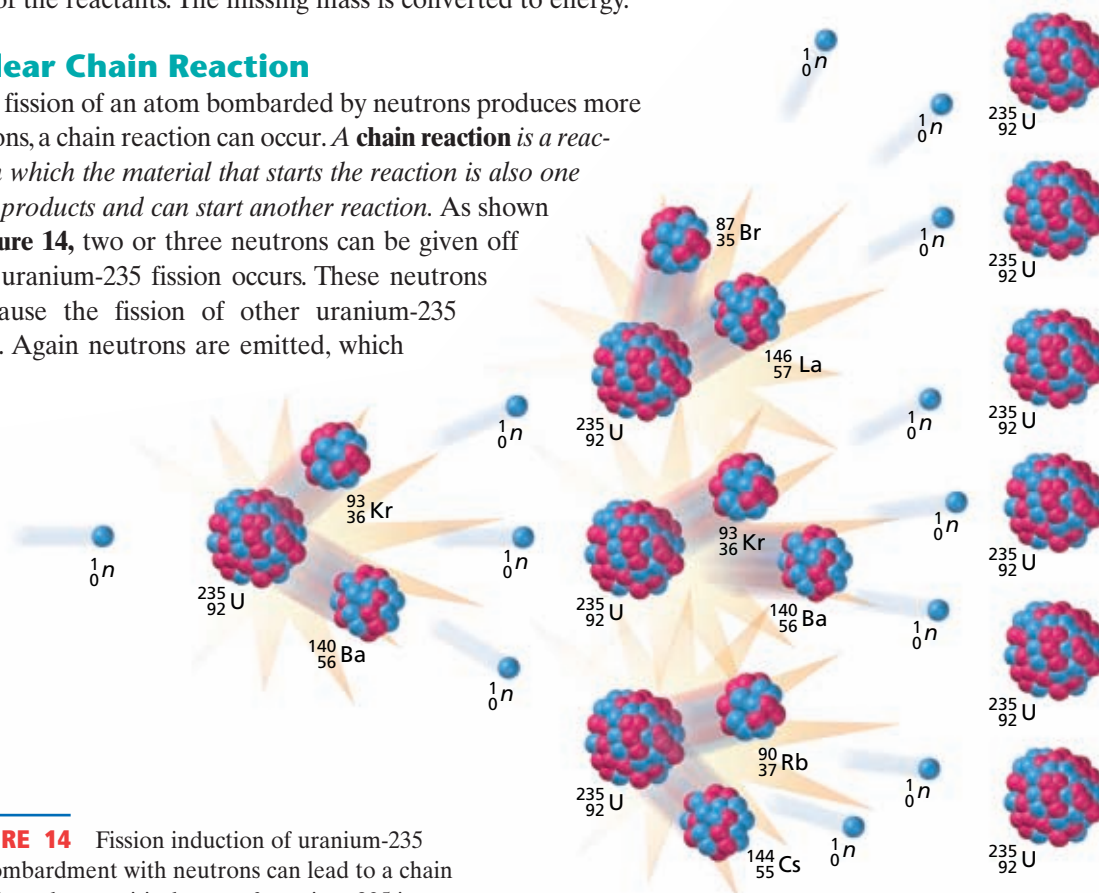
- Define *nuclear fission*, *chain reaction*, and *nuclear fusion*, and distinguish between them.
- Explain how a fission reaction is used to generate power.
- Discuss the possible benefits and the current difficulty of controlling fusion reactions.

## Nuclear Fission

Review **Figure 1**, which shows that nuclei of intermediate mass are the most stable. In **nuclear fission**, a very heavy nucleus splits into more-stable nuclei of intermediate mass. This process releases enormous amounts of energy. Nuclear fission can occur spontaneously or when nuclei are bombarded by particles. When uranium-235 is bombarded with slow neutrons, a uranium nucleus can capture one of the neutrons, making it very unstable. The nucleus splits into medium-mass nuclei with the emission of more neutrons. The mass of the products is less than the mass of the reactants. The missing mass is converted to energy.

## Nuclear Chain Reaction

When fission of an atom bombarded by neutrons produces more neutrons, a chain reaction can occur. A **chain reaction** is a reaction in which the material that starts the reaction is also one of the products and can start another reaction. As shown in **Figure 14**, two or three neutrons can be given off when uranium-235 fission occurs. These neutrons can cause the fission of other uranium-235 nuclei. Again neutrons are emitted, which



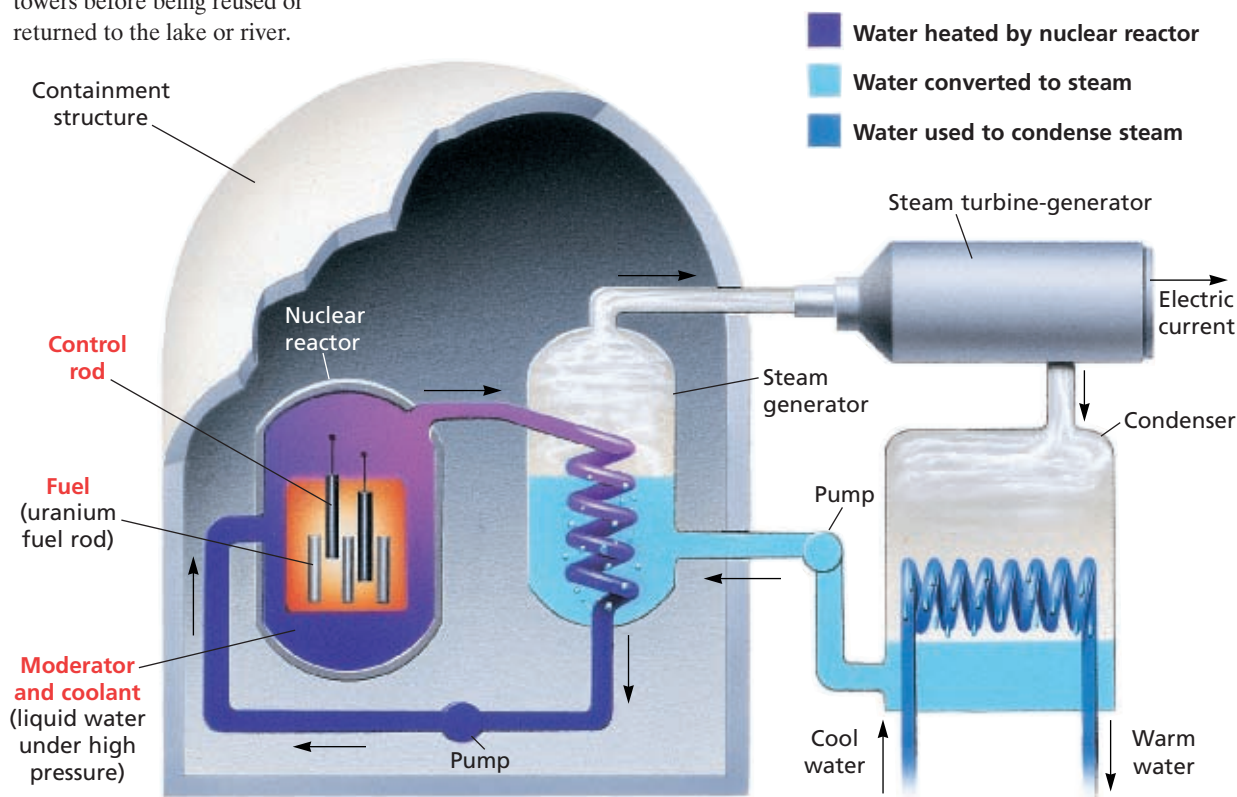
**FIGURE 14** Fission induction of uranium-235 by bombardment with neutrons can lead to a chain reaction when a critical mass of uranium-235 is present.

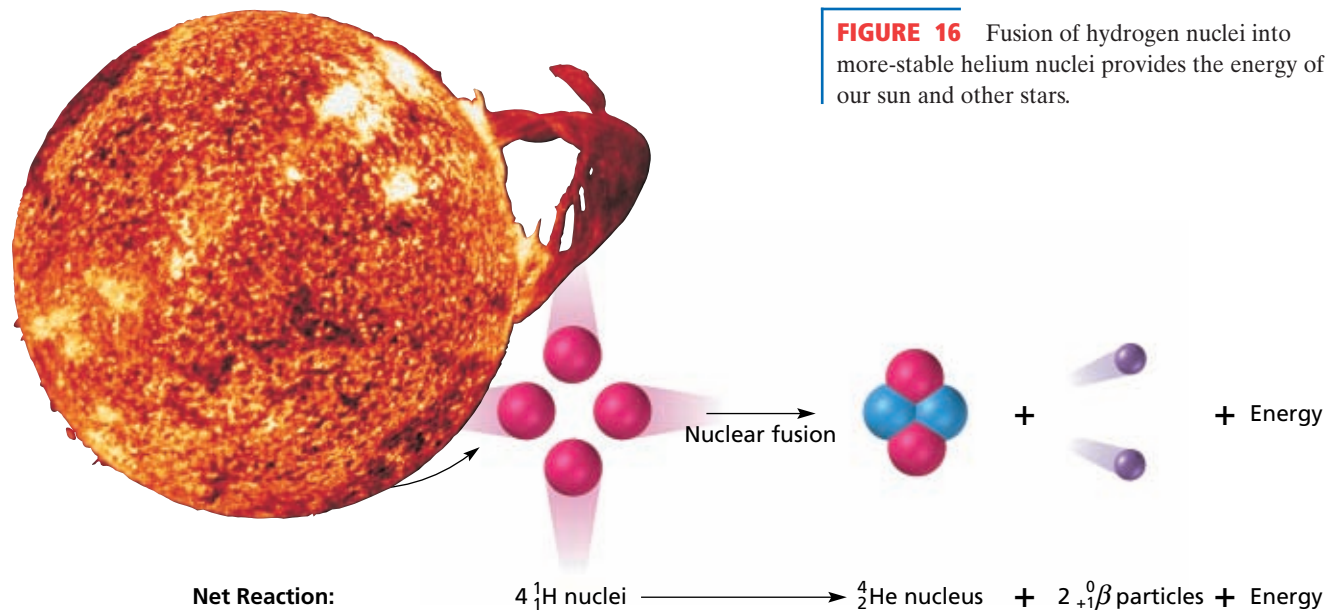
can cause the fission of still other uranium-235 nuclei. This chain reaction continues until all of the uranium-235 atoms have split or until the neutrons fail to strike uranium-235 nuclei. If the mass of the uranium-235 sample is below a certain minimum, too many neutrons will escape without striking other nuclei, and the chain reaction will stop. *The minimum amount of nuclide that provides the number of neutrons needed to sustain a chain reaction is called the **critical mass**.* Uncontrolled chain reactions provide the explosive energy of atomic bombs. **Nuclear reactors** use controlled-fission chain reactions to produce energy and radioactive nuclides.

## Nuclear Power Plants

**Nuclear power plants** convert the heat produced by nuclear fission into electrical energy. They have five main components: shielding, fuel, control rods, moderator, and coolant. The components, shown in **Figure 15**, are surrounded by shielding. **Shielding** is radiation-absorbing material that is used to decrease the emission of radiation, especially gamma rays, from nuclear reactors. Uranium-235 is typically used as the fissile fuel. Its fission produces energy as heat, which is absorbed by the coolant. **Control rods** are neutron-absorbing rods that help control the reaction by limiting the number of free neutrons. Because fission of uranium-235 is more efficiently induced by slow neutrons, a **moderator** is used to slow down the fast neutrons produced by fission. Nuclear power plants can provide competitively priced electricity without emitting greenhouse gases or particulates. Concerns about nuclear power include storage and disposal of spent radioactive fuel, as well as public perception.

**FIGURE 15** In this model of a nuclear power plant, pressurized water is heated by fission of uranium-235. This water is circulated to a steam generator. The steam drives a turbine to produce electricity. Cool water from a lake or river is then used to condense the steam into water. The warm water from the condenser may be cooled in cooling towers before being reused or returned to the lake or river.





**FIGURE 16** Fusion of hydrogen nuclei into more-stable helium nuclei provides the energy of our sun and other stars.

## Nuclear Fusion

The high stability of nuclei with intermediate masses can also be used to explain nuclear fusion. In **nuclear fusion**, *low-mass nuclei combine to form a heavier, more stable nucleus*. Nuclear fusion releases even more energy per gram of fuel than nuclear fission. In our sun and stars that are similar to the sun, hydrogen nuclei combine at extremely high temperature and pressure to form a helium nucleus with a loss of mass and release of energy. The net reaction is illustrated in **Figure 16**.

If fusion reactions can be controlled, they could be used for energy generation. Researchers are currently studying ways to contain the reacting plasma that is required for fusion. A plasma is an extremely hot mixture of positive nuclei and electrons. There is no known material that can withstand the initial temperatures, about  $10^8$  K, required to induce fusion. Scientists use strong magnetic fields to suspend the charged plasma inside a container but away from the walls. Additionally, a large amount of energy is needed to initiate fusion reactions. For fusion to be a practical energy source, more energy needs to be generated by the reaction than is put into the reaction.

### SECTION REVIEW

1. Distinguish between nuclear fission and nuclear fusion.
2. Define *chain reaction*.

3. List the five main components of a nuclear power plant.

#### Critical Thinking

4. **RELATING IDEAS** Explain how fusion is one of our sources of energy.



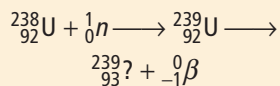
## An Unexpected Finding

### Historical Perspective

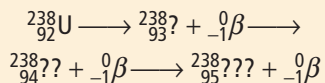
The discovery of the artificial transmutation of uranium in 1934 triggered great excitement in science. Chemists who were preoccupied with identifying what they thought were the final missing elements of the periodic table suddenly had to consider the existence of elements beyond atomic number 92. Physicists began to probe the stability of the nucleus more deeply. By 1939, nuclear investigators in both fields had collaborated to provide a stunning explanation for the mysterious results of uranium's forced transformation.

### Neutrons in Italy

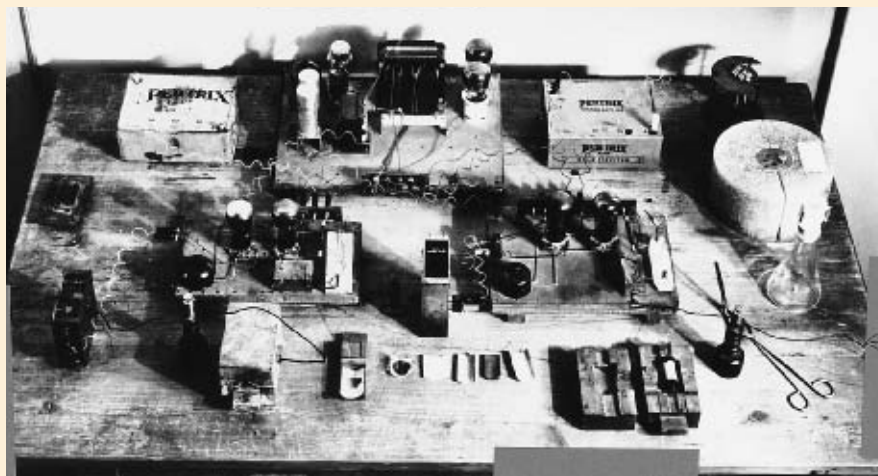
In 1934, uranium had the most protons, 92, of any known element. But that year, Italian physicist Enrico Fermi believed he had synthesized elements that have higher atomic numbers. After bombarding a sample of uranium with neutrons, Fermi and his co-workers recorded measurements that seemed to indicate that some uranium nuclei had absorbed neutrons and then undergone beta decay:



His report noted further, subsequent beta decays, by which he hypothesized the existence of a whole new series of "transuranic" elements, now called *transuranics*:



Unfortunately, Fermi and his group of scientists could not verify the existence of the transuranics because, according



▲ This apparatus from Otto Hahn's lab was used to produce fission reactions.

to Fermi, "We did not know enough chemistry to separate the products of uranium disintegration from one another."

### Curiosity in Berlin

Fermi's experiments caught the attention of a physicist in Berlin, Lise Meitner. Knowing that she could not perform the difficult task of chemically separating radionuclides either, Meitner persuaded a colleague, radiochemist Otto Hahn, to help her explain Fermi's results. Joined by expert chemical analyst Fritz Strassman, Meitner's team began investigating neutron-induced uranium decay at the end of 1934.

From the onset, Meitner's team, as well as all other scientists at the time, operated under two false assumptions. The first involved the makeup of the bombarded nuclei. In every nuclear reaction that had been observed, the resulting nucleus had never differed from the original by more than a few protons or neutrons. Thus, scientists assumed that the products of neutron bombardment were radioisotopes of elements that were at most a few places in the periodic table before or beyond the atoms being bombarded (as Fermi had presumed in hypothesizing the transuranics).

The second assumption concerned the periodicity of the transuranics. Because the elements Ac, Th, Pa, and U chemically resembled the transition elements in the third row of the periodic table, La, Hf, Ta, and W, scientists believed that elements beyond U would correspondingly resemble those following W. Thus, the transuranics were thought to be homologues of Re, Os, Ir, Pt, and the other transition elements in the third row. This belief was generally unquestioned and seemed to be confirmed. In fact, by 1937 Hahn was sure that the chemical evidence of transuranics confirmed their location in the periodic table.

### Meitner's Exile

By 1938, the political situation in Germany had become dangerous for Meitner. Because she was of Jewish descent, she was targeted by the Nazis and fled to Sweden to escape persecution. Meanwhile in Berlin, Hahn and Strassman, who were critical of the Nazis, had to be careful.

Despite being censored by the Nazis, Meitner's team continued to communicate through letters. Meitner could not formulate a satisfying physical explanation for the chemical results of Hahn and Strassman, and she insisted that her partners reexamine their findings. Because of her colleagues' great respect for her talent and expertise, they quickly performed control experiments to test their results.

### A Shocking Discovery

Prompted by Meitner, Hahn and Strassman realized they had been looking in the wrong place to find the cause of their results. In analyzing a fraction of a solution assay that they had previously ignored, they found the critical evidence they had been seeking.

The analysis indicated that barium appeared to be a result of neutron bombardment of uranium. Suspecting the spectacular truth but lacking confidence, Hahn wrote to Meitner for an explanation. After consultation with her nephew, Otto Frisch, Meitner proposed that the uranium nuclei had been broken apart into elemental fragments, one of which was barium, Ba. On January 3, 1939, she wrote to Hahn to congratulate him on the ground-breaking result.

Thus, the "transuranics" turned out to be merely radioisotopes of known elements—atomic fragments of uranium atoms that had burst apart when struck by neutrons.



▲ The politics of World War II prevented Lise Meitner from receiving the Nobel Prize in physics for explaining nuclear fission.

For the discovery of this unexpected phenomenon, which Meitner named nuclear fission, the talented Hahn was awarded the 1944 Nobel Prize in chemistry. Because of wartime politics, however, Lise Meitner did not receive the corresponding award in physics. She was not properly recognized until well after her death in 1968 for her role in clarifying the process that she first explained and named.

### Questions

1. What type of element did Fermi expect to find when uranium absorbed a neutron and then ejected a beta particle?
2. What were the products of uranium disintegrations?

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Topic: Enrico Fermi

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# CHAPTER HIGHLIGHTS

## The Nucleus

### Vocabulary

nucleon	nuclear shell model
nuclide	magic numbers
mass defect	nuclear reaction
nuclear binding energy	transmutation

- The difference between the sum of the masses of the nucleons and electrons in an atom and the actual mass of an atom is the mass defect, or nuclear binding energy.
- Nuclear stability tends to be greatest when nucleons are paired, when there are magic numbers of nucleons, and when there are certain neutron-proton ratios.
- Nuclear reactions, which are represented by nuclear equations, can involve the transmutation of nuclides.

## Radioactive Decay

### Vocabulary

radioactive decay	half-life
nuclear radiation	decay series
radioactive nuclide	parent nuclide
alpha particle	daughter nuclide
beta particle	artificial
positron	transmutation
electron capture	transuranium element
gamma ray	

- Radioactive nuclides become more stable by radioactive decay.
- Alpha, beta, positron, and gamma emission, and electron capture are types of radioactive decay. The type of decay is related to the nucleon content and the energy level of the nucleus.
- The half-life of a radioactive nuclide is the length of time that it takes for half of a given number of atoms of the nuclide to decay.
- Artificial transmutations are used to produce artificial radioactive nuclides, which include the transuranium elements.

## Nuclear Radiation

### Vocabulary

roentgen
rem
film badge
Geiger-Müller counter
scintillation counter
radioactive dating
radioactive tracer
nuclear waste

- Alpha particles, beta particles, and gamma rays have different penetrating abilities and shielding requirements.
- Film badges, Geiger-Müller counters, and scintillation detectors are used to detect radiation.
- Everyone is exposed to environmental background radiation.
- Radioactive nuclides have many uses, including radioactive dating, disease detection, and therapy.
- Nuclear waste must be contained, stored, and disposed of in a way that minimizes harm to people or the environment.

## Nuclear Fission and Nuclear Fusion

### Vocabulary

nuclear fission	shielding
chain reaction	control rod
critical mass	moderator
nuclear reactor	nuclear fusion
nuclear power plant	

- Nuclear fission and nuclear fusion are nuclear reactions in which the splitting and fusing of nuclei produce more stable nuclei and release enormous amounts of energy.
- Controlled fission reactions produce energy and radioactive nuclides.
- Fusion reactions produce the sun's energy in the form of heat and light. If fusion reactions could be controlled, they would produce more usable energy per gram of fuel than fission reactions.

# CHAPTER REVIEW

## The Nucleus

### SECTION 1 REVIEW

- How does mass defect relate to nuclear binding energy?
  - How does binding energy per nucleon vary with mass number?
  - How does binding energy per nucleon affect the stability of a nucleus?
- Describe three ways in which the number of protons and the number of neutrons in a nucleus affect the stability of the nucleus.

### PRACTICE PROBLEMS

- The mass of a  $^{20}_{10}\text{Ne}$  atom is 19.992 44 amu. Calculate the atom's mass defect.
- The mass of a  $^7_3\text{Li}$  atom is 7.016 00 amu. Calculate the atom's mass defect.
- Calculate the nuclear binding energy of one lithium-6 atom. The measured atomic mass of lithium-6 is 6.015 amu.
- Calculate the binding energies of the following two nuclei, and indicate which nucleus releases more energy when formed. You will need information from the periodic table and the text.
  - atomic mass 34.988011 amu,  $^{35}_{19}\text{K}$
  - atomic mass 22.989767 amu,  $^{23}_{11}\text{Na}$
- What is the binding energy per nucleon for each nucleus in the previous problem?
  - Which nucleus is more stable?
- The mass of  $^7_3\text{Li}$  is 7.016 00 amu. Calculate the binding energy per nucleon for  $^7_3\text{Li}$ .
- Calculate the neutron-proton ratios for the following nuclides:
  - $^{12}_6\text{C}$
  - $^3_1\text{H}$
  - $^{206}_{82}\text{Pb}$
  - $^{134}_{50}\text{Sn}$
- Locate the nuclides in problem 9 on the graph in **Figure 2**. Which ones lie within the band of stability?
  - For the stable nuclides, determine whether their neutron-proton ratio tends toward 1:1 or 1.5:1.

- Balance the following nuclear equations. (Hint: See Sample Problem A.)
  - $^{43}_{19}\text{K} \longrightarrow ^{43}_{20}\text{Ca} + \underline{\quad ?}$
  - $^{233}_{92}\text{U} \longrightarrow ^{229}_{90}\text{Th} + \underline{\quad ?}$
  - $^{11}_6\text{C} + \underline{\quad ?} \longrightarrow ^{11}_5\text{B}$
  - $^{13}_7\text{N} \longrightarrow ^0_{+1}\beta + \underline{\quad ?}$
- Write the nuclear equation for the release of an alpha particle by  $^{210}_{84}\text{Po}$ .
- Write the nuclear equation for the release of a beta particle by  $^{210}_{82}\text{Pb}$ .

## Radioactive Decay

### SECTION 2 REVIEW

- Where on the periodic table are most of the natural radioactive nuclides located?
- What changes in atomic number and mass number occur in each of the following types of radioactive decay?
  - alpha emission
  - beta emission
  - positron emission
  - electron capture
- Which types of radioactive decay cause the transmutation of a nuclide? (Hint: Review the definition of *transmutation*.)
- Explain how beta emission, positron emission, and electron capture affect the neutron-proton ratio.
- Write the nuclear reactions that show particle conversion for the following types of radioactive decay:
  - beta emission
  - positron emission
  - electron capture
- Compare electrons, beta particles, and positrons.
- What are gamma rays?
  - How do scientists think gamma rays are produced?
- How does the half-life of a nuclide relate to the stability of the nuclide?
- List the three parent nuclides of the natural decay series.

- How are artificial radioactive isotopes produced?
- Neutrons are more effective for bombarding atomic nuclei than protons or alpha particles are. Why?
- Why are all of the transuranium elements radioactive? (Hint: See Section 1.)

### PRACTICE PROBLEMS

- The half-life of plutonium-239 is 24 110 years. Of an original mass of 100.g, how much plutonium-239 remains after 96 440 years? (Hint: See Sample Problem B.)
- The half-life of thorium-227 is 18.72 days. How many days are required for three-fourths of a given amount of thorium-227 to decay?
- Exactly  $\frac{1}{16}$  of a given amount of protactinium-234 remains after 26.76 hours. What is the half-life of protactinium-234?
- How many milligrams of a 15.0 mg sample of radium-226 remain after 6396 years? The half-life of radium-226 is 1599 years.

## Nuclear Radiation

### SECTION 3 REVIEW

- Why can a radioactive material affect photographic film even though the film is well wrapped in black paper?
- How does the penetrating ability of gamma rays compare with that of alpha particles and beta particles?
- How does nuclear radiation damage biological tissue?
- Explain how film badges, Geiger-Müller counters, and scintillation detectors are used to detect radiation and measure radiation exposure.
- How is the age of an object that contains a radioactive nuclide estimated?

## Nuclear Fission and Nuclear Fusion

### SECTION 4 REVIEW

- How is the fission of a uranium-235 nucleus induced?
- How does the fission of uranium-235 produce a chain reaction?
- Describe the purposes of the five major components of a nuclear power plant.
- Describe the reaction that produces the sun's energy.
- What is one problem that must be overcome before controlled fusion reactions that produce energy are a reality?

### MIXED REVIEW

- Balance the following nuclear reactions:
  - ${}_{93}^{239}\text{Np} \longrightarrow {}_{-1}^0\beta + \underline{\quad ? \quad}$
  - ${}_4^9\text{Be} + {}_2^4\text{He} \longrightarrow \underline{\quad ? \quad}$
  - ${}_{15}^{32}\text{P} + \underline{\quad ? \quad} \longrightarrow {}_{15}^{33}\text{P}$
  - ${}_{92}^{236}\text{U} \longrightarrow {}_{36}^{94}\text{Kr} + \underline{\quad ? \quad} + 3{}_0^1n$
- After 4797 years, how much of the original 0.250 g of radium-226 remains? The half-life of radium-226 is 1599 years.
- The parent nuclide of the thorium decay series is  ${}_{90}^{232}\text{Th}$ . The first four decays are as follows: alpha emission, beta emission, beta emission, and alpha emission. Write the nuclear equations for this series of emissions.
- The half-life of radium-224 is 3.66 days. What was the original mass of radium-224 if 0.0500 g remains after 7.32 days?
- Calculate the neutron-proton ratios for the following nuclides, and determine where they lie in relation to the band of stability.
 

a. ${}_{92}^{235}\text{U}$	c. ${}_{26}^{56}\text{Fe}$
b. ${}_{8}^{16}\text{O}$	d. ${}_{60}^{156}\text{Nd}$
- Calculate the binding energy per nucleon of  ${}_{92}^{238}\text{U}$  in joules. The atomic mass of a  ${}_{92}^{238}\text{U}$  nucleus is 238.050 784 amu.



# Math Tutor

## CALCULATING WITH HALF-LIFE

The rate at which a sample of a radioactive nuclide decays is expressed in terms of half-life. This quantity is the time required for half of the atoms of a sample of a given nuclide to decay. For example, it takes 37.2 min for half of the nuclei of chlorine-38 to decay to argon-38. After 37.2 min, 0.50 g of a 1.0 g sample of chlorine-38 will remain and there will be 0.50 g of argon-38. After two half-lives (74.4 min), the fraction of chlorine-38 that remains will be  $\frac{1}{2}$  of  $\frac{1}{2}$ , or  $\frac{1}{4}$ .

After  $n$  half-lives, the fraction of a radioactive nuclide that remains is  $\left(\frac{1}{2}\right)^n$ , or  $2^{-n}$ .

If you know the amount of nuclide that was present initially and the amount of nuclide that remains, you can determine the number of half-lives that have passed.

### Problem-Solving TIPS

- Familiarize yourself with the values of some common powers of two ( $2^n$ ,  $n = 1, 2, 3, 4, 5, 6$ , etc). This will allow you to determine the number of half-lives quickly.

#### SAMPLE 1

**The half-life of polonium-218 is 3.04 min. A sample of polonium contains 0.00558 g of  $^{218}_{84}\text{Po}$ . What mass of  $^{218}_{84}\text{Po}$  will remain after 18.24 min?**

First, you must determine the number of half-lives that have passed in 18.24 min.

$$\begin{aligned}\text{number of half-lives} &= \frac{\text{time elapsed}}{\text{half-life}} = \\ &= \frac{18.24 \text{ min}}{3.04 \text{ min}} = 6.00 \text{ half-lives}\end{aligned}$$

Then, to determine the mass of polonium-218 remaining, apply the following relationship.

$$\text{mass remaining} = \text{starting mass} \times \text{fraction remaining}$$

$$\begin{aligned}\text{mass } ^{218}_{84}\text{Po remaining} &= 0.00558 \text{ g} \times \left(\frac{1}{2}\right)^6 = \\ &= 0.00558 \text{ g} \times \frac{1}{64} = 8.72 \times 10^{-5} \text{ g } ^{218}_{84}\text{Po}\end{aligned}$$

#### SAMPLE 2

**The half-life of potassium-40 is  $1.3 \times 10^9$  years. A volcanic rock contains  $\frac{1}{8}$  of the amount of potassium-40 found in newly formed rocks. When was the rock formed?**

First, determine the number of half-lives that have passed.

$$\text{fraction remaining} = \frac{1}{8} = \left(\frac{1}{2}\right) \times \left(\frac{1}{2}\right) \times \left(\frac{1}{2}\right)$$

Therefore, three half-lives have passed. The time since the rock was formed is 3 half-lives  $\times$   $\left(1.3 \times 10^9 \frac{\text{y}}{\text{half-life}}\right) = 3.9 \times 10^9 \text{ y}$ .

#### PRACTICE PROBLEMS

- A sample of chromium contains  $8.9 \times 10^{-5} \text{ g}$  of the radioactive nuclide chromium-51, which has a half-life of 28 days. What mass of chromium-51 will remain in the sample after 168 days?
- The half-life of lead-202 is 53 000 years. A sample of lead contains only  $\frac{1}{256}$  of the expected amount of lead-202. How old is the lead sample?

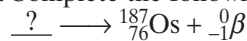


# Standardized Test Prep

Answer the following items on a separate piece of paper.

## MULTIPLE CHOICE

1. Complete the following nuclear equation:



- A.  ${}_{77}^{187}\text{Os}$
  - B.  ${}_{75}^{187}\text{Os}$
  - C.  ${}_{77}^{187}\text{Ir}$
  - D.  ${}_{75}^{187}\text{Re}$
2. The mass of the nucleus is
- A. greater than the mass of the protons and neutrons that make up the nucleus.
  - B. equal to the mass of the protons and neutrons that make up the nucleus.
  - C. less than the mass of the protons and neutrons that make up the nucleus.
  - D. converted to energy.
3. Which type of radiation has the most penetrating ability?
- A. an alpha particle
  - B. a beta particle
  - C. a gamma ray
  - D. a neutron
4. Which two particles have the same mass but opposite charge?
- A. a beta particle and a positron
  - B. a neutron and a proton
  - C. a proton and an electron
  - D. an alpha particle and a proton
5. Which of the following nuclear equations is correctly balanced?
- A.  ${}_{18}^{37}\text{Ar} + {}_{-1}^0e \longrightarrow {}_{17}^{37}\text{Cl}$
  - B.  ${}_{3}^6\text{Li} + 2{}_{0}^1n \longrightarrow {}_{2}^4\text{He} + {}_{1}^3\text{H}$
  - C.  ${}_{99}^{254}\text{Es} + {}_{2}^4\text{He} \longrightarrow {}_{101}^{258}\text{Md} + 2{}_{0}^1n$
  - D.  ${}_{7}^{14}\text{N} + {}_{2}^4\text{He} \longrightarrow {}_{8}^{17}\text{O} + {}_{1}^2\text{H}$
6. Gamma rays
- A. have the same energy as beta particles do.
  - B. are visible light.
  - C. have no charge and no mass.
  - D. are not a form of electromagnetic radiation.

7. Which of the following nuclides is radioactive?

- A.  ${}_{20}^{40}\text{Ca}$
- B.  ${}_{88}^{226}\text{Ra}$
- C.  ${}_{6}^{12}\text{C}$
- D.  ${}_{82}^{206}\text{Pb}$

8. The half-life of thorium-234 is 24 days. If you have a 42.0 g sample of thorium-24, how much will remain after 72 days?

- A. 42.0 g
- B. 21.0 g
- C. 10.5 g
- D. 5.25 g

9. It takes 5.2 min for a 4.0 g sample of francium-210 to decay until only 1.0 g is left. What is the half-life of francium-210?

- A. 1.3 min
- B. 2.6 min
- C. 5.2 min
- D. 7.8 min

## SHORT ANSWER

10. Write the nuclear equation that represents the process in which a neutron in the nucleus is changed to a proton with the emission of a beta particle.

11. Describe a positron, and write its nuclear symbol.

## EXTENDED RESPONSE

12. Explain the difference between nuclear fission and nuclear fusion, and explain the energy changes that accompany each process.

13. What is meant by the term *mass defect*?

## Test TIP

Keeping a positive attitude during any test will help you focus on the test and likely improve your score.

# Simulation of Nuclear Decay Using Pennies and Paper

## OBJECTIVES

- *Infer* that the rate of decay can be simulated by a random process.
- *Compare* the numbers of pennies that are showing heads with the number showing tails.
- *Create* a string plot that represents nuclear decay.
- *Relate* observations to the rate of nuclear decay.
- *Graph* the data.
- *Compare* the results of the two simulation procedures.

## MATERIALS

- colored paper or cloth strips, approximately 65 cm × 2.5 cm (2 strips)
- graph paper
- one sheet of stiff poster board, 70 cm × 60 cm
- pennies or other objects supplied by your teacher (100)
- scissors, tape, meter stick, pencil, and string
- shoe box with lid



## BACKGROUND

Radioactive isotopes are unstable. All radioactive matter decays, or breaks down, in a predictable pattern. Radioactive isotopes release radiation as they disintegrate into daughter isotopes.

The rate of decay is a measure of how fast an isotope changes into its daughter isotope. The rate of radioactive decay is conveniently characterized by the isotope's half-life, the period of time it takes one-half of the original material to decay. Half-lives vary from billions of years to fractions of a second.

## SAFETY



For a review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

## PREPARATION

1. For Part A, make a data table that has at least 10 rows in your lab notebook by using the format described by your teacher.

## PROCEDURE

### Part A: Simulating radioactive decay with pennies

1. Place 100 pennies into the shoe box so that the head sides are up. The pennies will represent atoms. Record 100 in the “Unchanged atoms” column and 0 in the “Changed atoms” column.
2. With the lid on the box, shake the box up and down 5 times. We will count each shaking period as being equivalent to 10 s.
3. Open the lid, and remove all of the pennies that have the tails side up. These pennies represent the changed atoms.

- Count the number of pennies remaining in the box. Record this number in the 10 s row of the “Unchanged atoms” column. Count the number of changed atoms (the pennies that you removed from the box), and record the number table in the 10 s row.
- Each lab partner should predict how many times steps 2–4 will need to be repeated until only one unchanged atom remains. Record the time that each lab partner predicted. Remember that each shaking period is counted as 10 s, so four shaking periods would be recorded as 40 s.
- Repeat steps 2–4 by counting and recording each time until only 1 (or 0) penny with the head side up remains.

#### Part B: Simulating decay with paper

- Draw an  $y$ -axis and  $x$ -axis on the poster board so that they are about 5 cm from the left side and the bottom edge respectively. Label the  $x$ -axis as “Time” and the  $y$ -axis as “Amount of material.”
- Along the  $x$ -axis, draw marks every 10 cm from the  $y$ -axis line. Label the first mark “0” and the next mark “1,” and so on. Each mark represents 1 minute.
- Place one of the colored strips vertically with its lower edge centered on the 0 mark of the  $x$ -axis. Tape the strip in place.
- Fold the other colored strip in half, and cut it in the middle. Place one-half of the strip so that it is centered on the next mark, and tape the strip in place.
- Fold the remaining piece of the strip in half, and cut it exactly in the middle.
- Place one of the pieces so that it is centered on the next mark, and tape the piece in place.
- Repeat steps 11 and 12, and each time, tape the first piece vertically at the next  $x$ -axis mark. Continue until you have at least 8 strips taped along the  $x$ -axis.
- Use the string to join the tops of each strip of paper to make a continuous curve.

## CLEANUP AND DISPOSAL

- Return the pennies and box to your teacher. Dispose of the poster board, strips, and string as instructed by your teacher. Clean up your lab station.



## ANALYSIS AND INTERPRETATION

### Part A

- Predicting Outcomes:** How long did it take to have only 1 penny (0 pennies) left in the box? How close was your prediction in step 5?
- Analyzing Data:** Make a graph of your data on a piece of graph paper. Label the  $x$ -axis “Time” and the  $y$ -axis “Unchanged atoms.” Plot the number of unchanged atoms versus time. Draw a smooth curve through the data points.
- Analyzing Results:** Each trial was comparable to a 10 s period of time. How long did it take for half of your pennies to be removed from the box? What is the half-life of the process?
- Interpreting Graphics:** Use your graph to determine the time it takes to have only 25% of the unchanged atoms remaining. In your experiment, how many pennies remained in the box at that time?

### Part B

- Analyzing Results:** How many half lives have passed after 4 minutes?
- Interpreting Graphics:** Using the string plot, determine how many minutes it took until only 20% of the original material remained.

## CONCLUSIONS

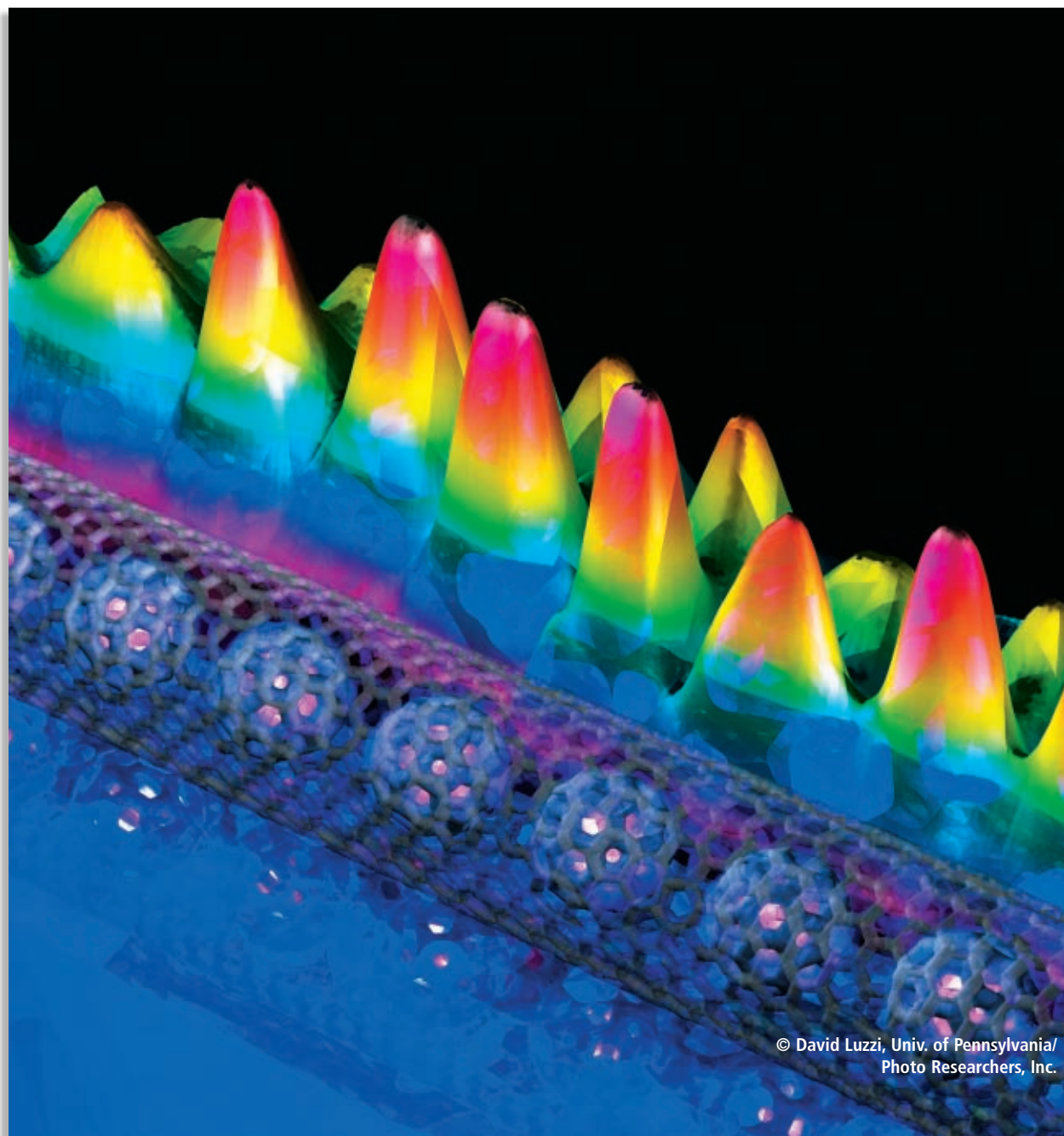
- Inferring Conclusions:** If you started with a paper strip that was twice as long, would the half-life change?
- Inferring Conclusions:** Is there a relationship between the graph from Part A and the string plot from Part B?



# Organic Chemistry

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*Nanotubes and buckminsterfullerene balls are molecules that contain repeated hexagons of carbon atoms. The rainbow pattern is a graph of electron distribution around the molecules.*



# Organic Compounds

## SECTION 1

### OBJECTIVES

- Explain how the structure and bonding of carbon lead to the diversity and number of organic compounds.
- Compare the use of molecular and structural formulas to represent organic compounds.
- Compare structural and geometric isomers of organic compounds.

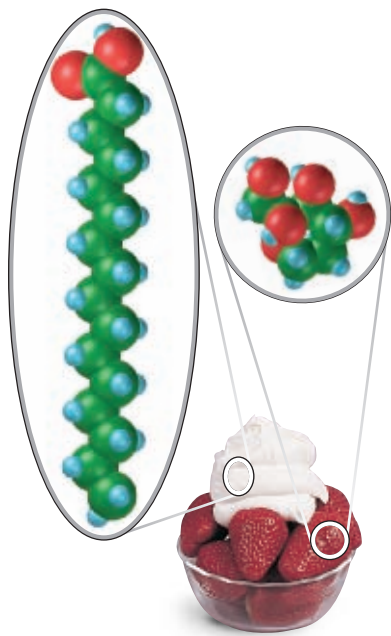
All organic compounds contain carbon atoms. However, not all carbon-containing compounds are classified as organic. There are a few exceptions, such as  $\text{Na}_2\text{CO}_3$ ,  $\text{CO}$ , and  $\text{CO}_2$ , that are considered inorganic. **Organic compounds**, then, can be defined as *covalently bonded compounds containing carbon, excluding carbonates and oxides*. **Figure 1** shows a few familiar items that contain organic compounds.

## Carbon Bonding and the Diversity of Organic Compounds

The diversity of organic compounds results from the uniqueness of carbon's structure and bonding. Carbon's electronic structure allows it to bind to itself to form chains and rings, to bind covalently to other elements, and to bind to itself and other elements in different arrangements.



**FIGURE 1** Aspirin, polyethylene in plastic bags, citric acid in fruit, and amino acids in animals are all examples of organic compounds.



**FIGURE 2** Compare the shape of a fatty acid found in cream with that of fructose, found in fruit. In the fatty acid, the carbon atoms are in chains. In fructose, carbon atoms form a ring.

## Carbon-Carbon Bonding

Carbon atoms are unique in their ability to form long chains and rings of covalently bonded atoms. This type of bonding is known as **catenation**, the covalent bonding of an element to itself to form chains or rings. In addition, carbon atoms in these structures can be linked by single, double, or triple covalent bonds. Examples of molecules containing carbon-atom rings and chains are shown in **Figure 2**.

## Carbon Bonding to Other Elements

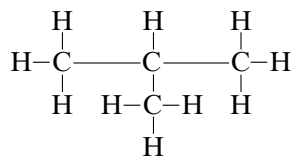
Besides binding to other carbon atoms, carbon atoms bind readily to elements with similar electronegativities. Organic compounds consist of carbon and these other elements. **Hydrocarbons** are composed of only carbon and hydrogen; they are the simplest organic compounds. Other organic compounds contain hydrocarbon backbones to which other elements, primarily O, N, S, and the halogens, are attached. **Figure 3** shows a molecule in which carbon atoms are bound to other elements.

## Arrangement of Atoms

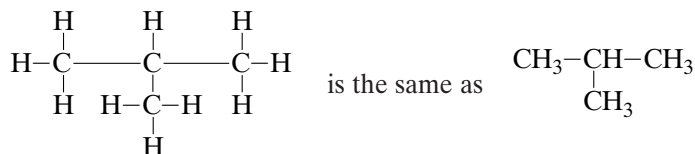
The bonding capabilities of carbon also allow for different arrangements of atoms. This means that some compounds may contain the same atoms but have different properties because the atoms are arranged differently. For example, the molecular formula  $C_2H_6O$  represents both ethanol and dimethyl ether. *Compounds that have the same molecular formula but different structures are called isomers.*

## Structural Formulas

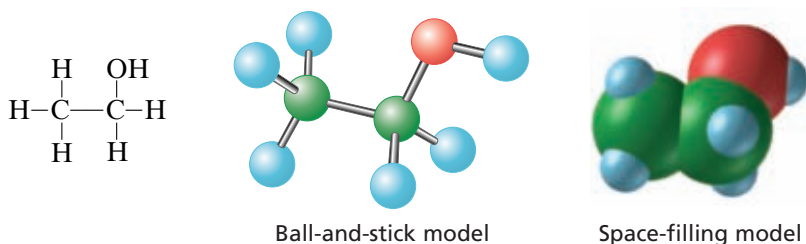
Organic chemists use structural formulas to represent organic compounds. A **structural formula** indicates the number and types of atoms present in a molecule and also shows the bonding arrangement of the atoms. An example of a structural formula for an isomer of  $C_4H_{10}$  is the following.



Structural formulas are sometimes condensed to make them easier to read. In one type of condensed structure, hydrogen single covalent bonds are not shown. The hydrogen atoms are understood to bind to the atom they are written beside. The following structural and condensed structural formulas represent the same molecule.



**FIGURE 3** In firefly luciferin, carbon atoms bind to hydrogen, oxygen, nitrogen, and sulfur. Luciferin is responsible for the light emitted from the tail of a firefly.



**FIGURE 4** The structure of ethanol can be represented in different ways. Ball-and-stick and space-filling models represent the three-dimensional shape of the molecule.

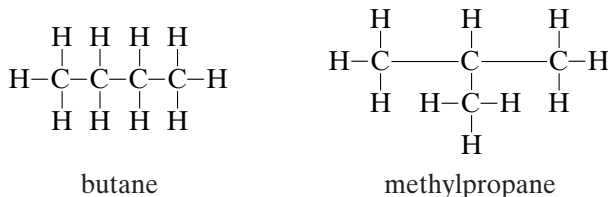
Remember that the structural formula does not accurately show the three-dimensional shape of the molecule. Three-dimensional shape is depicted with drawings or models, as shown for ethanol in **Figure 4**.

## Isomers

You have learned that isomers are compounds that have the same molecular formula but different structural formulas. Isomers can be further classified by structure and geometry.

### Structural Isomers

**Structural isomers**, also called “constitutional isomers,” are isomers in which the atoms are bonded together in different orders. For example, the atoms of the molecular formula  $C_4H_{10}$  can be arranged in two different ways.



Notice that the formula for butane shows a continuous chain of four carbon atoms. The chain may be bent or twisted, but it is continuous. The formula of methylpropane shows a continuous chain of three carbon atoms, with the fourth carbon atom attached to the second carbon atom of the chain.

Structural isomers can have different physical or chemical properties. For example, butane and methylpropane have different melting points, boiling points, and densities, as shown in **Table 1**.

**TABLE 1** Physical Properties of the Structural Isomers Butane and 2-Methylpropane

	Melting point (°C)	Boiling point (°C)	Density at 20°C (g/mL)
butane	-138.4	-0.5	0.5788
2-methylpropane	-159.4	-11.633	0.549

### extension

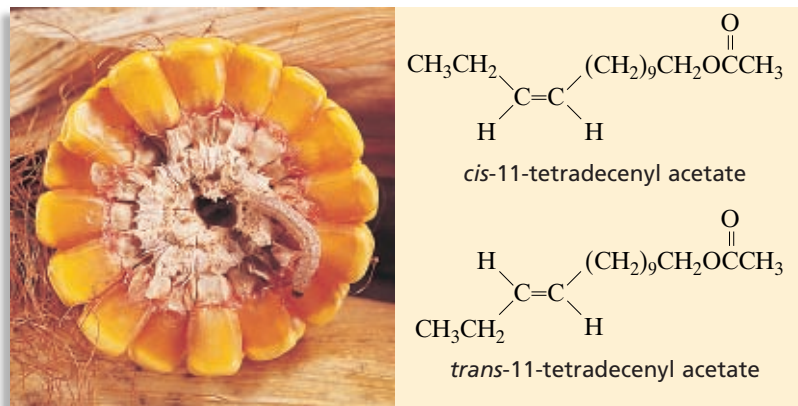
#### Chemical Content

Go to [go.hrw.com](http://go.hrw.com) for another version of this content. See the chapters “Carbon and Hydrocarbons” and “Other Organic Compounds.”



**Keyword:** HC6ORGX

**FIGURE 5** Males of the Iowa strain of the European corn borer respond most strongly to mixtures of the female sex attractant pheromone that are 96% *cis* isomer. But males of the New York strain respond most strongly to mixtures containing 97% *trans* isomer.



## Geometric Isomers

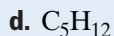
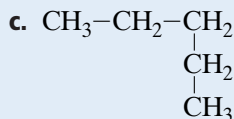
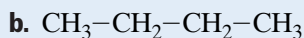
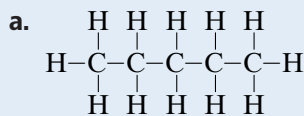
**Geometric isomers** are isomers in which the order of atom bonding is the same but the arrangement of atoms in space is different. Consider the molecule 1,2-dichloroethene, which contains a double bond. The double bond prevents free rotation and holds groups to either side of the molecule. This means there can be two different 1,2-dichloroethene geometric isomers, as shown below.



The geometric isomer with the two chlorine atoms on the same side of the molecule is called *cis*. The isomer with the chlorine atoms on opposite sides of the molecule is called *trans*. **Figure 5** shows an example of geometric isomers that occur in nature.

## SECTION REVIEW

1. Which of the following represent the same molecule?



- What are three characteristics of carbon that contribute to the diversity of organic compounds?
- Define the term *isomer*, and distinguish between structural and geometric isomers.
- Which of the following types of molecular representations can be used to show differences between isomers? Explain why each can or cannot.
  - molecular formula
  - structural formula
  - three-dimensional drawing or model

### Critical Thinking

- INTERPRETING CONCEPTS** Can molecules that have molecular formulas  $\text{C}_4\text{H}_{10}$  and  $\text{C}_4\text{H}_{10}\text{O}$  be isomers of one another? Why or why not?



## The Beginnings of Organic Chemistry

Today, “organic chemistry” refers to the study of carbon compounds. However, organic chemistry was originally thought of as unique among all chemical sciences, because it emphasized the study of compounds that came from living organisms. Chemists of the early 19th century viewed organic compounds as fundamentally different from inorganic compounds, such as rocks and minerals, because organic compounds seemed to come only from living things. But because living organisms are built from carbon-containing molecules, organic chemistry later came to mean the study of carbon compounds, including those that are manufactured artificially.

Like modern researchers, early chemists were interested in the workings of the body. They tried to characterize the chemicals in blood, tissue, and urine. Urea, the molecule that the human body excretes to get rid of ammonia waste, was first isolated from urine in 1773. Although scientists could successfully isolate urea and other organic compounds, they did not know how to synthesize—that is, make from other, simpler chemicals—organic compounds. This tended to confirm a commonly-held belief, called *vitalism*, that organic compounds could only be made inside living organisms with the help of a special life force known as the “vital force.”



▲ *Friedrich Wöhler was the first to synthesize an organic compound from inorganic chemicals.*

In 1828, the young German chemist Friedrich Wöhler announced that he had been able to make urea from inorganic chemicals. Wöhler had been attempting to prepare ammonium cyanate,  $\text{NH}_4\text{OCN}$ , from inorganic compounds, such as silver cyanate and ammonium chloride. But Wöhler unintentionally synthesized urea as a byproduct of the reactions that he carried out! Wöhler made an exciting discovery—the first example of organic synthesis—by using the same principles of qualitative analysis used by chemists today.

The remainder of the 19th century saw the syntheses of many other organic compounds. In 1845, acetic acid was prepared in several steps from charcoal. Many other organic molecules, such as dyes and glucose (blood sugar), were synthesized in the last half of the 19th century. Gradually, as more natural products were prepared in chemistry laboratories, the concept of vitalism was abandoned.

Today, urea, acetic acid, and many other organic chemicals are produced in huge quantities. Organic chemists can synthesize complex drug molecules, such as penicillin and taxol, which were once available only from natural sources. Using methods of organic synthesis, chemists can also prepare completely new drugs, polymers, flavors, and dyes that are not present in nature.

### Questions

1. What do people usually mean by the term *organic* when they use it to describe food, such as fruits and vegetables?
2. Think of three things that you use in your everyday life and that an organic chemist might be able to make.

## SECTION 2

### OBJECTIVES

- Distinguish among the structures of alkanes, alkenes, alkynes, and aromatic hydrocarbons.
- Write structural formulas and names for alkanes, alkenes, and alkynes.
- Relate properties of different types of hydrocarbons to their structures.

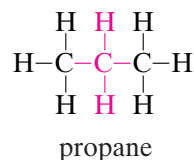
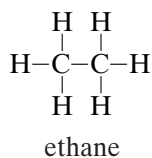
# Hydrocarbons

**H**ydrocarbons are compounds that contain only carbon and hydrogen. They make up the simplest class of organic compounds. All other organic compounds can be viewed as hydrocarbons in which one or more hydrogen atoms have been replaced by other atoms or groups of atoms.

Hydrocarbons are grouped mainly by the type of bonding between carbon atoms. **Saturated hydrocarbons** are hydrocarbons in which each carbon atom in the molecule forms four single covalent bonds with other atoms.

## Alkanes

Hydrocarbons that contain only single bonds are **alkanes**. In **Table 2**, the molecular formulas, structural formulas, and space-filling models are given for alkanes with one to four carbon atoms. If you examine the molecular formulas for successive alkanes in **Table 2**, you will see a clear pattern. Each member of the series differs from the preceding one by one carbon atom and two hydrogen atoms. For example, propane,  $C_3H_8$ , differs from ethane,  $C_2H_6$ , by one carbon atom and two hydrogen atoms, a  $-CH_2-$  group.



Compounds that differ in this fashion belong to a *homologous series*. A homologous series is one in which adjacent members differ by a constant unit. It is not necessary to remember the molecular formulas for all members of a homologous series. Instead, a general molecular formula can be used to determine the formulas. Look at the molecular formulas for ethane and propane,  $C_2H_6$  and  $C_3H_8$ . They both fit the formula  $C_nH_{2n+2}$ . For ethane,  $n = 2$ , so there are two carbon atoms and  $(2 \times 2) + 2 = 6$  hydrogen atoms. For propane,  $n = 3$ , so there are three carbon atoms and  $(2 \times 3) + 2 = 8$  hydrogen atoms. Now consider a molecule for which we do not know the molecular formula. Suppose a member of this series has 30 carbon atoms in its molecules. Then  $n = 30$ , and there are  $(2 \times 30) + 2 = 62$  hydrogen atoms. The formula is  $C_{30}H_{62}$ .


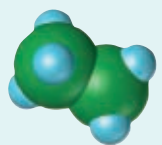
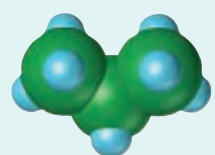
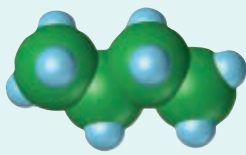
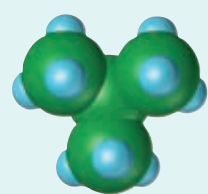
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Topic: Alkanes

Code: HC60045

**TABLE 2** Alkanes with One to Four Carbon Atoms

Molecular formulas	Structural formulas	Space-filling models
$\text{CH}_4$	$\begin{array}{c} \text{H} \\   \\ \text{H}-\text{C}-\text{H} \\   \\ \text{H} \end{array}$ <p>methane</p>	
$\text{C}_2\text{H}_6$	$\begin{array}{c} \text{H} \quad \text{H} \\   \quad   \\ \text{H}-\text{C}-\text{C}-\text{H} \\   \quad   \\ \text{H} \quad \text{H} \end{array}$ <p>ethane</p>	
$\text{C}_3\text{H}_8$	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\   \quad   \quad   \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$ <p>propane</p>	
$\text{C}_4\text{H}_{10}$	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \\   \quad   \quad   \quad   \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\   \quad   \quad   \quad   \\ \text{H} \quad \text{H} \quad \text{H} \quad \text{H} \end{array}$ <p>butane</p> $\begin{array}{c} \text{H} \quad \quad \text{H} \quad \quad \text{H} \\   \quad \quad   \quad \quad   \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\   \quad   \quad   \\ \text{H} \quad \text{H}-\text{C}-\text{H} \\ \quad \quad   \\ \quad \quad \text{H} \end{array}$ <p>methylpropane</p>	 

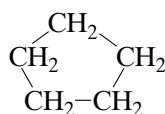
Notice that for alkanes with three or fewer carbon atoms, only one molecular structure is possible. However, in alkanes with more than three carbon atoms, the chains can be straight or branched. Thus, alkanes with four or more carbon atoms have structural isomers. There are two possible structural isomers for alkanes with four carbon atoms, butane and 2-methylpropane.

The number of structural isomers increases greatly as the number of carbon atoms in alkanes increases. There are three isomeric  $\text{C}_5\text{H}_{12}$  alkanes, five isomeric  $\text{C}_6\text{H}_{14}$  alkanes, and nine isomeric  $\text{C}_7\text{H}_{16}$  alkanes. There are nearly 37 million possible isomers of  $\text{C}_{25}\text{H}_{52}$ , though most have never been prepared or isolated.



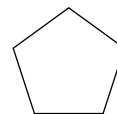
## Cycloalkanes

**Cycloalkanes** are alkanes in which the carbon atoms are arranged in a ring, or cyclic, structure. The structural formulas for cycloalkanes are often drawn in a simplified form. In these skeletal representations, such as the one below on the right, it is understood that there is a carbon atom at each corner and enough hydrogen atoms to complete the four bonds to each carbon atom.



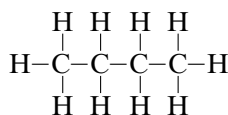
cyclopentane

or

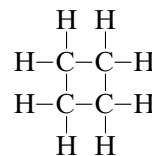


cyclopentane

The general structure for cycloalkanes,  $C_nH_{2n}$ , shows that they have  $2 \times n$  hydrogen atoms, two fewer hydrogen atoms than noncyclic alkanes,  $C_nH_{2n+2}$ , have. This is because cycloalkanes have no free ends where a carbon atom is attached to three hydrogen atoms. Another example, of a four-carbon alkane and cycloalkane, is shown below.



butane  
 $C_4H_{10}$



cyclobutane  
 $C_4H_8$

**TABLE 3** Carbon-Atom Chain Prefixes

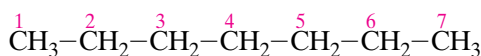
Number of carbon atoms	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-
9	non-
10	dec-

## Systematic Names of Alkanes

Historically, the names of many organic compounds were derived from the sources in which they were found. As more organic compounds were discovered, a systematic naming method became necessary. The systematic method used primarily in this book was developed by the International Union of Pure and Applied Chemistry, IUPAC.

### Unbranched-Chain Alkane Nomenclature

To name an unbranched alkane, find the prefix in **Table 3** that corresponds to the number of carbon atoms in the chain of the hydrocarbon. Then add the suffix *-ane* to the prefix. An example is shown below.

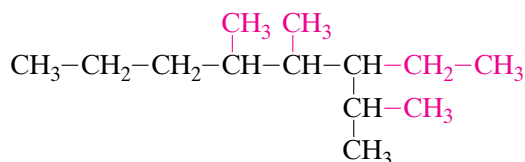


heptane

The molecule has a chain seven carbon atoms long, so the prefix *hept-* is added to the suffix *-ane* to form *heptane*.



To name the parent hydrocarbon, add the suffix *-ane* to the prefix *oct-* (for a carbon-atom chain with eight carbon atoms) to form *octane*. Now identify and name the alkyl groups.



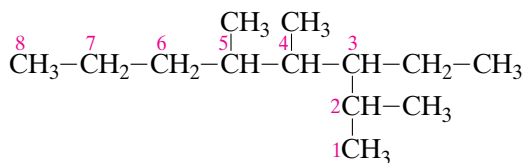
The three  $-\text{CH}_3$  groups are methyl groups. The  $-\text{CH}_2-\text{CH}_3$  group is an ethyl group. Arrange the names in alphabetical order in front of the name of the parent hydrocarbon.

### ethyl methyloctane

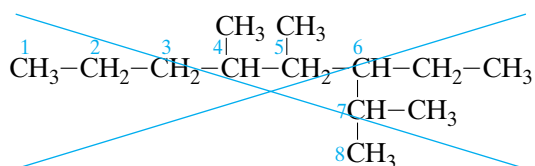
To show that there are three methyl groups present, attach the prefix *tri-* to the name *methyl* to form *trimethyl*.

### ethyl trimethyloctane

Now we need to show the locations of the alkyl groups on the parent hydrocarbon. Number the octane chain so that the alkyl groups have the lowest numbers possible.



**NOT**



Place the location numbers of *each* of the alkyl groups in front of its name. Separate the numbers from the names of the alkyl groups with hyphens. The ethyl group is on carbon 3.

### 3-ethyl trimethyloctane

Because there are three methyl groups, there will be three numbers, separated by commas, in front of *trimethyl*.

### 3-ethyl-2,4,5-trimethyloctane

The full name is 3-ethyl-2,4,5-trimethyloctane.

The procedure for naming simple branched-chain alkanes can be summarized as shown in the list on the next page.

## CAREERS in Chemistry

### Petroleum Engineer

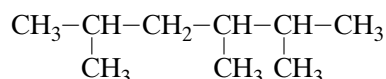
Petroleum engineers search for underground reservoirs of oil or natural gas. They then work with other scientists and engineers to develop and oversee drilling and production operations. Petroleum engineers are also responsible for developing technology and methods to increase the amount of oil or gas obtained and to lower the cost of drilling and production operations. In addition to possessing a broad range of engineering expertise (chemical, mechanical, or electrical) a petroleum engineer must also apply the basic sciences of chemistry, geology, physics, and mathematics.

## Alkane Nomenclature

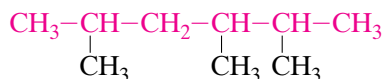
- 1. Name the parent hydrocarbon.** Find the longest continuous chain of carbon atoms that have straight-chain branches. Add the suffix *-ane* to the prefix corresponding to the number of carbon atoms in the chain.
- 2. Add the names of the alkyl groups.** Add the names of the alkyl groups in front of the name of the parent hydrocarbon in alphabetical order. When there is more than one branch of the same alkyl group present, attach the appropriate numerical prefix to the name: *di* = 2, *tri* = 3, *tetra* = 4, and so on. Do so after the names have been put in alphabetical order.
- 3. Number the carbon atoms in the parent hydrocarbon.** If one or more alkyl groups are present, number the carbon atoms in the continuous chain to give the lowest numbers possible in the name. If there are two equivalent lowest positions with two different alkyl groups, give the lowest number to the alkyl group that comes first in the name. (This will be the alkyl group that is first in alphabetical order, *before* any prefixes are attached.)
- 4. Insert position numbers.** Put the position numbers of each alkyl group in front of the name of that alkyl group.
- 5. Punctuate the name.** Use hyphens to separate the position numbers from the names. If there is more than one number in front of a name, use commas to separate the numbers.

### SAMPLE PROBLEM A

Name the following simple branched-chain alkane:

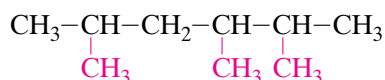


**SOLUTION** 1. Identify and name the parent hydrocarbon.



Because the longest continuous chain contains six carbon atoms, the parent hydrocarbon is *hexane*.

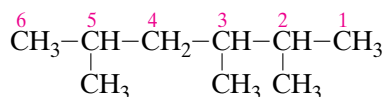
2. Identify and name the alkyl groups attached to the chain.



There is only one type of alkyl group, with one carbon atom. Alkyl groups with one carbon atom are methyl groups. Add the name *methyl* in front of the name of the continuous chain. Add the prefix *tri-* to show that there are three methyl groups present.

**trimethylhexane**

3. Number the carbon atoms in the continuous chain so that the alkyl groups have the lowest numbers possible.



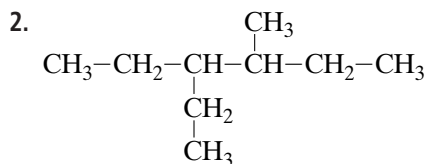
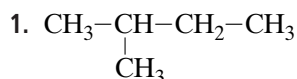
4. The methyl groups are on the carbon atoms numbered 2, 3, and 5. Put the numbers of the positions of the alkyl groups, separated by commas, in front of the name of the alkyl group. Separate the numbers from the name with a hyphen.

2,3,5-trimethylhexane

## PRACTICE

Answers in Appendix E

Name the following molecules:



### extension

Go to [go.hrw.com](http://go.hrw.com) for more practice problems that ask you to name alkanes.



Keyword: HC6ORGX

## Properties and Uses of Alkanes

Properties for some straight-chain alkanes are listed in **Table 5**. The trends in these properties can be explained by examining the structure of alkanes. The carbon-hydrogen bonds of alkanes are nonpolar. The only forces of attraction between nonpolar molecules are weak intermolecular forces, or London dispersion forces. The strength of London dispersion forces increases as the mass of a molecule increases.

**TABLE 5** Properties of Straight-Chain Alkanes

Molecular formula	IUPAC name	Boiling point (°C)	State at room temperature
CH <sub>4</sub>	methane	-164	gas
C <sub>2</sub> H <sub>6</sub>	ethane	-88.6	
C <sub>3</sub> H <sub>8</sub>	propane	-42.1	
C <sub>4</sub> H <sub>10</sub>	butane	-0.5	
C <sub>5</sub> H <sub>12</sub>	pentane	36.1	liquid
C <sub>8</sub> H <sub>18</sub>	octane	125.7	
C <sub>10</sub> H <sub>22</sub>	decane	174.1	
C <sub>17</sub> H <sub>36</sub>	heptadecane	301.8	solid
C <sub>20</sub> H <sub>42</sub>	eicosane	343	

## Physical States

The physical states at which some alkanes exist at room temperature and atmospheric pressure are found in **Table 5**. Alkanes that have the lowest molecular mass—those with one to four carbon atoms—are gases. **Natural gas** is a fossil fuel composed primarily of alkanes containing one to four carbon atoms. These alkanes are gases because they are very small molecules. Therefore, they have weak London dispersion forces between them and are not held together tightly. Larger alkanes are liquids. Gasoline and kerosene consist mostly of liquid alkanes. Stronger London dispersion forces hold these molecules close enough together to form liquids. Alkanes that have a very high molecular mass are solids, corresponding to a greater increase in London dispersion forces. Paraffin wax contains solid alkanes. It can be used in candles, as shown in **Figure 6**.

## Boiling Points

The boiling points of alkanes, also shown in **Table 5**, increase with molecular mass. As London dispersion forces increase, more energy is required to pull the molecules apart. This property is used in the separation of petroleum, a major source of alkanes. **Petroleum** is a complex mixture of different hydrocarbons that varies greatly in composition. The hydrocarbon molecules in petroleum contain from 1 to more than 50 carbon atoms. This range allows the separation of petroleum into different portions that have different boiling point ranges, as shown in **Table 6**. In *fractional distillation*, components of a mixture are separated on the basis of boiling point, by condensation of vapor in a fractionating column. **Figure 7** shows an example of refinery towers in which the process takes place. During its fractional distillation, petroleum is heated to about 370°C. Nearly all components of the petroleum are vaporized at this temperature. As the vapors rise in the fractionating column, or tower, they are gradually cooled. Alkanes that have higher boiling points have higher condensation temperatures and condense for collection lower in the tower. For example, lubricating oils, which have higher condensation temperatures than gasoline has, are collected lower in the fractionating tower.

**TABLE 6** Petroleum Fractions

Fraction	Size range of molecules	Boiling-point range (°C)
Gasoline	C <sub>4</sub> –C <sub>12</sub>	up to 200
Kerosene	C <sub>10</sub> –C <sub>14</sub>	180–290
Middle distillate, such as heating oil, gas-turbine fuel, diesel	C <sub>12</sub> –C <sub>20</sub>	185–345
Wide-cut gas oil, such as lubricating oil, waxes	C <sub>20</sub> –C <sub>36</sub>	345–540
Asphalt	above C <sub>36</sub>	residues



**FIGURE 6** Paraffin wax, used in candles, contains solid alkanes. Molecules of paraffin wax contain 26 to 30 carbon atoms.

**FIGURE 7** Fractional distillation takes place in petroleum refinery towers.



## Unsaturated Hydrocarbons

Hydrocarbons that do not contain the maximum amount of hydrogen are referred to as *unsaturated*. **Unsaturated hydrocarbons** are hydrocarbons in which not all carbon atoms have four single covalent bonds. An unsaturated hydrocarbon has one or more double bonds or triple bonds. Carbon atoms can easily form double and triple bonds to other carbon atoms, so multiple bonds between carbon atoms are common in organic compounds.

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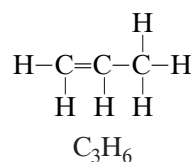
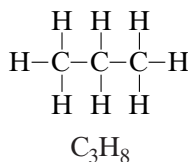
Topic: Alkenes

Code: HC60046

### Alkenes

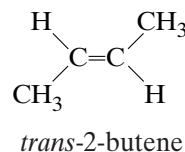
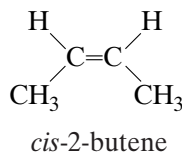
**Alkenes** are hydrocarbons that contain double covalent bonds. Some examples of alkenes are given in **Table 7**. Notice that because alkenes have a double bond, the simplest alkene, ethene, has two carbon atoms.

Carbon atoms linked by double bonds cannot bind as many atoms as those that are linked by only single bonds. An alkene with one double bond has two fewer hydrogen atoms than the corresponding alkane.

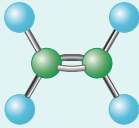
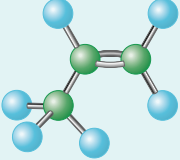
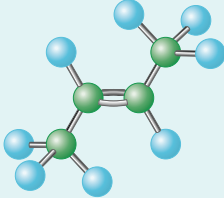
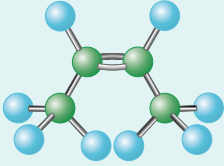


Thus, the general formula for noncyclic alkenes with one double bond is  $\text{C}_n\text{H}_{2n}$ .

Because alkenes have a double bond, they can have geometric isomers, as shown in the examples below.



**TABLE 7** Structures of Alkenes

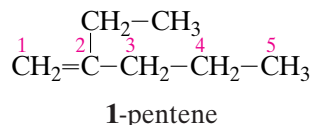
	ethene	propene	<i>trans</i> -2-butene	<i>cis</i> -2-butene
Structural formula	$\begin{array}{c} \text{H} \quad \quad \quad \text{H} \\ \quad \quad \quad \diagdown \quad \diagup \\ \quad \quad \quad \text{C}=\text{C} \\ \quad \quad \quad \diagup \quad \diagdown \\ \text{H} \quad \quad \quad \text{H} \end{array}$	$\begin{array}{c} \text{H} \quad \quad \quad \text{H} \\ \quad \quad \quad \diagdown \quad \diagup \\ \quad \quad \quad \text{C}=\text{C} \\ \quad \quad \quad \diagup \quad \diagdown \\ \text{CH}_3 \quad \quad \quad \text{H} \end{array}$	$\begin{array}{c} \text{H} \quad \quad \quad \text{CH}_3 \\ \quad \quad \quad \diagdown \quad \diagup \\ \quad \quad \quad \text{C}=\text{C} \\ \quad \quad \quad \diagup \quad \diagdown \\ \text{CH}_3 \quad \quad \quad \text{H} \end{array}$	$\begin{array}{c} \text{H} \quad \quad \quad \text{H} \\ \quad \quad \quad \diagdown \quad \diagup \\ \quad \quad \quad \text{C}=\text{C} \\ \quad \quad \quad \diagup \quad \diagdown \\ \text{CH}_3 \quad \quad \quad \text{CH}_3 \end{array}$
Ball-and-stick model				

## Systematic Names of Alkenes

The rules for naming a simple alkene are similar to those for naming an alkane. The parent hydrocarbon is the longest continuous chain of carbon atoms *that contains the double bond*.



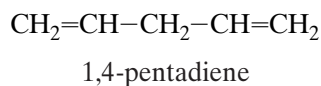
The carbon atoms in the chain are numbered so that the first carbon atom in the double bond has the lowest number.



The position number and name of the alkyl group are placed in front of the double-bond position number. This alkyl group has two carbon atoms, an ethyl group. It is on the second carbon atom of the parent hydrocarbon.

### 2-ethyl-1-pentene

If there is more than one double bond, the suffix is modified to indicate the number of double bonds: 2 = *-adiene*, 3 = *-atriene*, and so on.



The procedure for naming alkenes can be summarized as follows.

## Alkene Nomenclature

Use the rules for alkane nomenclature on page 721, with the following exceptions.

- 1. Name the parent hydrocarbon.** Locate the longest continuous chain that *contains the double bond(s)*. If there is only one double bond, add the suffix *-ene* to the prefix corresponding to the number of carbon atoms in this chain. If there is more than one double bond, modify the suffix to indicate the number of double bonds. For example, 2 = *-adiene*, 3 = *-atriene*, and so on.
- 2. Add the names of the alkyl groups.**
- 3. Number the carbon atoms in the parent hydrocarbon.** Number the carbon atoms in the chain so that the first carbon atom in the double bond nearest the end of the chain has the lowest number. If numbering from both ends gives equivalent positions for two double bonds, then number from the end nearest the first alkyl group.
- 4. Insert position numbers.** Place double-bond position numbers immediately before the name of the parent hydrocarbon alkene. Place alkyl group position numbers immediately before the name of the corresponding alkyl group.
- 5. Punctuate the name.**

## Chemistry in Action

### Carbon Allotropes

Carbon occurs in several different bonding patterns, called *allotropes*, which have very different properties. Diamond is a colorless, crystalline solid form of carbon, in which each atom is tetrahedrally bonded to four others in a network fashion. This three-dimensional bonding makes diamond the hardest material known.

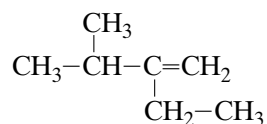
Graphite is a soft, black, crystalline form of carbon that is a fair conductor of electricity. The carbon atoms in graphite are bonded together in layers. Within each layer, each atom is bonded to three other carbon atoms. But because adjacent layers are held together only by very weak London dispersion forces, graphite is very soft.

Diamond and graphite have been known for thousands of years. In the mid-1980s, a new solid allotropic form of carbon known as a *fullerene*, or, informally, "Buckyball," was discovered in the soot that forms when carbon-containing materials are burned with limited oxygen. The structure of a fullerene consists of a near-spherical cage of carbon atoms. The most stable fullerene is  $\text{C}_{60}$ , the structure of which resembles the design of a soccer ball.

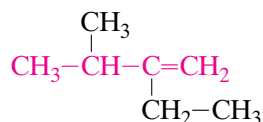


## SAMPLE PROBLEM B

Name the following alkene.

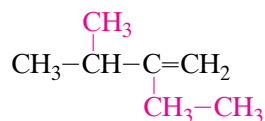


**SOLUTION** 1. Identify and name the parent hydrocarbon.



The parent hydrocarbon has four carbon atoms and one double bond, so it is named *butene*.

2. Identify and name the alkyl groups.

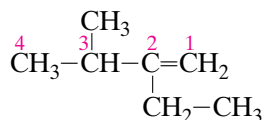


The alkyl groups are *ethyl* and *methyl*.

Place their names in front of the name of the parent hydrocarbon in alphabetical order.

**ethyl methyl butene**

3. Number the carbon chain to give the double bond the lowest position.

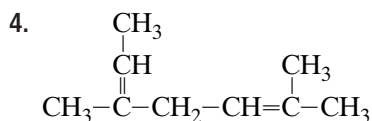
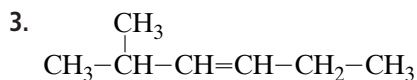
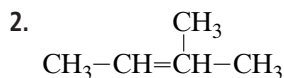


Place the position number of the double bond in front of butene. Place the position numbers of the alkyl groups in front of each alkyl group. Separate the numbers from the name with hyphens. The full name is 2-ethyl-3-methyl-1-butene.

## PRACTICE

*Answers in Appendix E*

Name the following alkenes:

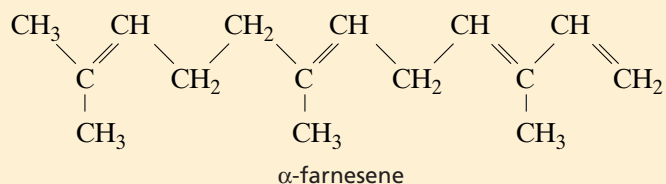


### extension

Go to [go.hrw.com](http://go.hrw.com) for more practice problems that ask you to name alkenes.

 **Keyword:** HC6ORGX

**FIGURE 8**  $\alpha$ -farnesene is a solid alkene found in the natural wax covering of apples. Can you determine the IUPAC name for this large alkene?



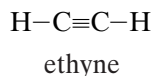
## Properties and Uses of Alkenes

Alkenes are nonpolar and show trends in properties similar to those of alkanes in boiling points and physical states. For example,  $\alpha$ -farnesene has 15 carbon atoms and 4 double bonds, as shown in **Figure 8**. This large alkene is a solid at room temperature and atmospheric pressure. It is found in the natural wax covering of apples. Ethene, the smallest alkene, is a gas. Ethene is commonly called *ethylene*.

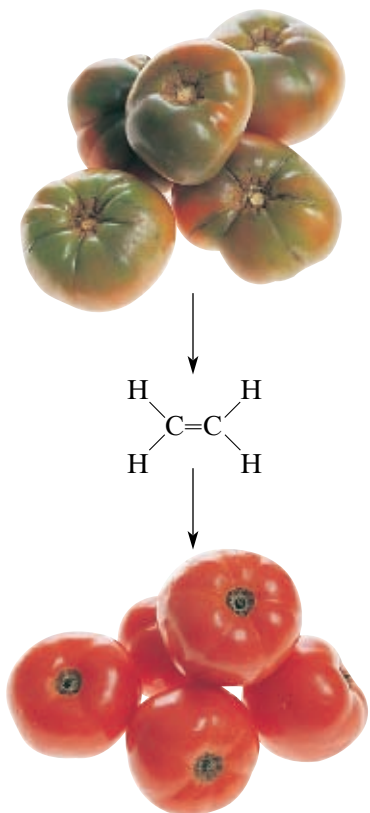
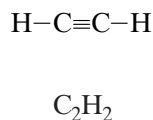
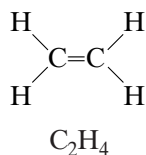
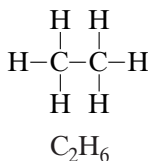
Ethene is the hydrocarbon commercially produced in the greatest quantity in the United States. It is used in the synthesis of many plastics and commercially important alcohols. Ethene is also an important plant hormone. Induction of flowering and fruit ripening, as shown in **Figure 9**, are effects of ethene hormone action that can be manipulated by commercial growers.

## Alkynes

Hydrocarbons with triple covalent bonds are **alkynes**. Like the double bond of alkenes, the triple bond of alkynes requires that the simplest alkyne has two carbon atoms.



The general formula for the alkynes is  $\text{C}_n\text{H}_{2n-2}$ . Alkynes have four fewer hydrogen atoms than the corresponding alkanes and two fewer hydrogen atoms than the corresponding alkenes. The simplest alkyne is ethyne, more commonly known as *acetylene*.



**FIGURE 9** Ethene is a plant hormone that triggers fruit ripening. At ordinary conditions, ethene is a gas.

## Systematic Naming of Alkynes

Alkyne nomenclature is almost the same as alkene nomenclature. The only difference is that the *-ene* suffix of the corresponding alkene is replaced with *-yne*. A complete list of rules follows.

### Alkyne Nomenclature

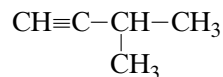
Use the rules for alkane nomenclature on page 721, with the following exceptions.

- 1. Name the parent hydrocarbon.** Locate the longest continuous chain that *contains the triple bond(s)*. If there is only one triple bond, add the suffix *-yne*. If there is more than one triple bond, modify the suffix. For example, 2 = *-adiyne*, 3 = *-atriyne*, and so on.
- 2. Add the names of the alkyl groups.**
- 3. Number the carbon atoms in the parent hydrocarbon.** Number the carbon atoms in the chain so that the first carbon atom in the triple bond nearest the end of the chain has the lowest number.
- 4. Insert position numbers.** Place the position numbers of the triple bond immediately before the name of the parent hydrocarbon alkyne. Place alkyl group position numbers immediately before the name of the corresponding alkyl group.
- 5. Punctuate the name.**

Two examples of correctly named alkynes are given below.



1-pentyne



3-methyl-1-butyne

## Properties and Uses of Alkynes

Alkynes are nonpolar and exhibit the same trends in boiling points and physical state as other hydrocarbons. The smallest alkyne, ethyne, is a gas. The combustion of ethyne when it is mixed with pure oxygen produces the intense heat of welding torches, as shown in **Figure 10**. The common name of ethyne is *acetylene*, and these welding torches are commonly called oxyacetylene torches.

**FIGURE 10** Ethyne is the fuel used in oxyacetylene torches. Oxyacetylene torches can reach temperatures of over 3000°C.



### extension

#### Chemistry in Action

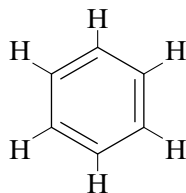
Go to [go.hrw.com](http://go.hrw.com) for a full-length article on synthetic diamonds.



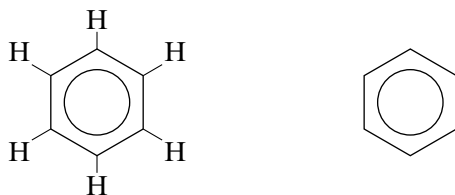
Keyword: HC6ORGX

## Aromatic Hydrocarbons

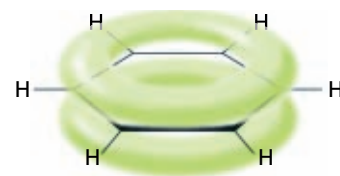
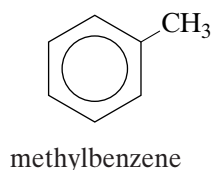
**Aromatic hydrocarbons** are hydrocarbons that have six-membered carbon rings and delocalized electrons. **Benzene** is the primary aromatic hydrocarbon. The molecular formula of benzene is  $C_6H_6$ . One possible structural formula is a six-carbon atom ring with three double bonds.



However, benzene does not behave chemically like an alkene. The entire molecule lies in the same plane, as shown in **Figure 11**. Benzene contains resonance hybrid bonds, and the structure of the benzene ring allows electrons to be spread through delocalized  $p$ -orbitals over the whole ring. The structural and skeletal formulas below show benzene as a resonance hybrid, representing the delocalization of electrons.



Aromatic hydrocarbons can be thought of as derivatives of benzene. The simplest have one benzene ring, as shown in the following example.



**FIGURE 11** Electron orbitals in benzene overlap to form continuous  $p$ -orbitals that allow the delocalized electrons to spread uniformly over the entire ring.

## SECTION REVIEW

- List the basic structural features that characterize each of the following hydrocarbons:
  - alkanes
  - alkenes
  - alkynes
  - aromatic hydrocarbons
- Draw all of the condensed structural formulas that can represent  $C_5H_{12}$ .
- Give the systematic name for each compound in your answers to item 2.
- Give examples of a property or use of three hydrocarbons.
- Name the following compounds:
  - $CH_3-CH_2-\overset{\overset{CH_3}{|}}{CH}-CH_2-CH_3$
  - $CH_2=CH-CH=CH_2$
  - $CH_3-C\equiv C-CH_2-CH_3$

### Critical Thinking

- ANALYZING INFORMATION** Write the structural formulas for an alkane, an alkene, and an alkyne that have five carbon atoms each. Why are these three hydrocarbons not considered isomers?

## SECTION 3

## Functional Groups

## OBJECTIVES

- Define “functional group” and explain why functional groups are important.
- Identify alcohols, alkyl halides, ethers, aldehydes, ketones, carboxylic acids, esters, and amines based on the functional group present in each.
- Explain the relationships between the properties and structures of compounds with various functional groups.

**A functional group** is an atom or group of atoms that is responsible for the specific properties of an organic compound. A given functional group undergoes the same types of chemical reactions in every molecule in which it is found. Therefore, compounds that contain the same functional group can be classified together.

## Classes of Organic Compounds

A functional group gives an organic compound properties that are very different from those of the corresponding hydrocarbon. The compounds in **Table 8** have four carbon atoms, but they have very different physical properties due to their different functional groups. Some functional groups and their characteristic general formulas are shown in **Table 9**.

TABLE 8 Comparing Classes of Organic Compounds

Name	Structural Formula	Melting Point (°C)	Boiling Point (°C)	Density (g/mol)
Butane	$  \begin{array}{cccc}  \text{H} & \text{H} & \text{H} & \text{H} \\    &   &   &   \\  \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\    &   &   &   \\  \text{H} & \text{H} & \text{H} & \text{H}  \end{array}  $	-138.4	-0.5	0.5788
1-Butanol	$  \begin{array}{cccc}  \text{H} & \text{H} & \text{H} & \text{H} \\    &   &   &   \\  \text{HO}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\    &   &   &   \\  \text{H} & \text{H} & \text{H} & \text{H}  \end{array}  $	-89.5	117.2	0.8098
Butanoic acid	$  \begin{array}{cccc}  & \text{O} & \text{H} & \text{H} & \text{H} \\  &    &   &   &   \\  \text{HO}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\  & &   &   &   \\  & & \text{H} & \text{H} & \text{H}  \end{array}  $	-4.5	163.5	0.9577
2-Butanone	$  \begin{array}{cccc}  \text{H} & \text{O} & \text{H} & \text{H} \\    &    &   &   \\  \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\    & &   &   \\  \text{H} & & \text{H} & \text{H}  \end{array}  $	-86.3	79.6	0.8054
Diethyl ether	$  \begin{array}{cccc}  \text{H} & \text{H} & & \text{H} & \text{H} \\    &   & &   &   \\  \text{H}-\text{C}-\text{C}-\text{O}-\text{C}-\text{C}-\text{H} \\    &   & &   &   \\  \text{H} & \text{H} & & \text{H} & \text{H}  \end{array}  $	-116.2	34.5	0.7138

**TABLE 9** Classes of Organic Compounds

Class	Functional group	General formula
alcohol	-OH	$R-OH$
alkyl halide	-X (X = F, Cl, Br, I)	$R-X$
ether	-O-	$R-O-R'$
aldehyde	$\begin{array}{c} O \\    \\ -C-H \end{array}$	$\begin{array}{c} O \\    \\ R-C-H \end{array}$
ketone	$\begin{array}{c} O \\    \\ -C- \end{array}$	$\begin{array}{c} O \\    \\ R-C-R' \end{array}$
amine	$\begin{array}{c} -N- \\   \end{array}$	$\begin{array}{c} R-N-R'' \\   \\ R' \end{array}$
carboxylic acid	$\begin{array}{c} O \\    \\ -C-OH \end{array}$	$\begin{array}{c} O \\    \\ R-C-OH \end{array}$
ester	$\begin{array}{c} O \\    \\ -C-O- \end{array}$	$\begin{array}{c} O \\    \\ R-C-O-R' \end{array}$

## Alcohols

**Alcohols** are organic compounds that contain one or more hydroxyl groups. The general formula for a class of organic compounds consists of the functional group and the letter *R*, which stands for the rest of the molecule. The general formula for alcohols is  $R-OH$ .

Hydrogen bonding in alcohols can also explain other properties and uses of alcohols. Cold creams, lipsticks, body lotions, and similar products generally include 1,2,3-propanetriol, commonly called glycerol, to keep them moist. A model for glycerol is shown in **Figure 12**. Multiple hydroxyl groups allow glycerol to form many hydrogen bonds with water molecules.

Alcohols are sometimes used today as alternative fuels and as octane enhancers in fuel for automobiles. Ethanol is combined with gasoline, for example, in a one-to-nine ratio to produce gasohol. Some experts have promoted the use of gasohol as a fuel for automobiles because it burns more cleanly and efficiently. However, there are also disadvantages. The combustion of ethanol produces only about 60% as much energy per gram as the combustion of gasoline does. The presence of ethanol also causes increased water absorption in the fuel.



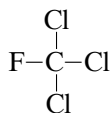
**FIGURE 12** Glycerol contains three hydroxyl groups. This structure allows it to form multiple hydrogen bonds with water. Glycerol is added as a moisturizer to skin products.



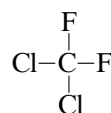
## Alkyl Halides

**Alkyl halides** are organic compounds in which one or more halogen atoms—fluorine, chlorine, bromine, or iodine—are substituted for one or more hydrogen atoms in a hydrocarbon. Because  $-X$  is often used to represent any halogen, an alkyl halide may be represented by the general formula  $R-X$ .

Alkyl halides are some of the most widely used organic chemicals. A family of alkyl halides that has received widespread attention in recent years is the chlorofluorocarbons, or CFCs. CFCs are alkyl halides that contain both chlorine and fluorine. The formulas for two widely used CFCs, Freon-11 and Freon-12, are shown below.

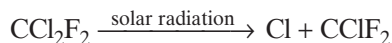


trichlorofluoromethane  
(Freon-11)

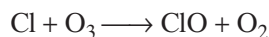


dichlorodifluoromethane  
(Freon-12)

CFCs, which have been used as liquid refrigerants, contribute to the destruction of ozone in the upper atmosphere. When released into the atmosphere, CFCs can break down and release free chlorine atoms.



The released chlorine atoms attack molecules of ozone,  $\text{O}_3$ , found in the upper atmosphere. The ozone is converted to diatomic oxygen.



Chlorine atoms are eventually regenerated in various ways, including the reaction of  $\text{ClO}$  with  $\text{O}$ .

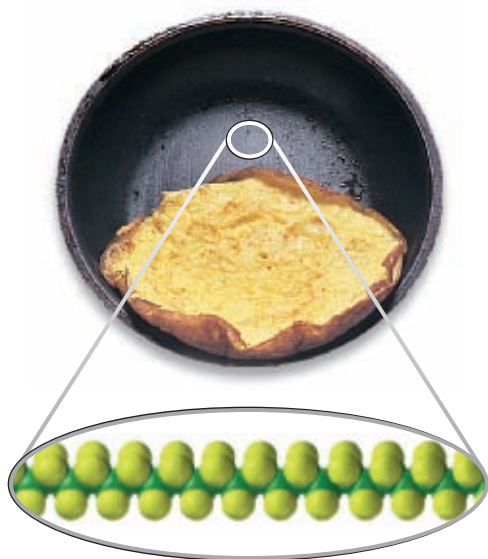


This makes it possible for a single chlorine atom to destroy thousands of ozone molecules. Because CFCs are a major cause of ozone depletion, more than 100 nations signed an agreement in 1987 to reduce the amount of CFCs produced.

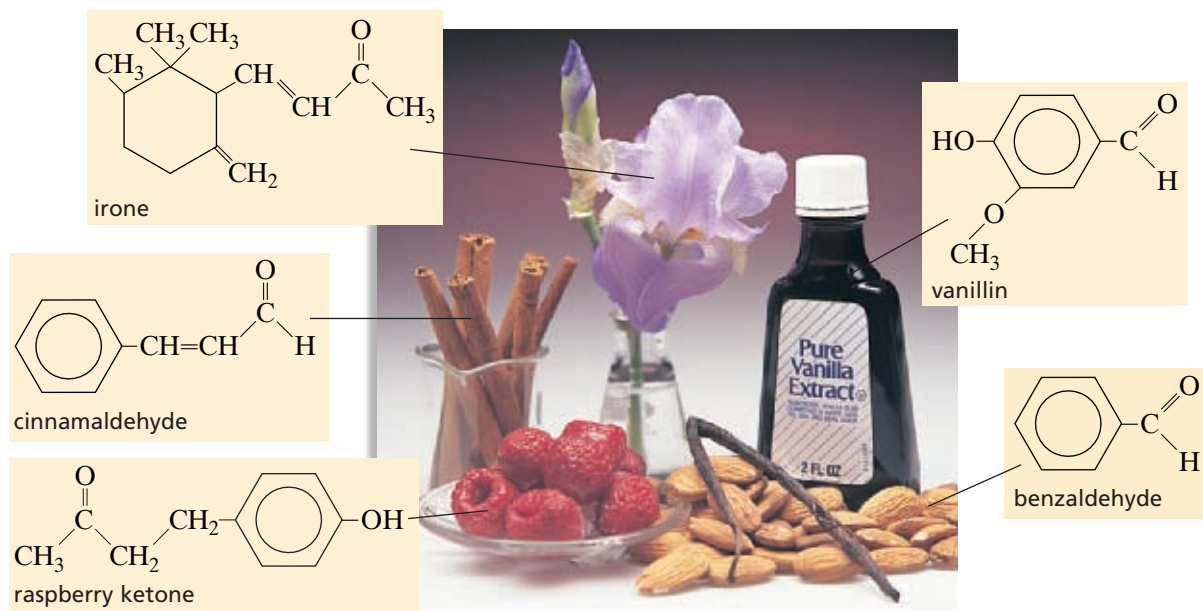
Another alkyl halide is tetrafluoroethene,  $\text{C}_2\text{F}_4$ . It is joined in long chains to make a material with the trade name Teflon<sup>®</sup>. Because of the unreactive carbon-fluorine bond, Teflon is inactive and stable to about  $325^\circ\text{C}$ . It also has a low coefficient of friction, which means that other objects slide smoothly over its surface. These properties enable Teflon to be used in heat-resistant machine parts that cannot be lubricated. It is also used in making utensils with “nonstick” surfaces, such as the frying pan in **Figure 13**.

## Ethers

**Ethers** are organic compounds in which two hydrocarbon groups are bonded to the same atom of oxygen. They can be represented by the general formula  $R-O-R'$ . In this formula,  $R'$  may be the same hydrocarbon group as  $R$  or a different one. Like alkanes, ethers are not very reactive compounds, so they are commonly used as solvents.



**FIGURE 13** The nonstick coating on this pan is made of Teflon, an alkyl halide.



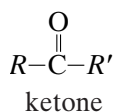
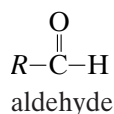
**FIGURE 14** Many common odors and flavors come from aldehydes and ketones.

## Aldehydes and Ketones

Aldehydes and ketones contain the *carbonyl group*, shown below.



The difference between aldehydes and ketones is the location of the carbonyl group. **Aldehydes** are organic compounds in which the carbonyl group is attached to a carbon atom at the end of a carbon-atom chain. **Ketones** are organic compounds in which the carbonyl group is attached to carbon atoms within the chain. These differences can be seen in their general formulas, shown below.

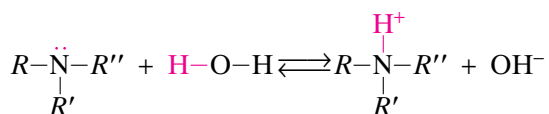


Aldehydes and ketones are often responsible for odors and flavors. **Figure 14** gives some examples.

## Amines

**Amines** are organic compounds that can be considered to be derivatives of ammonia,  $\text{NH}_3$ . Amines are common in nature. They often form during the breakdown of proteins in animal cells.

The unshared electron pair on the nitrogen atom makes amines act as weak bases in aqueous solutions, as shown below.



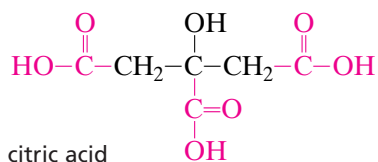
### extension

#### Chemical Content

Go to [go.hrw.com](http://go.hrw.com) for more information on functional groups.

Keyword: HC60RGX

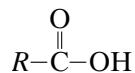




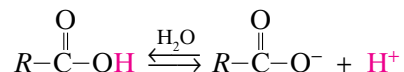
**FIGURE 15** Citric acid, found in citrus fruits, contains three carboxylic acid groups shown in red on the structural formula.

## Carboxylic Acids

**Carboxylic acids** are organic compounds that contain the carboxyl functional group. A member of this class of organic compounds can be represented by the general formula shown below.



Carboxylic acids, like inorganic acids, react to lose a hydrogen ion and become a negatively charged ion in water.

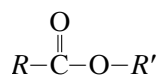


Carboxylic acids are much weaker than many inorganic acids, such as hydrochloric, sulfuric, and nitric acids. Acetic acid, the weak acid in vinegar, is a carboxylic acid.

A number of carboxylic acids occur naturally in plants and animals. For example, citrus fruits, shown in **Figure 15**, contain citric acid. Benzoic, propanoic, and sorbic acids are used as preservatives. All three acids kill microorganisms that cause foods to spoil.

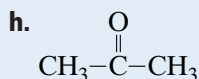
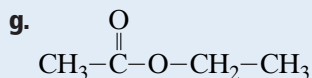
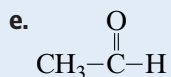
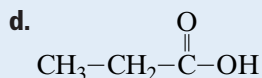
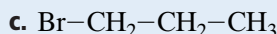
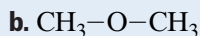
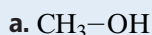
## Esters

**Esters** are organic compounds that have carboxylic acid groups in which the hydrogen of the hydroxyl group has been replaced by an alkyl group. Esters are considered *derivatives* of carboxylic acids because of their structural similarity to carboxylic acids. The general formula for an ester is given below.



## SECTION REVIEW

1. Give the general formula and class of organic compounds for each of the following:



- Compare the boiling points of alcohols, ethers, and alkanes, and explain one reason for the differences.
- How are aldehydes and ketones alike? How do they differ?
- How do the strengths of organic acids compare with the strengths of most inorganic acids?

### Critical Thinking

- APPLYING MODELS** Identify the functional groups in vanillin, shown in Figure 14.

# Organic Reactions

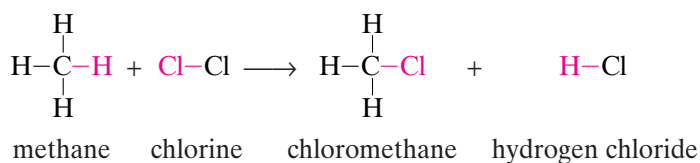
## SECTION 4

### OBJECTIVES

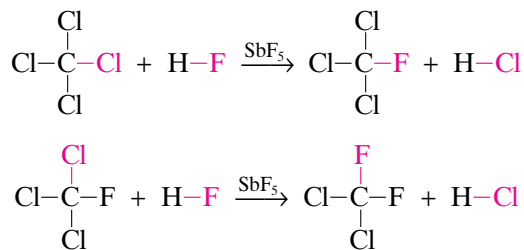
- Describe and distinguish between the organic reactions: substitution, addition, condensation, and elimination.
- Relate some functional groups to some characteristic reactions.
- Identify the two main types of polymers and the basic reaction mechanisms by which they are made.

## Substitution Reactions

A **substitution reaction** is one in which one or more atoms replace another atom or group of atoms in a molecule. The reaction between an alkane, such as methane, and a halogen, such as chlorine, to form an alkyl halide is an example of a substitution reaction. Notice that in this reaction, a chlorine atom replaces a hydrogen atom on the methane molecule.



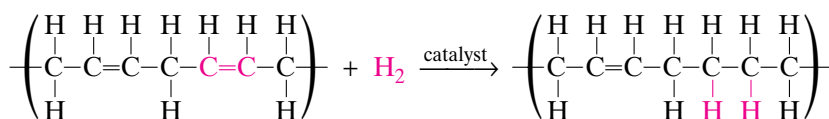
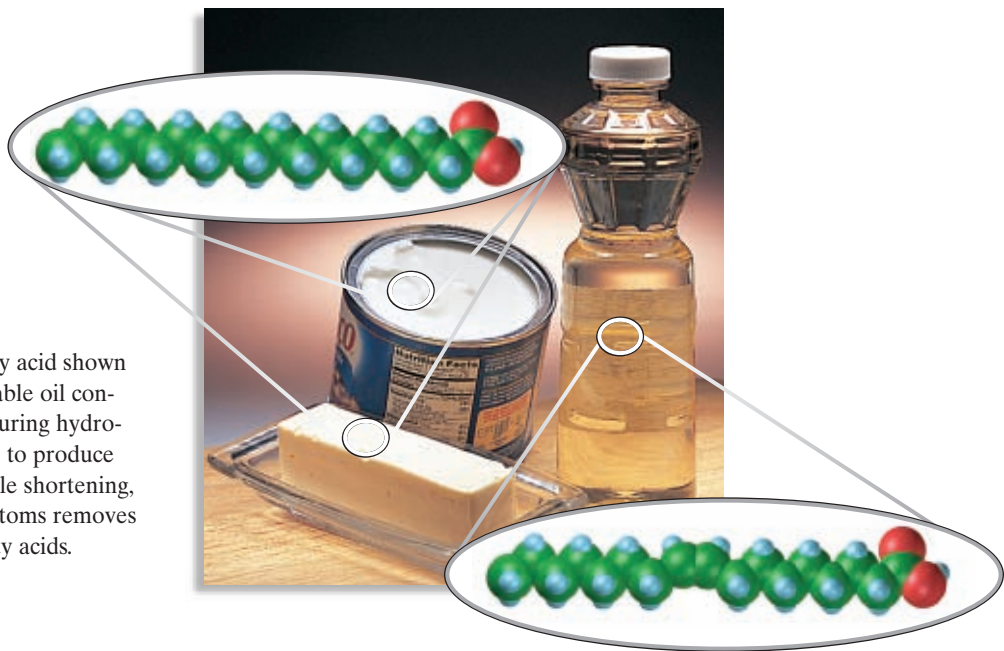
Additional compounds can be formed by replacing the other hydrogen atoms remaining in the methane molecule. The products are dichloromethane, trichloromethane, and tetrachloromethane. Trichloromethane is also known as chloroform, and tetrachloromethane is also known as carbon tetrachloride. CFCs are formed by further substitution reactions between chloroalkanes and HF.



## Addition Reactions

An **addition reaction** is one in which two parts of a molecule are added to an unsaturated molecule, increasing the saturation of the molecule. A common type of addition reaction is hydrogenation. In *hydrogenation*, hydrogen atoms are added to an unsaturated molecule. Vegetable oils are triesters of unsaturated fatty acids, long chains of carbon atoms that have many double bonds. The following equation shows just one portion of an oil molecule. When hydrogen gas is blown through an oil, hydrogen atoms may add to the double bonds in the oil molecule.

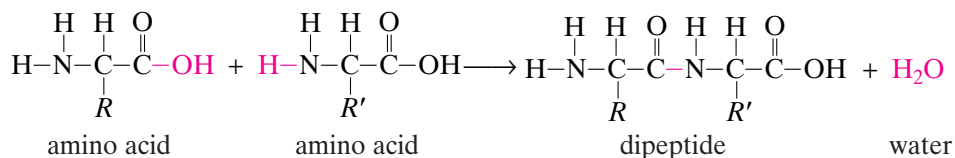
**FIGURE 16** The fatty acid shown in the model for vegetable oil contains a double bond. During hydrogenation, which is used to produce margarine and vegetable shortening, addition of hydrogen atoms removes double bonds from fatty acids.



The molecule still consists of long chains of carbon atoms, but it contains far fewer double bonds. The conversion of these double bonds to single bonds changes the material from an oil, which is a liquid, into a fat, which is a solid. When you see the word *hydrogenated* on a food product, you know that an oil has been converted to a fat by this process. Examples of an oil and hydrogenated fats are shown in **Figure 16**.

## Condensation Reactions

A **condensation reaction** is one in which two molecules or parts of the same molecule combine. A small molecule, such as water, is usually removed during the reaction. An example is the reaction between two amino acids, which contain both amine and carboxyl groups. One hydrogen from the amine group of one amino acid combines with the hydroxyl from the carboxyl group of the other amino acid to form a molecule of water. When repeated many times, this reaction forms a protein molecule.

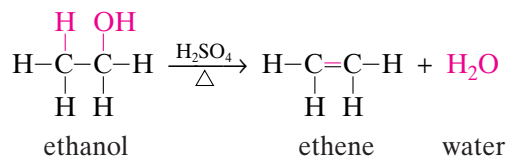




**FIGURE 17** Sucrose is dehydrated when it reacts with concentrated sulfuric acid. Elimination of water produces a compound that is mostly carbon.

## Elimination Reactions

An **elimination reaction** is one in which a simple molecule, such as water or ammonia, is formed from adjacent carbon atoms of a larger molecule. A simple example of an elimination reaction is the heating of ethanol in the presence of concentrated sulfuric acid. Under these conditions, a hydrogen atom bonded to one carbon atom and a hydroxyl group bonded to the second carbon atom are removed from the ethanol molecule. A molecule of water is formed as a result.



Another example of an elimination reaction is the dehydration of sucrose with concentrated sulfuric acid, shown in **Figure 17**.

## Polymers

**Polymers** are large molecules made of many small units joined to each other through organic reactions. The small units are **monomers**. A polymer can be made from identical or different monomers. A polymer made from two or more different monomers is a **copolymer**.

Polymers are all around us. The foods we eat and clothes we wear are made of polymers. Some of the most common natural polymers include starch, cellulose, and proteins. Some synthetic polymers may be familiar to you as plastics and synthetic fibers.



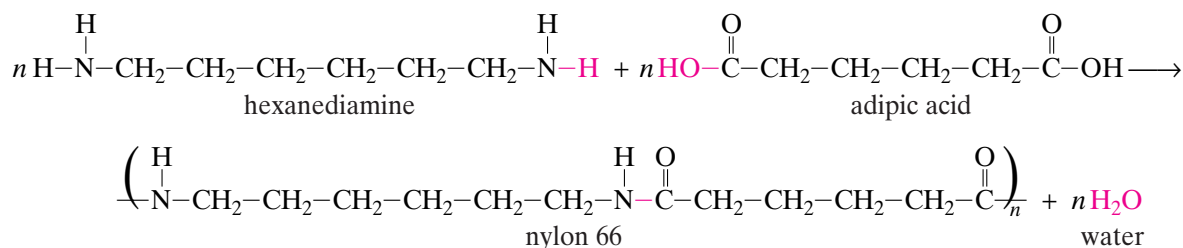


## Condensation Polymers

A *condensation polymer* is a polymer formed by condensation reactions. Monomers of condensation polymers must contain two functional groups. This allows each monomer to link with two other monomers by condensation reactions. Condensation polymers are usually copolymers with two monomers in an alternating order.

### Polyamides and Polyesters

One example of a condensation polymer is shown below. A carboxylic acid with two carboxyl groups, adipic acid, and an amine with two amine groups, hexanediamine, react with each other to form water.



The product is a polyamide and contains the adipic acid monomer and the hexanediamine monomer. This copolymer is known as nylon 66 because each of the monomers contains six carbon atoms. Nylon 66 is one of the most widely used of all synthetic polymers.

Polyesters are another common type of condensation polymer. They are formed from dialcohols and dicarboxylic acids, which undergo a condensation reaction to form an ester group, linking the alcohol end of a monomer to the acid end of a monomer. Polyesters have many uses, such as in tires, in food packaging, and as fibers in permanent press fabrics.

## SECTION REVIEW

1. Can an addition reaction occur between chlorine and ethane? Why or why not?
2. Does an addition reaction increase or decrease the saturation of a molecule?
3. The condensation reaction between two amino acids has a water molecule as one product. Which functional groups are the sources of the atoms in the water molecule?
4. Explain how elimination reactions could be considered the opposite of addition reactions.
5. Why can a molecule that has only one functional group *not* undergo a condensation reaction to form a polymer?

6. Would it be possible to have an addition polymer synthesized from a monomer that has only single bonds? Why or why not?

### Critical Thinking

7. **APPLYING MODELS** Polyvinyl chloride (PVC) is a polymer that is widely used in pipes and flooring. It is an addition polymer made from chloroethene, commonly known as *vinyl chloride*.
  - a. Draw the structure of vinyl chloride. Then, look up the structure or check it with your teacher.
  - b. Write the reaction for the polymerization of vinyl chloride to form polyvinyl chloride (PVC).

# CHAPTER HIGHLIGHTS

## Organic Compounds

### Vocabulary

organic compound  
catenation  
hydrocarbon  
isomer  
structural formula  
structural isomer  
geometric isomer

- The ability of carbon to bond to other elements and to allow different arrangements of atoms contributes to the diversity of carbon compounds.
- Isomers are compounds that have the same molecular formula but different structures. In structural isomers, the atoms are bonded together in different orders. In geometric isomers, the order of atom bonding is the same, but the atoms are oriented differently in space.

## Hydrocarbons

### Vocabulary

saturated hydrocarbon	unsaturated hydrocarbon
alkane	alkene
cycloalkane	alkyne
alkyl group	aromatic hydrocarbon
natural gas	benzene
petroleum	

- Alkanes are saturated hydrocarbons; they contain only single bonds. Physical trends in alkanes correspond to trends in alkane size and amount of branching.
- Organic compounds are named according to a systematic method.
- Unsaturated hydrocarbons have one or more multiple carbon-carbon bonds: these include alkenes, alkynes, and aromatic hydrocarbons.

## Functional Groups

### Vocabulary

functional group  
alcohol  
alkyl halide  
ether  
aldehyde  
ketone  
amine  
carboxylic acid  
ester

- Functional groups are responsible for the properties of the organic compound that contains the functional group.
- Alcohols contain the hydroxyl functional group.
- Alkyl halides contain one or more halogen atoms.
- Two alkyl groups are joined to an oxygen atom in ethers.
- Both aldehydes and ketones contain the carbonyl group.
- Amines are derivatives of ammonia.
- Carboxylic acids contain carboxyl groups.
- In esters, the hydrogen atom of a carboxylic acid group has been replaced with an alkyl group.

## Organic Reactions

### Vocabulary

substitution reaction  
addition reaction  
condensation reaction  
elimination reaction  
polymer  
monomer  
copolymer

- In substitution reactions, an atom or group of atoms is replaced. In addition reactions, an atom or group of atoms is added to a double or triple bond.
- In a condensation reaction, two molecules combine. In an elimination reaction, a small molecule forms from a large molecule.
- Polymers are large molecules made of many repeating units called *monomers*. A copolymer consists of two or more different monomers.

# CHAPTER REVIEW

## Organic Compounds

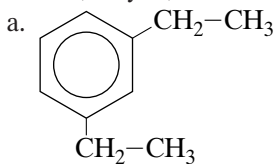
### SECTION 1 REVIEW

- What is catenation?
  - How does catenation contribute to the diversity of organic compounds?
- What information about a compound is provided by a structural formula?
  - How are structural formulas used in organic chemistry?
- Can molecules with the molecular formulas  $C_4H_{10}$  and  $C_4H_{10}O$  be structural isomers of one another? Why or why not?
- Can molecules with only single bonds (and no rings) have geometric isomers? Why or why not?

## Hydrocarbons

### SECTION 2 REVIEW

- What are hydrocarbons, and what is their importance?
- What do the terms *saturated* and *unsaturated* mean when applied to hydrocarbons?
  - What other meanings do these terms have in chemistry?
  - Classify alkenes, alkanes, alkynes, and aromatic hydrocarbons as either saturated or unsaturated.
- Classify each of the following as an alkane, alkene, alkyne, or aromatic hydrocarbon.



- $CH_3-CH=CH_2$
  - $$\begin{array}{c} CH_3 \\ | \\ CH \equiv C - CH - CH_2 - CH_3 \end{array}$$
  - $$\begin{array}{c} CH_3 - CH - CH_2 - CH_2 - CH_2 - CH_2 - CH_3 \\ | \\ CH_3 \end{array}$$
- Give the general formula for the members of the following:
    - alkane series

b. alkene series

c. alkyne series

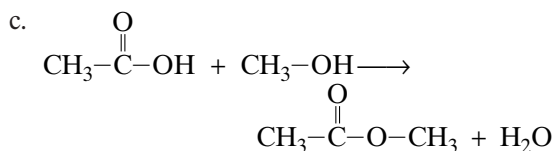
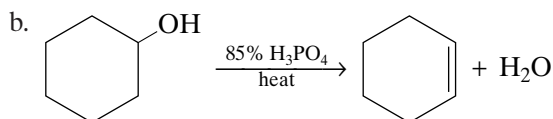
- What is a homologous series?
  - By what method are straight-chain hydrocarbons named?
  - Name the straight-chain alkane with the molecular formula  $C_{10}H_{22}$
- What are cycloalkanes?
- What trend occurs in the boiling points of alkanes?
  - How would you explain this trend?
  - How is the trend in alkane boiling points used in petroleum fractional distillation?
- Give examples of ethene's commercial uses.
- Give one use for ethyne.
- What is the name of the parent hydrocarbon of simple aromatic hydrocarbons?

### PRACTICE PROBLEMS

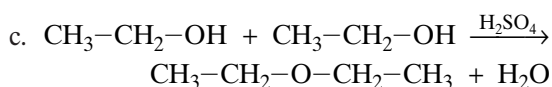
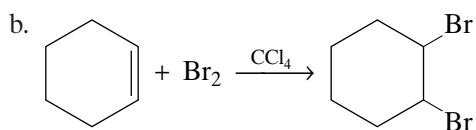
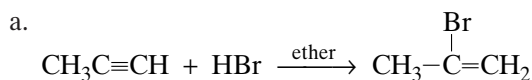
- Name the following molecules. (Hint: See Sample Problem A.)
  - $CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$
  - $$\begin{array}{c} CH_3 \\ | \\ CH_3 - C - CH_2 - CH - CH - CH_3 \\ | \quad | \\ CH_3 \quad CH_3 \end{array}$$
  - $$\begin{array}{c} CH_3 \quad \quad \quad CH_2-CH_2-CH_2-CH_3 \\ | \quad \quad \quad | \\ CH_3 - C - CH_2 - CH_2 - CH - CH_2 - CH_2 - CH_2 - CH_3 \\ | \\ CH_3 \end{array}$$
- Give the complete, uncondensed, structural formula for each of the following alkanes. (Hint: See Sample Problem A.)
  - decane
  - 3,3-dimethylpentane
- Give the condensed structural formula for 2,2,4,4-tetramethylpentane.
- For each of the following, determine whether the alkane is named correctly. If it is not, give the correct name.
  - 1-methylpropane
  - nonane
  - 4-methylhexane
  - 4-ethyl-2-methylhexane



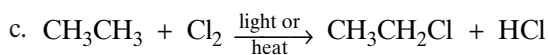
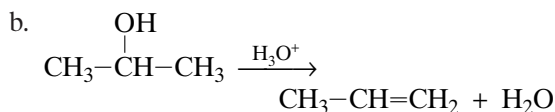
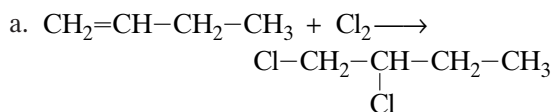




38. Which of the following reactions is a condensation reaction?

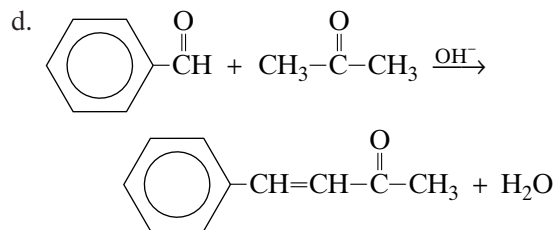
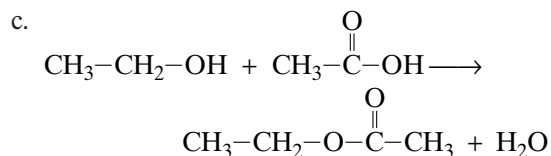
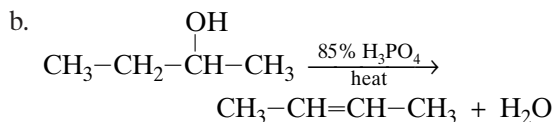
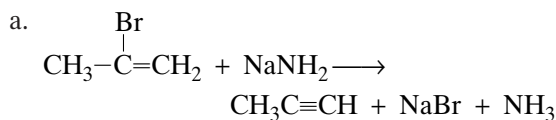


39. Which of the following reactions is an elimination reaction?

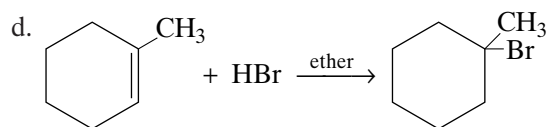
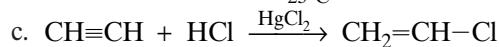
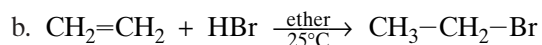
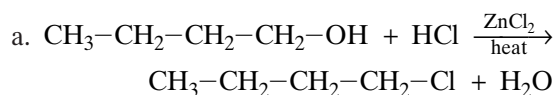


## MIXED REVIEW

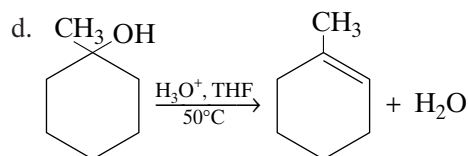
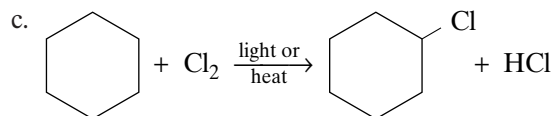
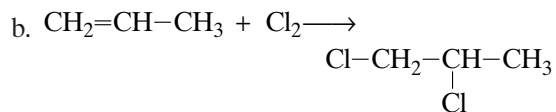
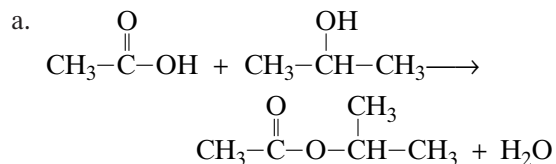
40. Classify each of the following reactions as an elimination reaction or a condensation reaction:



41. Classify each of the following reactions as a substitution reaction or an addition reaction:



42. Identify each of the following reactions as an addition, substitution, elimination, or condensation reaction:



43. Give the molecular formula for each type of hydrocarbon if it contains seven carbon atoms.
- an alkane
  - an alkene
  - an alkyne
44. a. Alkyne nomenclature is very similar to the nomenclature of what other group of hydrocarbons?  
b. How do these nomenclatures differ?
45. a. What are delocalized electrons?  
b. What is their effect on the reactivity of aromatic hydrocarbons?

### CRITICAL THINKING

46. **Inferring Conclusions** Why are organic compounds with covalent bonds usually less stable when heated than inorganic compounds with ionic bonds?
47. **Inferring Relationships** The element that appears in the greatest number of compounds is hydrogen. The element found in the second greatest number of compounds is carbon. Why are there more hydrogen compounds than carbon compounds?
48. **Relating Ideas** As the number of carbon atoms in an alkane molecule increases, does the percentage by mass of hydrogen increase, decrease, or remain the same?
49. **Applying Ideas** How does ethylene glycol protect radiator fluid in an automobile from both freezing in the winter and boiling over in the summer?



### USING THE HANDBOOK

50. The top 10 chemicals produced in the United States are listed in Table 7B of the *Elements Handbook*. Review this material, and answer the following:
- Which of the top ten compounds are organic?
  - Write structural formulas for the compounds you listed in item (a).
- c. To what homologous series do each of these compounds belong?
51. The reaction of methane with oxygen produces two different oxides of carbon. Review this material in the *Elements Handbook*, and answer the following:
- What condition determines whether the product of the methane reaction is  $\text{CO}_2$  or  $\text{CO}$ ?
  - If a home heating system is fueled by natural gas, what difference does it make if the combustion produces  $\text{CO}_2$  or  $\text{CO}$ ?
52. Silicon is similar to carbon in forming long-chain compounds. Review the material on silicon in the *Elements Handbook* and answer the following.
- How does a long-chain silicon compound differ in composition from a long-chain carbon compound?
  - The simplest alkane is methane. Methyl groups are found in all alkanes. What is a common subunit of a silicate? What is the geometry of that subunit?
53. Mercury in the environment poses a hazard to living things. Review the section on mercury poisoning in the *Elements Handbook*.
- Draw a structural formula for the organic mercury compound described in that section.
  - What is the IUPAC name for this compound?

### RESEARCH & WRITING

54. *Chemical and Engineering News* publishes a list once a year of the top 50 chemicals. Find out which chemicals on the current year's list are hydrocarbons, and report your findings to the class.
55. Consult reference materials at the library, and read about products made from hydrocarbons. Keep a list of the number of petroleum-related products you use in a single day.
56. The widespread use of synthetic polymers in modern society has led to a number of new environmental problems. Find out what some of these problems are and what can be done to reduce them.

## ALTERNATIVE ASSESSMENT

- 57. Performance** Models are often used to visualize the three-dimensional shape of molecules. Using gumdrops as atoms and toothpicks to bond them together, construct models of different hydrocarbons. Use large gumdrops for carbon and smaller gumdrops for hydrogen.
- 58. Performance** Using your gumdrop models, demonstrate why alkenes can have geometric isomers, while alkanes (except cycloalkanes) cannot.
- 59. Performance** Devise a set of experiments to study how well biodegradable plastics break down. If your teacher approves your plan, conduct an experiment to test the procedure on products labeled “biodegradable.”
- 60. Performance** Your teacher will make available unlabeled samples of benzoic acid, ethyl alcohol, and hexanediamine. Develop an experiment to identify each. If your teacher approves your plan, identify the unknown substances.
- 61.** Keep a list of the food you consume in a single day. Compare the content labels from those foods, and then list the most commonly used chemicals in them. With the aid of your teacher and some reference books, try to classify the organic chemicals by their functional groups.
- 62.** As a class or small group, research the preservatives used in various foods. Examine their chemical structures. Determine a way to test for organic functional groups of possibly hazardous preservatives.

## extension



## Graphing Calculator

## Formulas of Hydrocarbons

Go to [go.hrw.com](http://go.hrw.com) for a graphing calculator exercise that asks you to find the formula of any straight-chain hydrocarbon, given the number of carbon atoms and the number of double bonds in the compound.



Keyword: HC6ORGX

# Math Tutor

## CALCULATING EMPIRICAL FORMULAS

An empirical formula shows the simplest whole-number ratio among the elements in a compound. For example, the simplest ratio among the atoms in benzene,  $C_6H_6$ , can be expressed by the empirical formula CH (or  $C_1H_1$ ). Empirical formulas may be calculated from simple analytical data, as shown in the example below.

### Problem-Solving TIPS

- Sometimes, you can deduce the mole ratios of the elements in a compound just by examining the moles of each element. If not, divide the moles of each element by the moles of the least abundant element in the compound.

### SAMPLE

**Find the empirical formula of acetone, a common organic solvent, whose composition by mass is 62.04% carbon, 10.41% hydrogen, and 27.55% oxygen.**

The easiest way to calculate an empirical formula from percentage composition is to consider a 100.00 g sample of the compound. In this case, a 100.00 g sample would contain 62.04 g of carbon, 10.41 g of hydrogen, and 27.55 g of oxygen. Convert each mass to moles so that you can compare the mole ratio of the three elements.

$$62.04 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.166 \text{ mol C}$$

$$10.41 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 10.33 \text{ mol H}$$

$$27.55 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.722 \text{ mol O}$$

$$\frac{5.166 \text{ mol C}}{1.722} : \frac{10.33 \text{ mol H}}{1.722} : \frac{1.722 \text{ mol O}}{1.722}$$

$$3.000 \text{ mol C} : 5.999 \text{ mol H} : 1.000 \text{ mol O}$$

As you can see, the empirical formula of this compound is  $C_3H_6O$ .

### PRACTICE PROBLEMS

1. Urea was the first organic compound to be synthesized in the laboratory. Urea's composition by mass is 20.00% carbon, 6.71% hydrogen, 46.65% nitrogen, and 26.64% oxygen. What is the empirical formula of urea?
2. An organic compound sometimes used in the manufacture of perfumes is 29.78% carbon, 4.17% hydrogen, and 66.05% bromine, by mass. What is the empirical formula of this compound?
3. The composition by mass of lactic acid is 40.00% carbon, 6.71% hydrogen, and 53.28% oxygen. What is the empirical formula of lactic acid?



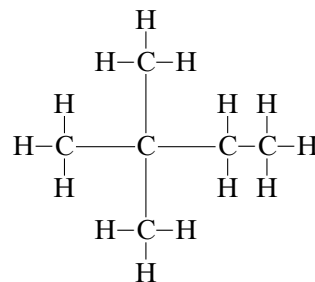
# Standardized Test Prep

Answer the following items on a separate piece of paper.

## MULTIPLE CHOICE

- Which of the following hydrocarbons must be an alkane?
  - $C_2H_2$
  - $C_5H_{10}$
  - $C_7H_{12}$
  - $C_{14}H_{30}$
- A hydrocarbon with the formula  $C_8H_{18}$  is called.
  - octene.
  - octyne.
  - octane.
  - propane.
- During a condensation polymerization reaction,
  - single bonds replace all double bonds that are present in the monomer.
  - water is often produced.
  - alcohol groups are formed.
  - an aldehyde group is changed to a ketone group.
- In naming an organic compound, we
  - should remember that naming the locations of all functional groups is optional.
  - do not consider the number of carbon atoms in the molecule as a factor.
  - begin by identifying and naming the longest hydrocarbon chain.
  - ignore side chains when we name the molecule.
- Which of the following compounds cannot have different isomers?
  - $C_7H_{16}$
  - $C_5H_{10}$
  - $C_3H_8$
  - $C_6H_{12}O_6$
- Compounds that differ only in the order in which the atoms are bonded together are called
  - condensation polymers.
  - addition polymers.
  - structural isomers.
  - geometric isomers.

7. Examine the following structural formula:



The correct name for this compound is

- 2,2-dimethylbutane.
  - 1,1,1-trimethylpropane.
  - 2-ethyl-2-methylpropane.
  - 3,3-dimethylbutane.
8. Organic functional groups
- give characteristic properties to compounds that contain them.
  - always contain oxygen.
  - always contain a double or triple bond.
  - are present in every organic compound.

## SHORT ANSWER

- What feature must be present in a compound for it to undergo an addition reaction?
- Organic compounds of what class contain only two elements?

## EXTENDED RESPONSE

- Describe the difference between substitution and addition reactions. Your answer should mention the degree of saturation of the organic compound.
- Explain why some alkanes are gases, others are liquids, and still others are solids at room temperature.

### Test TIP

If you become short on time, quickly scan the unanswered questions to see which might be easiest to answer.

# Polymers and Toy Balls

## OBJECTIVES

- *Synthesize* two different polymers.
- *Prepare* a small toy ball from each polymer.
- *Observe* the similarities and differences between the two types of balls.
- *Measure* the density of each polymer.
- *Compare* the bounce height of the two balls.

## MATERIALS

- 2 L beaker, or plastic bucket or tub
- 3 mL 50% ethanol solution
- 5 oz paper cups, 2
- 10 mL 5% acetic acid solution (vinegar)
- 25 mL graduated cylinder
- 10 mL graduated cylinder
- 10 mL liquid latex
- 12 mL sodium silicate solution
- distilled water
- gloves
- meterstick
- paper towels
- wooden stick



## BACKGROUND

What polymers make the best toy balls? Two possibilities are latex rubber and a polymer produced from ethanol and sodium silicate. Latex rubber is a polymer of covalently bonded atoms.

The polymer formed from ethanol,  $C_2H_5OH$ , and a solution of sodium silicate,  $Na_2Si_3O_7$ , also has covalent bonds. It is known as *water glass* because it dissolves in water.

In this experiment, you will synthesize rubber and the ethanol silicate polymer and test their properties.

## SAFETY



For review of safety, please see **Safety in the Chemistry Laboratory** in the front of your book.

## PREPARATION

1. **Organizing Data:** Create a data table in your lab notebook in which the columns are labeled “Trial,” “Height (cm),” “Mass (g),” and “Diameter (cm).” Create three rows (Trials 1, 2, and 3) in your table.

## PROCEDURE

1. Fill the 2 L beaker, bucket, or tub about half-full with distilled water.
2. Using a clean 25 mL graduated cylinder, measure 10 mL of liquid latex and pour it into one of the paper cups.
3. Thoroughly clean the 25 mL graduated cylinder with soap and water, and then rinse it with distilled water.

- Measure 10 mL of distilled water. Pour it into the paper cup with the latex.
- Measure 10 mL of the 5% acetic acid solution, and pour it into the paper cup with the latex and water.
- Immediately stir the mixture with the wooden stick.
- As you continue stirring, a polymer lump will form around the wooden stick. Pull the stick with the polymer lump from the paper cup, and immerse the lump in the 2 L beaker, bucket, or tub.
- While wearing gloves, gently pull the lump from the wooden stick. Be sure to keep the lump immersed under the water.
- Keep the latex rubber underwater, and use your gloved hands to mold the lump into a ball. Then, squeeze the lump several times to remove any unused chemicals. You may remove the latex rubber from the water as you roll it in your hands to smooth the ball.
- Set aside the latex rubber ball to dry. While it is drying, proceed to step 11.
- In a clean 25 mL graduated cylinder, measure 12 mL of sodium silicate solution, and pour it into the other paper cup.
- In a clean 10 mL graduated cylinder, measure 3 mL of 50% ethanol. Pour the ethanol into the paper cup with the sodium silicate, and mix with the wooden stick until a solid substance is formed.
- While wearing gloves, remove the polymer that forms and place it in the palm of one hand. Gently press it with the palms of both hands until a ball that does not crumble is formed. This step takes a little time and patience. The liquid that comes out of the ball is a combination of ethanol and water. Occasionally, moisten the ball by letting a small amount of water from a faucet run over it. When the ball no longer crumbles, you are ready to go to the next step.
- Observe as many physical properties of the balls as possible, and record your observations in your lab notebook.
- Drop each ball several times, and record your observations.
- Drop one ball from a height of 1 m, and measure its bounce. Perform three trials for each ball.
- Measure the diameter and mass of each ball.

## CLEANUP AND DISPOSAL

- Dispose of any extra solutions in the containers indicated by your teacher. Clean up your lab area. Remember to wash your hands thoroughly when your lab work is finished.



## ANALYSIS AND INTERPRETATION

- Analyzing Information:** List at least three of your observations of the properties of the two balls.
- Organizing Data:** Calculate the average height of the bounce for each type of ball.
- Organizing Data:** Calculate the volume for each ball. Even though the balls may not be perfectly spherical, assume that they are. (Hint: The volume of a sphere is equal to  $\frac{4}{3} \times \pi \times r^3$ , where  $r$  is the radius of the sphere, which is one-half of the diameter.) Then, calculate the density of each ball, using your mass measurements.

## CONCLUSIONS

- Inferring Conclusions:** Which polymer would you recommend to a toy company for making new toy balls? Explain your reasoning.
- Evaluating Viewpoints:** What are some other possible practical applications for each of the polymers you made?

## EXTENSIONS

- Predicting Outcomes:** Explain why you would not be able to measure the volumes of the balls by submerging them in water.