

► **Learning Check 9.13** Write balanced equations to represent the preparation of AlCl_3 , using Reactions 9.26 through 9.30. Use molecular equations to emphasize the salt formation.

We have expressed the amounts of materials involved in chemical reactions primarily in terms of mass (grams) or number of moles. However, in certain applications, it is important to express the amount of salt in a solution in terms of the amount of electrical charge represented by the ions of the salt. This is especially true in medical applications, where the levels and balance of electrolytes in various body fluids are extremely important.

A unit that expresses the amount of ionic electrical charge for salts is the equivalent. One **equivalent of salt** is the amount that will produce 1 mol of positive (or negative) charges when dissolved and dissociated. To determine the amount of salt that represents 1 equivalent (eq), you must know what ions are produced when the salt dissociates. For example, potassium chloride (KCl), an electrolyte often administered to patients following surgery, dissociates as follows: $\text{KCl}(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{Cl}^-(\text{aq})$. Thus, 1 mol, or 74.6 g, of solid KCl provides 1 mol of positive charges (1 mol of K^+ ions) when it dissolves and dissociates. Thus, 1 eq of KCl is equal to 1 mol, or 74.6 g, of KCl.

equivalent of salt The amount that will produce 1 mol of positive electrical charge when dissolved and dissociated.

► Example 9.10

Determine the number of equivalents and milliequivalents (meq) of salt contained in the following:

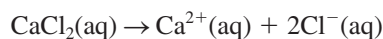
- 0.050 mol KCl
- 0.050 mol CaCl_2

Solution

- As shown above, 1 mol KCl = 1 eq KCl. Therefore, 0.050 mol KCl = 0.050 eq KCl. Also, because 1 eq = 1000 meq,

$$0.050 \text{ eq} \times \frac{1000 \text{ meq}}{1 \text{ eq}} = 50 \text{ meq}$$

- The dissociation reaction is



Thus, we see that 1 mol CaCl_2 produces 1 mol Ca^{2+} or 2 mol of positive charges. Thus, 1 mol CaCl_2 = 2 eq CaCl_2 . Therefore,

$$0.050 \text{ mol } \text{CaCl}_2 \times \frac{2 \text{ eq } \text{CaCl}_2}{1 \text{ mol } \text{CaCl}_2} = 0.10 \text{ eq } \text{CaCl}_2$$

Also, because 1 eq = 1000 meq,

$$0.10 \text{ eq} \times \frac{1000 \text{ meq}}{1 \text{ eq}} = 1.0 \times 10^2 \text{ meq}$$

Notice that we could have focused on the negative charges in either part above and still arrived at the same answers. In part a, 0.050 mol KCl produces 0.050 mol Cl^- , or 0.050 mol of negative charge. Similarly, in part b, 0.050 mol CaCl_2 produces 2×0.050 , or 0.10 mol, Cl^- ions, or 0.10 mol of negative charge. We have arbitrarily chosen to use the positive charges, but either will work. Just remember that you do not count both the negative and positive charges for a salt.

► **Learning Check 9.14** Determine the number of equivalents and milliequivalents in each of the following:

- a. 0.10 mol NaCl
- b. 0.10 mol $\text{Mg}(\text{NO}_3)_2$

► **Example 9.11**

A sample of blood serum contains 0.139 equivalents of Na^+ ion per liter of serum. Assume the Na^+ comes from dissolved NaCl, and calculate the number of equivalents, number of moles and number of grams of NaCl in 250 mL of the serum.

Solution

The Na^+ ion has a single charge, so we may write the following relationships:

$$1.00 \text{ mol NaCl} = 1.00 \text{ mol Na}^+ = 1.00 \text{ eq. Na}^+ = 1.00 \text{ eq NaCl}$$

These relationships provide numerous factors that can be used in the factor-unit method to solve problems. Two that will be used to solve these problems are:

$$\frac{1.00 \text{ eq NaCl}}{1.00 \text{ eq Na}^+} \quad \text{and} \quad \frac{1.00 \text{ mol NaCl}}{1.00 \text{ eq NaCl}}$$

The known quantity is 250 mL of serum. This is converted into equivalents of Na^+ by using the given concentration of $\frac{0.139 \text{ eq Na}^+}{\text{L serum}}$ as the factor.

$$(0.250 \text{ L serum}) \times \frac{0.139 \text{ eq Na}^+}{1.00 \text{ L serum}} = 0.0347 \text{ eq Na}^+$$

This number of equivalents of Na^+ will now be converted into the quantities asked for by using the two factors given above, and a factor obtained from the formula weight for NaCl of 58.44 u.

$$(0.0347 \text{ eq Na}^+) \times \frac{1.00 \text{ eq NaCl}}{1.00 \text{ eq Na}^+} = 0.0347 \text{ eq NaCl}$$

$$(0.0347 \text{ eq NaCl}) \times \frac{1.00 \text{ mol NaCl}}{1.00 \text{ eq NaCl}} = 0.0347 \text{ mol NaCl}$$

$$(0.0347 \text{ mol NaCl}) \times \frac{58.44 \text{ g NaCl}}{1.00 \text{ mol NaCl}} = 2.03 \text{ g NaCl}$$

► **Learning Check 9.15** A sample of blood serum contains 0.103 equivalents of Cl^- ions per liter. Assume the Cl^- comes from dissolved NaCl and calculate the number of equivalents, number of moles, and number of grams of NaCl in 250 mL of serum.

9.9 The Strengths of Acids and Bases

Learning Objective

9. Demonstrate an understanding of the words *weak* and *strong* as applied to acids and bases.

When salts dissolve in water, they generally dissociate completely, but this is not true for all acids and bases. The acids and bases that do dissociate almost completely are classified as **strong acids** and **strong bases** (they are also strong electrolytes). Those that dissociate to a much smaller extent are called **weak** or **moderately weak**, depending on the degree of dissociation (they are also weak or moderately weak electrolytes). Examples of strong and weak acids are given in ► Table 9.7.

strong acids and strong bases Acids and bases that dissociate (ionize) completely when dissolved to form a solution.

weak (or moderately weak) acids and bases Acids and bases that dissociate (ionize) less than completely when dissolved to form a solution.

Chemistry Around Us 9.1

Beware the Negative Effects of Acids on Teeth



Dental cavities and associated tooth decay are among the most common of worldwide health problems. They are especially common in children and young adults, but anyone with teeth can get cavities.

A typical tooth consists of several layers of different materials. The visible, outer, exposed layer is called the enamel, and is the hardest substance found in the human body. Enamel is composed primarily of a compound called hydroxyapatite with the formula $\text{Ca}_{10}(\text{PO}_4)_6(\text{OH})_2$. Most of the mass of a tooth is made up of dentin, a substance chemically identical to bone. The dentin forms a layer under the enamel. Near the center of the dentin layer is the pulp cavity of the tooth which contains arteries, veins and nerves.

Dental cavities develop in teeth when the enamel is penetrated and the underlying dentin layer is exposed. The primary cause of cavities and tooth decay is thought to be acids that can attack the hydroxyapatite and other minerals of tooth enamel. Once the enamel layer is penetrated, the dentin is reached by the acids. Dentin is softer and less resistant to acid than enamel, and the process of decay speeds up. If not stopped, the decay can reach the pulp of the tooth with its contained blood vessels and nerves. Severe pain often results when decay becomes this advanced.

With acid identified as the culprit, the next steps in preventing or minimizing tooth decay are to identify the sources of acid and devise ways to minimize the exposure of the teeth to acid from these sources. The mouth naturally contains many kinds of bacteria. Some of these bacteria thrive on sugars and starches that are known as fermenting carbohydrates, and convert them into numerous products, including acids. This conversion process can begin in as little as 20 minutes after eating. The bacteria, produced acids, food particles and saliva form a sticky film called dental plaque that coats the teeth. If not removed by the standard methods of thorough brushing and flossing, the acids in the dental plaque attack the enamel and begin the process of decay.

Sugar and other carbohydrates are not the only sources of acid in the mouth. Some foods and drinks contain significant amounts of acid. Examples are carbonated soft drinks, fruit juices, wine, pickles, sauerkraut, buttermilk, fresh citrus fruits such as grapefruit, oranges and lemons, and even sour hard candies. The timing of the consumption of such foods can help minimize acid damage. Acidic foods should not be eaten alone, but should be eaten as a part of a meal. When this is done, the non-acidic foods and saliva help neutralize the acids. The worst time to consume acidic foods is just before bedtime, because the production of acid-neutralizing saliva decreases during

sleep. Even the method used to drink liquids can help minimize the contact of acids with the teeth. For example, the use of a straw delivers liquids past the teeth with minimal contact. Also, drinking liquids quickly rather than sipping them over long periods helps reduce the time of contact with the teeth. Chewing sugar-free gum after eating stimulates the flow of acid-neutralizing saliva.

In summary, minimizing the contact time of the teeth with acidic materials, including those produced by bacteria, is an effective way to help prevent or minimize tooth decay. If some of the techniques and procedures described above seem too hard or inconvenient to be followed, remember that one of the easiest and most effective ways to accomplish the prevention or reduction of tooth decay is to follow the advice always given by dental health professionals: brush and floss teeth on a frequent and regular basis.

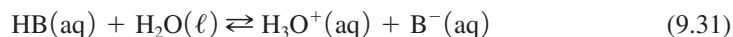


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Oranges and other citrus fruits are sources of acid in the mouth.

A 0.10 M solution of hydrochloric acid could be prepared by dissolving 0.10 mol (3.7 g) of HCl gas in enough water to give 1.0 L of solution. According to Table 9.7, 100% of the dissolved gas would dissociate into H^+ and Cl^- . Thus, the concentration of H^+ in a 0.10 M HCl solution is 0.10 mol/L, and the pH is 1.00.

The strength of acids and bases is shown quantitatively by the value of the equilibrium constant for the dissociation reaction in water solutions. Equation 9.31 represents the dissociation of a general acid, HB, in water, where B^- represents the conjugate base of the acid:



The equilibrium expression for this reaction is

$$K = \frac{[\text{H}_3\text{O}^+][\text{B}^-]}{[\text{HB}][\text{H}_2\text{O}]} \quad (9.32)$$

Table 9.7 Some Common Strong and Weak Acids

Name	Formula	% Dissociation ^a	K_a	Classification
Hydrochloric acid	HCl	100	Very large	Strong
Hydrobromic acid	HBr	100	Very large	Strong
Nitric acid	HNO ₃	100	Very large	Strong
Sulfuric acid	H ₂ SO ₄	100	Very large	Strong
Phosphoric acid	H ₃ PO ₄	28	7.5×10^{-3}	Moderately weak
Sulfurous acid ^b	H ₂ SO ₃	34	1.5×10^{-2}	Moderately weak
Acetic acid	HC ₂ H ₃ O ₂	1.3	1.8×10^{-5}	Weak
Boric acid	H ₃ BO ₃	0.01	7.3×10^{-10}	Weak
Carbonic acid ^b	H ₂ CO ₃	0.2	4.3×10^{-7}	Weak
Nitrous acid ^b	HNO ₂	6.7	4.6×10^{-4}	Weak

^aBased on dissociation of one proton in 0.1 M solution at 25°C.

^bUnstable acid.

In Equation 9.32, the brackets, again, represent molar concentrations of the materials in the solution. Only a tiny amount of the water in the solution actually enters into the reaction, so the concentration of water is considered to be constant (see Section 9.4). We can then write Equation 9.33, where K_a is a new constant called the **acid dissociation constant**:

$$K[\text{H}_2\text{O}] = K_a = \frac{[\text{H}_3\text{O}^+][\text{B}^-]}{[\text{HB}]} \quad (9.33)$$

In Equation 9.33, $[\text{H}_3\text{O}^+]$ and $[\text{B}^-]$ are, respectively, the equilibrium concentrations of the hydronium ion and the anion conjugate base that is characteristic of the acid. The $[\text{HB}]$ represents the concentration of that part of the dissolved acid that remains undissociated in the equilibrium mixture. In solutions of strong acids, $[\text{H}_3\text{O}^+]$ and $[\text{B}^-]$ values are quite large, while $[\text{HB}]$ has a value near 0, so K_a is quite large. In weak acids, $[\text{HB}]$ has larger values, while $[\text{H}_3\text{O}^+]$ and $[\text{B}^-]$ are smaller, so K_a is smaller. Thus, the larger a K_a value, the stronger the acid it represents. This is illustrated by the K_a values given in Table 9.7. If we simplify by substituting H^+ for H_3O^+ , introduced in Section 9.5, we obtain

$$K_a = \frac{[\text{H}^+][\text{B}^-]}{[\text{HB}]} \quad (9.34)$$

It is important to remember that the terms *weak* and *strong* apply to the extent of dissociation and not to the concentration of an acid or base. For example, gastric juice (0.05% HCl) is a dilute (not weak) solution of a strong acid.

acid dissociation constant The equilibrium constant for the dissociation of an acid.

Example 9.12

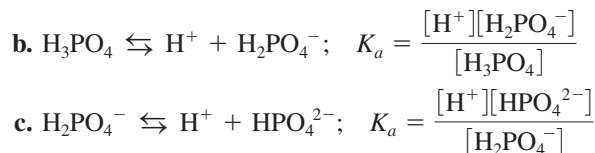
Write dissociation reactions and expressions for K_a for each of the following weak acids:

- Hydrocyanic acid (HCN)
- Phosphoric acid (H₃PO₄) (1st H only)
- Dihydrogen phosphate ion (H₂PO₄⁻) (1st H only)

Solution

In each case, H^+ has been substituted for H_3O^+ .

- $\text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^-$; $K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]}$



► **Learning Check 9.16** Write dissociation reactions and K_a expressions for the following weak acids:

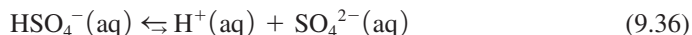
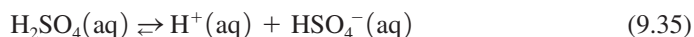
- Hydrogen phosphate ion (HPO_4^{2-})
- Nitrous acid (HNO_2)
- Hydrofluoric acid (HF)

monoprotic acid An acid that gives up only one proton (H^+) per molecule when dissolved.

diprotic acid An acid that gives up two protons (H^+) per molecule when dissolved.

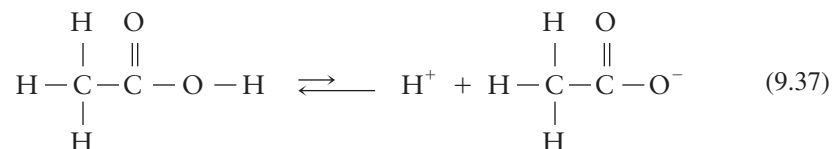
triprotic acid An acid that gives up three protons (H^+) per molecule when dissolved.

Acid behavior is linked to the loss of protons. Thus, acids must contain hydrogen atoms that can be removed to form H^+ . **Monoprotic acids** can lose only one proton per molecule, whereas **diprotic** and **triprotic** acids can lose two and three, respectively. For example, HCl is monoprotic, H_2SO_4 is diprotic, and H_3PO_4 is triprotic. Di- and triprotic acids dissociate in steps, as shown for H_2SO_4 in Equations 9.35 and 9.36:



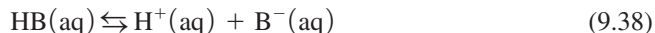
The second proton is not as easily removed as the first because it must be pulled away from a negatively charged particle, HSO_4^- . Accordingly, HSO_4^- is a weaker acid than H_2SO_4 .

The number of ionizable hydrogens cannot always be determined from the molecular formula for an acid. For example, acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) is monoprotic even though the molecule contains four hydrogen atoms. The dissociation of acetic acid is represented by Equation 9.37, where structural formulas are used to emphasize the different H atoms in the molecule:



Only the hydrogen bound to the oxygen is ionizable. Those hydrogens bound to C are too tightly held to be removed. ► Table 9.8 contains other examples, with the ionizable hydrogens shown in color.

We have focused our attention on the strength of acids, using the extent of dissociation as a basis. However, all acid dissociations are reversible to some degree, and in the reverse reactions, anions produced by the forward reaction behave as Brønsted bases. What can be said about the strength of these bases? Because Brønsted acid–base behavior is really just competition for protons, we can answer this question by looking again at the simplified form of the equation for dissociation of a general acid:



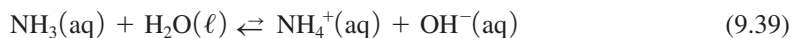
We see from this equation that the strength of HB as an acid depends on how tightly the conjugate base B^- holds onto the proton. If HB is a strong acid, the conjugate base holds onto the proton only weakly. If HB is a weak acid, the conjugate base holds on more strongly, depending on the strength of HB. Thus, we have answered our earlier question. If HB is a strong acid, the H^+ is held only weakly by the B^- . We can conclude then that B^- is not strongly attracted to protons—it is a weak base. Conversely, if HB is a weak acid, the H^+ is held tightly by the B^- , and we conclude that B^- is more strongly attracted to protons—it is a stronger base than the B^- from a strong acid.

Table 9.8 Examples of Monoprotic, Diprotic, and Triprotic Acids

Name	Formula	Structural Formula	Classification
Butyric acid	$\text{HC}_4\text{H}_7\text{O}_2$	$ \begin{array}{ccccccc} & \text{H} & & \text{H} & & \text{H} & & \text{O} \\ & & & & & & & \\ \text{H} & - \text{C} & - & \text{C} & - & \text{C} & - & \text{C} - \text{O} - \text{H} \\ & & & & & & & \\ & \text{H} & & \text{H} & & \text{H} & & \end{array} $	Monoprotic
Carbonic acid	H_2CO_3	$ \begin{array}{c} \text{O} \\ \\ \text{H} - \text{O} - \text{C} - \text{O} - \text{H} \end{array} $	Diprotic
Formic acid	HCHO_2	$ \begin{array}{c} \text{O} \\ \\ \text{H} - \text{C} - \text{O} - \text{H} \end{array} $	Monoprotic
Nitric acid	HNO_3	$ \begin{array}{c} \text{O} \\ \\ \text{N} - \text{O} - \text{H} \\ \\ \text{O} \end{array} $	Monoprotic
Phosphoric acid	H_3PO_4	$ \begin{array}{c} \text{O} \\ \\ \text{H} - \text{O} - \text{P} - \text{O} - \text{H} \\ \\ \text{O} \\ \\ \text{H} \end{array} $	Triprotic
Phosphorous acid	H_3PO_3	$ \begin{array}{c} \text{O} \\ \\ \text{H} - \text{O} - \text{P} - \text{O} - \text{H} \\ \\ \text{H} \end{array} $	Diprotic

In general, the conjugate base anions produced by the dissociation of strong Brønsted acids are weak Brønsted bases. The conjugate base anions of weak acids are stronger bases, with their strengths dependent on the strength of the parent acid.

Ammonia (NH_3) is the weak base most often encountered in addition to the anions of strong acids. The dissociation reaction of gaseous NH_3 in water, given earlier in Example 9.1, is



The most common strong bases are the hydroxides of group IA(1) metals (NaOH , KOH , etc.) and the hydroxides of group IIA(2) metals ($\text{Mg}(\text{OH})_2$, $\text{Ca}(\text{OH})_2$, etc.).

Example 9.13

Classify each of the following pairs by identifying the stronger of the pair according to the indicated behavior. Information from Tables 9.7 and 9.8 may be used.

- H_3PO_4 and H_2PO_4^- (acid)
- H_2PO_4^- and HPO_4^{2-} (base)
- HNO_3 and HNO_2 (acid)

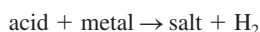
Study Skills 9.1 Writing Reactions of Acids

As you study this chapter, you will acquire a knowledge of the characteristic reactions of acids and the ability to write balanced equations for the reactions. In the Key Equations section at the end of the chapter, five characteristic reactions are summarized in item 3. One useful way to remember the reactions is to learn them as general word equations, such as

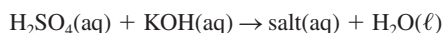
“An acid plus a metal gives a salt plus hydrogen gas”

rather than as specific equations such as $\text{H}_2\text{SO}_4(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g})$. If you learn the general word equations and you recognize the starting materials for a reaction (such as an acid and a metal), you will know what the products will be (a salt and H_2 gas).

A second approach is to remember that all five of the general reactions of acids are either single-replacement

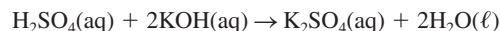


or double-replacement (the remaining four acid reactions in the Key Equations section). The formulas of the products of double-replacement reactions can be predicted by simply breaking each reactant into its positive and negative parts and recombining the parts in the other possible way (the positive part of one reactant with the negative part of the other reactant). For example, let's determine the products and the balanced equation for a reaction between H_2SO_4 and KOH . If you remember that H_2SO_4 is an acid and KOH is a base, the general word equation says the products will be a salt and water:

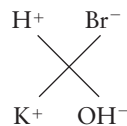


H_2SO_4 breaks apart to give H^+ and SO_4^{2-} ; KOH breaks apart to give K^+ and OH^- . If we combine the positive part of the acid (H^+) with the negative part of the base (OH^-), we get the water (H_2O). A similar combination of the positive part of the base (K^+) with the

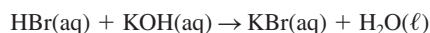
negative part of the acid (SO_4^{2-}) gives the salt. We must remember that the total charges of the combined parts must add up to 0. Thus, two K^+ will combine with one SO_4^{2-} to give the salt K_2SO_4 . The balanced equation is



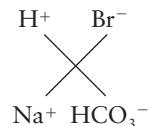
As another example, consider a reaction between HBr and KOH . KBr breaks into H^+ and Br^- . KOH breaks into K^+ and OH^- . Now, switch the parts and recombine:



The products are KBr and H_2O , and the balanced equation is



For a final example, let's try a reaction between HBr and NaHCO_3 . The word equation predicts that the products should be a salt, water, and carbon dioxide gas. HBr breaks into H^+ and Br^- ; NaHCO_3 breaks into Na^+ and HCO_3^- . Now, switch the parts and recombine:



The products are NaBr and H_2CO_3 . However, H_2CO_3 is not stable; it decomposes to give H_2O and CO_2 , the products predicted earlier. The balanced equation is



Solution

- H_3PO_4 is stronger as an acid. In di- and triprotic acids, each proton in the removal sequence is harder to remove, so the corresponding acid is weaker.
- The stronger acid produces the weaker anion base. Thus, H_2PO_4^- , the anion of the stronger acid H_3PO_4 , would be a weaker base than HPO_4^{2-} , the anion of the weaker acid H_2PO_4^- . So HPO_4^{2-} is a stronger base than H_2PO_4^- .
- HNO_3 is the stronger acid according to Table 9.7. Generally, when related acids (same atoms, etc.) are compared for strength, the one containing more oxygen atoms will be the stronger.

► **Learning Check 9.17** Classify each of the following according to strength for the indicated behavior. If more than two are compared, list them with the strongest at the top and the weakest at the bottom. Use Tables 9.7 and 9.8 as needed.

- HClO , HClO_3 , HClO_2 (acid)
- NO_2^- and NO_3^- (base)
- $\text{HC}_2\text{H}_3\text{O}_2$ and $\text{C}_2\text{H}_3\text{O}_2^-$ (acid)

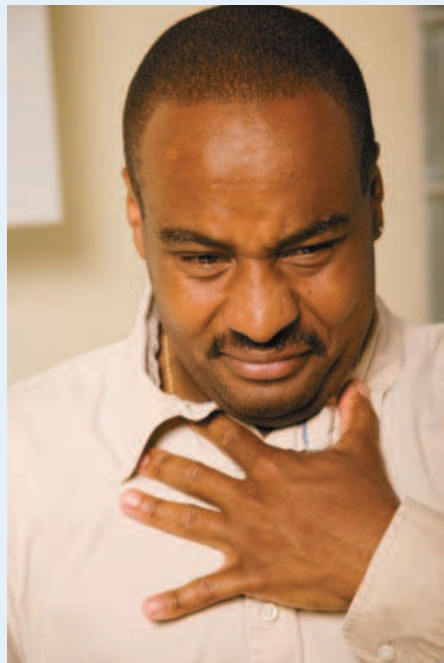


The medical name for acid reflux disease is *gastroesophageal reflux disease*, which is often abbreviated and referred to as GERD. The disease is often mistaken for occasional heartburn and treated with over-the-counter remedies (See At the Counter 9.1). GERD is the result of a malfunctioning muscle (the lower esophageal sphincter (LES) muscle) located at the bottom of the esophagus, just above the stomach. When operating normally, this muscle relaxes and opens to allow food to pass from the esophagus down into the stomach, then contracts to close the opening and prevent the acidic contents of the stomach from backing up into the esophagus.

When the muscle relaxes at inappropriate times, the acidic stomach contents get into the esophagus and cause the burning chest pain called heartburn. However, when this occurs repeatedly and frequently, the acidic stomach contents can also erode the lining of the esophagus. GERD is a complex condition with many degrees of severity, ranging from only frequent heartburn symptoms to erosive esophagitis, in which the esophagus can suffer different degrees of damage. In extreme cases of erosive esophagitis, ulcers develop in the esophagus and lead to esophageal bleeding that, if persistent and undetected, can lead to iron deficiency and anemia as well as extreme pain and weight loss. In some cases, severe GERD can lead to other serious medical conditions that require hospitalization and even surgery to correct.

Doctors often recommend lifestyle and dietary changes for most GERD patients, including the avoidance of foods and beverages that weaken the LES muscle. These foods include chocolate, peppermint, fatty foods, coffee, and alcoholic beverages. The use of foods and beverages that can irritate a damaged esophageal lining, such as acidic fruits and juices, pepper, and tomato products is also discouraged. GERD symptoms in overweight individuals often diminish when some weight is lost. Smokers who quit also generally gain some relief. Prescription medications that reduce the amount of acid in the stomach are also available. Two types of medication are available. Both types, called H₂ blockers and proton (acid) pump inhibitors, decrease the amount of acid secreted into the stomach, but by different mechanisms.

The main symptom that may indicate the presence of GERD is frequent, persistent heartburn that occurs two or more times a week. Other symptoms include difficulty in swallowing and frequent belching and regurgitation. Less common symptoms that occur in some people resemble respiratory conditions and include a persistent sore throat, wheezing, chronic coughing, and hoarseness. Individuals who suspect they might be suffering from some degree of GERD should consult a physician to determine the extent of the disease and proper treatment.



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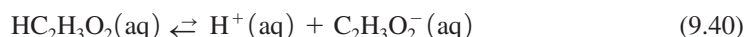
GERD has many degrees of severity.

9.10 Analyzing Acids and Bases

Learning Objective

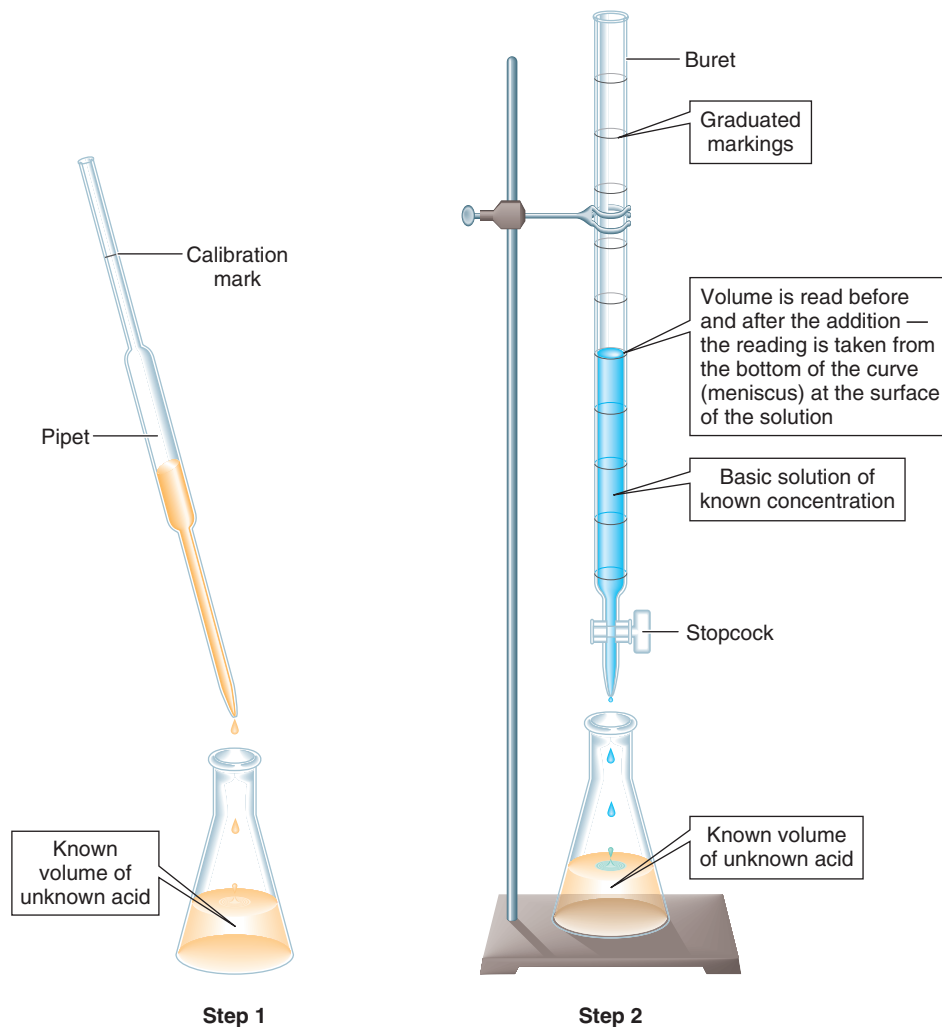
10. Demonstrate an understanding of the titration technique used to analyze acids and bases.

The analysis of solutions for the total amount of acid or base they contain is a regular activity in many laboratories. The total amount of acid in a solution is indicated by its capacity to neutralize a base. The pH is related to the acidity or concentration of H⁺ in solution, while the capacity to neutralize a base depends on the total amount of H⁺ available. For example, a 0.10 M acetic acid solution has an H⁺ concentration of about 1.3×10^{-3} M (pH = 2.89). However, 1 L of the solution can neutralize 0.10 mol of OH⁻, not just 1.3×10^{-3} mol. The reason is that the dissociation equilibrium of acetic acid is



As OH⁻ is added, H⁺ reacts to form water (see Equation 9.24). The removal of H⁺ causes the equilibrium of Reaction 9.40 to shift right in accordance with Le Châtelier's principle. The continued addition of OH⁻ and removal of H⁺ will eventually cause all of the acetic acid molecules to dissociate and react.

Figure 9.4 Titration.



titration An analytical procedure in which one solution (often a base) of known concentration is slowly added to a measured volume of an unknown solution (often an acid). The volume of the added solution is measured with a buret.

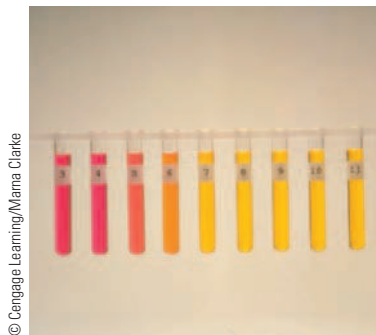
equivalence point of a titration The point at which the unknown solution has exactly reacted with the known solution. Neither is in excess.

end point of a titration The point at which the titration is stopped on the basis of an indicator color change or pH meter reading.

A common procedure often used to analyze acids and bases is called **titration** (Figure 9.4). Suppose the total acidity of an unknown acid solution needs to be determined. A known volume of the acidic solution is first measured out by drawing it up to the calibration mark of a pipet. This solution is placed in a container (Step 1). A basic solution of known concentration (a standard solution) is added to the acid solution in the container until the **equivalence point** is reached. This is the point where the unknown acid is completely reacted with base. The volume of base needed to reach the equivalence point is obtained from the buret readings.

To successfully complete a titration, the point at which the reaction is completed must somehow be detected. One way to do this is to add an indicator to the solution being titrated. An indicator is an organic compound that changes to different colors depending on the pH of its surroundings. An indicator is selected that will change color at a pH as close as possible to the pH the solution will have at the equivalence point. If the acid and base are both strong, the pH at that point will be 7. However, for reasons discussed later, the pH is not always 7 at the equivalence point. The point at which the indicator changes color and the titration is stopped is called the titration **endpoint**. Figure 9.5 gives the colors shown by a number of indicators at various pH values.

Indicator papers (litmus paper, pH paper, etc.) are often used to make routine pH measurements. These papers are impregnated with one or more indicators that change to a variety of colors depending on the pH. However, indicators may not be practical under certain conditions. For example, there might not be any indicator that changes color close enough to the equivalence point, or the solution being titrated might be so highly colored that an indicator color change cannot be detected. Under such circumstances, a pH meter



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- 1 Methyl red goes from red at low pH to orange.



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- 2 Bromthymol blue from low pH to high pH.



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- 3 Phenolphthalein goes from colorless to pink.

Figure 9.5 Indicators change color with changes in pH (the numbers on the tubes). Would phenolphthalein be a useful indicator to differentiate between two solutions with pH values of 5 and 7? Explain.

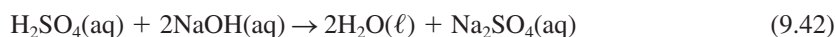
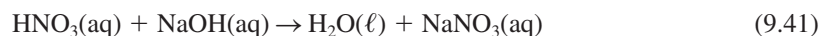
can be used. The electrodes of the meter are placed in the solution being titrated, and the pH is read directly (see ► Figure 9.6). The titration is continued until the meter reading matches the pH of the equivalence point.

9.11 Titration Calculations

Learning Objective

11. Do calculations related to the analysis of acids and bases by titration.

Equations 9.41 and 9.42 represent the reactions that occur when solutions of nitric and sulfuric acid are titrated with sodium hydroxide:



One mole of HNO_3 requires 1 mol of NaOH for a complete reaction, but 1 mol of H_2SO_4 requires 2 mol of NaOH . Because these are solution reactions, stoichiometric calculations can be done using the methods described in Section 7.6.



© Spencer L. Seager

- 1 At the beginning, the pH meter gives the pH of the acid solution being titrated.



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- 2 Partway through the titration, the pH meter reading is of a solution of unreacted acid and the salt produced by the reaction.



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- 3 At the end of the titration, the pH meter gives the pH of the salt solution formed by the complete reaction of acid with base.

Figure 9.6 An acid–base titration using a pH meter to detect the equivalence point.

Example 9.14

Calculate the molarity of the HNO_3 and H_2SO_4 solutions involved in the following titrations:

- A 25.0-mL sample of an HNO_3 solution requires the addition of 16.3 mL of 0.200 M NaOH to reach the equivalence point.
- A 25.0-mL sample of an H_2SO_4 solution requires the addition of 32.6 mL of 0.200 M NaOH to reach the equivalence point.

Solution

In each case we know the volume of acid solution reacted. If we also knew the number of moles of acid in the volume of reacted solution, we could calculate the solution molarity by using Equation 7.5:

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Therefore, our task will be to calculate the number of moles of acid reacted in each case. The methods described in Section 7.6 will be used. We will begin with the volume of NaOH solution used in each titration, and convert that volume to moles of acid using the factor-unit method.

a. Nitric acid

Step 1. 0.0163 L NaOH solution

Step 2. 0.0163 L NaOH solution = mol HNO_3

Step 3. $0.0163 \text{ L NaOH solution} \times \frac{0.200 \text{ mol NaOH}}{1 \text{ L NaOH solution}} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol NaOH}}$
= mol HNO_3

The factor $\frac{0.200 \text{ mol NaOH}}{1 \text{ L NaOH solution}}$ came from the molarity of the NaOH solution, and the factor $\frac{1 \text{ mol HNO}_3}{1 \text{ mol NaOH}}$ came from the 1 to 1 reaction ratio of HNO_3 with NaOH.

Step 4. $0.0163 \times \frac{0.200}{1} \times \frac{1 \text{ mol HNO}_3}{1} = 0.00326 \text{ mol HNO}_3$

This result tells us that the 25.0 mL titrated sample of HNO_3 contained 0.00326 moles of HNO_3 . After converting the HNO_3 volume to liters, the molarity can be calculated:

$$M = \frac{\text{moles of HNO}_3}{\text{liters of HNO}_3 \text{ solution}} = \frac{0.00326 \text{ mol HNO}_3}{0.0250 \text{ L HNO}_3 \text{ solution}} = \frac{0.130 \text{ mol HNO}_3}{\text{L HNO}_3 \text{ solution}} = 0.130 \text{ M HNO}_3$$

b. Sulfuric acid

This calculation follows the same pattern as the HNO_3 calculation, except for the fact that the reaction ratio of NaOH and H_2SO_4 is 2 mol of NaOH reacts with 1 mol of H_2SO_4 . This gives a factor of $\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}$ that will be used in step 3.

Step 1. 0.0326 L NaOH solution

Step 2. 0.0326 L NaOH solution = mol H_2SO_4

Step 3. $0.0326 \text{ L NaOH solution} \times \frac{0.200 \text{ mol NaOH}}{1 \text{ L NaOH solution}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}$
= mol H_2SO_4

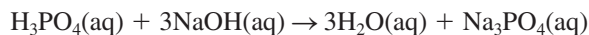
Step 4. $0.0326 \times \frac{0.200}{1} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2} = 0.00326 \text{ mol H}_2\text{SO}_4$

The molarity of the solution is calculated as before:

$$M = \frac{\text{moles of H}_2\text{SO}_4}{\text{liters of H}_2\text{SO}_4 \text{ solution}} = \frac{0.0326 \text{ mol H}_2\text{SO}_4}{0.0250 \text{ L H}_2\text{SO}_4 \text{ solution}} = \frac{0.130 \text{ mol H}_2\text{SO}_4}{\text{L H}_2\text{SO}_4 \text{ solution}} \\ = 0.130 \text{ M H}_2\text{SO}_4$$

Thus, we see that the two acid solutions have the same molar concentrations. Why did the second require twice the volume of NaOH solution for the titration?

► **Learning Check 9.18** Calculate the molarity of a solution of phosphoric acid (H_3PO_4) if a 25.0-mL sample of acid solution requires 14.1 mL of a 0.250 M NaOH solution to titrate to the equivalence point. The equation for the reaction is



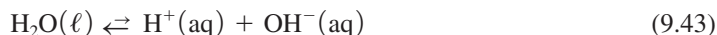
9.12 Hydrolysis Reactions of Salts

Learning Objective

12. Explain the concept of salt hydrolysis, and write equations to illustrate the concept.

In general, a **hydrolysis reaction** is a reaction with water. Many types of hydrolysis reactions are known, but at this point we will discuss only one, the hydrolysis of salts. In Section 9.10, we said that the pH at the equivalence point is not 7 for all acid–base titrations. However, we pointed out earlier that the only products of acid–base reactions are water and a salt. Therefore, it seems reasonable to conclude that some salts and water must interact (a hydrolysis reaction) and cause the solution pH to differ from that of pure water.

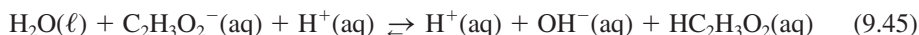
Suppose three solutions are prepared by dissolving equal molar amounts of sodium acetate, ammonium chloride, and sodium chloride in identical volumes of pure water. Measurement of pH shows that the sodium acetate solution is alkaline (pH higher than 7), the ammonium chloride solution is acidic (pH lower than 7), and the sodium chloride solution is neutral (pH = 7). The alkaline sodium acetate solution must contain more OH^- than H^+ ions. However, the only source of both ions in a solution of salt and water is the dissociation of water, so both should be present in equal amounts. The equation for the water dissociation is



The OH^- excess would result if something removed H^+ from this equilibrium, causing it to shift right and produce more OH^- . This is exactly what happens. The acetate anion is the conjugate base of a weak acid (acetic acid), and it is a strong enough Brønsted base to react with the H^+ ions of the above equilibrium:



The overall result of this reaction can be seen by adding Equations 9.43 and 9.44:

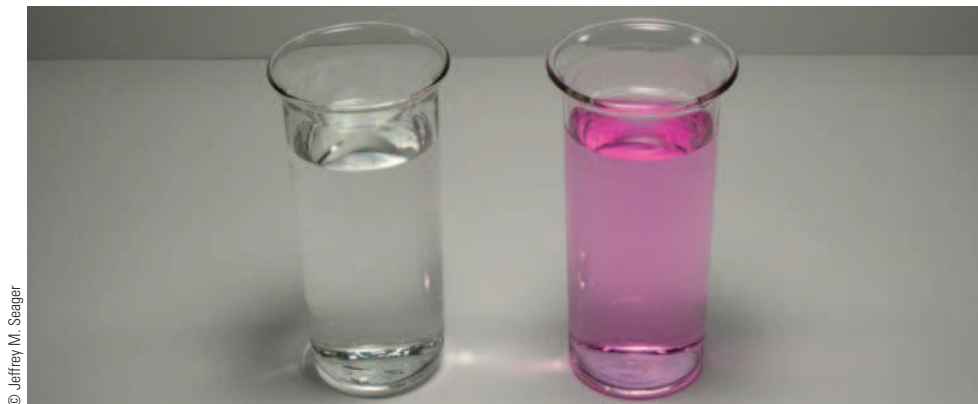


Now it can be seen that the excess OH^- comes from a reaction between the acetate ion and water. But what about the Na^+ ion that is also in the solution? Does it react with water? The Na^+ cation is the conjugate acid of the strong base NaOH. As a result, Na^+ is an extremely weak Brønsted acid and makes no contribution to the pH of the solution.

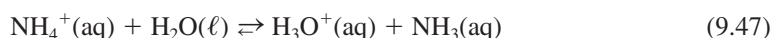
In general, salts influence the pH of water solutions as follows: (1) Salts containing an anion of a weak acid and a cation of a strong base will form alkaline water solutions. (2) Salts containing an anion of a strong acid and a cation of a weak base will form acidic solutions. (3) Salts containing an anion and a cation from equal-strength acids and bases (both weak or both strong) will form neutral solutions (see ► Figure 9.7).

hydrolysis reaction Any reaction with water. For salts it is a reaction of the acidic cation and/or basic anion of the salt with water.

Figure 9.7 Samples of pure water (left) and sodium acetate dissolved in water (right) behave differently when phenolphthalein indicator is added. The acetate ion hydrolyzes in water to form a basic solution that turns phenolphthalein to a pink color. Why didn't the Na^+ ion from the sodium acetate hydrolyze?



The acidic ammonium chloride solution is an example of the second category of salts, where the NH_4^+ cation is the conjugate acid of the weak base NH_3 , and the Cl^- anion is the conjugate base of the strong acid HCl . The hydrolysis reaction is given in Equation 9.47, where H_3O^+ is used to emphasize the reaction with water:



The neutral sodium chloride solution is an example of the third category. Both NaOH and HCl are strong, and an NaCl solution is neutral.

Learning Check 9.19 Predict the relative pH value (higher than 7, lower than 7, etc.) for water solutions of the following salts. Table 9.7 may be useful.

- Sodium nitrate (NaNO_3)
- Sodium nitrite (NaNO_2)
- Potassium borate (K_3BO_3)

9.13 Buffers

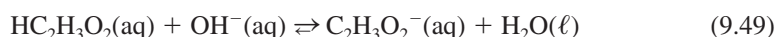
Learning Objective

- 13.** Explain how buffers work, and write equations to illustrate their action.

Buffers are solutions that have the ability to resist changes in pH when acids or bases are added (see Figure 9.8). Most buffers consist of a pair of compounds, one with the ability to react with H^+ and the other with the ability to react with OH^- . An example is a mixture of acetic acid and its salt, sodium acetate. The acetate ion from sodium acetate is the conjugate base of acetic acid and reacts with any added acid (H^+ ions):

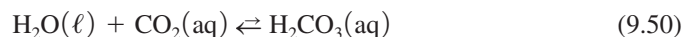


Any added base (OH^- ions) reacts with nonionized acetic acid molecules:



In this way, the buffer solution resists changes in pH. The amount of H^+ or OH^- that a buffer system can absorb without allowing significant pH changes to occur is called its **buffer capacity**.

Buffered solutions are found extensively in the body, where they serve to protect us from the disastrous effects of large deviations in pH. An important buffer system in blood is composed of carbonic acid (H_2CO_3) and bicarbonate salts such as NaHCO_3 . The unstable carbonic acid results when dissolved CO_2 reacts with water in the blood:

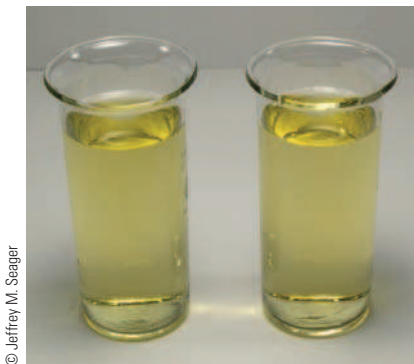


Added H^+ is neutralized by reacting with HCO_3^- from the dissolved bicarbonate salts:



buffer A solution with the ability to resist changing pH when acids (H^+) or bases (OH^-) are added.

buffer capacity The amount of acid (H^+) or base (OH^-) that can be absorbed by a buffer without causing a significant change in pH.



1

The solution on the left is not buffered; the one on the right is; universal indicator has been added to each solution.



2

Sodium hydroxide has been added to each solution.

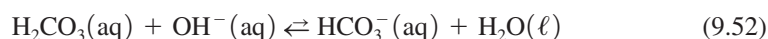


3

Hydrochloric acid has been added to two fresh samples that originally looked like the first pair of samples.

Figure 9.8 Buffered solutions resist changing pH when bases and acids are added. What observation of this experiment indicates that the pH of the buffered solutions did not change significantly when base or acid was added?

The carbonic acid protects against added OH^- :



If large amounts of H^+ or OH^- are added to a buffer, the buffer capacity can be exceeded, the buffer system is overwhelmed, and the pH changes. For example, if large amounts of H^+ were added to the bicarbonate ion–carbonic acid buffer, Reaction 9.51 would take place until the HCO_3^- was depleted. The pH would then drop as additional H^+ ions were added. In blood, the concentration of HCO_3^- is ten times the concentration of H_2CO_3 . Thus, this buffer has a greater capacity against added acid than against bases. This is consistent with the normal functions of the body that cause larger amounts of acidic than basic substances to enter the blood.

The pH of buffers can be calculated by using a form of Equation 9.34 that was given earlier and is repeated here:

$$K_a = \frac{[\text{H}^+][\text{B}^-]}{[\text{HB}]}$$

This equation can be rearranged to give

$$[\text{H}^+] = K_a \frac{[\text{HB}]}{[\text{B}^-]} \quad (9.53)$$

Application of the logarithm concept to Equation 9.53 gives

$$\text{pH} = \text{p}K_a + \log \frac{[\text{B}^-]}{[\text{HB}]} \quad (9.54)$$

In this equation, $\text{pH} = -\log[\text{H}^+]$ (Equation 9.11), and $\text{p}K_a = -\log K_a$. Equation 9.54 is known as the **Henderson–Hasselbalch equation**; it is often used by biologists, biochemists, and others who frequently work with buffers. We see from Equation 9.54 that when the concentrations of a weak acid and its conjugate base (anion) are equal in a solution, the pH of the solution is equal to $\text{p}K_a$. When it is desired to produce a buffer with a pH different from the exact $\text{p}K_a$, Equation 9.54 indicates it can be done. An acid with a $\text{p}K_a$ near the desired pH is selected, and the ratio of the concentrations of conjugate base (anion of the acid) and acid is adjusted to give the desired pH. When a weak acid and its salt (the source of the acid's conjugate base anion) are mixed in solution, you can generally assume that the amount of acid that dissociates is small and may be neglected. This means that the buffer concentrations of the acid and anion (the salt of acid) are “equal” to the made-up concentrations. Table 9.9 lists some weak acids, together with values of K_a and $\text{p}K_a$.

pKa The negative logarithm of K_a .

Henderson–Hasselbalch equation
A relationship between the pH of a buffer, $\text{p}K_a$, and the concentrations of acid and salt in the buffer.

Table 9.9 K_a and pK_a Values for Selected Weak Acids

Name	Formula	K_a	pK_a
Acetic acid ^a	CH_3COOH	1.8×10^{-5}	4.74
Ammonium ion	NH_4^+	5.6×10^{-10}	9.25
Boric acid	H_3BO_3	7.3×10^{-10}	9.14
Dihydrogen borate ion	H_2BO_3^-	1.8×10^{-13}	12.75
Hydrogen borate ion	HBO_3^{2-}	1.6×10^{-14}	13.80
Carbonic acid	H_2CO_3	4.3×10^{-7}	6.37
Bicarbonate ion	HCO_3^-	5.6×10^{-11}	10.25
Citric acid ^a	$\text{C}_3\text{H}_4(\text{OH})(\text{COOH})_3$	8.4×10^{-4}	3.08
Dihydrogen citrate ion ^a	$\text{C}_3\text{H}_4(\text{OH})(\text{COOH})_2\text{COO}^-$	1.8×10^{-5}	4.74
Hydrogen citrate ion ^a	$\text{C}_3\text{H}_4(\text{OH})(\text{COOH})(\text{COO})_2^{2-}$	4.0×10^{-6}	5.40
Formic acid ^a	HCOOH	1.8×10^{-4}	3.74
Lactic acid ^a	$\text{C}_2\text{H}_4(\text{OH})\text{COOH}$	1.4×10^{-4}	3.85
Nitrous acid	HNO_2	4.6×10^{-4}	3.33
Phosphoric acid	H_3PO_4	7.5×10^{-3}	2.12
Dihydrogen phosphate ion	H_2PO_4^-	6.2×10^{-8}	7.21
Hydrogen phosphate ion	HPO_4^{2-}	2.2×10^{-13}	12.66
Sulfurous acid	H_2SO_3	1.5×10^{-2}	1.82
Bisulfite ion	HSO_3^-	1.0×10^{-7}	7.00

^aThe hydrogen that ionizes in organic acids and ions is a part of a carboxylic acid group, represented by COOH .

Example 9.15

- Calculate the pH of a buffer solution that contains 0.10 mol acetic acid (CH_3COOH) and 0.10 mol sodium acetate (CH_3COONa) per liter.
- What is the pH of a buffer in which the concentration of NaH_2PO_4 is 0.10 M and that of Na_2HPO_4 is 0.50 M?
- A buffer system consisting of the ions H_2PO_4^- and HPO_4^{2-} helps control the pH of urine. What value of the ratio $[\text{HPO}_4^{2-}]/[\text{H}_2\text{PO}_4^-]$ would be required to maintain a pH of 6.00, the average value for normal urine?

Solution

- The acetic acid concentration is 0.10 M and that of its conjugate base, the acetate ion from the sodium acetate salt, is also 0.10 M. The pK_a for acetic acid is 4.74:

$$\begin{aligned}\text{pH} &= pK_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 4.74 + \log \left(\frac{0.10 \text{ mol/L}}{0.10 \text{ mol/L}} \right) \\ &= 4.74 + \log(1) = 4.74 + 0 = 4.74\end{aligned}$$

We see that, as mentioned earlier, when the acid and conjugate base concentrations are equal, $\text{pH} = pK_a$.

- In this case, both compounds are salts that produce the H_2PO_4^- and HPO_4^{2-} ions. In such situations, the ion with more hydrogens will be the acid, and the one with fewer will be the anion or conjugate base. Therefore, the acid is H_2PO_4^- at a concentration of

Heartburn Remedies: Something Old, Something New



Every day, the average person's stomach produces about a quart of hydrochloric acid as a part of the digestive gastric juices. The resulting 0.1 M solution of HCl does not cause any significant problems for someone with a healthy, normally functioning stomach. However, it causes extreme discomfort for anyone suffering from a peptic ulcer or a gastroesophageal reflux condition. This latter condition is characterized by a tendency for the acidic contents of the stomach to be forced back up into the esophagus, causing what is commonly called *heartburn* or, in extreme cases, *gastroesophageal reflux disease (GERD)*.

For many years the traditional nonprescription treatment for heartburn has been antacids, which are stocked by most drugstores in a variety of forms. Some of them fizz when placed in water, others relieve headache as well as heartburn, several will relieve gas pains, and still others provide a laxative action to clean out the intestinal tract. They come in various forms and flavors. Some are soft and chewy, others are crunchy, and at least one is a thick liquid that "coats" a distressed stomach.

Regardless of their form, taste, or added side effects, most of these conventional antacids act on stomach acid by either neutralizing it with a hydroxide-containing base or reacting it with a carbonate or bicarbonate. A quick survey of 12 antacid products resulted in the list of active ingredients below. The number in parentheses is the number of different antacids that list it as an ingredient. Some products contained more than one active ingredient, so the numbers in parentheses will add up to more than 12.

- Sodium bicarbonate (5), NaHCO_3
- Potassium bicarbonate (1), KHCO_3
- Calcium carbonate (4), CaCO_3
- Magnesium carbonate (2), MgCO_3
- Magnesium hydroxide (4), $\text{Mg}(\text{OH})_2$
- Aluminum hydroxide (2), $\text{Al}(\text{OH})_3$

For years, compounds known collectively as H2 blockers had been available by prescription for use as maintenance therapy for peptic ulcers. H2 blockers are named for the fact that they block histamine signals in the body that cause the stomach to produce acid. However, the discovery that most ulcers are caused by the bacterium *Helicobacter pylori* and can be completely cured with antibiotics essentially eliminated this use for H2 blockers. The manufacturers quickly responded by making their products available in OTC strengths to compete with the traditional antacids. These new products have been heavily advertised in the media under such names as Pepcid AC[®], Zantac 75[®], Tagamet HB[®], and Axid AR[®]. A disadvantage of H2 blockers is that they must be taken up to an hour before a meal in order to be effective, while traditional antacids may be taken when symptoms first appear. However, H2 blockers are effective for several hours, whereas the effects of regular antacids may wear off in as little as a few minutes. Another advantage of the traditional antacids that use CaCO_3 as an ingredient is that they provide a source of calcium for the body.



Antacids neutralize stomach acid; H2 blockers limit its production.

0.10 M, and the conjugate base or anion is HPO_4^{2-} at a concentration of 0.50 M. The $\text{p}K_a$ for H_2PO_4^- is 7.21:

$$\begin{aligned}\text{pH} &= \text{p}K_a + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} = 7.21 + \log \left(\frac{0.50 \text{ mol/L}}{0.10 \text{ mol/L}} \right) \\ &= 7.21 + \log(5.00) = 7.21 + 0.70 = 7.91\end{aligned}$$

- c. As in part (b), the acid is H_2PO_4^- and the conjugate base is HPO_4^{2-} . The desired pH is 6.00. Substitution into Equation 9.54 gives

$$\begin{aligned}\text{pH} &= \text{p}K_a + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \\ 6.00 &= 7.21 + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \\ 6.00 - 7.21 &= \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \\ -1.21 &= \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}\end{aligned}$$

At this point, we know the value of the log of the desired ratio. To get the ratio, we must evaluate the antilog of -1.21 . Use Step 3 given in Table 9.3 to get 0.062 as the value of

the ratio. This result says the ratio $[\text{HPO}_4^{2-}]/[\text{H}_2\text{PO}_4^-]$ must be 0.062. In other words, the HPO_4^{2-} concentration must be only about 0.06 times the value of the H_2PO_4^- concentration. Any concentrations giving this ratio would be satisfactory. For example, if $[\text{H}_2\text{PO}_4^-] = 0.50 \text{ M}$, then $[\text{HPO}_4^{2-}]$ would have to equal $0.062 \times 0.5 \text{ M}$, or 0.031 M .

Learning Check 9.20

- What is the pH of a buffer solution in which formic acid (HCOOH) and sodium formate (HCOONa) are both at a concentration of 0.22 M ?
- What is the pH of a buffer solution that is 0.25 M in sulfurous acid (H_2SO_3) and 0.10 M in sodium bisulfite (NaHSO_3)?
- A buffer system in the blood consists of the bicarbonate ion, HCO_3^- , and carbonic acid, H_2CO_3 . What value of the ratio $[\text{HCO}_3^-]/[\text{H}_2\text{CO}_3]$ would be required to maintain blood at the average normal pH of 7.40?

Concept Summary

The Arrhenius Theory. Svante Arrhenius defined acids as substances that dissociate in water to provide hydrogen ions (H^+), and bases as substances that dissociate in water to provide hydroxide ions, (OH^-).

Objective 1, Exercise 9.2

The Brønsted Theory. Johannes Brønsted and Thomas Lowry proposed a theory in which acids are defined as any hydrogen-containing substances capable of donating protons to other substances. Bases are substances that accept and form covalent bonds with protons. When a substance behaves as a Brønsted acid by donating a proton, the substance becomes a conjugate base.

Objective 2, Exercises 9.6 and 9.10

Naming Acids. Two types of acids are named differently. Water solutions of binary covalent compounds containing hydrogen and a nonmetal are named following the pattern hydro(stem)ic acid, where (stem) is the stem of the name of the nonmetal bonded to hydrogen. Acids in which hydrogen is bonded to polyatomic ions have names based on the name of the polyatomic ion to which hydrogen is bonded.

Objective 3, Exercise 9.22

The Self-Ionization of Water. Water, a hydrogen-containing material, is able to behave as both a Brønsted acid and a Brønsted base. In pure water, a small number of water molecules (10^{-7} mol/L) donate protons to other water molecules.

Objective 4, Exercises 9.28a & b and 9.30a & b

The pH Concept. The pH is the negative logarithm of the molar H^+ concentration of a solution. Solutions with pH values lower than 7 are acidic, those with pH values higher than 7 are basic or alkaline, and those with a pH value of 7 are neutral.

Objective 5, Exercises 9.36 and 9.40

Properties of Acids. All acids have certain characteristic properties that include (1) a sour taste; (2) a reaction with water to produce H_3O^+ ; (3) reactions with solid metallic oxides, hydroxides, carbonates, and bicarbonates; and (4) a reaction with certain metals to give hydrogen gas.

Objective 6, Exercise 9.50

Properties of Bases. Basic solutions feel soapy or slippery and change the color of litmus from red to blue. Their most characteristic chemical property is a neutralization reaction with acids to produce water and a salt.

Objective 7, Exercise 9.60

Salts. At room temperature, salts are solid crystalline substances that contain the cation of a base and the anion of an acid. Hydrated salts contain specific numbers of water molecules as a part of their crystalline structures. Salts can be prepared by reacting an appropriate acid with one of a number of other materials.

Objective 8, Exercises 9.68 and 9.74

The Strengths of Acids and Bases. Acids and bases that dissociate completely when dissolved in solution are called strong, and those that do not are called weak or moderately weak, depending on the degree of dissociation they undergo. Acid strength is indicated by the value of K_a , the acid dissociation constant. In general, polyprotic acids become weaker as an acid during each successive dissociation reaction. The Brønsted base produced by the dissociation of an acid has a strength opposite that of the acid. The acid strengths of cations produced by the dissociation of bases follow a similar pattern.

Objective 9, Exercise 9.86

Analyzing Acids and Bases. The neutralization reaction of acids and bases is used in a process called titration to analyze acids and bases. During a typical titration, a base solution of known concentration is added slowly to an acid solution of unknown concentration. The titration is stopped at the endpoint when a color change occurs in an indicator. The volumes of acid and base required are used to calculate the acid concentration.

Objective 10, Exercise 9.92

Titration Calculations. Titrations are used to determine the total amount of acid or base in solutions. Data collected from titrations are treated like other stoichiometric data in calculations.

Objective 11, Exercises 9.98 and 9.100a

Hydrolysis Reactions of Salts. The cation and anion of a salt may have the same or different Brønsted acid and base strengths. When they are the same, solutions of the salt have a pH of 7. When they are different, salt solutions have H^+ and OH^- concentrations that are unequal and pH values different than 7. The reactions of salt ions with water that cause these results are called hydrolysis reactions.

Objective 12, Exercise 9.108

Buffers. Solutions with the ability to maintain essentially constant pH values when acid (H^+) or base (OH^-) are added are called buffers. All buffers have a limit to the amount of acid or base they can absorb without changing pH. This limit is called the buffer capacity.

Objective 13, Exercise 9.116


Key Terms and Concepts

Acid dissociation constant (9.9)	Conjugate acid–base pair (9.2)	Neutral (9.4)
Acidic solution (9.4)	Conjugate base (9.2)	Neutralization reaction (9.7)
Activity series (9.6)	Diprotic acid (9.9)	pH (9.5)
Anion (9.8)	Endpoint of a titration (9.10)	pK_a (9.13)
Arrhenius acid (9.1)	Equivalence point of a titration (9.10)	Salt (9.8)
Arrhenius base (9.1)	Equivalent of salt (9.8)	Strong acids and strong bases (9.9)
Basic or alkaline solution (9.4)	Henderson–Hasselbalch equation (9.13)	Titration (9.10)
Brønsted acid (9.2)	Hydrate (9.8)	Triprotic acid (9.9)
Brønsted base (9.2)	Hydrolysis reaction (9.12)	Water of hydration (9.8)
Buffer (9.13)	Ion product of water (9.4)	Weak (or moderately weak) acids and bases (9.9)
Buffer capacity (9.13)	Monoprotic acid (9.9)	
Cation (9.8)		

Key Equations

1. Relationship between $[H^+]$ and $[OH^-]$ in water solutions (Section 9.4):	$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ (mol/L)}^2$	Equation 9.10
2. Relationships between pH and $[H^+]$ (Section 9.5):	$pH = -\log[H^+]$	Equation 9.11
	$[H^+] = 1 \times 10^{-pH}$	Equation 9.12
3. General reactions of acids (Section 9.8):	acid + metal \rightarrow salt + H_2	Equation 9.26
	acid + metal oxide \rightarrow salt + H_2O	Equation 9.27
	acid + metal hydroxide \rightarrow salt + H_2O	Equation 9.28
	acid + metal carbonate \rightarrow salt + H_2O + CO_2	Equation 9.29
	acid + metal bicarbonate \rightarrow salt + H_2O + CO_2	Equation 9.30
4. Dissociation constant for acids (Section 9.9):	$K_a = \frac{[H^+][B^-]}{[HB]}$	Equation 9.34
5. Calculation of buffer pH, Henderson–Hasselbalch equation (Section 9.13):	$pH = pK_a + \log \frac{[B^-]}{[HB]}$	Equation 9.54

Exercises

 **OWL** Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

The Arrhenius Theory (Section 9.1)

9.1 Write the dissociation equations for the following that emphasize their behavior as Arrhenius acids:

- HI
- HBrO

c. HCN

d. $HClO_2$

9.2 Write the dissociation equations for the following that emphasize their behavior as Arrhenius acids:

- HF
- $HClO_3$
- H_3BO_3 (show only 1st H)
- HSe^-

- 9.3** Each of the following produces a basic solution when dissolved in water. Identify those that behave as Arrhenius bases and write dissociation equations to illustrate that behavior.
- CsOH
 - CH_3NH_2
 - NH_3
 - $\text{Ca}(\text{OH})_2$
- 9.4** Each of the following produces a basic solution when dissolved in water. Identify those that behave as Arrhenius bases and write dissociation equations to illustrate that behavior.
- NaNH_2
 - RbOH
 - $\text{C}_3\text{H}_7\text{NH}_2$
 - $\text{Ba}(\text{OH})_2$

The Brønsted Theory (Section 9.2)

- 9.5** Identify each Brønsted acid and base in the following equations. Note that the reactions are assumed to be reversible.
- $\text{HBr}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Br}^-(\text{aq})$
 - $\text{H}_2\text{O}(\ell) + \text{N}_3^-(\text{aq}) \rightleftharpoons \text{HN}_3(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{H}_2\text{S}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HS}^-(\text{aq})$
 - $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HSO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{HCN}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CN}^-(\text{aq})$
- 9.6** Identify each Brønsted acid and base in the following equations. Note that the reactions are assumed to be reversible.
- $\text{OCl}^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HOCl}(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{H}_2\text{C}_2\text{O}_4(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HC}_2\text{O}_4^-(\text{aq})$
 - $\text{HPO}_4^{2-}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{PO}_4^{3-}(\text{aq})$
 - $\text{C}_2\text{O}_4^{2-}(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HC}_2\text{O}_4^-(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{H}_3\text{AsO}_4(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{H}_2\text{AsO}_4^-(\text{aq})$
- 9.7** Identify each conjugate acid–base pair in the equations of Exercise 9.5.
- 9.8** Identify each conjugate acid–base pair in the equations of Exercise 9.6.
- 9.9** Write equations to represent the Brønsted acid behavior for each of the following acids in water solution. Remember to represent the reactions as being reversible.
- HI
 - HBrO
 - HCN
 - HSe^-
- 9.10** Write equations to represent the Brønsted acid behavior for each of the following acids in water solution. Remember to represent the reactions as being reversible.
- HF
 - HClO_3
 - HClO
 - HS^-

- 9.11** Write a formula for the conjugate base formed when each of the following behaves as a Brønsted acid:
- HSO_3^-
 - HPO_4^{2-}
 - HClO_3
 - CH_3NH_3^+
 - $\text{H}_2\text{C}_2\text{O}_4$
- 9.12** Write a formula for the conjugate base formed when each of the following behaves as a Brønsted acid:
- H_2BO_3^-
 - $\text{C}_6\text{H}_5\text{NH}_3^+$
 - HS^-
 - HC_2O_4^-
 - HClO_4
- 9.13** Write a formula for the conjugate acid formed when each of the following behaves as a Brønsted base:
- NH_2^-
 - CO_3^{2-}
 - OH^-
 - $(\text{CH}_3)_2\text{NH}$
 - NO_2^-
- 9.14** Write a formula for the conjugate acid formed when each of the following behaves as a Brønsted base:
- $\text{C}_6\text{H}_5\text{NH}_2$
 - $\text{S}_2\text{O}_3^{2-}$
 - CN^-
 - HAsO_4^{2-}
 - F^-
- 9.15** The following reactions illustrate Brønsted acid–base behavior. Complete each equation.
- $\text{HI}(\text{aq}) + ? \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{I}^-(\text{aq})$
 - $\text{NH}_3(\ell) + ? \rightarrow \text{NH}_4^+ + \text{NH}_2^-$
 - $\text{H}_2\text{C}_2\text{O}_4(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow ? + \text{HC}_2\text{O}_4^-(\text{aq})$
 - $\text{H}_2\text{N}_2\text{O}_2(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + ?$
 - $? + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$
- 9.16** The following reactions illustrate Brønsted acid–base behavior. Complete each equation.
- $\text{H}_2\text{AsO}_4^-(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + ?$
 - $? + \text{H}_2\text{O}(\ell) \rightarrow \text{C}_6\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$
 - $\text{S}^{2-}(\text{aq}) + ? \rightarrow \text{HS}^-(\text{aq}) + \text{OH}^-(\text{aq})$
 - $? + \text{HBr}(\text{aq}) \rightarrow (\text{CH}_3)_2\text{NH}_2^+(\text{aq}) + \text{Br}^-(\text{aq})$
 - $\text{CH}_3\text{NH}_2(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow ? + \text{Cl}^-(\text{aq})$

- 9.17** Write equations to illustrate the acid–base reaction when each of the following pairs of Brønsted acids and bases are combined:

<i>Acid</i>	<i>Base</i>
a. HOCl	H ₂ O
b. HClO ₄	NH ₃
c. H ₂ O	NH ₂ [−]
d. H ₂ O	OCI [−]
e. HC ₂ O ₄ [−]	H ₂ O

- 9.18** Write equations to illustrate the acid–base reaction when each of the following pairs of Brønsted acids and bases are combined:

<i>Acid</i>	<i>Base</i>
a. H ₃ O ⁺	NH ₂ [−]
b. H ₂ PO ₄ [−]	NH ₃
c. HS ₂ O ₃ [−]	OCI [−]
d. H ₂ O	ClO ₄ [−]
e. H ₂ O	NH ₃

Naming Acids (Section 9.3)

- 9.19** A water solution of HF gas is used to etch glass. Name the water solution as an acid.
- 9.20** Hydrogen cyanide, HCN, behaves in a water solution very much like the binary covalent compounds of hydrogen, but it liberates the cyanide ion, CN[−]. Name the acidic water solution by following the rules for binary covalent compounds of hydrogen.
- 9.21** Name the following acids. Refer to Table 4.7 as needed.
- H₂Se(aq)
 - HClO₃
 - H₂SO₄
 - HNO₃
- 9.22** Name the following acids. Refer to Table 4.7 as needed.
- H₂Te(aq)
 - HClO
 - H₂SO₃
 - HNO₂
- 9.23** The acid H₃C₆H₅O₇ forms the citrate ion, C₆H₅O₇^{3−}, when all three hydrogens are removed. This acid is involved in an important energy-storing process in the body. Name the acid.
- 9.24** The acid H₂C₄H₄O₄ forms the succinate ion, C₄H₄O₄^{2−}, when both hydrogens are removed. This acid is involved in the same energy-storing process as the acid of Exercise 9.23. Name H₂C₄H₄O₄ as an acid.
- 9.25** Refer to Table 4.7, and write the formula for chromic acid.
- 9.26** Refer to Table 4.7, and write the formula for permanganic acid.

The Self-Ionization of Water (Section 9.4)

- 9.27** Calculate the molar concentration of OH[−] in water solutions with the following H₃O⁺ molar concentrations:
- 1.0×10^{-7}
 - 3.2×10^{-3}
 - 4.7×10^{-11}

- 1.2
- 0.043

- 9.28** Calculate the molar concentration of OH[−] in water solutions with the following H₃O⁺ molar concentrations:

- 1.2×10^{-5}
- 0.27
- 0.031
- 3.6×10^{-9}
- 5.3×10^{-2}

- 9.29** Calculate the molar concentration of H₃O⁺ in water solutions with the following OH[−] molar concentrations:

- 1.0×10^{-7}
- 5.2×10^{-4}
- 9.9×10^{-10}
- 0.092
- 3.7

- 9.30** Calculate the molar concentrations of H₃O⁺ in water solutions with the following OH[−] molar concentrations:

- 0.0071
- 4.2×10^{-4}
- 2.8
- 7.9×10^{-10}
- 9.1×10^{-6}

- 9.31** Classify the solutions represented in Exercises 9.27 and 9.29 as acidic, basic, or neutral.

- 9.32** Classify the solutions represented in Exercises 9.28 and 9.30 as acidic, basic, or neutral.

The pH Concept (Section 9.5)

- 9.33** Classify solutions with the following characteristics as acidic, basic, or neutral:

- pH = 10
- pH = 4
- pH = 7.3
- pH = 6

- 9.34** Classify solutions with the following characteristics as acidic, basic, or neutral:

- pH = 2.8
- pH = 8
- pH = 6.9
- pH = 12

- 9.35** Determine the pH of water solutions with the following characteristics. Classify each solution as acidic, basic, or neutral.

- $[H^+] = 1.0 \times 10^{-5}$
- $[OH^-] = 6.0 \times 10^{-3}$
- $[H^+] = [OH^-]$
- $[H^+] = 9.0 \times 10^{-4}$
- $[OH^-] = 3.0 \times 10^{-9}$

Heartburn Remedies: Something Old, Something New



Every day, the average person's stomach produces about a quart of hydrochloric acid as a part of the digestive gastric juices. The resulting 0.1 M solution of HCl does not cause any significant problems for someone with a healthy, normally functioning stomach. However, it causes extreme discomfort for anyone suffering from a peptic ulcer or a gastroesophageal reflux condition. This latter condition is characterized by a tendency for the acidic contents of the stomach to be forced back up into the esophagus, causing what is commonly called *heartburn* or, in extreme cases, *gastroesophageal reflux disease (GERD)*.

For many years the traditional nonprescription treatment for heartburn has been antacids, which are stocked by most drugstores in a variety of forms. Some of them fizz when placed in water, others relieve headache as well as heartburn, several will relieve gas pains, and still others provide a laxative action to clean out the intestinal tract. They come in various forms and flavors. Some are soft and chewy, others are crunchy, and at least one is a thick liquid that "coats" a distressed stomach.

Regardless of their form, taste, or added side effects, most of these conventional antacids act on stomach acid by either neutralizing it with a hydroxide-containing base or reacting it with a carbonate or bicarbonate. A quick survey of 12 antacid products resulted in the list of active ingredients below. The number in parentheses is the number of different antacids that list it as an ingredient. Some products contained more than one active ingredient, so the numbers in parentheses will add up to more than 12.

- Sodium bicarbonate (5), NaHCO_3
- Potassium bicarbonate (1), KHCO_3
- Calcium carbonate (4), CaCO_3
- Magnesium carbonate (2), MgCO_3
- Magnesium hydroxide (4), $\text{Mg}(\text{OH})_2$
- Aluminum hydroxide (2), $\text{Al}(\text{OH})_3$

For years, compounds known collectively as H2 blockers had been available by prescription for use as maintenance therapy for peptic ulcers. H2 blockers are named for the fact that they block histamine signals in the body that cause the stomach to produce acid. However, the discovery that most ulcers are caused by the bacterium *Helicobacter pylori* and can be completely cured with antibiotics essentially eliminated this use for H2 blockers. The manufacturers quickly responded by making their products available in OTC strengths to compete with the traditional antacids. These new products have been heavily advertised in the media under such names as Pepcid AC[®], Zantac 75[®], Tagamet HB[®], and Axid AR[®]. A disadvantage of H2 blockers is that they must be taken up to an hour before a meal in order to be effective, while traditional antacids may be taken when symptoms first appear. However, H2 blockers are effective for several hours, whereas the effects of regular antacids may wear off in as little as a few minutes. Another advantage of the traditional antacids that use CaCO_3 as an ingredient is that they provide a source of calcium for the body.



Antacids neutralize stomach acid; H2 blockers limit its production.

0.10 M, and the conjugate base or anion is HPO_4^{2-} at a concentration of 0.50 M. The $\text{p}K_a$ for H_2PO_4^- is 7.21:

$$\begin{aligned}\text{pH} &= \text{p}K_a + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} = 7.21 + \log \left(\frac{0.50 \text{ mol/L}}{0.10 \text{ mol/L}} \right) \\ &= 7.21 + \log(5.00) = 7.21 + 0.70 = 7.91\end{aligned}$$

- c. As in part (b), the acid is H_2PO_4^- and the conjugate base is HPO_4^{2-} . The desired pH is 6.00. Substitution into Equation 9.54 gives

$$\begin{aligned}\text{pH} &= \text{p}K_a + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \\ 6.00 &= 7.21 + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \\ 6.00 - 7.21 &= \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]} \\ -1.21 &= \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}\end{aligned}$$

At this point, we know the value of the log of the desired ratio. To get the ratio, we must evaluate the antilog of -1.21 . Use Step 3 given in Table 9.3 to get 0.062 as the value of

- 9.50** Write balanced molecular equations to illustrate the following characteristic reactions of acids, using sulfuric acid (H_2SO_4).
- Reaction with water to form hydronium ions
 - Reaction with the solid oxide, CaO
 - Reaction with the solid hydroxide, $\text{Mg}(\text{OH})_2$
 - Reaction with the solid carbonate, CuCO_3
 - Reaction with the solid bicarbonate, KHCO_3
 - Reaction with Mg metal
- 9.51** Write each molecular equation of Exercise 9.49 in total ionic and net ionic form. Use Table 7.4 to decide which products will be soluble.
- 9.52** Write each molecular equation of Exercise 9.50 in total ionic and net ionic form. Use Table 7.4 to decide which products will be soluble.
- 9.53** Write balanced molecular equations to illustrate five different reactions that could be used to prepare BaCl_2 from hydrochloric acid (HCl) and other appropriate substances.
- 9.54** Write balanced molecular equations to illustrate five different reactions that could be used to prepare SrCl_2 from hydrochloric acid (HCl) and other appropriate substances.
- 9.55** Write balanced molecular, total ionic, and net ionic equations to illustrate each of the following reactions. All the metals form $2+$ ions.
- Zinc with H_2SO_4
 - Magnesium with HCl
 - Calcium with $\text{HC}_2\text{H}_3\text{O}_2$
- 9.56** Write balanced molecular, total ionic, and net ionic equations to illustrate each of the following reactions. All the metals form $2+$ ions.
- Tin with H_2SO_3
 - Magnesium with H_3PO_4
 - Calcium with HBr

Properties of Bases (Section 9.7)

- 9.57** Write balanced molecular, total ionic, and net ionic equations to represent neutralization reactions between RbOH and the following acids. Use all $\text{H}'\text{s}$ possible for each acid.
- HCl
 - HNO_3
 - H_2SO_4
- 9.58** Write balanced molecular, total ionic, and net ionic equations to represent neutralization reactions between RbOH and the following acids. Use all $\text{H}'\text{s}$ possible for each acid.
- H_3PO_4
 - $\text{H}_2\text{C}_2\text{O}_4$ (oxalic acid)
 - $\text{HC}_2\text{H}_3\text{O}_2$
- 9.59** Some polyprotic acids can form more than one salt, depending on the number of $\text{H}'\text{s}$ that react with base. Write balanced molecular, total ionic, and net ionic equations to represent the following neutralization reactions between KOH and
- H_2SO_4 (only react one H)
 - H_2SO_4 (react both $\text{H}'\text{s}$)
 - H_3PO_4 (only react one H)
- 9.60** Some polyprotic acids can form more than one salt, depending on the number of $\text{H}'\text{s}$ that react with base. Write balanced molecular, total ionic, and net ionic equations to represent the following neutralization reactions between KOH and
- H_3PO_4 (react two $\text{H}'\text{s}$)
 - H_3PO_4 (react three $\text{H}'\text{s}$)
 - $\text{H}_2\text{C}_2\text{O}_4$ (react one H)

Salts (Section 9.8)

- 9.61** Identify with ionic formulas the cations and anions of the following salts:
- LiCl
 - $\text{Cu}(\text{NO}_3)_2$
 - SrSO_4
 - K_3PO_4
 - K_2HPO_4
 - CaCO_3
- 9.62** Identify with ionic formulas the cations and anions of the following salts:
- NH_4NO_3
 - CaCl_2
 - $\text{Mg}(\text{HCO}_3)_2$
 - $\text{KC}_2\text{H}_3\text{O}_2$
 - LiHSO_3
- 9.63** Identify with formulas the acid and base from which the anion and cation of each salt in Exercise 9.61 was derived. Pay special attention to salts derived from polyprotic acids and be sure to list the acid formula with all $\text{H}'\text{s}$.
- 9.64** Identify with formulas the acid and base from which the anion and cation of each salt in Exercise 9.62 was derived. Pay special attention to salts derived from polyprotic acids and be sure to list the acid formula with all $\text{H}'\text{s}$.
- 9.65** Calculate the mass of water that would be released if the water of hydration were completely driven off 1.0 mol of (a) plaster of Paris and (b) gypsum (see Table 9.6). How would the products of these reactions compare?
- 9.66** Calculate the mass of water that would be released if the water of hydration were completely driven off 1.0 mol of (a) Epsom salts and (b) borax (see Table 9.6). How would the products of these reactions compare?
- 9.67** Write formulas for the acid and indicated solid that could be used to prepare each of the following salts:
- CuCl_2 (solid is an oxide)
 - MgSO_4 (solid is a carbonate)
 - LiBr (solid is a hydroxide)
- 9.68** Write formulas for the acid and indicated solid that could be used to prepare each of the following salts:
- KNO_3 (solid is a bicarbonate)
 - ZnCl_2 (solid is a metal)
 - LiBr (solid is an oxide)

- 9.69** Write balanced molecular equations to illustrate each salt preparation described in Exercise 9.67.
- 9.70** Write balanced molecular equations to illustrate each salt preparation described in Exercise 9.68.
- 9.71** Determine the number of moles of each of the following salts that would equal 1 eq of salt:
- KNO_3
 - Li_2CO_3
 - SrCl_2
- 9.72** Determine the number of moles of each of the following salts that would equal 1 eq of salt:
- MgCO_3
 - $\text{Zn}(\text{HCO}_3)_2$
 - FeCl_3
- 9.73** Determine the number of equivalents and milliequivalents in each of the following:
- 0.10 mol KI
 - 0.25 mol MgCl_2
 - 4.73×10^{-2} mol AgNO_3
- 9.74** Determine the number of equivalents and milliequivalents in each of the following:
- 0.22 mol ZnCl_2
 - 0.45 mol CsCl
 - 3.12×10^{-2} mol $\text{Fe}(\text{NO}_3)_2$
- 9.75** Determine the number of equivalents and milliequivalents in 5.00 g of each of the following salts. Include any waters of hydration given in the salt formula when you calculate salt formula weights.
- NaCl
 - NaNO_3
 - Na_3PO_4
 - $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$
- 9.76** Determine the number of equivalents and milliequivalents in 5.00 g of each of the following salts. Include any waters of hydration given in the salt formula when you calculate salt formula weights.
- $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$
 - $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
 - Li_2CO_3
 - NaH_2PO_4
- 9.77** A sample of intracellular fluid contains 45.1 meq/L of Mg^{2+} ion. Assume the Mg^{2+} comes from dissolved MgCl_2 , and calculate the number of moles and number of grams of MgCl_2 that would be found in 250 mL of the intracellular fluid.
- 9.78** A sample of intracellular fluid contains 133 meq/L of K^+ ion. Assume the K^+ comes from dissolved K_2SO_4 , and calculate the number of moles and number of grams of K_2SO_4 that would be found in 150 mL of the intracellular fluid.

The Strengths of Acids and Bases (Section 9.9)

- 9.79** Illustrate the difference between weak, moderately strong, and strong acids by writing dissociation reactions for the hypothetical acid HB, using arrows of various lengths.
- 9.80** The K_a values have been determined for four acids and are listed below. Arrange the acids in order of increasing acid strength (weakest first, strongest last).
- acid A ($K_a = 5.6 \times 10^{-5}$)
- acid B ($K_a = 1.8 \times 10^{-5}$)
- acid C ($K_a = 1.3 \times 10^{-4}$)
- acid D ($K_a = 1.1 \times 10^{-3}$)
- 9.81** Arrange the four acids classified as weak in Table 9.7 in order of increasing strength (weakest first, strongest last).
- 9.82** K_a values for four weak acids are given below:
- acid A ($K_a = 2.6 \times 10^{-4}$)
- acid B ($K_a = 3.7 \times 10^{-5}$)
- acid C ($K_a = 5.8 \times 10^{-4}$)
- acid D ($K_a = 1.5 \times 10^{-3}$)
- Arrange the four acids in order of increasing acid strength (weakest first, strongest last).
 - Arrange the conjugate bases of the acids (identify as base A, etc.) in order of increasing base strength (weakest base first, strongest last).
- 9.83** Write dissociation reactions and K_a expressions for the following weak acids:
- Hypobromous acid, HBrO
 - Sulfurous acid, H_2SO_3 (1st H only)
 - Hydrogen sulfite ion, HSO_3^-
 - Hydroselenic acid, H_2Se (1st H only)
 - Arsenic acid, H_3AsO_4 (1st H only)
- 9.84** Write dissociation reactions and K_a expressions for the following weak acids:
- hydrogen selenide ion, HSe^-
 - dihydrogen borate ion, H_2BO_3^- (1st H only)
 - hydrogen borate ion, HBO_3^{2-}
 - hydrogen arsenate ion, HAsO_4^{2-}
 - hypochlorous acid, HClO
- 9.85** Equal molar solutions are made of three monoprotic acids HA, HB, and HC. The pHs of the solutions are, respectively, 4.82, 3.16, and 5.47. Rank the acids in order of increasing acid strength and explain your reasoning.
- 9.86** If someone asked you for a weak acid solution, which of the following would you provide according to definitions in this chapter?
- 0.05 M HCl
 - 20% acetic acid

If the individual really wanted the other solution, what term should have been used instead of *weak*?

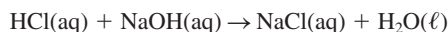
- 9.87** Arsenic acid (H_3AsO_4) is a moderately weak triprotic acid. Write equations showing its stepwise dissociation. Which of the three anions formed in these reactions will be the strongest Brønsted base? Which will be the weakest Brønsted base? Explain your answers.

Analyzing Acids and Bases (Section 9.10)

- 9.88** Explain the purpose of doing a titration.
- 9.89** Describe the difference between the information obtained by measuring the pH of an acid solution and by titrating the solution with base.
- 9.90** Suppose a student is going to titrate an acidic solution with a base and just picks an indicator at random. Under what circumstances will (a) the endpoint and equivalence point be the same? (b) The endpoint and equivalence point be different?
- 9.91** Determine the number of moles of NaOH that could be neutralized by each of the following:
- 1.00 L of 0.200 M HCl
 - 500 mL of 0.150 M HNO_3
- 9.92** Determine the number of moles of NaOH that could be neutralized by each of the following:
- 500 mL of 0.300 M $\text{HC}_2\text{H}_3\text{O}_2$ (monoprotic)
 - 1.50 L of 0.200 M HBr

Titration Calculations (Section 9.11)

- 9.93** Write a balanced molecular equation to represent the neutralization reaction between NaOH and each of the following acids. React all of the acid H^+ 's.
- Molybdic acid, H_2MoO_4
 - Permanganic acid, HMnO_4
 - Phosphoric acid, H_3PO_4
- 9.94** Write a balanced molecular equation to represent the neutralization reaction between NaOH and each of the following acids. React all of the acid H^+ 's.
- trichloroacetic acid, $\text{HC}_2\text{O}_2\text{Cl}_3$
 - dithionic acid, $\text{H}_2\text{S}_2\text{O}_6$
 - hypophosphorous acid, $\text{H}_4\text{P}_2\text{O}_6$
- 9.95** Write a balanced molecular equation to represent the neutralization reaction between HCl and each of the following bases:
- $\text{Cd}(\text{OH})_2$
 - $\text{Cr}(\text{OH})_3$
 - $\text{Fe}(\text{OH})_2$
- 9.96** Write a balanced molecular equation to represent the neutralization reaction between HCl and each of the following bases:
- $\text{Zn}(\text{OH})_2$
 - $\text{Tl}(\text{OH})_3$
 - CsOH
- 9.97** A 25.00-mL sample of gastric juice is titrated with a 0.0210 M NaOH solution. The titration to the equivalence point requires 29.8 mL of NaOH solution. If the equation for the reaction is



what is the molarity of HCl in the gastric juice?

- 9.98** A 25.0-mL sample of H_2SO_4 solution required 34.7 mL of 0.0400 M NaOH solution to titrate it to the equivalence point. Calculate the molarity of the H_2SO_4 solution.

- 9.99** A 20.00-mL sample of each of the following acid solutions is to be titrated to the equivalence point using a 0.120 M NaOH solution. Determine the number of milliliters of NaOH solution that will be needed for each acid sample.

- 0.180 M HCl
- 0.180 M H_2SO_4
- 0.100 M HCl
- 10.00 g of H_3PO_4 in 250 mL of solution
- 0.150 mol H_2MoO_4 in 500 mL of solution
- 0.215 mol H_2MoO_4 in 700 mL of solution

- 9.100** A 20.00-mL sample of each of the following acid solutions is to be titrated to the equivalence point using a 0.120 M NaOH solution. Determine the number of milliliters of NaOH solution that will be needed for each acid sample.

- 0.200 M HClO_4
- 0.125 M H_2SO_4
- 0.150 M $\text{H}_4\text{P}_2\text{O}_6$
- 0.120 mol H_3PO_4 in 500 mL of solution
- 6.25 g of H_2SO_4 in 250 mL of solution
- 0.500 mol HClO_3 in 1.00 L of solution

- 9.101** The following acid solutions were titrated to the equivalence point with the base listed. Use the titration data to calculate the molarity of each acid solution.

- 25.00 mL of HI solution required 27.15 mL of a 0.250 M NaOH solution.
- 20.00 mL of H_2SO_4 solution required 11.12 mL of a 0.109 M KOH solution.
- 25.00 mL of gastric juice (HCl) required 18.40 mL of a 0.0250 M NaOH solution.

- 9.102** The following acid solutions were titrated to the equivalence point with the base listed. Use the titration data to calculate the molarity of each acid solution.

- 5.00 mL of dilute H_2SO_4 required 29.88 mL of a 1.17 M NaOH solution.
- 10.00 mL of vinegar (acetic acid) required 35.62 mL of a 0.250 M KOH solution.
- 10.00 mL of muriatic acid (HCl) used to clean brick and cement required 20.63 mL of a 6.00 M NaOH solution.

- 9.103** A 20.00-mL sample of diprotic oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$) solution is titrated with a 0.250 M NaOH solution. A total of 27.86 mL of NaOH is required. Calculate:

- The number of moles of oxalic acid in the 20.00-mL sample.
- The molarity of the oxalic acid solution.
- The number of grams of oxalic acid in the 20.00-mL sample.

- 9.104** A sample of monoprotic benzoic acid weighing 0.5823 g is dissolved in about 25 mL of water. The solution is titrated to the equivalence point using 0.1021 M NaOH. The volume of base required is 46.75 mL. Calculate the molecular weight of the solid acid.

Hydrolysis Reactions of Salts (Section 9.12)

- 9.105** A solution of solid NH_4Cl dissolved in pure water is acidic (the pH is less than 7). Explain.
- 9.106** A solution of solid Na_3PO_4 dissolved in pure water is basic (the pH is greater than 7). Explain.
- 9.107** Predict the relative pH (greater than 7, less than 7, etc.) for water solutions of the following salts. Table 9.9 may be useful. For each solution in which the pH is greater or less than 7, explain why and write a net ionic equation to justify your answer.
- Potassium sulfite, K_2SO_3
 - Lithium nitrite, LiNO_2
 - Sodium carbonate, Na_2CO_3
 - Methylammonium chloride, $\text{CH}_3\text{NH}_3\text{Cl}$ (CH_3NH_2 is a weak base)
- 9.108** Predict the relative pH (greater than 7, less than 7, etc.) for water solutions of the following salts. Table 9.9 may be useful. For each solution in which the pH is greater or less than 7, explain why and write a net ionic equation to justify your answer.
- Sodium hypochlorite, NaOCl (HOCl is a weak acid)
 - Sodium formate, NaCHO_2
 - Potassium nitrate, KNO_3
 - Sodium phosphate, Na_3PO_4
- 9.109** A chemist has 20.00-mL samples of 0.100 M acid A and 0.100 M acid B in separate flasks. Both acids are monoprotic. Unfortunately, the flasks were not labeled, so the chemist doesn't know which sample is in which flask. But fortunately, it is known that acid A is strong and acid B is weak. Before thinking about the problem, the chemist adds 20.00 mL of 0.100 M NaOH solution to each flask. Explain how the chemist could use a pH meter (or pH paper) to determine which flask originally contained which acid.
- 9.110** Explain why the hydrolysis of salts makes it necessary to have available in a laboratory more than one acid–base indicator for use in titrations.
- 9.111** How would the pH values of equal molar solutions of the following salts compare (highest, lowest, etc.)? NaH_2PO_4 , Na_2HPO_4 , and Na_3PO_4 .

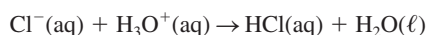
Buffers (Section 9.13)

- 9.112** Write equations similar to Equations 9.48 and 9.49 of the text to illustrate how a mixture of sodium hydrogen phosphate (Na_2HPO_4) and sodium dihydrogen phosphate (NaH_2PO_4) could function as a buffer when dissolved in water. Remember that phosphoric acid (H_3PO_4) ionizes in three steps.
- 9.113** Could a mixture of ammonia (NH_3), a weak base, and ammonium chloride (NH_4Cl) behave as a buffer when dissolved in water? Use reaction equations to justify your answer.
- 9.114** Some illnesses lead to a condition of excess acid (acidosis) in the body fluids. An accepted treatment is to inject solutions containing bicarbonate ions (HCO_3^-) directly into the bloodstream. Write an equation to show how this treatment would help combat the acidosis.

- 9.115** Calculate the pH of a buffer made by dissolving 1 mol formic acid (HCOOH) and 1 mol sodium formate (HCOONa) in 1 L of solution (see Table 9.9).
- 9.116**
- Calculate the pH of a buffer that is 0.1 M in lactic acid ($\text{C}_2\text{H}_4(\text{OH})\text{COOH}$) and 0.1 M in sodium lactate, $\text{C}_2\text{H}_4(\text{OH})\text{COONa}$.
 - What is the pH of a buffer that is 1 M in lactic acid and 1 M in sodium lactate?
 - What is the difference between the buffers described in parts a and b?
- 9.117** Which of the following acids and its conjugate base would you use to make a buffer with a pH of 3.00? Explain your reasons: formic acid, lactic acid, nitrous acid.
- 9.118** Calculate the pH of buffers that contain the acid and conjugate base concentrations listed below.
- $[\text{CH}_3\text{COOH}] = 0.40 \text{ M}$, $[\text{CH}_3\text{COO}^-] = 0.25 \text{ M}$
 - $[\text{H}_2\text{PO}_4^-] = 0.10 \text{ M}$, $[\text{HPO}_4^{2-}] = 0.40 \text{ M}$
 - $[\text{HSO}_3^-] = 1.50 \text{ M}$, $[\text{SO}_3^{2-}] = 0.20 \text{ M}$
- 9.119** Calculate the pH of buffers that contain the acid and conjugate base concentrations listed below.
- $[\text{HPO}_4^{2-}] = 0.33 \text{ M}$, $[\text{PO}_4^{3-}] = 0.52 \text{ M}$
 - $[\text{HNO}_2] = 0.029 \text{ M}$, $[\text{NO}_2^-] = 0.065 \text{ M}$
 - $[\text{HCO}_3^-] = 0.50 \text{ M}$, $[\text{CO}_3^{2-}] = 0.15 \text{ M}$
- 9.120** What ratio of concentrations of NaH_2PO_4 and Na_2HPO_4 in solution would give a buffer with $\text{pH} = 7.65$?
- 9.121** A citric acid–citrate buffer has a pH of 3.20. You want to increase the pH to a value of 3.35. Would you add citric acid or sodium citrate to the solution? Explain.

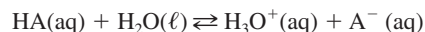
Additional Exercises

- 9.122** Consider the following reaction equation:



Draw Lewis structures (electron-dot formulas) for all the substances in the reaction equation, identify each reactant as a Brønsted acid or base, and then propose a definition for an acid and a base based on the ability of the substance to accept or donate a pair of electrons to form a covalent bond.

- 9.123** Consider the following dissociation reaction of a weak acid, HA:



It was determined that 2.63% of the acid in a 0.150 M solution of HA in water was dissociated. Calculate the pH of the 0.150 M solution.

- 9.124** The value of K_w for water varies with temperature. At 25°C , $K_w = 1.0 \times 10^{-14}$. At 50°C , $K_w = 5.5 \times 10^{-14}$. Calculate the pH of water at 50°C .
- 9.125** When sodium metal, Na, reacts with water, hydrogen gas, H_2 , and sodium hydroxide, NaOH , in solution are produced. Write a balanced equation for the reaction, and explain how a basic solution is produced.

Allied Health Exam Connection

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9.126 An acid is a substance that dissociates in water into one or more _____ ions and one or more _____.

- a. hydrogen . . . anions
- b. hydrogen . . . cations
- c. hydroxide . . . anions
- d. hydroxide . . . cations

9.127 A base is a substance that dissociates in water into one or more _____ ions and one or more _____.

- a. hydrogen . . . anions
- b. hydrogen . . . cations
- c. hydroxide . . . anions
- d. hydroxide . . . cations

9.128 Which is true of alkaline solutions?

- a. More H^+ ion than OH^- ion
- b. Same amount of H^+ ion + OH^- ion
- c. More OH^- ion than H^+ ion
- d. None of the above

9.129 Which of the following is NOT an acid/conjugate base pair?

- a. HCN/CN^-
- b. H_2CO_3/OH^-
- c. H_2SO_4/HSO_4^-
- d. $H_3PO_4/H_2PO_4^-$

9.130 What is the formula of the hydronium ion?

- a. H^+
- b. NH_4^+
- c. H_3O^+
- d. H_2O^+

9.131 Which of the following substances has a pH closest to 7?

- a. ammonia
- b. blood
- c. lemon juice
- d. vinegar

9.132 Dissolving H_2SO_4 in water creates an acid solution by increasing the:

- a. sulfate ions
- b. water ions
- c. hydrogen ions
- d. oxygen ions

9.133 When a solution has a pH of 7, it is:

- a. a strong base
- b. a strong acid
- c. a weak base
- d. neutral

9.134 A common detergent has a pH of 11.0, so the detergent is:

- a. neutral
- b. acidic
- c. alkaline
- d. none of the above

9.135 In a 0.001 M solution of HCl, the pH is:

- a. 2
- b. -3
- c. 1
- d. 3

9.136 The pH of a blood sample is 7.40 at room temperature. The pOH is therefore:

- a. 6.60
- b. 7.40
- c. 6×10^{-6}
- d. 4×10^{-7}

9.137 As the concentration of hydrogen ions in a solution decreases:

- a. the pH numerically decreases
- b. the pH numerically increases
- c. the product of the concentrations $[H^+] \times [OH^-]$ comes closer to 1×10^{-14}
- d. the solution becomes more acidic

9.138 Blood with a pH of 7.4 indicates that the blood sample is:

- a. strongly acidic
- b. strongly basic
- c. weakly acidic
- d. weakly basic

9.139 Atmospheric moisture (H_2O) combines with oxides of carbon, nitrogen, and sulfur (CO_2 , NO_3 and SO_2) to produce:

- a. alkaline precipitation
- b. acid rain
- c. alcohol
- d. none of the above

9.140 Which of the following substances is most likely to taste sour?

- a. NaOH
- b. NaCl
- c. NH_3
- d. $\text{HC}_2\text{H}_3\text{O}_2$

9.141 Which of the following compounds would be classified as a salt?

- a. Na_2CO_3
- b. $\text{Ca}(\text{OH})_2$
- c. H_2CO_3
- d. CH_3OH

9.142 A substance that functions to prevent rapid, drastic changes in the pH of a body fluid by changing strong acids and bases into weak acids and bases is called a(n):

- a. salt
- b. buffer
- c. enzyme
- d. coenzyme

9.143 Which one of the following equations represents neutralization?

- a. $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$
- b. $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$
- c. $\text{HNO}_3 + \text{KOH} \rightarrow \text{KNO}_3 + \text{H}_2\text{O}$
- d. $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2$

9.144 Which reaction below demonstrates a neutralization reaction?

- a. ${}^{232}_{90}\text{Th} \rightarrow {}^{228}_{88}\text{Ra} + {}^4_2\text{He}$
- b. $\text{NaCl} + \text{H}_2\text{O} \rightarrow \text{HCl} + \text{NaOH}$
- c. $2\text{HNO}_3 + \text{Mg}(\text{OH})_2 \rightarrow 2\text{H}_2\text{O} + \text{Mg}(\text{NO}_3)_2$
- d. $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

9.145 In a titration of 40.0 mL of 0.20 M NaOH with 0.4 M HCl, what will be the final volume of the solution when the sodium hydroxide is completely neutralized?

- a. 42 mL
- b. 20 mL
- c. 60 mL
- d. 80 mL

9.146 When titrating 50 mL of 0.2 M HCl, what quantity of 0.5 M NaOH is needed to bring the solution to the equivalence point?

- a. 80 mL
- b. 40 mL
- c. 20 mL
- d. 10 mL

Chemistry for Thought

9.147 In an early industrial method, H_2SO_4 was manufactured in lead-lined chambers. Propose an explanation for this.

9.148 A saturated solution of solid $\text{Ca}(\text{OH})_2$ in water has a $[\text{OH}^-]$ of only 2.50×10^{-2} M, and yet $\text{Ca}(\text{OH})_2$ is a strong base. Explain this apparent contradiction.

9.149 Imagine that a solution of weak acid is being titrated with a strong base. Describe the substances present in the solution being titrated when it has been titrated halfway to the equivalence point. How is the pH of this half-titrated solution related to the $\text{p}K_a$ for the weak acid?

9.150 Calculate K_a for the following weak acids based on the equilibrium concentrations and pH values given:

a. Benzoic acid, represented by HBz:

$$\text{pH} = 2.61, [\text{Bz}^-] = 2.48 \times 10^{-3}, [\text{HBz}] = 0.0975$$

b. Abietic acid, represented by HAB:

$$\text{pH} = 4.16, [\text{Ab}^-] = 6.93 \times 10^{-5}, [\text{HAB}] = 0.200$$

c. Cacodylic acid, represented by HCc:

$$\text{pH} = 3.55, [\text{Cc}^-] = 0.000284, [\text{HCc}] = 0.150$$

9.151 Bottles of ketchup are routinely left on the counters of cafés, yet the ketchup does not spoil. Why not?

9.152 Refer to Figure 9.2 and answer the question. What implications does this reaction have for the long-term durability of marble structures when exposed to acid rain?

9.153 Refer to Figure 9.3 and answer the question. Do you think the results would be the same if sulfuric acid (H_2SO_4) were substituted for the HCl? Explain.

9.154 Refer to Figure 9.5 and answer the question. Answer the same question for the other two indicators.

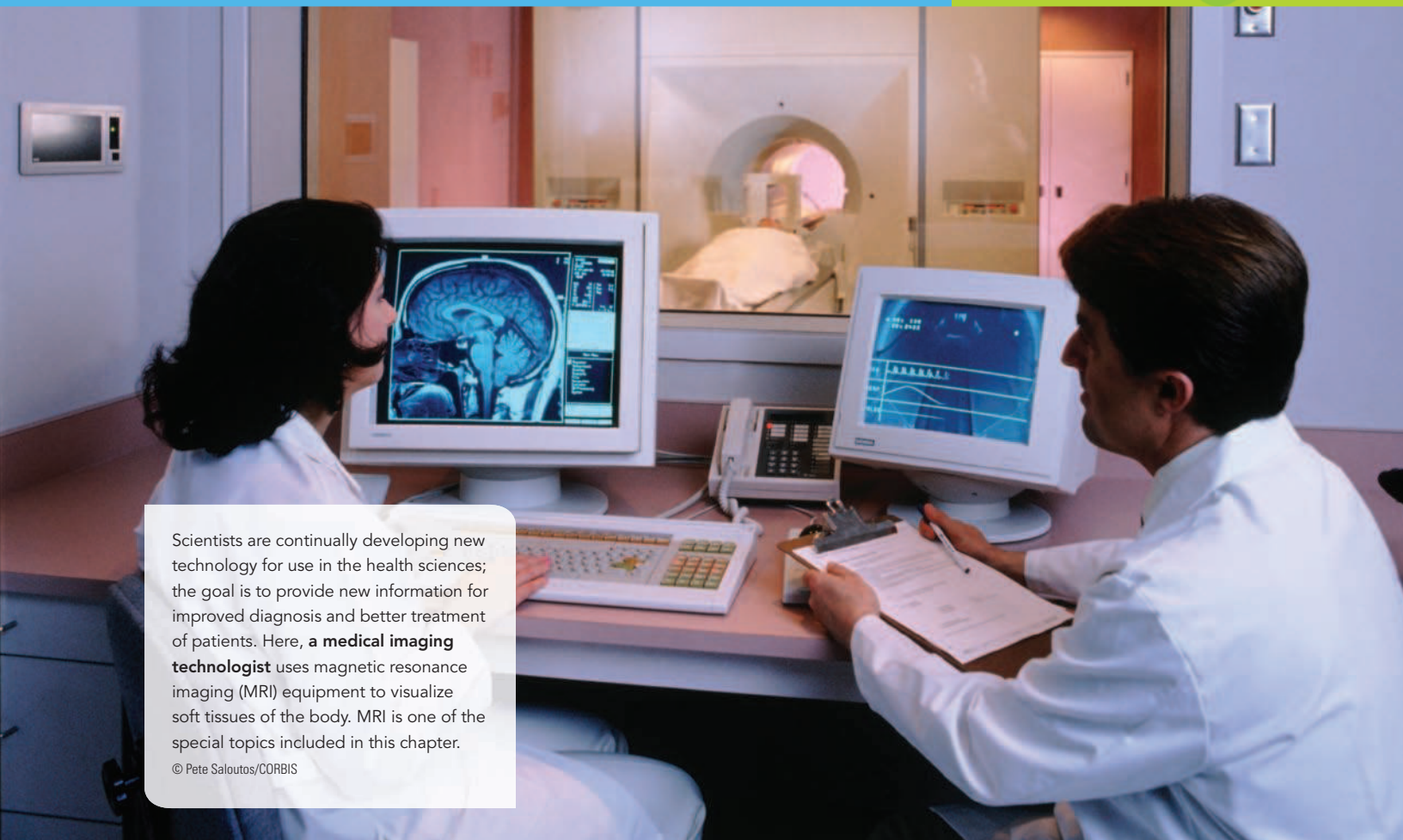
9.155 Refer to Figure 9.6 and explain why the pH reading of the meter might not be 7.00 at the equivalence point in a titration. Would the pH at the equivalence point be greater or less than 7 if hydrochloric acid were titrated with aqueous ammonia?

9.156 Refer to Figure 9.7 and answer the question. List at least two other ions that would behave like the Na^+ ion.

9.157 Refer to Figure 9.8 and answer the question. What color is universal indicator in acidic solutions? In basic solutions? Explain how you arrived at your conclusions.

Radioactivity and Nuclear Processes

10



Scientists are continually developing new technology for use in the health sciences; the goal is to provide new information for improved diagnosis and better treatment of patients. Here, a **medical imaging technologist** uses magnetic resonance imaging (MRI) equipment to visualize soft tissues of the body. MRI is one of the special topics included in this chapter.

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Learning Objectives

When you have completed your study of this chapter, you should be able to:

- 1 Describe and characterize the common forms of radiation emitted during radioactive decay and other nuclear processes. (**Section 10.1**)
- 2 Write balanced equations for nuclear reactions. (**Section 10.2**)
- 3 Solve problems using the half-life concept. (**Section 10.3**)
- 4 Describe the effects of radiation on health. (**Section 10.4**)
- 5 Describe and compare the units used to measure quantities of radiation. (**Section 10.5**)
- 6 Describe, with examples, medical uses of radioisotopes. (**Section 10.6**)
- 7 Describe, with examples, nonmedical uses of radioisotopes. (**Section 10.7**)
- 8 Show that you understand the concept of induced nuclear reactions. (**Section 10.8**)
- 9 Describe the differences between nuclear fission and nuclear fusion reactions. (**Section 10.9**)

OWL Online homework for this chapter may be assigned in OWL.

In Chapters 3 and 4, we learned that the arrangements of electrons around the nuclei of atoms determine the chemical behavior of elements and compounds. We saw that chemical reactions take place between atoms and molecules as electrons rearrange to achieve noble gas configurations for the atoms involved. These noble gas configurations represent a stable arrangement for the electrons. In this chapter, we will see that the achievement of stability is also a driving force that causes certain atomic nuclei to undergo changes that release small subnuclear particles and energy.

10.1 Radioactive Nuclei

Learning Objective

1. Describe and characterize the common forms of radiation emitted during radioactive decay and other nuclear processes.

In 1896 Henri Becquerel, a French physicist, discovered that uranium compounds emitted rays that could expose and fog photographic plates wrapped in lightproof paper. Research conducted since that time has shown that the penetrating rays originate from changes that occur in the nuclei of some atoms. We saw earlier (Section 2.3) that isotopes of the same element differ only in the number of neutrons present in their nuclei. Thus, each of the three isotopes of hydrogen, represented by the symbols ${}^1_1\text{H}$, ${}^2_1\text{H}$, and ${}^3_1\text{H}$, contains one proton in the nucleus and one electron outside the nucleus. However, besides the one nuclear proton, ${}^2_1\text{H}$ and ${}^3_1\text{H}$ contain one and two nuclear neutrons, respectively. It has been found that some combinations of neutrons and protons in the nucleus are stable and do not change spontaneously, but some combinations are not stable. It is these unstable nuclei that emit radiation as they become more stable. For example, both ${}^1_1\text{H}$ and ${}^2_1\text{H}$ are stable, but ${}^3_1\text{H}$ emits radiation. Nuclei that emit radiation are said to be **radioactive nuclei**.

Henri Becquerel discovered natural radioactivity by chance. Through research he found that radiation was emitted by any compound of uranium. He further found that the intensity of the radiation was unaffected by factors that normally influence the rates of chemical reactions: the temperature, pressure, and type of uranium compound used.

Later studies by other investigators showed that the radiation emitted by uranium, and by other radioactive elements discovered later, could be separated into three types by an electrical or magnetic field (see ▶ Active Figure 10.1). The three types had different electrical charges: One was positive, one was negative, and one carried no charge. The types of radiation were given names that are still used today: alpha rays (positive), beta rays (negative), and gamma rays (uncharged).

Today it is known that other types of radiation, such as neutrons and positrons, are also emitted by radioactive nuclei. However, alpha, beta, and gamma are the most common. ▶ Table 10.1 summarizes the characteristics of these forms. In this book, we usually use the symbol given first for each form; the symbols in parentheses are alternatives you might find elsewhere. Except for gamma radiation, which is very high-energy

radioactive nuclei Nuclei that undergo spontaneous changes and emit energy in the form of radiation.

Active Figure 10.1 The three types of emitted radiation. Go to www.cengage.com/chemistry/seager or OWL to explore an interactive version of this figure.

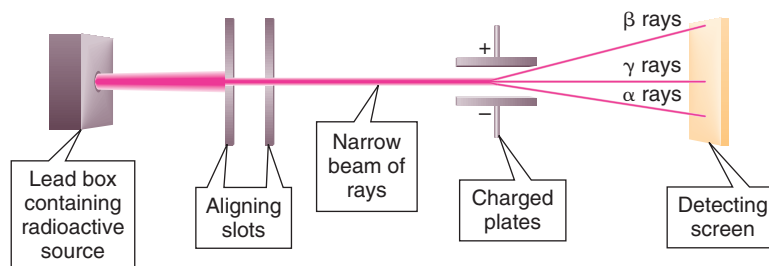


Table 10.1 Characteristics of Nuclear Radiation

Type of Radiation	Symbols	Mass Number	Charge	Composition
Alpha	${}^4_2\alpha$ (α , ${}^4_2\text{He}$, He^{2+})	4	+2	Helium nuclei, 2 protons + 2 neutrons
Beta	${}^0_{-1}\beta$ (β , β^- , ${}^0_{-1}e$, e^-)	0	-1	Electrons produced in nucleus and ejected
Gamma	γ (${}^0_0\gamma$)	0	0	Electromagnetic radiation
Neutron	${}_0^1n$ (n)	1	0	Neutrons
Positron	${}^0_1\beta$ (β^+ , 0_1e , e^+)	0	+1	Positive electrons

radiation somewhat like X-rays, the radiation emitted by radioactive nuclei consists of streams of particles. The emission of radiation by unstable nuclei is often called **radioactive decay**.

We note from Table 10.1 that the particles that make up alpha rays are identical to the nuclei of helium atoms; that is, they are clusters containing two protons and two neutrons. Because they are positively charged, alpha particles are attracted toward the negatively charged plate in Figure 10.1. **Alpha particles** move at a speed of nearly one-tenth the speed of light. They are the most massive particles emitted by radioactive materials, and they have the highest charge. Because of their mass and charge, they collide often with molecules of any matter through which they travel. As a result of these collisions, their energy is quickly dissipated, and they do not travel far. They cannot even penetrate through a few sheets of writing paper or the outer cells of the skin. However, exposure to intense alpha radiation can result in severe burns of the skin.

Beta particles are actually electrons, but they are produced within the nucleus and then emitted. They are not part of the group of electrons moving around the nucleus that is responsible for chemical characteristics of the atoms. Beta particles have a smaller charge than alpha particles and are also 7000 times less massive than alpha particles. They undergo far fewer collisions with molecules of matter through which they pass and therefore are much more penetrating than alpha particles. Because of their negative charge and tiny mass, beta particles are attracted toward the positive plate of Figure 10.1, and are deflected more than alpha particles.

Gamma rays are not streams of particles but rays of electromagnetic radiation similar to X-rays that have higher energies. Like X-rays, gamma radiation is very penetrating and dangerous to living organisms. Adequate shielding from gamma rays is provided only by thick layers of heavy materials like metallic lead or concrete. Because they are not charged, gamma rays are not attracted to either charged plate in Figure 10.1.

radioactive decay A process in which an unstable nucleus changes energy states and in the process emits radiation.

alpha particle The particle that makes up alpha rays. It is identical to the helium nucleus and is composed of two protons and two neutrons.

beta particle The particle that makes up beta rays. It is identical to an electron but is produced in the nucleus when a neutron is changed into a proton and an electron.

gamma ray A high-energy ray that is like an X-ray, but with a higher energy.

10.2 Equations for Nuclear Reactions

Learning Objective

- Write balanced equations for nuclear reactions.

Isotopes of elements that emit nuclear radiation are called **radioisotopes**. In nuclear reactions, a specific isotope of an element may behave differently from another isotope of the same element. Thus, all particles involved in nuclear reactions are designated by a symbol, a mass number (the sum of protons and neutrons in the particle), and an atomic number (or charge for electrons or positrons). The symbolism used was introduced in Section 2.3. It is ${}_Z^AX$, where X is the symbol for the particle, A is the mass number, and

radioisotope An isotope of an element that emits nuclear radiation.

Z is the atomic number or charge. When X represents the nucleus of an isotope of an element, the chemical symbol for the element is generally used.

Example 10.1

Write appropriate symbols for the following particles using the A_ZX symbolism:

- A lead-214 nucleus
- An alpha particle
- A deuteron, a particle containing one proton and one neutron

Solution

- Lead-214 is an isotope of lead, so the symbol X is replaced by Pb . The mass number is 214 according to rules learned earlier for designating isotopes, so $A = 214$. The atomic number of lead obtained from the periodic table is 82, so $Z = 82$, and the symbol is ${}^{214}_{82}\text{Pb}$.
- An alpha particle is represented by the symbol α . Since it contains 2 protons and 2 neutrons, $A = 4$ and $Z = 2$. The symbol is ${}^4_2\alpha$, as shown in Table 10.1.
- A deuteron is not familiar to you, so we will replace its symbol X with D . It contains one proton and one neutron, so $A = 2$ and $Z = 1$. The symbol is 2_1D . This particle is the nucleus of a deuterium or heavy hydrogen atom and so could also be represented as ${}^2_1\text{H}$.

Learning Check 10.1 Write appropriate symbols for the following particles using the A_ZX symbolism:

- An iodine-131 nucleus
- A beta particle
- A hypothetical particle composed of four protons
- A hypothetical particle composed of four electrons

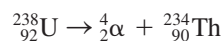
A nuclear equation is balanced when the sum of the atomic numbers on the left side equals the sum of the atomic numbers on the right side, and the sum of the mass numbers on the left equals the sum of the mass numbers on the right. Thus, we focus on mass numbers and atomic numbers in nuclear reactions. As shown in the following example, a nucleus of an element that undergoes a nuclear reaction by emitting a particle is changed to a nucleus of a different element.

Example 10.2

Uranium-238 nuclei undergo radioactive decay by emitting an alpha particle and changing to the nucleus of another element. Write a balanced nuclear equation for the process.

Solution

Because the uranium-238 nucleus gives up an alpha particle, ${}^4_2\alpha$, the mass number of the uranium-238 must decrease by 4 and the atomic number must decrease by 2. As a result, the new nucleus, called a **daughter nucleus**, has a mass number of 234 and an atomic number of 90 (remember, the atomic number of uranium can be obtained from a periodic table). The balanced nuclear equation is



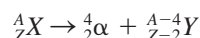
Note that this equation is balanced because

$$238 = 4 + 234 \quad \text{and} \quad 92 = 2 + 90$$

The symbol for the daughter (thorium) was obtained by locating element number 90 in the periodic table.

daughter nuclei The new nuclei produced when unstable nuclei undergo radioactive decay.

Processes such as this where one alpha particle is ejected can be represented by the general equation



where Y is the symbol of the daughter nucleus.

► **Learning Check 10.2** Thorium, element number 90, exists in a number of isotopic forms. One form, thorium-234, decays by emitting a beta particle. Write a balanced nuclear equation for the process and write a general equation for decays in which one beta particle is ejected.

Because gamma rays have no mass or atomic numbers, they do not enter into the balancing process of nuclear reactions. However, they do represent energy and should be included in balanced equations when they are known to be emitted.

As described earlier, beta particles are electrons ejected from the nucleus of a radioactive atom, and yet, we know from our discussions of atomic theory in Chapter 2 that atomic nuclei contain only protons and neutrons. Where then do the beta particles originate? According to a simplified theory of nuclear behavior, a beta particle (or electron) is produced in the nucleus when a neutron changes into a proton. The balanced equation for the nuclear reaction is



Note that the criteria for balance in this equation are satisfied; that is, $1 = 1 + 0$ and $0 = 1 + (-1)$. Thus, a nuclear neutron is converted to a proton during beta emission, and the daughter has an atomic number higher by 1 than the decaying nucleus. You found this result if you did Learning Check 10.2 correctly. Similar changes in nuclear particles occur during two other types of nuclear reactions. Some nuclei decay by positron emission. A **positron** is a positively charged electron with the symbol ${}^0_1\beta$ (see Table 10.1). When a nucleus emits a positron, a nuclear proton is changed to a neutron. Positron-emitting isotopes are used in diagnostic PET scans of the brain (see ► Figure 10.2).

positron A positively charged electron.

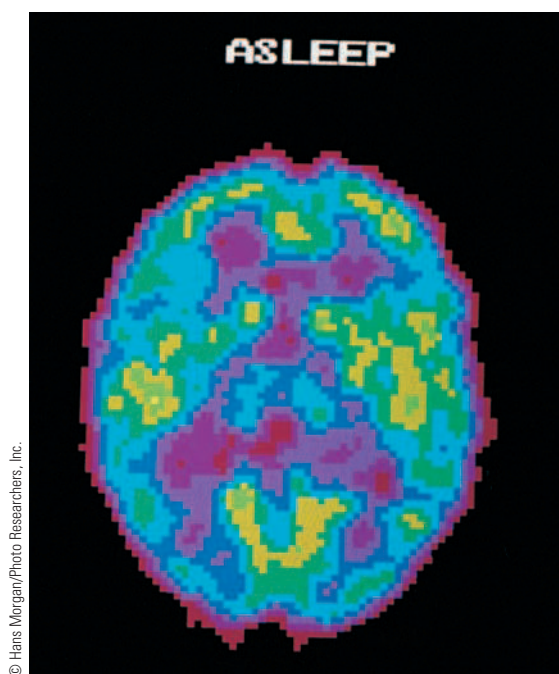


Figure 10.2 Positron emission tomography (PET) scan shows normal brain activity during sleep. A radioactive isotope that emits positrons is made into a chemical compound that is absorbed by active areas of the brain. The emitted positrons collide with nearby electrons and produce gamma rays that pass through the skull to detectors surrounding the patient's head. A computer uses the detector data to construct the image.

electron capture A mode of decay for some unstable nuclei in which an electron from outside the nucleus is drawn into the nucleus, where it combines with a proton to form a neutron.

► **Learning Check 10.3** Write a general balanced equation for the nuclear process that occurs during positron emission. How do the atomic numbers of daughter nuclei compare with those of the decaying nucleus?

Certain nuclear changes occur when an electron from outside the nucleus is drawn into the nucleus, where it reacts with a proton and converts it to a neutron. This process, called **electron capture**, is not as common as other decay processes.

► **Learning Check 10.4** Write a general balanced equation for the nuclear process that occurs during electron capture. How do the atomic numbers of daughter nuclei compare with those of the decaying nucleus?

► Example 10.3

Write a balanced nuclear equation for the decay of each of the following. The mode of decay is indicated in parentheses.

- ${}^{14}_6\text{C}$ (beta emission)
- ${}^{122}_{53}\text{I}$ (positron emission)
- ${}^{55}_{26}\text{Fe}$ (electron capture)
- ${}^{212}_{84}\text{Po}$ (alpha emission)

Solution

In each case, the atomic number of the daughter is obtained by balancing atomic numbers in the equation. The symbol for the daughter is obtained by matching atomic numbers to symbols in the periodic table.

- ${}^{14}_6\text{C} \rightarrow {}^0_{-1}\beta + {}^{14}_7\text{N}$
- ${}^{122}_{53}\text{I} \rightarrow {}^0_1\beta + {}^{122}_{52}\text{Te}$
- ${}^{55}_{26}\text{Fe} + {}^0_{-1}e \rightarrow {}^{55}_{25}\text{Mn}$
- ${}^{212}_{84}\text{Po} \rightarrow {}^4_2\alpha + {}^{208}_{82}\text{Pb}$

► **Learning Check 10.5** Write a balanced nuclear equation for the decay of each of the following. The mode of decay is indicated in parentheses.

- ${}^{50}_{25}\text{Mn}$ (positron emission)
- ${}^{54}_{25}\text{Mn}$ (electron capture)
- ${}^{56}_{25}\text{Mn}$ (beta emission)
- ${}^{224}_{88}\text{Ra}$ (alpha emission)

10.3 Isotope Half-Life

Learning Objective

3. Solve problems using the half-life concept.

Some radioactive isotopes are more stable than others. The more stable isotopes undergo radioactive decay more slowly than the less stable isotopes. The **half-life** of an isotope is used to indicate stability, and it is equal to the time required for one-half (50%) of the atoms of a sample of the isotope to decay. ► Table 10.2 contains examples showing the wide range of half-lives that have been determined.

half-life The time required for one-half the unstable nuclei in a sample to undergo radioactive decay.

Table 10.2 Examples of Half-Lives

Isotope	Half-Life	Source
${}_{92}^{238}\text{U}$	4.5×10^9 years	Naturally occurring
${}_{19}^{40}\text{K}$	1.3×10^9 years	Naturally occurring
${}_{88}^{226}\text{Ra}$	1600 years	Naturally occurring
${}_{6}^{14}\text{C}$	5600 years	Naturally occurring
${}_{94}^{239}\text{Pu}$	24,000 years	Synthetically produced
${}_{38}^{90}\text{Sr}$	28 years	Synthetically produced
${}_{53}^{131}\text{I}$	8 days	Synthetically produced
${}_{11}^{24}\text{Na}$	15 hours	Synthetically produced
${}_{8}^{15}\text{O}$	2 minutes	Synthetically produced
${}_{3}^5\text{Li}$	10^{-21} seconds	Synthetically produced

Example 10.4

Rubidium-79 decays by positron emission and forms krypton-79, which is a gas. A weighed 100.00-mg sample of solid rubidium-79 was allowed to decay for 42 minutes, then weighed again. Its mass was 25.00 mg. What is the half-life of rubidium-79?

Solution

We assume all the gaseous krypton-79 that was formed escaped into the surrounding air and so was not weighed. In the first half-life, one-half the original 100.00-mg sample of rubidium would have been lost, so the sample would have a mass of 50.00 mg. During the second half-life, this 50.00-mg sample would again lose half its mass and would be reduced to 25.00 mg. Thus, we can conclude that two half-lives passed during the 42 minutes that elapsed between the weighings. Because two half-lives equal 42 minutes, one half-life equals 21 minutes.

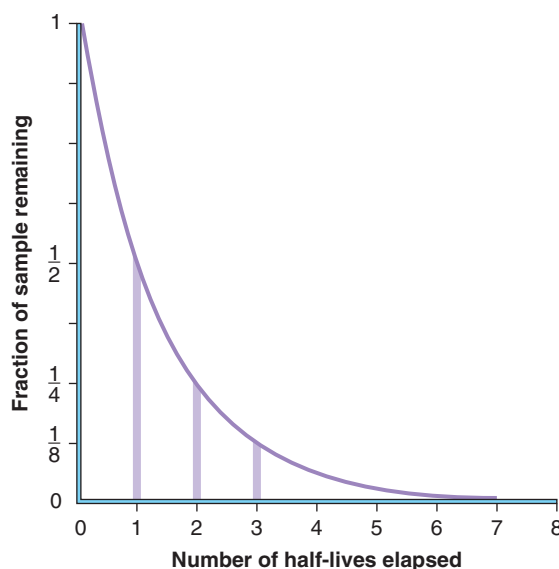
Learning Check 10.6 Potassium-38 decays by positron emission to argon-38, a gas. Potassium-38 has a half-life of 7.7 minutes. How long will it take a 200.00-mg sample of potassium-38 to be reduced to 25.00 mg?

Because radioisotopes are continuously decaying, any found in nature must belong to one of three categories. They may have very long half-lives and decay very slowly (${}_{92}^{238}\text{U}$), they may be daughters produced by the decay of long-lived isotopes (${}_{88}^{226}\text{Ra}$), or they may result from natural processes such as cosmic-ray bombardment of stable nuclei (${}_{6}^{14}\text{C}$).

The fraction of original radioactive atoms remaining in a sample after a specific time has passed can be determined from the half-life. After one half-life has passed, $\frac{1}{2}$ the original number have decayed, so $\frac{1}{2}$ remain. During the next half-life, $\frac{1}{2}$ the remaining $\frac{1}{2}$ decay, so $\frac{1}{4}$ of the original atoms remain. After three half-lives, $\frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} = \frac{1}{8}$ of the original atoms remain undecayed (see Figure 10.3).

The amount of radiation given off by a sample of radioactive material is proportional to the number of radioactive atoms present in the sample. The amount of radiation given off is called the *activity* of the sample and is measured by devices such as the

Figure 10.3 A radioactive decay curve (bars give values at 1, 2, and 3 half-lives).



Geiger–Müller counter described in Section 10.5. Data such as the sample fractions shown in Figure 10.3 are obtained quite easily. For example, suppose a sample gave an initial reading of 120 units of activity, and it took 35 minutes for the reading to drop to 60 units. The reading of 60 units means that the sample contained just half as many radioactive atoms as it contained when the reading was 120 units. Thus, the 35 minutes represents one half-life, and the fraction remaining from the time of the initial measurement is $\frac{1}{2}$.

Example 10.5

The activity (amount of radiation emitted) of a radioactive substance is measured at 9 A.M. At 5 P.M. the activity is found to be only $\frac{1}{16}$ the original (9 A.M.) value. What is the half-life of the radioactive material?

Solution

Because only $\frac{1}{16}$ the original radioactive material is present, four half-lives have passed: $\frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} \times \frac{1}{2} = \frac{1}{16}$. The time between measurements is 8 hours (9 A.M. to 5 P.M.). Therefore, 8 hours is four half-lives, and the half-life is 2 hours.

Learning Check 10.7 Iodine-123 is a radioisotope used to diagnose the function of the thyroid gland. It has a half-life of 13.3 hours. What fraction of a diagnostic dose of iodine-123 would be present in a patient 79.8 hours (a little over 3 days) after it was administered?

10.4 The Health Effects of Radiation

Learning Objective

- Describe the effects of radiation on health.

You are probably aware of the fact that radiation is hazardous to living organisms. But why? What does radiation do that is so dangerous? The greatest danger to living organisms results from the ability of high-energy or ionizing radiation to knock electrons out of compounds and generate highly reactive particles called **radicals** or **free radicals** in tissue through which the radiation passes. These electron-deficient radicals are very reactive. They cause reactions to occur among more stable materials in the cells of living organisms. When these reactions involve genetic materials such as genes and chromosomes, the changes might lead to genetic mutations, cancer, or other serious consequences.

radical or free radical An electron-deficient particle that is very reactive.

Table 10.3 The Effects on Humans of Short-Term, Whole-Body Radiation Exposure

Dose (rems) ^a	Effects
0–25	No detectable clinical effects.
25–100	Slight short-term reduction in number of some blood cells, disabling sickness not common.
100–200	Nausea and fatigue, vomiting if dose is greater than 125 rems, longer-term reduction in number of some blood cells.
200–300	Nausea and vomiting first day of exposure, up to a 2-week latent period followed by appetite loss, general malaise, sore throat, pallor, diarrhea, and moderate emaciation. Recovery in about 3 months unless complicated by infection or injury.
300–600	Nausea, vomiting, and diarrhea in first few hours. Up to a 1-week latent period followed by loss of appetite, fever, and general malaise in the second week, followed by hemorrhage, inflammation of mouth and throat, diarrhea, and emaciation. Some deaths in 2–6 weeks. Eventual death for 50% if exposure is above 450 rems; others recover in about 6 months.
600 or more	Nausea, vomiting, and diarrhea in first few hours. Rapid emaciation and death as early as second week. Eventual death of nearly 100%.

^aThe rem, a biological unit of radiation, is defined in Section 10.5.

Source: U.S. Atomic Energy Commission.

Long-term exposure to low-level radiation is more likely to cause the problems just described than are short exposures to intense radiation. Short-term intense radiation tends to destroy tissue rapidly in the area exposed and can cause symptoms of so-called **acute radiation syndrome**. Some of these symptoms are given in Table 10.3. It is this rapid destruction of tissue that makes relatively intense radiation a useful tool in the treatment of some cancers.

The health hazards presented by radiation make it imperative that exposures be minimized. This is especially true for individuals with occupations that continually present opportunities for exposure. One approach is to absorb radiation by shielding. As shown in Table 10.4, alpha and beta rays are the easiest to stop. Gamma radiation and X-rays require the use of very dense materials such as lead to provide protection. If the type of radiation is known, a careful choice of shielding materials can provide effective protection.

Another protection against radiation is distance. Because radiation spreads in all directions from a source, the amount falling on a given area decreases the farther that area is from the source. Imagine a light bulb at the center of a balloon that is being filled slowly with air. The intensity of light on each square centimeter of the balloon will decrease as

acute radiation syndrome The condition associated with and following short-term exposure to intense radiation.

Table 10.4 Penetrating Abilities of Alpha, Beta, and Gamma Rays

Type of Radiation	Depth of Penetration Into		
	Dry Air	Tissue	Lead
Alpha	4 cm	0.05 mm	0
Beta ^a	6–300 cm	0.06–4 mm	0.005–0.3 mm
Gamma ^b	400 m	50 cm	300 mm

^aDepth depends on energy of rays.

^bDepth listed is that necessary to reduce initial intensity by 10%.

inverse square law of radiation

A mathematical way of saying that the intensity of radiation is inversely proportional to the square of the distance from the source of the radiation.

the surface of the balloon moves away from the light. It has been shown that the intensity is inversely proportional to the square of the distance of the surface from the radiation source. This **inverse square law of radiation** is true only for radiation traveling in a vacuum, but it gives pretty good results when the radiation travels through air. The law is represented by Equation 10.2, where I_x is the intensity at distance d_x and I_y is the intensity at distance d_y :

$$\frac{I_x}{I_y} = \frac{d_y^2}{d_x^2} \quad (10.2)$$

Example 10.6

The intensity of radiation is 18.5 units at a distance of 50 cm from a source. What is the intensity at a distance of 100 cm?

Solution

Let $I_x = 18.5$ units and $d_x = 50$ cm. Then I_y is wanted at $d_y = 100$ cm. Substitution into Equation 10.2 gives

$$\frac{18.5}{I_y} = \frac{(100 \text{ cm})^2}{(50 \text{ cm})^2}$$

or

$$I_y = \frac{(50 \text{ cm})^2(18.5)}{(100 \text{ cm})^2} = 4.63 \text{ units}$$

Thus, doubling the distance cut the intensity to one-fourth the initial intensity.

Learning Check 10.8 The intensity of a radioactive source is measured at a distance of 25 feet and equals 10.0 units. What is the intensity 5 feet from the source?

physical unit of radiation A radiation measurement unit indicating the activity of the source of the radiation; for example, the number of nuclear decays per minute.

biological unit of radiation

A radiation measurement unit indicating the damage caused by radiation in living tissue.

curie A physical unit of radiation measurement corresponding to 3.7×10^{10} nuclear disintegrations per second.

becquerel A physical unit of radiation measurement corresponding to one nuclear disintegration per second.

roentgen A biological unit of radiation measurement used with X-rays and gamma rays; the quantity of radiation that generates 2.1×10^9 ion pairs per 1 cm³ of dry air or 1.8×10^{12} ion pairs per 1 g of tissue.

10.5 Measurement Units for Radiation

Learning Objective

- Describe and compare the units used to measure quantities of radiation.

Two methods, physical and biological, are used to describe quantities of radiation. **Physical units** indicate the activity of a source of radiation, typically in terms of the number of nuclei that decay per unit of time. **Biological units** are related to the damage caused by radiation and account for the fact that a given quantity (or number of particles) of one type of radiation does not have the same damaging effect on tissue as the same quantity of another type of radiation.

The curie and the becquerel are physical units. The **curie** is most widely used at the present time and is equal to 3.7×10^{10} nuclear disintegrations per second. One curie (Ci) represents a large unit, so fractions such as the millicurie (mCi), the microcurie (μ Ci), and the picocurie (pCi) are often used. The **becquerel** (Bq) is a relatively new unit that has achieved little use among U.S. scientists, even though it is smaller and more manageable than the curie. One becquerel is equal to one disintegration per second.

A number of biological units are used, including the roentgen (R), rad (D), gray (Gy), and rem. The **roentgen** is a biological unit used with X-rays or gamma rays. One roentgen is the quantity of radiation that will generate 2.1×10^9 ion pairs in 1 cm³ of dry air at normal temperature and pressure. The ionization effect of X-rays and gamma rays is greater when the radiation passes through tissue; one roentgen generates about 1.8×10^{12} ion pairs per gram of tissue.

Chemistry Around Us 10.1

Medical Imaging



Through the first half of the 20th century, physicians often used exploratory surgery to diagnose abnormalities of internal organs. The technique allowed them to locate and diagnose tumors or malfunctioning organs. The development of various imaging techniques has greatly reduced the number of such surgeries done each year. Medical imaging has its roots in a discovery made in 1895 when Wilhelm Roentgen devised a method to generate invisible rays that could pass through solid matter. Because their nature was unknown at the time, these rays were called *X-rays*. Today, X-ray-generated images are indispensable aids in the practice of medicine and dentistry. However, the penetrating ability of X-rays makes them most effective at imaging dense materials such as bones or teeth, and their usefulness in making soft tissues visible is limited.

The usefulness of X-ray imaging was greatly enhanced in the early 1970s by integrating X-rays with computers. In the resulting technique, called *computed tomography* or CT scanning, the intensity of X-rays that pass through the body is analyzed by a powerful computer. The X-ray intensity depends on the density and opacity of the material through which the X-rays pass. In soft tissues, these differences are very small and cannot be detected by ordinary X-ray techniques. However, the computer is able to distinguish between these small differences and construct an image of the structures through which the X-rays pass. A cross-sectional image is created by moving the X-ray source around the body in a circle. The computer interprets the intensities of the X-rays that pass through and creates a cross-sectional image of organs, bones, and tissues that were in the path of the X-rays. The image is projected onto a TV-like screen where it can be photographed or studied. By moving the X-ray source along the length of the body, an image can be obtained at any point from the head to the feet. The resulting images are like a series of orange slices, where each one shows a cross section of the orange at a different point.

A discovery made in 1946 supplied the means for another step forward in medical imaging. E. M. Purcell and Felix Bloch independently developed a method of detecting hydrogen atoms in matter. They discovered that when atoms of specific isotopes of certain elements (including hydrogen) were placed in a strong magnetic field, the atoms

became capable of absorbing, then re-emitting, radio waves. Many years passed before this discovery was applied to medicine, but today *magnetic resonance imaging* (MRI) scans are routinely used to observe soft tissues without exposing patients to X-rays.

MRI scans are produced using a method similar to CT scans. The patient is placed in a strong magnetic field, which causes the hydrogen atoms in water and other molecules of the body to become radio-wave absorbers. Radio waves are then used like the X-rays of a CT scan. The radio waves that are absorbed and re-emitted by the hydrogen atoms are detected, then analyzed by a computer. The intensity of the radio waves allows the computer to construct an image of the tissue from which the radio waves come. All soft tissue and organs contain large amounts of hydrogen, which makes it possible to obtain MRI scans of essentially any organ in the body. This makes the detection of soft-tissue disorders such as brain tumors much safer and much more comfortable for the patient than was true when exploratory surgery was the only way to get a look inside.



Magnetic resonance imaging (MRI) scans enable medical personnel to view soft tissues in detail.

The rad and the gray are biological units used to describe the effects of radiation in terms of the amount of energy transferred from the radiation to the tissue through which it passes. One **rad** transfers 1×10^{-2} J or 2.4×10^{-3} cal of energy to 1 kg of tissue. One **gray**, a larger unit, transfers 1 J per 1 kg of tissue. From a health point of view, the roentgen and rad are essentially equal. One roentgen of X-rays or gamma rays equals 0.96 rad.

The rem is a biological unit devised to account for health differences in various types of radiation. A 1-rad dose of gamma radiation does not produce the same health effects as a 1-rad dose of alpha radiation. An additive unit was needed so that one unit of alpha had the same health effects as one unit of gamma. The devised unit is the rem, which stands for roentgen equivalent in man. One **rem** of any type of radiation has the same health effect as 1 roentgen of gamma rays or X-rays. The advantage of such a unit is that individuals exposed to low levels of various types of radiation have to keep track only of the rems absorbed from each type and add them up to get the total health effect. If other units were used, the number of units from each type of radiation would

rad A biological unit of radiation measurement corresponding to the transfer of 1×10^{-2} J or 2.4×10^{-3} cal of energy to 1 kg of tissue.

gray A biological unit of radiation measurement corresponding to the transfer of 1 J of energy to 1 kg of tissue.

rem A biological unit of radiation measurement corresponding to the health effect produced by 1 roentgen of gamma or X-rays regardless of the type of radiation involved.



Figure 10.4 Film badges provide a convenient way to monitor the total amount of radiation received during a specific time period.

Table 10.5 Units for Measuring Radiation

Unit	Type of Unit	Relationships Between Units
Curie (Ci)	Physical	$1 \text{ Ci} = 3.7 \times 10^{10} \text{ Bq}$
Becquerel (Bq)	Physical	$1 \text{ Bq} = 2.7 \times 10^{-11} \text{ Ci}$
Roentgen (R)	Biological	$1 \text{ R} = 0.96 \text{ D (tissue)}$
Rad (D)	Biological	$1 \text{ D} = 1 \times 10^{-2} \text{ Gy}$
Gray (Gy)	Biological	$1 \text{ Gy} = 1 \times 10^2 \text{ D}$
Rem	Biological	$1 \text{ rem} = 1 \text{ R X-ray or gamma ray in health effect}$

have to be determined and the health effects calculated for each type. The units used for measuring radiation are summarized in Table 10.5, together with relationships among the units.

Radiation can be detected in a number of ways. Film badges are commonly used by people working in environments where they might be exposed to radiation. The badges contain photographic film that becomes exposed when subjected to radiation. The extent of exposure of the film increases with the amount of radiation. The film is developed after a specific amount of time, and the degree of exposure indicates the radiation dose that has been absorbed during that time (see Figure 10.4).

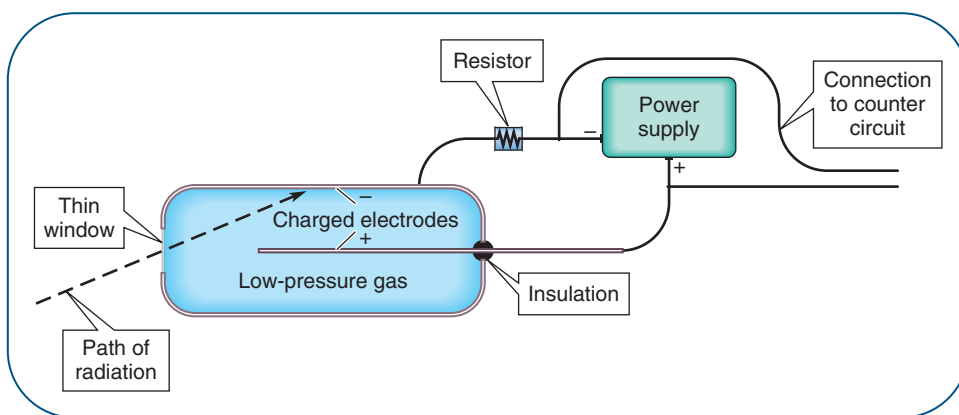
Radiation causes certain substances, called phosphors, to give off visible light. The light appears as brief flashes when radiation strikes a surface coated with a thin layer of phosphors. The number of flashes (called *scintillations*) per unit time is proportional to the amount of radiation striking the surface. Devices employing this approach are called **scintillation counters**.

A third type of detector, called a **Geiger–Müller tube**, is based on the fact that electricity is conducted through a gas that contains charged particles. A Geiger–Müller tube is represented in Figure 10.5. The tube essentially consists of two electrodes charged to high voltages of opposite sign. The tube is filled with a gas at very low pressure. When radiation passes through the tube, it creates a path of charged particles (ions) by knocking electrons off gas molecules. They conduct a brief pulse of current between the electrodes. This brief current is amplified and triggers a counting circuit to record one “count.” The number of counts per unit time is proportional to the amount of radiation passing through the Geiger–Müller tube.

scintillation counter A radiation-detection device operating on the principle that phosphors give off light when struck by radiation.

Geiger–Müller tube A radiation-detection device operating on the principle that ions form when radiation passes through a tube filled with low-pressure gas.

Figure 10.5 A cross section of a detection tube used in Geiger–Müller counters.



10.6 Medical Uses of Radioisotopes

Learning Objective

6. Describe, with examples, medical uses of radioisotopes.

Chemically, radioisotopes undergo the same reactions as nonradioactive isotopes of the same element. For example, radioactive iodine-123 is absorbed by the thyroid gland and used to produce the hormone thyroxine just as if it were ordinary nonradioactive iodine. This characteristic, together with the fact that the location of radioisotopes in the body can be readily detected, makes them useful for diagnostic and therapeutic medical applications. In diagnostic applications, radioisotopes are used as **tracers** whose progress through the body or localization in specific organs can be followed.

To minimize the risks to patients from exposure to radiation, radioisotopes used as diagnostic tracers should have as many of the following characteristics as possible:

1. Tracers should have short half-lives so they will decay while the diagnosis is being done but will give off as little radiation as possible after the diagnostic procedure is completed.
2. The daughter produced by the decaying radioisotope should be nontoxic and give off little or no radiation of its own. Ideally, it will be stable.
3. The radioisotope should have a long enough half-life to allow it to be prepared and administered conveniently.
4. The radiation given off by the radioisotope should be penetrating gamma rays, if possible, to ensure that they can be detected readily by detectors located outside the body.
5. The radioisotope should have chemical properties that make it possible for the tissue being studied to either concentrate it in diseased areas and form a **hot spot** or essentially reject it from diseased areas to form a **cold spot**.

Iodine-123 is used diagnostically to determine whether a thyroid gland is properly functioning. As mentioned earlier, the thyroid gland concentrates iodine and uses it in the production of thyroxine, an iodine-containing hormone. Iodine is an excellent tracer to use because the thyroid is the only user of iodine in the body. Iodine-123 is a good tracer because it is a gamma-emitter with a half-life of 13.3 hours. Iodine-131 is also radioactive and has a short half-life (8 days), but it emits gamma and beta radiation. The beta radiation cannot penetrate through the tissue to be detected diagnostically and therefore is simply an added radiation risk to the patient. For this reason, iodine-123 is preferred. An overactive thyroid absorbs more iodine and forms a hot spot, whereas an underactive or nonactive gland absorbs less iodine and may show a cold spot on a thyroid scan (see Figure 10.6).

Radioisotopes administered internally for therapeutic use should ideally have the following characteristics:

1. The radioisotope should emit less penetrating alpha or beta radiation to restrict the extent of damage to the desired tissue.
2. The half-life should be long enough to allow sufficient time for the desired therapy to be accomplished.
3. The decay products should be nontoxic and give off little or no radiation.
4. The target tissue should concentrate the radioisotope to restrict the radiation damage to the target tissue.

Iodine-131 is used therapeutically to treat thyroid cancer and hyperthyroidism. The primary advantage of this radioisotope is the ability of thyroid tissue to absorb it and localize its effects.

Until the mid-1950s, when improvements in the generation of penetrating X-rays made them the therapy of choice, cobalt-60 was widely used to treat cancer. This radioisotope emits beta and gamma radiation and has a half-life of 5.3 years. When used therapeutically

tracer A radioisotope used medically because its progress through the body or localization in specific organs can be followed.

hot spot Tissue in which a radioactive tracer concentrates.

cold spot Tissue from which a radioactive tracer is excluded or rejected.

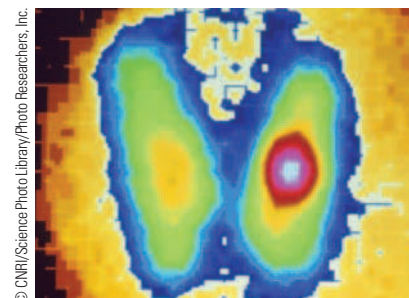


Figure 10.6 A thyroid scan produced after the administration of a radioactive iodine isotope. Is either a hot spot or a cold spot present?

Table 10.6 Examples of Medically Useful Radioisotopes

Isotope	Emission	Half-Life	Applications
^3_1H	Beta	12.3 years	To measure water content of the body
$^{32}_{15}\text{P}$	Beta	14.3 days	Detection of tumors, treatment of a form of leukemia
$^{51}_{24}\text{Cr}$	Gamma	27.8 days	Diagnosis: size and shape of spleen, gastrointestinal disorders
$^{59}_{26}\text{Fe}$	Beta	45.1 days	Diagnosis: anemia, bone marrow function
$^{60}_{27}\text{Co}$	Beta, gamma	5.3 years	Therapy: cancer treatment
$^{67}_{31}\text{Ga}$	Gamma	78.1 hours	Diagnosis: various tumors
$^{75}_{34}\text{Se}$	Beta	120.4 days	Diagnosis: pancreatic scan
$^{81}_{36}\text{Kr}$	Gamma	2.1×10^5 years	Diagnosis: lung ventilation scan
$^{85}_{38}\text{Sr}$	Gamma	64 days	Diagnosis: bone scan
$^{99}_{43}\text{Tc}$	Gamma	6 hours	Diagnosis: brain, liver, kidney, bone, and heart muscle scans
$^{123}_{53}\text{I}$	Gamma	13.3 hours	Diagnosis: thyroid cancer
$^{131}_{53}\text{I}$	Beta, gamma	8.1 days	Diagnosis and treatment: thyroid cancer
$^{197}_{80}\text{Hg}$	Gamma	65 hours	Diagnosis: kidney scan

today, a sample of cobalt-60 is placed in a heavy lead container with a window aimed at the cancerous site. The beam of gamma radiation that exits through the window is focused on a small area of the body where the tumor is located. Table 10.6 contains other examples of radioisotopes used diagnostically and therapeutically.

10.7 Nonmedical Uses of Radioisotopes

Learning Objective

- Describe, with examples, nonmedical uses of radioisotopes.

The study of the photosynthetic process in plants has been made much easier because of the use of radioactive carbon-14 as a tracer. The overall photosynthetic process is represented by Equation 10.3:



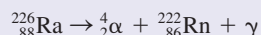
However, the conversion of water (H_2O) and carbon dioxide (CO_2) into carbohydrates ($\text{C}_6\text{H}_{12}\text{O}_6$) involves numerous chemical reactions and carbon-containing intermediate compounds. By using CO_2 that contains radioactive carbon-14, scientists have been able to follow the chemical path of CO_2 through the various intermediate compounds.

Interesting uses for radioactive tracers are found in industry. In the petroleum industry, for example, radioisotopes are used in pipelines to indicate the boundary between different products moving through the lines (see Figure 10.7). In another application, tracer added to fluids moving in pipes makes it easy to detect leaks. Also, the effectiveness of lubricants has been studied by the use of radioactive metal isotopes as components of metal parts; as the metal part wears, the isotope shows up in the lubricant, and the amount of wear can be determined.

Radon: A Chemically Inert Health Risk

Substances that are chemically unreactive are generally not thought to be dangerous to one's health. However, radon, a colorless, odorless, chemically inert noble gas, is considered to be a health risk because it is radioactive. It is believed to be the leading cause of lung cancer among nonsmokers. It has been estimated by U.S. officials that up to 8% of the country's annual lung cancer deaths can be attributed to indoor exposure to radon-222.

Radon-222 is formed during a multistep process in which naturally occurring uranium-238 decays to stable lead-206. During one step of the process, radium-226, with a half-life of 1600 years, emits alpha and gamma radiation to produce radon-222:



Radon-222, with a half-life of 4 days, is produced in rocks and soil that contain even minute amounts of uranium-238. Because it is a gas, radon migrates readily from the soil into the surrounding air. It seeps into houses and other buildings through openings around pipes, and through cracks in basement floors and walls.

The average level of radon in homes is lower than 1 picocurie/L of air. A picocurie (10^{-12} curie) is equivalent to the radioactive decay of 2 radon nuclei per minute. The Environmental Protection Agency recommends an upper limit of 4 picocuries/L for indoor air, and it has been estimated that 99% of the homes in the United States have levels below this value. However, much higher levels have been found. The highest recorded level of radioactivity from radon in a home was 2700 picocuries/L of air. This was found in a home in Pennsylvania

that was built on soil rich in uranium-238. The level of radon in home air can be minimized by sealing the entry points and providing good ventilation.

The hazards associated with breathing radon-222 come not only from the alpha radiation given off as the radon decays but also from the daughter products of that decay. None of these decay products is a gas, but all are radioactive. Because they are not gases, they are not exhaled and become lodged in the lungs, where their radiation can cause serious damage.

It is difficult to decide just how much concern people should have about radon in their homes. Some experts in the field agree with the EPA recommendations, but others question the data on which those recommendations are based. It is quite easy and relatively inexpensive to test for radon in a home, and homes built in locations where the gas has been identified as a problem should be tested.



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Radon detectors are available for home use.

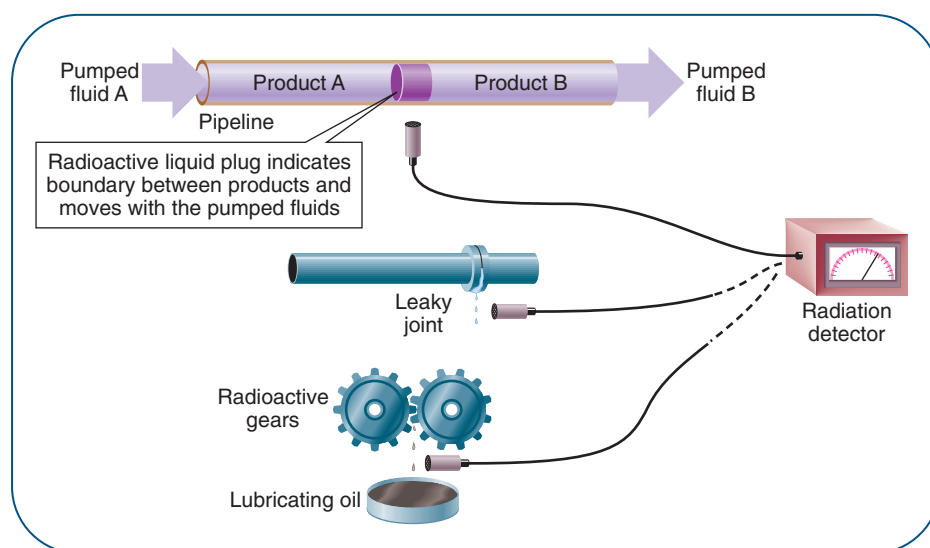


Figure 10.7 Radioactive tracers used in industry.

Radioisotopes are also useful in forms other than tracers. The thickness of manufactured materials, such as metal foils or sheets, can be determined continuously by the use of a radioisotope. The amount of radiation that penetrates the material to the detector is related to the thickness of the material.

An interesting application of radioisotopes involves the use of naturally occurring radioactive materials to determine the ages of artifacts and rocks. Perhaps the most

radioactive dating A process for determining the age of artifacts and rocks, based on the amount and half-life of radioisotopes contained in the object.

widely known example of **radioactive dating** is the use of carbon-14, which is produced in the atmosphere when cosmic-ray neutrons strike nitrogen nuclei. The equation for the reaction is



The resulting ${}^{14}_6\text{C}$ is converted to carbon dioxide ${}^{14}_6\text{CO}_2$ through a chemical reaction with oxygen. In this form, the radioactive carbon is absorbed by plants and converted into cellulose or other carbohydrates during photosynthesis. As long as the plant lives, it takes in ${}^{14}_6\text{CO}_2$. As a result, an equilibrium is established in which the plant contains the same fraction of carbon-14 as the surrounding air. Thus, as long as the plant lives, a constant fraction of the total carbon present is carbon-14. When the plant dies, ${}^{14}_6\text{CO}_2$ intake stops, and the fractional amount of carbon-14 begins to decrease. From a measurement of this amount in, say, a wooden object, and a knowledge of the half-life of carbon-14 (5600 years), it is possible to calculate the age of the object—how long since the tree was cut down to make the object. An object containing only about one-eighth as much carbon-14 as a fresh wood sample from a living tree would be about three half-lives, or 16,800 years old.

This method is limited to objects less than about 50,000 years old because of difficulties in measuring the small amount of carbon-14 present. For this reason, other dating methods have been developed. One method involves potassium-40 and argon-40. The half-life of potassium-40, which absorbs one of its own inner electrons to produce argon-40, is 1.3×10^9 years. By determining the amount of argon-40 in a potassium-containing mineral, it is possible to estimate the age of the mineral. The equation for the reaction is



10.8 Induced Nuclear Reactions

Learning Objective

8. Show that you understand the concept of induced nuclear reactions.

Before 1934, the study of radioactivity was limited to reactions of the relatively few radioisotopes found in nature. In that year, Irene and Frederic Joliot-Curie, French physical chemists, found that radioactivity could be induced in nonradioactive nuclei by bombarding them with small, subatomic particles. They produced an artificial radioactive isotope, nitrogen-13, by bombarding boron-10 with alpha particles from a natural radioisotope:



Notice that a neutron, 1_0n , is produced in addition to ${}^{13}_7\text{N}$. Fifteen years earlier, British physicist Ernest Rutherford had used a similar reaction to produce nonradioactive oxygen-17 from nitrogen-14. A small particle, a proton, was produced in addition to oxygen-17:



A great deal of research has been done since these first experiments were performed, and today it is known that a variety of nuclear reactions can be induced by bombarding either stable or naturally radioactive nuclei with high-energy particles. These artificially induced reactions may lead to one of four results. The first possibility is that the bombarded nucleus may be changed into a different nucleus that is stable. A second possibility is that a new nucleus may be formed that is unstable and undergoes radioactive decay. The other two possibilities are that nuclear fission or nuclear fusion will take place. These are discussed in more detail in Section 10.9. The result of bombardment depends on the kind of nuclei bombarded and the nature and energy of the bombarding particles.

Charged or uncharged particles may be used to bombard nuclei, and each type presents unique problems to overcome. It is believed that a bombarding particle will be drawn

into the nucleus of an atom by strong nuclear forces if it can get to within 10^{-12} cm of the nucleus. Neutrons are captured readily if their speeds are reduced enough to allow the nuclear forces to work. Neutrons are slowed by passing them through materials known as **moderators**. Graphite is an example of a neutron moderator. Charged particles present a different problem. If they are negatively charged ($_{-1}^0\beta$), they are repelled by the electron cloud around the nucleus; for a reaction to take place they must have sufficient energy to overcome this repulsion. Similarly, positive particles must be energetic enough to overcome repulsion by the positively charged nucleus.

Neutrons for bombardment reactions are obtained from other nuclear reactions and, together with proper moderators, are used as they come from the sources. Charged particles are obtained from radioactive materials or from ionization reactions. For example, alpha particles are produced by an ionization reaction in which the two electrons are removed from helium atoms. These charged particles are generally accelerated to high speeds (and energies) before bombarding the intended target nuclei. Two types of particle accelerators, cyclic and linear, are used. The **cyclotron**, one type of cyclic accelerator, is represented in Figure 10.8.

The charged particles to be accelerated enter the evacuated chamber at the center of the cyclotron and, because of the magnetic field, move in a circular path toward the gap between the Dees. Just as the particles reach the gap, the electrical charge on the Dees is adjusted so the particles are repelled by the Dee they are leaving and attracted to the other one. The particles then coast inside the Dee until they again reach a gap, at which point the charges are again adjusted to cause acceleration. This process continues, and the particles are accelerated each time they pass through the gap. As the speed and energy of the particles increase, so does the radius of their circular path, until they finally leave the Dees and strike the target.

In a **linear accelerator**, the particles are accelerated through a series of axially aligned, charged tubes located in an evacuated chamber (see Figure 10.9). Each time a particle passes through a gap and goes from one tube to another, it is accelerated in the same way as particles passing between the Dees of a cyclotron. To make the particle accelerate, the charges on the tubes are reversed at the proper times. The tubes get successively longer to allow the particles the same residence or coasting time as they move toward the end at ever-increasing speeds. After acceleration, the particles exit and strike the target.

moderators Materials capable of slowing down neutrons that pass through them.

cyclotron A cyclic particle accelerator that works by changing electrical polarities as charged particles cross a gap. The particles are kept moving in a spiral path by a strong magnetic field.

linear accelerator A particle accelerator that works by changing electrical polarities as charged particles cross gaps between segments of a long tube.

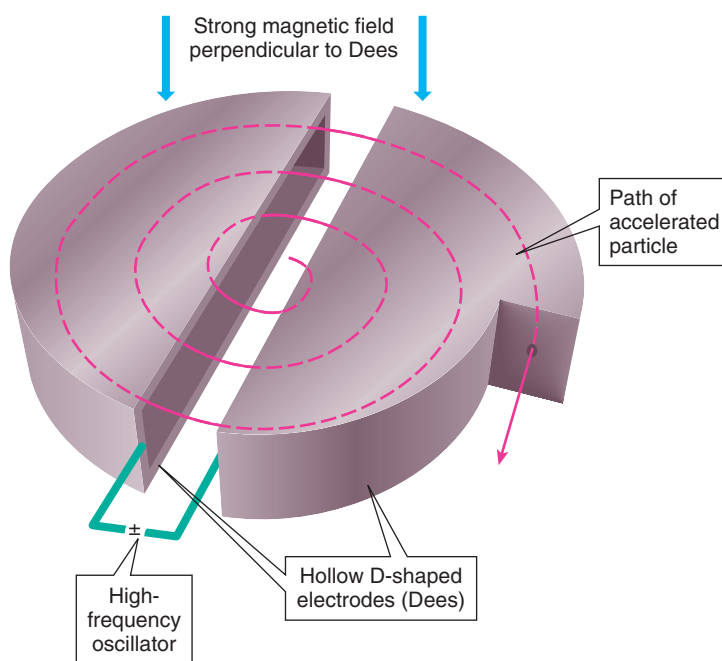
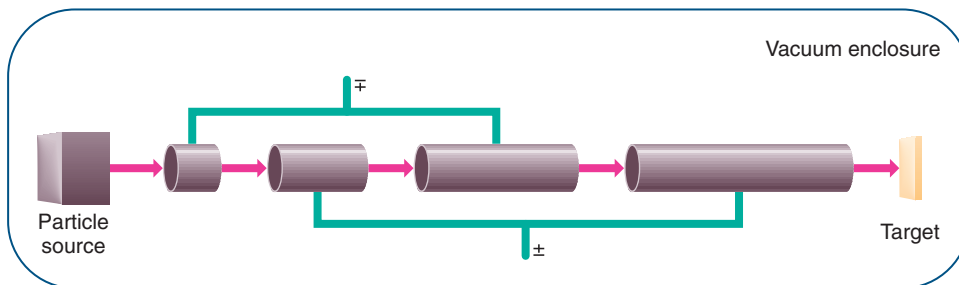
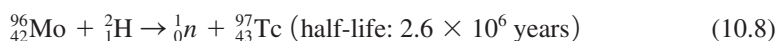


Figure 10.8 The cyclotron. Early versions of the cyclotron were about the size of a dinner plate.

Figure 10.9 A linear particle accelerator.



One of the most interesting results of bombardment is the creation of completely new elements. Four of these elements, produced between 1937 and 1941, filled gaps in the periodic table for which no naturally occurring element had been found. These four are technetium (Tc, number 43), promethium (Pm, number 61), astatine (At, number 85), and francium (Fr, number 87). The equations for the reactions for their production are:



transuranium elements Synthetic elements with atomic numbers greater than that of uranium.

Synthetic elements heavier than uranium, the heaviest naturally occurring element, have also been produced. All these elements, called **transuranium elements**, are radioactive. The first, neptunium (Np, number 93), was produced in 1940; since then at least 19 more (numbers 94 through 112) have been synthesized.

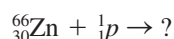
Significant quantities of only a few of these elements have been produced. Plutonium is synthesized in quantities large enough to make it available for use in atomic weapons and reactors. In 1968 the entire world's supply of californium (10^{-4} g) was gathered and made into a target for bombardment by heavy, accelerated particles. These experiments led to the discovery of several elements heavier than californium. Usually, the amounts of the new elements produced are extremely small—sometimes only hundreds of atoms. When the half-life is short, this small amount quickly disappears. Thus, these elements have to be identified not by chemical properties but by the use of instruments that analyze the characteristic radiation emitted by each new element.

Example 10.7

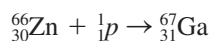
An isotope of gallium (Ga) useful in medical diagnostic work is formed when zinc-66 is bombarded with accelerated protons. Write the equation for this nuclear reaction and determine what isotope of gallium is produced.

Solution

The equation for the reaction is written



where the product is to be determined. By balancing the mass numbers and atomic numbers, we see that the product has a mass number of 67 and an atomic number of 31. Reference to the periodic table shows element number 31 to be gallium, Ga. Therefore, the isotope produced is gallium-67, and the equation for the reaction is:



► **Learning Check 10.9** When manganese-55 is bombarded by protons, a neutron is one product. Write the equation for the process and identify the other product.

10.9 Nuclear Energy

Learning Objective

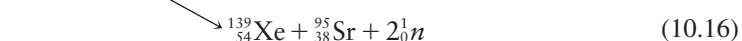
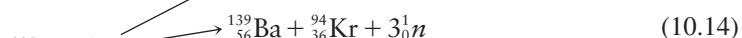
9. Describe the differences between nuclear fission and nuclear fusion.

In 1903, Ernest Rutherford and Frederick Soddy made observations that led them to conclude that the nuclei of all atoms, not just radioactive ones, must contain large quantities of energy. However, the form it took and the means for releasing it remained a mystery. In 1905, Albert Einstein, who was then working as a patent clerk, provided a partial answer. According to his theory of relativity, matter and energy were equivalent and one could be converted into the other. This fact was expressed in his now-famous equation,

$$E = mc^2 \quad (10.12)$$

According to his equation, a small amount of mass, m , would, upon conversion, yield a huge amount of energy, E . The amount of energy liberated is huge because of the size of the multiplier c^2 , the velocity of light squared. The velocity of light is a very large number (3.0×10^8 m/s, or 1.9×10^5 mi/s). The verification of Einstein's theory was not made until 1939, when the work of Otto Hahn, Lise Meitner, Fritz Strassman, and Enrico Fermi led to the discovery of **nuclear fission**. Attempts to produce new heavier transuranium elements by bombarding uranium-235 with neutrons led instead to the production of elements much less massive than uranium-235. The uranium-235 nuclei had undergone fission and split into smaller fragments. In the process, a small amount of matter was converted into energy, just as Einstein had predicted.

Today, we know that uranium-235 is the only naturally occurring isotope that will undergo fission, and a number of different products can result. Equations 10.13–10.16 represent some of the numerous fission reactions that can occur.



For every uranium-235 nucleus undergoing fission, more than one neutron is generated, which opens up the possibility for a chain reaction to take place. If only one of the neutrons produced each time reacts with another uranium-235 nucleus, the process will become a **chain reaction** that continues at a constant rate. If more than one of the neutrons generated per fission reaction produces another reaction, the process becomes an **expanding** or **branching chain reaction** that leads to an explosion. A constant-rate chain reaction is called a **critical reaction**. A branching chain reaction that will lead to an explosion is called **supercritical** (see ► Figure 10.10).

History's first critical chain reaction was carried out in an atomic pile, which was literally a stack or pile of uranium-containing graphite blocks and blocks of pure graphite that served as a neutron moderator (see ► Figure 10.11). The pile was constructed in a squash court beneath the stands of an athletic fieldhouse at the University of Chicago, and the experiments were carried out by a team of scientists led by Italian physicist Enrico Fermi. In December 1942, the team observed that nuclear reactions in the pile had become self-sustaining, or critical. The control rods, composed of strong neutron absorbers, were pushed into the pile to stop the reaction.

Once it had been demonstrated that uranium-235 would undergo self-sustaining fission, the construction of an atomic bomb was fairly simple in principle, but it depended on the concept of critical mass. Neutrons released inside a small piece of fissionable material

nuclear fission A process in which large nuclei split into smaller, approximately equal-sized nuclei when bombarded by neutrons.

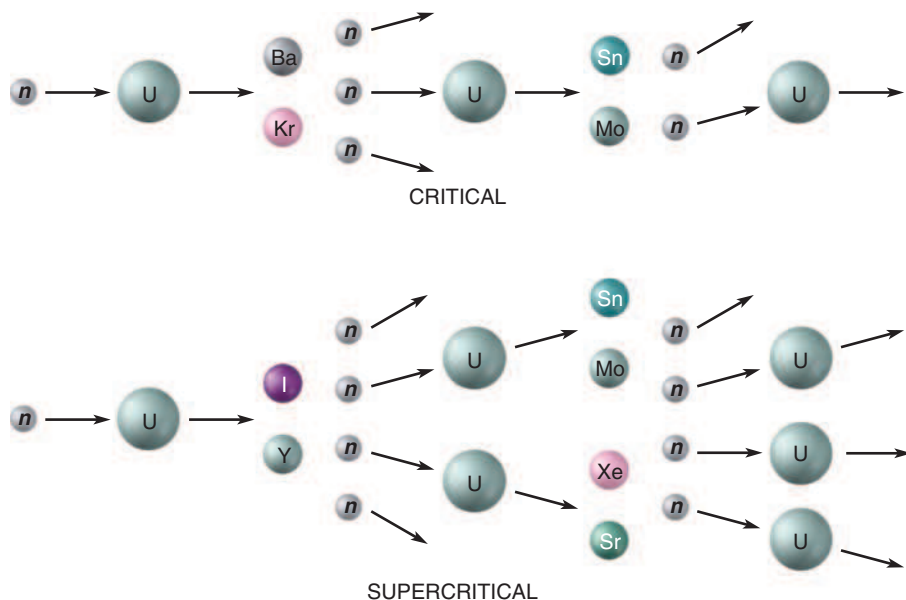
chain reaction A nuclear reaction in which the products of one reaction cause a repeat of the reaction to take place. In the case of uranium fission, neutrons from fission reactions cause other fission reactions to occur.

expanding or branching chain reaction A reaction in which the products of one reaction cause more than one more reaction to occur.

critical reaction A constant-rate chain reaction.

supercritical reaction A branching chain reaction.

Figure 10.10 Nuclear chain reactions.



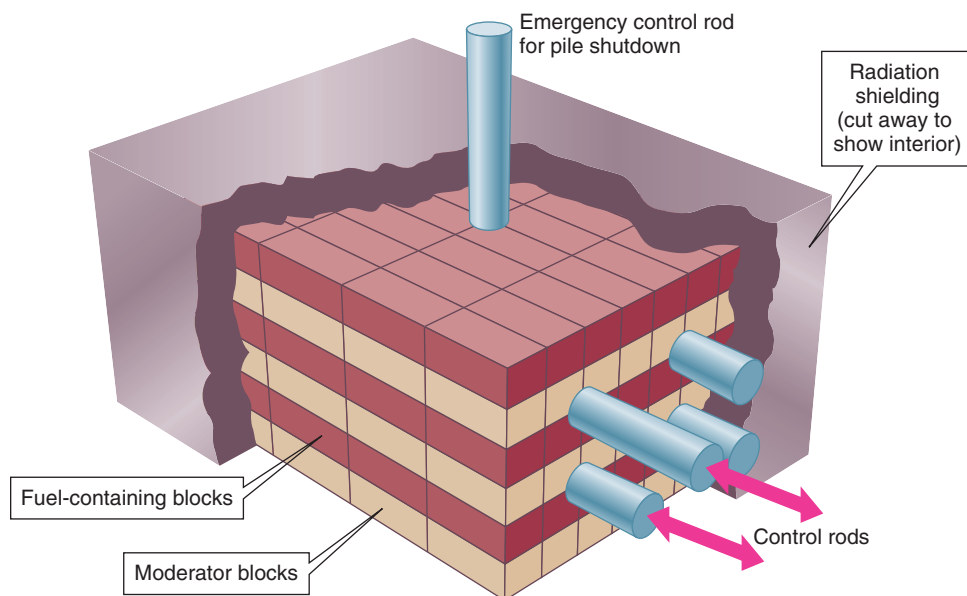
critical mass The minimum amount of fissionable material needed to sustain a critical chain reaction at a constant rate.

supercritical mass The minimum amount of fissionable material that must be present to cause a branching chain reaction to occur.

might escape to the outside before they collide with and split another atom. If the majority of neutrons are lost this way, a chain reaction will not take place. However, as the size of the piece of fissionable material is increased, the chance for neutrons to escape without hitting other atoms decreases. The amount of fissionable material needed to cause a critical reaction to occur is the **critical mass**. The amount needed to cause a branching chain reaction and explosion to occur is called a **supercritical mass**.

During World War II, one major difficulty encountered in building a nuclear bomb was the size of the device needed to bring together subcritical masses to form a supercritical mass. Any bomb used in combat would have to fit into a B-29, the largest bomber available at the time, and could not be heavier than the load capacity of the airplane. The configurations of the first two devices actually used are represented in Figure 10.12. The spherical configuration resulted in a short, wide bomb, while the tubular configuration gave rise to a long, narrow bomb. In each configuration, explosive charges were used to drive subcritical

Figure 10.11 An atomic pile, the world's first nuclear reactor.



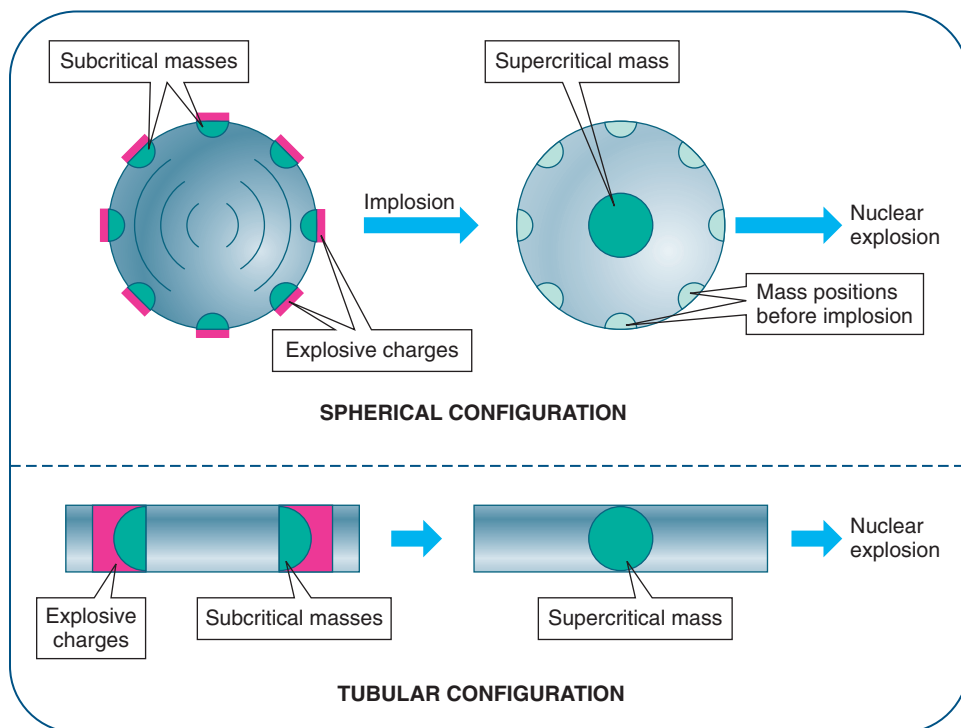


Figure 10.12 Configurations used in fission-type nuclear bombs.

masses together to form a supercritical mass. These nuclear bombs were the first applications of the concepts of subcritical and supercritical mass.

On August 6, 1945, the United States became the first nation to use a nuclear weapon in combat. This bomb used a tubular configuration. Approximately 70,000 inhabitants of Hiroshima, Japan, were killed in the explosion; an equal number were seriously injured. A few days later a second bomb, with a spherical configuration, was dropped on the city of Nagasaki. The casualty figures were similar, and World War II came to an end. The Nagasaki bomb was made of plutonium, an artificial fissionable material produced in atomic piles by breeding reactions.

Breeder reactions are induced reactions in which the product nuclei undergo fission reactions. The plutonium used in the Nagasaki bomb was made by bombarding uranium-238 with neutrons in the first atomic pile. Uranium-238 is the most abundant isotope of uranium found in nature, but it will not undergo fission, as will the much less abundant uranium-235. However, uranium-238 reacts as follows when placed in a neutron-rich environment.



The plutonium-239 will undergo fission and, when collected into a supercritical mass, will cause an explosion. Reactions such as 10.17–10.19 are called breeding reactions because a fissionable material is “bred” from a nonfissionable material.

Curiosity about the seemingly unlimited energy supply of the stars (especially the sun) eventually led scientists to an understanding of a second nuclear process that releases vast amounts of energy. In 1920, Sir Arthur Eddington suggested that the energy of the stars is a byproduct of a reaction in which hydrogen is changed into helium. This was followed

breeder reaction A nuclear reaction in which isotopes that will not undergo spontaneous fission are changed into isotopes that will.

Chemistry and Your Health 10.1

Is Irradiated Food Safe?



In 1993, four deaths and hundreds of illnesses were linked to the eating of undercooked ground beef that was contaminated with *E. coli* bacteria. In 1997 the largest recall of meat products in U.S. history up to that time took place when a producer voluntarily recalled 25 million pounds of ground beef suspected of being contaminated with the same type of bacteria. In 2002, a new meat product recall record was established when 27.4 million pounds of poultry products was recalled because of possible contamination by *Listeria*, a potentially deadly type of bacteria.

It is a well-established fact that treating food by exposing it to radiation destroys disease-causing organisms, kills insects, delays maturation and sprouting of seeds, and slows spoiling. But is irradiated food safe? After years of scientific study, the U.S. government began answering this question with a somewhat-qualified *yes* in 1963, when the Food and Drug Administration (FDA) first allowed radiation-treated wheat and wheat flour to be marketed. Ongoing studies have focused on such areas as the chemical effects of radiation on food, the impact of radiation on nutrient content, and toxicity concerns. Food irradiation for the control of disease-causing organisms has received endorsements and praise from numerous food industry and health organizations.

Since 1963, the list of foods approved for radiation treatment has grown steadily. The effect of the radiation on specific foods is determined by controlling the dose of radiation used. The following table shows the types of foods approved for radiation at this time, along with the approved dose and the desired effect the dose has on the food. The doses are measured in units of kilograys (kGy), which were defined in Table 10.5.

Food	Effect	Dose (kGy)
Spices and dry vegetable seasonings	Decontaminates, controls insects and harmful bacteria	30
Dry or dehydrated enzymes	Controls insects and harmful bacteria	10
All foods	Controls insects	1
Fresh foods	Delays maturation	1
Poultry	Controls harmful bacteria	3
Red meat	Controls spoilage and harmful bacteria	4.5 (fresh) 7 (frozen)

In spite of scientific studies, endorsements from health organizations, and repeated assurances about the safety of irradiated foods, most consumers do not intentionally purchase and use them. Some

irradiated foods are used unintentionally as a result of the labeling rules. For example, irradiated spices or fresh strawberries sold in a grocery store must have labels that indicate they have been irradiated. However, when such irradiated foods are used as ingredients in other foods, the other foods are not required to have irradiation labels. Also, irradiation labeling does not apply to foods served in restaurants.

Two factors that currently contribute to low consumer acceptance of irradiated food are continuing misconceptions about the effects of radiation on food and the low availability of irradiated foods for sale. Some misconceptions that are difficult to eliminate despite reassuring results from many studies are that irradiation causes food to become radioactive, irradiation produces harmful compounds in food, and irradiation reduces the amounts of nutrients in food. Studies indicate that continued education about the safety of irradiated food will eventually overcome this problem. The availability problem stems primarily from the lack of food-irradiation facilities. However, this factor is rapidly becoming less important as more facilities are being built in response to the acceptance of irradiation as a protection against harmful bacteria in food, especially in meat products. Some large meat-packing companies are adding irradiation facilities to existing packing plants, and including irradiation capabilities in plans for new plants, in order to eliminate the costs of shipping their products away from the packing plants to be irradiated.

It appears that food irradiation is becoming more and more acceptable to consumers, and the acceptance will continue to grow. It is likely that in the future, food irradiation will be considered by most consumers to be a normal practice in providing a safe food supply to the consuming public.



The irradiated strawberries on the left have a much longer shelf life than the untreated berries on the right.

thermonuclear reactions Nuclear fusion reactions that require a very high temperature to start.

in 1929 by the concept of **thermonuclear reactions**, nuclear reactions that are started by very high temperatures. According to this idea, at high temperatures the nuclei of very lightweight elements will combine to form nuclei of heavier elements, and in the process some mass is converted into energy. This idea was consistent with Eddington’s suggestion

that hydrogen is converted to helium in the sun and other stars. It was not until nuclear fission reactions became available that high enough temperatures could be achieved to test the theory of **nuclear fusion**. The theory proved to be correct, and today most nuclear weapons in the world are based on fusion reactions.

The fusion reactions responsible for the energy output of the sun are thought to be



The net overall reaction is



which is what Eddington proposed; hydrogen ($\text{}^1_1\text{H}$) is converted to helium ($\text{}^4_2\text{He}$).

A goal of many of the early developers of nuclear energy was to harness it and make it a virtually unlimited source of useful energy. Today, that goal has been partially achieved. Fission reactors generate a small percentage of the electricity used in the world, and researchers are making progress in their attempts to harness fusion reactions.

The essential components of a fission-powered generating plant are shown in Figure 10.13. The use of nuclear power plants has created much controversy. The plants release large amounts of waste heat and thus cause thermal pollution of natural waters. In addition, radioactive wastes create disposal problems—where do you dump radioactive materials that will be a health hazard for thousands of years? Electricity consumers and producers have lost some of their confidence in the safety of nuclear power plants as a result of accidents such as the well-known incidents at Three-Mile Island (1979) and Chernobyl (1986). Another problem is the relatively short supply of fissionable uranium-235. Breeder reactors could solve this problem, since they can convert abundant supplies of uranium-238 and thorium-232 into fissionable fuels, but the breeder reactor program in the United States is at a standstill. As a result of these factors, the future of fission power generation is uncertain at the present time.

In principle, fusion reactors can be used as controlled energy sources just as fission reactors are. Fusion reactors would have many advantages over fission reactors. The deuterium fuel (hydrogen-2) is readily available since it is a minor constituent of all water on Earth, no significant amounts of radioactive wastes would be produced during operation, and the possibility of a serious accident is much smaller. Unfortunately, the feasibility of fusion reactors has not yet been established, although scientists are getting closer through continuing research.

nuclear fusion A process in which small nuclei combine or fuse to form larger nuclei.

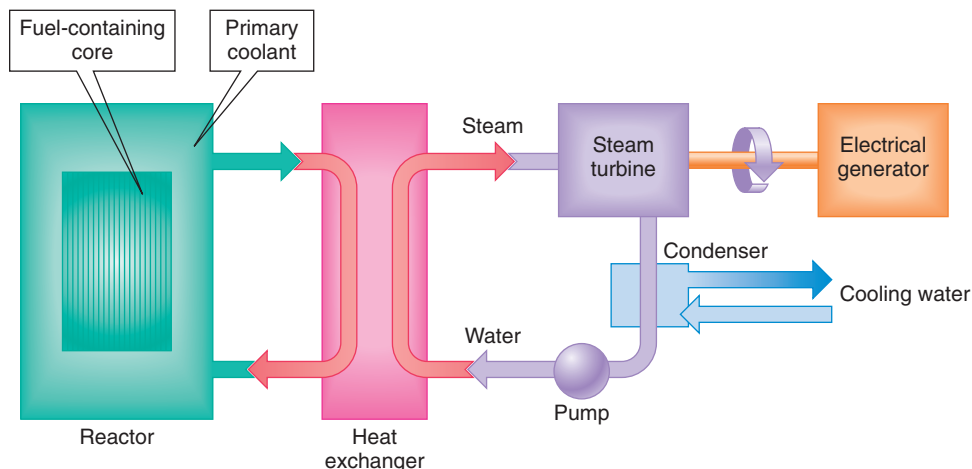


Figure 10.13 Components of a nuclear electricity power plant.

The Do's and Don'ts of Buying Prescription Drugs Online



In our world of modern technology, consumers are relying more and more on the internet to make purchases. Wedding gifts are ordered online, gift wrapped, and shipped directly to the bride and groom. Movies are downloaded to computers for viewing or are selected online and mailed to our homes, making a trip to the video store unnecessary. Often, both time and money are saved by “shopping around” for the best value in the virtual world without ever leaving the comfort of our home or office. However, remember *caveat emptor*—let the buyer beware. When shopping for prescription drugs, your health and maybe even your life are potentially on the line. Although there are many reputable pharmacies offering online convenience, others are questionable. Examples of the questionable practices followed by such pharmacies include distributing expired or illegal drugs, or drugs not approved by the FDA, not checking for drug interactions and even supplying medications other than those specified by a physician.

If medications are ordered online, a few simple guidelines should be followed. First, a professional health care provider should be consulted before ordering or using any medications. This helps eliminate purchasing medications with false claims or those that have harmful interactions with medications currently being taken. Second, a written prescription for any controlled substances should be obtained and the online provider should require a copy of the valid prescription. Third, the standing of the online pharmacies' license should be checked with the National Association of Boards of Pharmacy. Approved sites often bear the seal of VIPPS (Verified Internet Pharmacy Practice Sites). Fourth, if the motivation for ordering online is to save money, the online prices should be checked with local pharmacies to see if the online prices are actually less than the local pharmacy prices. Fifth, before any personal information is divulged, the online provider's website

policies regarding security and the selling of information to other companies should be checked. Sixth, be sure the online pharmacy provides access to a licensed pharmacist as well as contact information such as a phone number and/or physical address for the pharmacy itself. Finally, if the online pharmacy is based out of a country other than the United States, be especially cautious about the quality and legality of the medications, the language on the labels and the potential differences in the names of the provided medications.



Tim Boyle/Getty Images

Buying prescriptions online might save money, but precautions should be followed.

Concept Summary

Radioactive Nuclei. Some nuclei are unstable and undergo radioactive decay. The common types of radiation emitted during decay processes are alpha, beta, and gamma, which can be characterized by mass and charge values.

Objective 1, Exercise 10.2

Equations for Nuclear Reactions. Nuclear reactions can be represented by balanced equations in which the focus is on mass number and atomic number balance on each side. Symbols are used in nuclear equations that make it convenient to balance mass and atomic numbers.

Objective 2, Exercise 10.12

Isotope Half-Life. Different radioisotopes generally decay at different rates, which are indicated by half-lives. One half-life is the time required for one-half of the unstable nuclei in a sample to undergo radioactive decay.

Objective 3, Exercise 10.16

The Health Effects of Radiation. Radiation generates free radicals in tissue as it passes through. Radiation is hazardous even at low intensity if there is long-term exposure. Radiation sickness is caused by short-term intense radiation. Those working around radioactive sources can minimize exposure by using shielding or distance as a protection.

Objective 4, Exercise 10.22

Measurement Units for Radiation. Two systems, physical and biological, are used to describe quantities of radiation. Physical units indicate the number of nuclei of radioactive material that decay per unit of time. Biological units are related to the damage caused by radiation in living tissue. The common physical units are the curie and its fractions and the becquerel. Biological units include the roentgen (for gamma and X-rays), the rad, the gray, and the rem.

Objective 5, Exercise 10.24

Medical Uses of Radioisotopes. Radioisotopes behave chemically like nonradioisotopes of the same element and can be used diagnostically and therapeutically. Diagnostically, radioisotopes are used as tracers whose movement or localization in the body can be followed. Therapeutic radioisotopes localize in diseased areas of the body, where their radiation can destroy diseased tissue.

Objective 6, Exercise 10.30

Nonmedical Uses of Radioisotopes. Nonmedical uses of radioisotopes include (1) tracers in chemical reactions, (2) boundary markers between liquids in pipelines, (3) tracers to detect leaks, (4) metal-wear indicators, and (5) thickness indicators for foils and sheet metal. An especially interesting application is the determination of the ages of artifacts and rocks.

Objective 7, Exercise 10.36

Induced Nuclear Reactions. The bombardment of nuclei with small high-energy particles causes the bombarded nuclei to (1) change into stable nuclei, (2) change into radioactive nuclei, (3) break into smaller pieces (fission), or (4) form larger nuclei (fusion). Most bombarding particles are accelerated by using cyclic or linear accelerators. Bombardment reactions have been used to produce elements that are not found in nature.

Objective 8, Exercise 10.38

Nuclear Energy. During spontaneous nuclear fission reactions, heavy nuclei split when bombarded by neutrons, and release large

amounts of energy. This process was used to produce two atomic bombs whose use ended World War II. A second energy-releasing nuclear process, fusion, is the basis for today's hydrogen bombs. Nuclear fission is in limited use as a source of electrical power; however, this use is controversial. Nuclear fusion has not yet proved feasible as a controlled source of power, but research toward this end continues.

Objective 9, Exercise 10.48


Key Terms and Concepts

Acute radiation syndrome (10.4)	Expanding or branching chain reaction (10.9)	Rad (10.5)
Alpha particle (10.1)	Gamma ray (10.1)	Radical or free radical (10.4)
Becquerel (10.5)	Geiger–Müller tube (10.5)	Radioactive dating (10.7)
Beta particle (10.1)	Gray (10.5)	Radioactive decay (10.1)
Biological unit of radiation (10.5)	Half-life (10.3)	Radioactive nuclei (10.1)
Breeder reaction (10.9)	Hot spot (10.6)	Radioisotope (10.2)
Chain reaction (10.9)	Inverse square law of radiation (10.4)	Rem (10.5)
Cold spot (10.6)	Linear accelerator (10.8)	Roentgen (10.5)
Critical mass (10.9)	Moderators (10.8)	Scintillation counter (10.5)
Critical reaction (10.9)	Nuclear fission (10.9)	Supercritical mass (10.9)
Curie (10.5)	Nuclear fusion (10.9)	Supercritical reaction (10.9)
Cyclotron (10.8)	Physical unit of radiation (10.5)	Thermonuclear reaction (10.9)
Daughter nuclei (10.2)	Positron (10.2)	Tracer (10.6)
Electron capture (10.2)		Transuranium elements (10.8)

Key Equations

- Symbol for isotopes (Section 10.2): A_ZX
- Conversion of a neutron to a proton and electron in the nucleus, leading to beta emission (Section 10.2): ${}_0^1n \rightarrow {}_1^1p + {}_{-1}^0\beta$ Equation 10.1
- Variation of radiation intensity with distance (Section 10.4): $\frac{I_x}{I_y} = \frac{d_y^2}{d_x^2}$ Equation 10.2
- Fission of uranium-235 (Section 10.9): See also Equations 10.14, 10.15, and 10.16. ${}_{92}^{235}\text{U} + {}_0^1n \rightarrow {}_{53}^{135}\text{I} + {}_{39}^{97}\text{Y} + 4{}_0^1n$ Equation 10.13
- Hydrogen fusion reaction of the sun (Section 10.9): $4{}_1^1\text{H} \rightarrow {}_2^4\text{He} + 2{}_1^0\beta + 2\gamma$ Equation 10.23

Exercises

 **OWL** Interactive versions of these problems are assignable in OWL. Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

Radioactive Nuclei (Section 10.1)

- 10.1** Define the term *radioactive*; then criticize the following statements:
- Beta rays are radioactive.
 - Radon is a stable radioactive element.
- 10.2** Group the common nuclear radiations (Table 10.1) into the following categories:
- Those with a mass number of 0
 - Those with a positive charge
 - Those with a charge of 0
- 10.3** Group the common nuclear radiations (Table 10.1) into the following categories:
- Those with a negative charge
 - Those with a mass number greater than 0
 - Those that consist of particles
- 10.4** Characterize the following nuclear particles in terms of the fundamental particles—protons, neutrons, and electrons:
- A beta particle
 - An alpha particle
 - A positron
- 10.5** Discuss how the charge and mass of particles that comprise radiation influence the range or ability of the radiation to penetrate matter.

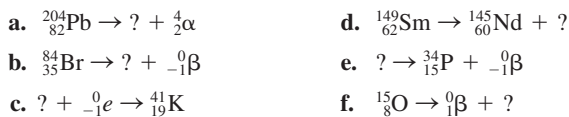
Equations for Nuclear Reactions (Section 10.2)

- 10.6** Summarize how the atomic number and mass number of daughter nuclei compare with the original nuclei after:
- An alpha particle is emitted.
 - A beta particle is emitted.
 - An electron is captured.
 - A gamma ray is emitted.
 - A positron is emitted.
- 10.7** Write appropriate symbols for the following particles using the A_ZX symbolism:
- A tin-117 nucleus
 - A nucleus of the chromium (Cr) isotope containing 26 neutrons
 - A nucleus of element number 20 that contains 24 neutrons
- 10.8** Write appropriate symbols for the following particles using the A_ZX symbolism:
- A nucleus of the element in period 5 and group VB(5) with a mass number of 96
 - A nucleus of element number 37 with a mass number of 80
 - A nucleus of the calcium (Ca) isotope that contains 18 neutrons

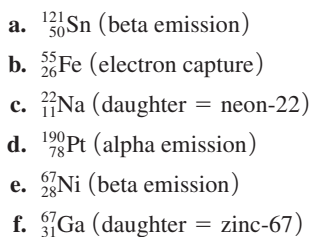
- 10.9** Complete the following equations, using appropriate notations and formulas:



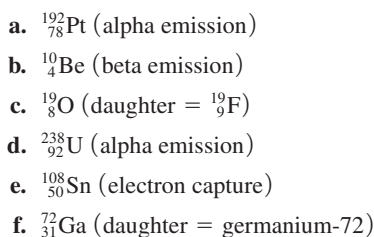
- 10.10** Complete the following equations, using appropriate notations and formulas:



- 10.11** Write balanced equations to represent decay reactions of the following isotopes. The decay process or daughter isotope is given in parentheses.



- 10.12** Write balanced equations to represent decay reactions of the following isotopes. The decay process or daughter isotope is given in parentheses.



Isotope Half-Life (Section 10.3)

- 10.13** What is meant by a half-life?
- 10.14** Describe half-life in terms of something familiar, such as a cake or cookies or your checking account.
- 10.15** An isotope of lead, ${}^{194}_{82}\text{Pb}$ has a half-life of 11 minutes. What fraction of the lead-194 atoms in a sample would remain after 44 minutes had elapsed?
- 10.16** Technetium-99 has a half-life of 6 hours. This isotope is used diagnostically to perform brain scans. A patient is given a 6.0-ng dose. How many nanograms will be present in the patient 30 hours later?
- 10.17** An archaeologist sometime in the future analyzes the iron used in an old building. The iron contains tiny amounts of nickel-63, with a half-life of 92 years. On the basis of the amount of nickel-63 and its decay products found, it is estimated that about 0.78% (1/128) of the original nickel-63 remains. If the building was constructed in 1980, in what year did the archaeologist make the discovery?

- 10.18** An archaeologist unearths the remains of a wooden box, analyzes for the carbon-14 content, and finds that about 93.75% of the carbon-14 initially present has decayed. Estimate the age of the box. The half-life of carbon-14 is 5600 years.
- 10.19** Germanium-66 decays by positron emission, with a half-life of 2.5 hours. What mass of germanium-66 remains in a sample after 10.0 hours if the sample originally weighed 50.0 mg?
- 10.20** A grain sample was found in a cave. The ratio of $^{14}\text{C}/^{12}\text{C}$ was $\frac{1}{8}$ the value in a fresh grain sample. How old was the grain in the cave?

The Health Effects of Radiation (Section 10.4)

- 10.21** A source of radiation has an intensity of 120 units at a distance of 15 feet. How far away from the source would you have to be to reduce the intensity to 20 units?
- 10.22** Compare and contrast the general health effects of long-term exposure to low-level radiation and short-term exposure to intense radiation.
- 10.23** A radiologist found that in a 20-minute period, the dose from a radioactive source was 80 units at a distance of 10.0 m. What would the dose be in the same time period at a distance of 2.00 m from the source?

Measurement Units for Radiation (Section 10.5)

- 10.24** Explain the difference between physical and biological units of radiation and give an example of each type.
- 10.25** Explain why the rem is the best unit to use when evaluating the radiation received by a person working in an area where exposure to several types of radiation is possible.
- 10.26** An individual receives a short-term whole-body dose of 2.8 rads of beta radiation. How many roentgens of X-rays would represent the same health hazard?
- 10.27** Describe how scintillation counters and Geiger–Müller counters detect radiation.
- 10.28** One Ci corresponds to 3.7×10^{10} nuclear disintegrations per second. How many disintegrations per second would take place in a sample containing 3.2 μCi of radioisotope?

Medical Uses of Radioisotopes (Section 10.6)

- 10.29** Explain what a diagnostic tracer is and list the ideal characteristics one should have.
- 10.30** Describe the importance of hot and cold spots in diagnostic work using tracers.
- 10.31** List the ideal characteristics of a radioisotope that is to be administered internally for therapeutic use.
- 10.32** Chromium-51 is used medically to monitor kidney activity. Chromium-51 decays by electron capture. Write a balanced equation for the decay process and identify the daughter that is produced.
- 10.33** Gold-198 is a β emitter used to treat leukemia. It has a half-life of 2.7 days. The dosage is 1.0 mCi/kg body weight. How long would it take for a 70-mCi dose to decay to an activity of 2.2 mCi?

Nonmedical Uses of Radioisotopes (Section 10.7)

- 10.34** A mixture of water (H_2O) and hydrogen peroxide (H_2O_2) will give off oxygen gas when solid manganese dioxide is added as a catalyst. Describe how you could use a tracer to determine if the oxygen comes from the water or the peroxide.
- 10.35** Suppose you are planning a TV commercial for motor oil. Describe how you would set it up to show that your oil prevents engine wear better than a competing brand.
- 10.36** Propose a method for measuring the volume of water in an irregularly shaped swimming pool. You have 1 gallon of water that contains a radioisotope with a long half-life, and a Geiger–Müller counter.
- 10.37** Explain why carbon-14, with a half-life of about 5600 years, is not a good radioisotope to use if you want to determine the age of a coal bed thought to be several million years old.

Induced Nuclear Reactions (Section 10.8)

- 10.38** Write a balanced equation to represent the synthesis of silicon-27 that takes place when magnesium-24 reacts with an accelerated alpha particle. A neutron is also produced.
- 10.39** Americium-241 (Am) results when plutonium-239 captures two neutrons. A beta particle is also produced. Write a balanced equation for the process.
- 10.40** Write a balanced equation to represent the net breeding reaction that occurs when uranium-238 reacts with a neutron to form plutonium-239 and two beta particles.
- 10.41** Explain why neutrons cannot be accelerated in a cyclotron or linear accelerator.
- 10.42** Describe the role of a moderator in nuclear reactions involving neutrons.
- 10.43** To make indium-111 for diagnostic work, silver-109 is bombarded with α particles. What isotope forms when a silver-109 nucleus captures an α particle? Write an equation to represent the process. What two particles must be emitted by the isotope to form indium-111?
- 10.44** To make gallium-67 for diagnostic work, zinc-66 is bombarded with accelerated protons. Write a balanced equation to represent the process when a zinc-66 nucleus captures a single proton.

Nuclear Energy (Section 10.9)

- 10.45** Write one balanced equation that illustrates nuclear fission.
- 10.46** Write one balanced equation that illustrates nuclear fusion.
- 10.47** Explain why critical mass is an important concept in neutron-induced fission chain reactions.
- 10.48** Describe the difference between nuclear fission and nuclear fusion processes. Use nuclear equations to illustrate the difference.
- 10.49** Complete the following equations, which represent two additional ways uranium-235 can undergo nuclear fission.
- a. $^{235}_{92}\text{U} + {}^1_0\text{n} \rightarrow {}^{160}_{62}\text{Sm} + ? + 4{}_^1_0\text{n}$
- b. $^{235}_{92}\text{U} + {}^1_0\text{n} \rightarrow {}^{87}_{35}\text{Br} + ? + 3{}_^1_0\text{n}$

10.50 Plutonium-238 is used to power batteries for pacemakers. It decays by α emission. Write a balanced equation for the process.

Additional Exercises

10.51 Polonium-210 decays by alpha emission and has a half-life of 138 days. Suppose a sample of Po-210 contains 6.02×10^{20} Po-210 atoms and undergoes decay for 138 days. What was the mass in grams of the original sample? Assume all alpha particles formed by the decay escape, and the daughter nuclei remain in the sample as nonradioactive isotopes; then calculate the mass of the sample in grams after 138 days.

10.52 Propose an explanation for the fact that the fusion reactions occurring in the sun only take place at extremely high temperatures.

10.53 Explain how you would make a 10.0-mL sample of 0.100 M NaI that contained iodine in the form of the $^{123}_{53}\text{I}$ isotope.

10.54 Zinc-71 decays by beta emission, and has a half-life of 2.4 minutes. Suppose a 0.200-g sample of Zn-71 was put inside the tube of a Geiger-Müller counter that counted every beta particle emitted by the sample for 2.4 minutes. What would the average reading of the Geiger-Müller counter be in counts per minute during the 2.4 minutes? How many curies is this?

10.55 At one time, nuclear bombs were tested by exploding them above ground. The fallout from such tests contained some Sr-90, a radioactive isotope of strontium. If Sr-90 gets into the food supply, it can become incorporated into the bones of humans and other animals. Explain how this can happen even though strontium is not a normal component of bones.

Allied Health Exam Connection

The following questions are from these sources:

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10.56 What happens when a radioactive element decays?

- a. The nucleus gives off particles.
- b. The ribosomes give off particles.
- c. The mitochondria give off particles.
- d. The Golgi apparatus gives off particles.

10.57 Which of the following particles/rays is not a type of radioactive emanation?

- a. alpha particle
- b. beta particle
- c. gamma ray
- d. carbon-12

10.58 Which of the following forms of radiation contains the least energy?

- a. radio waves
- b. gamma rays
- c. ultraviolet radiation
- d. visible light

10.59 Which of the following is NOT a form of radioactive decay?

- a. electron capture
- b. beta emission
- c. alpha emission
- d. proton emission

10.60 The least penetrating radiation given off by a radioactive substance consists of:

- a. alpha particles
- b. beta particles
- c. gamma rays
- d. X-rays

10.61 Alpha, beta, and gamma radiation are the three major products of natural radioactive decay. Which of the following is (are) the most penetrating radiation?

- a. alpha rays
- b. beta rays
- c. gamma rays
- d. they are all equally penetrating

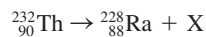
10.62 The isotope $^{234}_{92}\text{U}$ decays by releasing an alpha particle. What is the resulting isotope?

- a. $^{234}_{93}\text{Np}$
- b. $^{233}_{91}\text{Pa}$
- c. $^{234}_{91}\text{Pa}$
- d. $^{230}_{90}\text{Th}$

10.63 Carbon dating involves the decay of a carbon-14 isotope by beta emission. Which of the following equations describes this decay?

- a. $^{14}_6\text{C} \rightarrow ^{13}_5\text{B} + ^1_1\text{H}$
- b. $^{14}_6\text{C} \rightarrow ^{14}_7\text{N} + ^0_{-1}\beta$
- c. $^{14}_6\text{C} \rightarrow ^{13}_5\text{B} + ^1_0\text{n}$
- d. $^{14}_7\text{N} + ^1_0\text{n} \rightarrow ^{14}_6\text{C} + ^1_1\text{H}$

10.64 What is the name of particle X in the following reaction?



- a. deuterium
- b. gamma radiation
- c. beta particle
- d. alpha particle

- 10.65** If element A below gives off an alpha particle, what is the atomic number and mass of the resulting element B? ${}^{210}_{83}\text{A}$
- ${}^{210}_{81}\text{B}$
 - ${}^{206}_{81}\text{B}$
 - ${}^{206}_{83}\text{B}$
 - ${}^{204}_{81}\text{B}$
- 10.66** If element B below gives off a beta particle and gamma rays, what is the resulting element? ${}^{238}_{92}\text{B}$
- ${}^{238}_{93}\text{B}$
 - ${}^{234}_{90}\text{B}$
 - ${}^{239}_{92}\text{B}$
 - ${}^{239}_{91}\text{B}$
- 10.67** The loss of an alpha particle from the radioactive atom ${}^{228}_{88}\text{X}$ would leave:
- ${}^{224}_{86}\text{Rn}$
 - ${}^{222}_{86}\text{Rn}$
 - ${}^{224}_{88}\text{Ra}$
 - ${}^{230}_{90}\text{Th}$
- 10.68** What is the missing product?
- $${}^{42}_{17}\text{X} \rightarrow {}^{42}_{18}\text{Y} + ?$$
- ${}^4_2\text{He}$
 - ${}^0_0\gamma$
 - 0_1e
 - ${}^0_{-1}\beta$
- 10.69** What is the missing product?
- $${}^{60}_{24}\text{A} \rightarrow {}^{60}_{24}\text{B} + ?$$
- ${}^4_2\text{He}$
 - ${}^0_0\gamma$
 - ${}^0_{-1}e$
 - ${}^0_1\beta$
- 10.70** The half-life of a certain radioactive isotope is 5 days. How much of a 100 g sample of this radioactive isotope will remain after 10 days?
- 25
 - 50
 - 75
 - 100
- 10.71** The radioactive isotope I-131 has a half-life of 8 days. What fraction of a sample of I-131 will remain after 24 days?
- $\frac{1}{4}$
 - $\frac{1}{16}$
 - $\frac{1}{8}$
 - $\frac{1}{2}$
- 10.72** If 40 g of a radioactive substance naturally decays to 10 g after 16 days, what is the half-life of the substance?
- 40 days
 - 8 days
 - 16 days
 - 10 days
- 10.73** A dating technique involves electron capture by the potassium-40 isotope according to the following equation: ${}^{40}_{19}\text{K} + {}^0_{-1}e \rightarrow {}^{40}_{18}\text{Ar}$. If the half-life is 1.2×10^9 years, how long does it take for only 10 g to remain of the original 40 g of potassium-40 in a rock sample?
- 1.2×10^9 years
 - 0.6×10^9 years
 - 2.4×10^9 years
 - 1.8×10^9 years
- 10.74** If the half-life of a certain isotope is 5 years, what fraction of a sample of that isotope will remain after 15 years?
- $\frac{1}{5}$
 - $\frac{1}{8}$
 - $\frac{1}{15}$
 - $\frac{1}{20}$
- 10.75** The half-life of a given element is 70 years. How long will it take 5.0 g of this element to be reduced to 1.25 g?
- 70 years
 - 140 years
 - 210 years
 - 35 years
- 10.76** I^{131} has a half-life of eight days. A 100-milligram sample of this radioactive element would decay to what amount after eight days?
- 50 milligrams
 - 40 milligrams
 - 30 milligrams
 - 20 milligrams
- 10.77** If a radioactive element with a half-life ($t^{1/2}$) of 100 years has 31.5 kg remaining after 400 years of decay, the amount in the original sample was close to:
- 2,500 kg
 - 500 kg
 - 50 kg
 - 5,000 kg

- 10.78** Carbon-14 has a half-life of 5.73×10^3 years. If a sample contained 1 gram of C-14, the time required to decay to only 0.0625 g would be:
- 11.46×10^3 years
 - 5.73×10^3 years
 - 22.92×10^3 years
 - none of the above
- 10.79** The time required for $\frac{1}{2}$ of the atoms in a sample of a radioactive element to disintegrate is known as the element's:
- decay period
 - lifetime
 - radioactive period
 - half-life

Chemistry for Thought

- 10.80** One (unrealized) goal of ancient alchemists was to change one element into another (such as lead to gold). Do such changes occur naturally? Explain your reasoning.
- 10.81** Refer to Figure 10.2. Some isotopes used as positron emitters in PET scans are fluorine-18, oxygen-15, and nitrogen-13. What element results in each case when a positron is emitted?
- 10.82** Consider the concept of half-life and decide if, in principle, a radioactive isotope ever completely disappears by radioactive decay. Explain your reasoning.
- 10.83** Do you think Earth is more or less radioactive than it was when it first formed? Explain your answer.
- 10.84** Nuclear wastes typically have to be stored for at least 20 half-lives before they are considered safe. This can be a time of hundreds or thousands of years for some isotopes. With that in mind, would you consider sending such wastes into outer space a responsible solution to the nuclear waste disposal problem? Explain your answer.
- 10.85** Read Chemistry Around Us 10.2 and write an equation to represent the radioactive decay of radon-222. Then, write an equation to represent the decay of the daughter produced by the radon decay. The daughter decays by alpha emission. Then, write an equation for the decay of the daughter produced by this second reaction, which decays by beta emission. In each of the three reactions, assume that only one particle in addition to the daughter is produced. What element is radon converted into by this series of three decays?
- 10.86** Uranium-238 is the most abundant naturally occurring isotope of uranium. It undergoes radioactive decay to form other isotopes. In the first three steps of this decay, uranium-238 is converted to thorium-234, which is converted to protactinium-234, which is converted to uranium-234. Assume that only one particle in addition to the daughter is produced in each step, and use equations for the decay processes to determine the type of radiation emitted during each step.

Organic Compounds: Alkanes

11




Pharmacists are responsible for the appropriate dispensing of drugs and other medications. Pharmacists must have a broad knowledge of the many drugs and medications available, as well as the effects they have on the human body when administered individually and in combinations. The great majority of drugs are organic compounds. This and the next several chapters provide the basis for understanding the structure, properties, and physiological effects of organic compounds.

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Learning Objectives

When you have completed your study of this chapter, you should be able to:

- 1 Show that you understand the general importance of organic chemical compounds. (**Section 11.1**)
- 2 Be able to recognize the molecular formulas of organic and inorganic compounds. (**Section 11.1**)
- 3 Explain some general differences between inorganic and organic compounds. (**Section 11.2**)
- 4 Be able to use structural formulas to identify compounds that are isomers of each other. (**Section 11.3**)
- 5 Write condensed or expanded structural formulas for compounds. (**Section 11.4**)
- 6 Classify alkanes as normal or branched. (**Section 11.5**)
- 7 Use structural formulas to determine whether compounds are structural isomers. (**Section 11.6**)
- 8 Assign IUPAC names and draw structural formulas for alkanes. (**Section 11.7**)
- 9 Assign IUPAC names and draw structural formulas for cycloalkanes. (**Section 11.8**)
- 10 Name and draw structural formulas for geometric isomers of cycloalkanes. (**Section 11.9**)
- 11 Describe the key physical properties of alkanes. (**Section 11.10**)
- 12 Write alkane combustion reactions. (**Section 11.11**)

 Online homework for this chapter may be assigned in OWL.

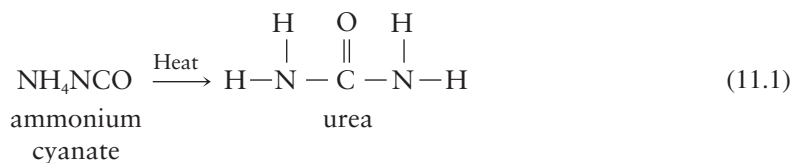
The word *organic* is used in several different contexts. Scientists of the 18th and 19th centuries studied compounds extracted from plants and animals and labeled them “organic” because they had been obtained from organized (living) systems. Organic fertilizer is organic in the original sense that it comes from a living organism. There is no universal definition of organic foods, but the term is generally taken to mean foods grown without the application of pesticides or synthetic fertilizers. When referring to organic chemistry, however, we mean the chemistry of carbon-containing compounds.

11.1 Carbon: The Element of Organic Compounds

Learning Objectives

1. Show that you understand the general importance of organic chemical compounds.
2. Be able to recognize the molecular formulas of organic and inorganic compounds.

Early chemists thought organic compounds could be produced only through the action of a “vital force,” a special force active only in living organisms. This idea was central to the study of organic chemistry until 1828, because up to that time, no one had been able to synthesize an organic compound from its elements or from naturally occurring minerals. In that year, Friedrich Wöhler, a German chemist, heated an inorganic salt called ammonium cyanate and produced urea. This compound, normally found in blood and urine, was unquestionably organic, and it had come from an inorganic source. The reaction is



Wöhler’s urea synthesis discredited the “vital force” theory, and his success prompted other chemists to attempt to synthesize organic compounds. Today, organic compounds are being synthesized in thousands of laboratories, and most of the synthetics have never been isolated from natural sources.

Organic compounds share one unique feature: They all contain carbon. Therefore, **organic chemistry** is defined as the study of carbon-containing compounds. There are a few exceptions to this definition; a small number of carbon compounds—such as CO, CO₂, carbonates, and cyanides—were studied before Wöhler’s urea synthesis. These were classified as inorganic because they were obtained from nonliving systems, and even though they contain carbon, we still consider them to be a part of **inorganic chemistry**.

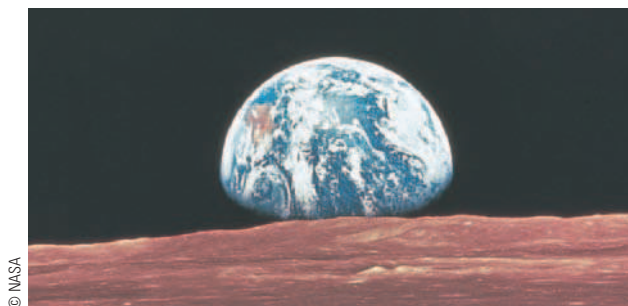
The importance of carbon compounds to life on Earth cannot be overemphasized. If all carbon compounds were removed from Earth, its surface would be somewhat like the barren surface of the moon (see ► Figure 11.1). There would be no animals, plants, or

organic compound A compound that contains the element carbon.

organic chemistry The study of carbon-containing compounds.

inorganic chemistry The study of the elements and all noncarbon compounds.

Figure 11.1 Organic chemistry makes a tremendous difference when comparing the physical makeup of the Earth and the moon.



any other form of life. If carbon-containing compounds were removed from the human body, all that would remain would be water, a very brittle skeleton, and a small residue of minerals. Many of the essential constituents of living matter—such as carbohydrates, fats, proteins, nucleic acids, enzymes, and hormones—are organic chemicals.

The essential needs of daily human life are food, fuel, shelter, and clothing. The principal components of food (with the exception of water) are organic. The fuels we use (e.g., wood, coal, petroleum and natural gas) are mixtures of organic compounds. Our homes typically incorporate wood construction, and our clothing, whether made of natural or synthetic fibers, is organic.

Besides the major essentials, many of the smaller everyday things often taken for granted are also derived from carbon and its compounds. Consider an ordinary pencil. The “lead” (actually graphite), the wood, the rubber eraser, and the paint on the surface are all either carbon or carbon compounds. The paper in this book, the ink on its pages, and the glue holding it all together are also made of carbon compounds.

11.2 Organic and Inorganic Compounds Compared

Learning Objective

3. Explain some general differences between inorganic and organic compounds.

It is interesting that the subdivision of chemistry into its organic and inorganic branches results in one branch that deals with compounds composed mainly of one element and another branch that deals with compounds formed by the more than 100 remaining elements. However, this classification seems more reasonable when we recognize that known organic compounds are much more numerous than inorganic compounds. An estimated 500,000 inorganic compounds have been identified, but more than 9 million organic compounds are known, and thousands of new ones are synthesized or isolated each year.

One of the reasons for the large number of organic compounds is the unique ability of carbon atoms to form stable covalent bonds with other carbon atoms and with atoms of other elements. The resulting covalently bonded molecules may contain as few as one or more than a million carbon atoms.

In contrast, inorganic compounds are often characterized by the presence of ionic bonding. Covalent bonding also may be present, but it is less common. These differences generally cause organic and inorganic compounds to differ physically (see ► Figure 11.2) and chemically, as shown in ► Table 11.1.

Table 11.1 Properties of Typical Organic and Inorganic Compounds

Property	Organic Compounds	Inorganic Compounds
Bonding within molecules	Covalent	Often ionic
Forces between molecules	Generally weak	Quite strong
Normal physical state	Gases, liquids, or low-melting-point solids	Usually high-melting-point solids
Flammability	Often flammable	Usually nonflammable
Solubility in water	Often low	Often high
Conductivity of water solutions	Nonconductor	Conductor
Rate of chemical reactions	Usually slow	Usually fast

Figure 11.2 Many organic compounds, such as ski wax, have relatively low melting points. What does this fact reveal about the forces between organic molecules?



► **Learning Check 11.1** Classify each of the following compounds as organic or inorganic:

- | | |
|----------------------------------|--------------------------------------|
| a. NaCl | d. NaOH |
| b. CH ₄ | e. CH ₃ OH |
| c. C ₆ H ₆ | f. Mg(NO ₃) ₂ |

► **Learning Check 11.2** Decide whether each of the following characteristics most likely describes an organic or inorganic compound:

- | | | |
|--------------|----------------------|---------------------|
| a. Flammable | b. Low boiling point | c. Soluble in water |
|--------------|----------------------|---------------------|

Study Skills 11.1 Changing Gears for Organic Chemistry

You will find that organic chemistry is very different from general or inorganic chemistry. By quickly picking up on the changes, you will help yourself prepare for quizzes and exams.

There is almost no math in these next six chapters or in the biochemistry section. Very few mathematical formulas need to be memorized. The problems you will encounter fall mainly into four categories: naming compounds and drawing structures, describing physical properties of substances, writing reactions, and identifying typical uses of compounds. This pattern holds true for all six of the organic chemistry chapters.

The naming of compounds is introduced in this chapter, and the rules developed here will serve as a starting point in the next

five chapters. Therefore, it is important to master naming in this chapter. A well-developed skill in naming will help you do well on exams covering the coming chapters.

Only a few reactions are introduced in this chapter, but many more will be in future chapters. Writing organic reactions is just as important (and challenging) as naming, and Study Skills 12.1 will help you. Identifying the uses of compounds can best be handled by making a list as you read the chapter or by highlighting compounds and their uses so that they are easy to review. All four categories of problems are covered by numerous end-of-chapter exercises to give you practice.

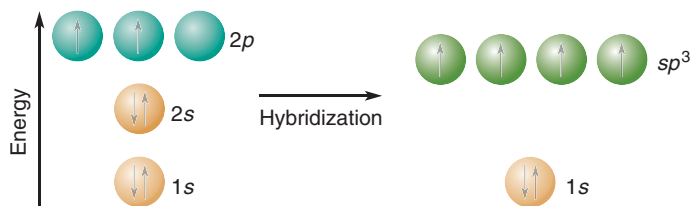


Figure 11.3 The mixing of 2s and three 2p orbitals of a carbon atom to generate four sp^3 carbon orbitals, each with energy intermediate between 2s and 2p orbitals.

11.3 Bonding Characteristics and Isomerism

Learning Objective

- Be able to use structural formulas to identify compounds that are isomers of each other.

There are two major reasons for the astonishing number of organic compounds: the bonding characteristics of carbon atoms, and the isomerism of carbon-containing molecules. As a group IVA(14) element, a carbon atom has four valence electrons. Two of these outermost-shell electrons are in an s orbital, and two are in p orbitals (see Section 3.4):



With only two unpaired electrons, we might predict that carbon would form just two covalent bonds with other atoms. Yet, we know from the formula of methane (CH_4) that carbon forms four bonds.

Linus Pauling (1901–1994), winner of the Nobel Prize in chemistry (1954) and Nobel Peace Prize (1963), developed a useful model to explain the bonding characteristics of carbon. Pauling found that a mathematical mixing of the 2s and three 2p orbitals could produce four new, equivalent orbitals (see ▶ Figure 11.3). Each of these **hybrid orbitals** has the same energy and is designated sp^3 . An sp^3 orbital has a two-lobed shape, similar to the shape of a p orbital but with different-sized lobes (see ▶ Figure 11.4). Each of the four sp^3 hybrid orbitals contains a single unpaired electron available for covalent bond formation. Thus, carbon forms four bonds.

Each carbon–hydrogen bond in methane arises from an overlap of a C (sp^3) and an H (1s) orbital. The sharing of two electrons in this overlap region creates a sigma (σ) bond. The four equivalent sp^3 orbitals point toward the corners of a regular tetrahedron (see ▶ Figure 11.5).

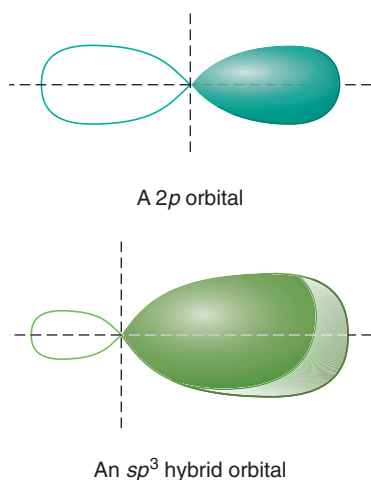


Figure 11.4 A comparison of unhybridized p and sp^3 hybridized orbital shapes. The atomic nucleus is at the junction of the lobes in each case.

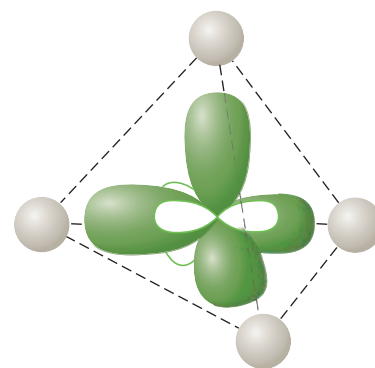
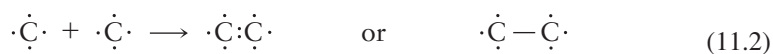
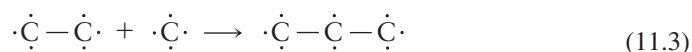


Figure 11.5 Directional characteristics of sp^3 hybrid orbitals of carbon and the formation of C—H bonds in methane (CH_4). The hybrid orbitals point toward the corners of a regular tetrahedron. Hydrogen 1s orbitals are illustrated in position to form bonds by overlap with the major lobes of the hybrid orbitals.

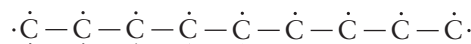
Carbon atoms also have the ability to bond covalently to other carbon atoms to form chains and networks. This means that two carbon atoms can join by sharing two electrons to form a single covalent bond:



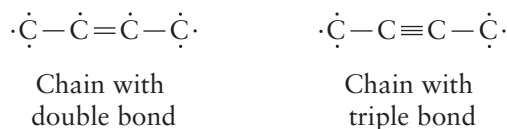
A third carbon atom can join the end of this chain:



This process can continue and form carbon chains of almost any length, such as



The electrons not involved in forming the chain can be shared with electrons of other carbon atoms (to form chain branches) or with electrons of other elements such as hydrogen, oxygen, or nitrogen. Carbon atoms may also share more than one pair of electrons to form multiple bonds:



In principle, there is no limit to the number of carbon atoms that can bond covalently. Thus, organic molecules range from the simple molecules such as methane (CH_4) to very complicated molecules containing over a million carbon atoms.

The variety of possible carbon atom arrangements is even more important than the size range of the resulting molecules. The carbon atoms in all but the very simplest organic molecules can bond in more than one arrangement, giving rise to different compounds with different structures and properties. This property, called **isomerism**, is characterized by compounds that have identical molecular formulas but different arrangements of atoms. One type of isomerism is characterized by compounds called **structural isomers**. Another type of isomerism is introduced in Section 11.9.

isomerism A property in which two or more compounds have the same molecular formula but different arrangements of atoms.

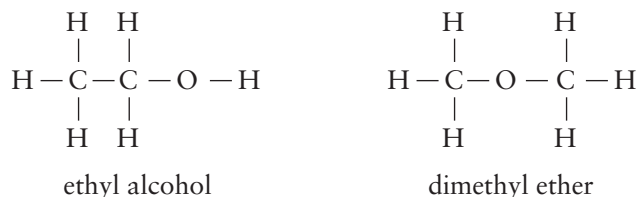
structural isomers Compounds that have the same molecular formula but in which the atoms bond in different patterns.

Example 11.1

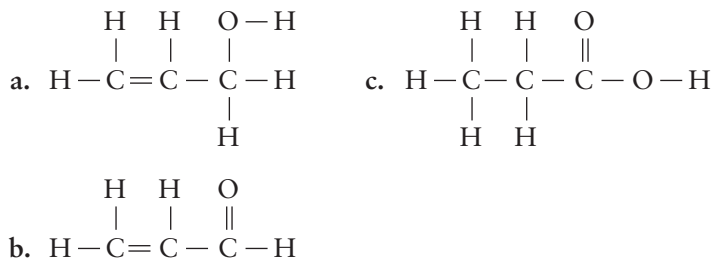
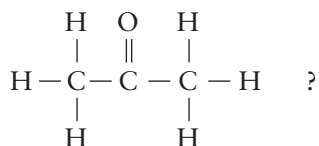
Use the usual rules for covalent bonding to show that a compound with the molecular formula $\text{C}_2\text{H}_6\text{O}$ demonstrates the property of isomerism. Draw formulas for the isomers, showing all covalent bonds.

Solution

Carbon forms four covalent bonds by sharing its four valence-shell electrons. Similarly, oxygen should form two covalent bonds, and hydrogen a single bond. On the basis of these bonding relationships, two structural isomers are possible:



► **Learning Check 11.3** Which one of the structures below represents a structural isomer of



The two isomers of Example 11.1 are quite different. Ethyl alcohol (grain alcohol) is a liquid at room temperature, whereas dimethyl ether is a gas. As we've seen before, the structural differences exert a significant influence on properties. From this example, we can see that molecular formulas such as $\text{C}_2\text{H}_6\text{O}$ provide much less information about a compound than do structural formulas. ► Figure 11.6 shows ball-and-stick models of these two molecules.

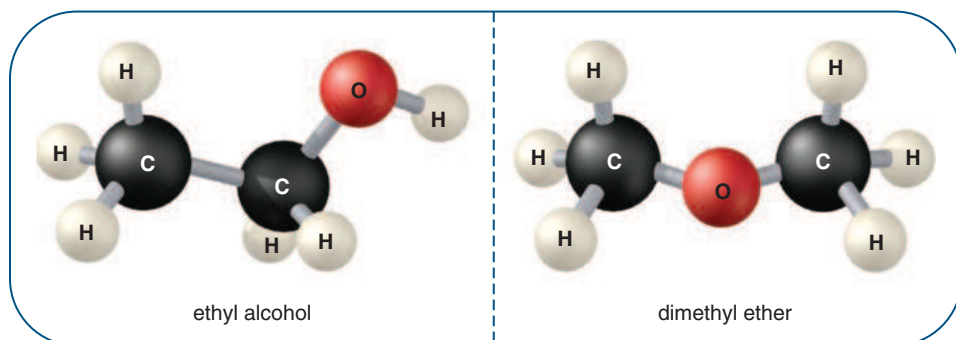
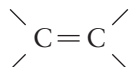
As the number of carbon atoms in the molecular formula increases, the number of possible isomers increases dramatically. For example, 366,319 different isomers are possible for a molecular formula of $\text{C}_{20}\text{H}_{42}$. No one has prepared all these isomers or even drawn their structural formulas, but the number helps us understand why so many organic compounds have been either isolated from natural sources or synthesized.

11.4 Functional Groups: The Organization of Organic Chemistry

Learning Objective

5. Write condensed or expanded structural formulas for compounds.

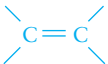

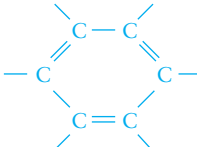

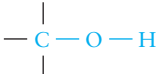
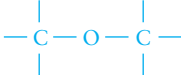
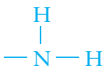
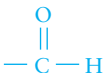
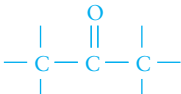
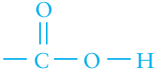
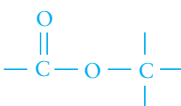

Because of the enormous number of possible compounds, the study of organic chemistry might appear to be hopelessly difficult. However, the arrangement of organic compounds into a relatively small number of classes can simplify the study a great deal. This organization is done on the basis of characteristic structural features called **functional groups**. For example, compounds with a carbon-carbon double bond



functional group A unique reactive combination of atoms that differentiates molecules of organic compounds of one class from those of another.

Figure 11.6 Ball-and-stick models of the isomers of $\text{C}_2\text{H}_6\text{O}$. Ethyl alcohol is a liquid at room temperature and completely soluble in water, whereas dimethyl ether is a gas at room temperature and only partially soluble in water.

Table 11.2 Classes and Functional Groups of Organic Compounds

Class	Functional Group	Example of Expanded Structural Formula	Example of Condensed Structural Formula	Common Name
Alkane	None	$ \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} $	CH_3CH_3	ethane
Alkene		$ \begin{array}{c} \text{H} \quad \text{H} \\ \diagdown \quad \diagup \\ \text{C}=\text{C} \\ \diagup \quad \diagdown \\ \text{H} \quad \text{H} \end{array} $	$\text{H}_2\text{C}=\text{CH}_2$	ethylene
Alkyne		$ \begin{array}{c} \text{H}-\text{C}\equiv\text{C}-\text{H} \end{array} $	$\text{HC}\equiv\text{CH}$	acetylene
Aromatic		$ \begin{array}{c} \text{H} \quad \text{H} \\ \diagdown \quad \diagup \\ \text{C}=\text{C} \\ \diagup \quad \diagdown \\ \text{H}-\text{C} \quad \text{C}-\text{H} \\ \diagdown \quad \diagup \\ \text{C}=\text{C} \\ \diagup \quad \diagdown \\ \text{H} \quad \text{H} \end{array} $		benzene
Alcohol		$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array} $	$\text{CH}_3\text{CH}_2-\text{OH}$	ethyl alcohol
Ether		$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{O}-\text{C}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array} $	$\text{CH}_3-\text{O}-\text{CH}_3$	dimethyl ether
Amine		$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{N}-\text{H} \\ \\ \text{H} \end{array} $	CH_3-NH_2	methylamine
Aldehyde		$ \begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{H} \\ \\ \text{H} \end{array} $	$ \begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{H} \end{array} $	acetaldehyde
Ketone		$ \begin{array}{c} \text{H} \quad \text{O} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $	$ \begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{CH}_3 \end{array} $	acetone
Carboxylic acid		$ \begin{array}{c} \text{H} \quad \text{O} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} \\ \\ \text{H} \end{array} $	$ \begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{OH} \end{array} $	acetic acid
Ester		$ \begin{array}{c} \text{H} \quad \text{O} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $	$ \begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{O}-\text{CH}_3 \end{array} $	methyl acetate
Amide		$ \begin{array}{c} \text{H} \quad \text{O} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{N}-\text{H} \\ \\ \text{H} \end{array} $	$ \begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{NH}_2 \end{array} $	acetamide

are classified as alkenes. The major classes and functional groups are given in Table 11.2. Notice that each functional group in Table 11.2 (except for alkanes) contains a multiple bond or at least one oxygen or nitrogen atom.

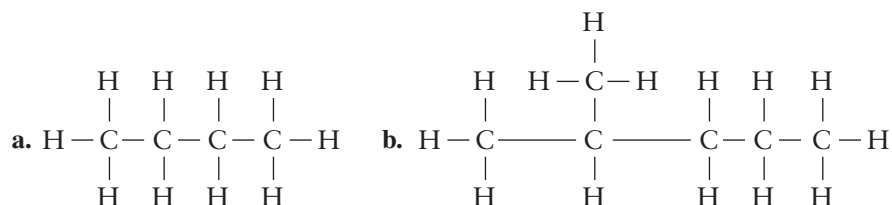
In Table 11.2, we have used both expanded and condensed structural formulas for the compounds. **Expanded structural formulas** show all covalent bonds, whereas **condensed structural formulas** show only specific bonds. You should become familiar with both types, but especially with condensed formulas because they will be used often.

expanded structural formula A structural molecular formula showing all the covalent bonds.

condensed structural formula A structural molecular formula showing the general arrangement of atoms but without showing all the covalent bonds.

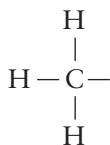
Example 11.2

Write a condensed structural formula for each of the following compounds:



Solution

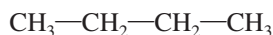
a. Usually the hydrogens belonging to a carbon are grouped to the right. Thus, the group



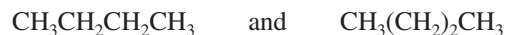
condenses to CH_3- , and



condenses to $-\text{CH}_2-$. Thus, the formula condenses to

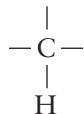


Other acceptable condensations are



Parentheses are used here to denote a series of two $-\text{CH}_2-$ groups.

b. The group



condenses to $-\text{CH}-$. The condensed formula is therefore

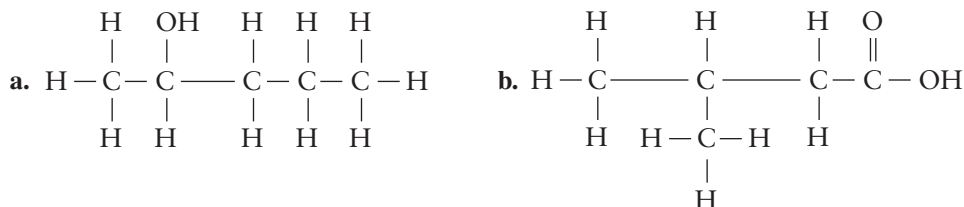


or



In the last form, the first set of parentheses indicates two identical CH_3 groups attached to the same carbon atom, and the second set of parentheses denotes two CH_2 groups.

► **Learning Check 11.4** Write a condensed structural formula for each of the following compounds. Retain the bonds to and within the functional groups.



11.5 Alkane Structures

Learning Objective

6. Classify alkanes as normal or branched.

hydrocarbon An organic compound that contains only carbon and hydrogen.

saturated hydrocarbon Another name for an alkane.

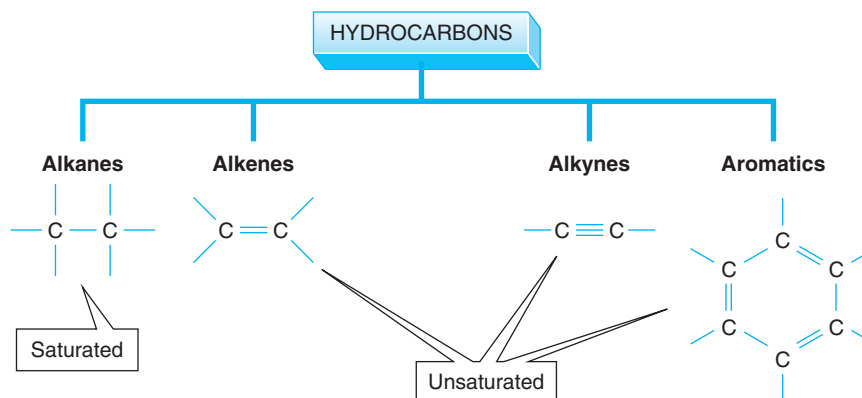
alkane A hydrocarbon that contains only single bonds.

Hydrocarbons, the simplest of all organic compounds, contain only two elements, carbon and hydrogen. **Saturated hydrocarbons** or **alkanes** are organic compounds in which carbon is bonded to four other atoms by single bonds; there are no double or triple bonds in the molecule. Unsaturated hydrocarbons, studied later, are called alkenes, alkynes, and aromatics and contain double bonds, triple bonds, or six-carbon rings, as shown in ► Figure 11.7.

Most life processes are based on the reactions of functional groups. Since alkanes have no functional group, they are not abundant in the human body. However, most compounds in human cells contain parts consisting solely of carbon and hydrogen that behave very much like hydrocarbons. Thus, to understand the chemical properties of the more complex biomolecules, it is useful to have some understanding of the structure, physical properties, and chemical behavior of hydrocarbons.

Another important reason for becoming familiar with the characteristics of hydrocarbons is the crucial role they play in modern industrial society. We use naturally occurring hydrocarbons as primary sources of energy and as important sources of raw materials for the manufacture of plastics, synthetic fibers, drugs, and hundreds of other compounds used daily (see ► Figure 11.8).

Figure 11.7 Classification of hydrocarbons.



Organic Foods: Are They Safer? More Nutritious?



The answer to the question posed by the title is mixed. The United States Department of Agriculture (USDA) makes no claim that organically produced food is safer or more nutritious than conventionally produced food. Experts in the organic food producing industry seem to agree with the USDA. They point out that the word *organic* refers only to a method of food production. In December 2001, the USDA standardized the way the word *organic* can be used in food labeling. According to the USDA, a food product labeled *100% organic* must contain only ingredients that meet the following requirements: No genetic engineering, ionizing radiation, sewage-sludge fertilizer, or synthesized antibiotics, pesticides, hormones, or fertilizers can be used in their production. In order for a food product to be labeled *95% organic*, at least 95% of the ingredients must meet this definition, and the label *made with organic ingredients* can only be used on food products that contain a minimum of 70% organic ingredients.

Consumers of organic foods have a different answer to the question. Some say they use organic foods because of a combination of environmental and personal health concerns, while a larger number use good flavor as their primary reason. Those who are concerned about their health feel that organic products are better for them because no pesticides, growth hormones, antibiotics, or other synthesized chemicals were used in the food production and so cannot remain in the food as a residue.

While it is true that conventionally produced food may contain residues of such things as pesticides that are known to be toxic in high doses, there is no scientific evidence that they cause health problems when ingested in the quantities found on conventional food products. Some researchers feel that the concern over pesticide residues is misplaced because food-borne bacteria are a much greater health hazard than pesticide residues, and organic farming techniques that use no antibiotics are more likely to produce food carrying disease-causing organisms than are conventional techniques.

There is supporting evidence for those who say organic foods taste better. Organically-grown fruits and vegetables are allowed to ripen

naturally on the tree or vine, a practice generally recognized to improve flavor over produce that is picked green and ripened artificially. Also, such produce must be transported to market quickly to avoid spoiling, and so tends to be fresher when consumed. Proponents of organic food also point out that the lack of pesticide and antibiotic use in organically grown foods helps slow down the development of resistant strains of bacteria, weeds, and insects. One characteristic of organic foods on which everyone agrees is that they are generally more expensive than conventional foods.

It appears that the answer to the original question about organic foods versus conventional foods is going to continue to be based on who is answering, but it is important to note that all foods have to meet the same USDA standards of safety and quality. As a result, all consumers can be confident that they are benefitting from a safe, high-quality food supply.



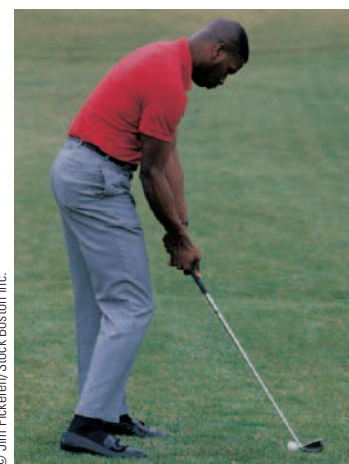
Masterfile (Royalty-Free Div.)

Organically-grown fruits and vegetables are now an option in most grocery stores.

Alkanes can be represented by the general formula C_nH_{2n+2} , where n is the number of carbon atoms in the molecule. The simplest alkane, methane, contains one carbon atom and therefore has the molecular formula CH_4 . The carbon atom is at the center, and the four bonds of the carbon atom are directed toward the hydrogen atoms at the corners of a regular tetrahedron; each hydrogen atom is geometrically equivalent to the other three in the molecule (see Figure 11.9). A tetrahedral orientation of bonds with bond angles of 109.5° is typical for carbon atoms that form four single bonds. Methane is the primary compound in natural gas. Tremendous quantities of natural gas are consumed worldwide because methane is an efficient, clean-burning fuel. It is used to heat homes, cook food, and power factories.

The next alkane is ethane, which has the molecular formula C_2H_6 and the structural formula CH_3-CH_3 . This molecule may be thought of as a methane molecule with one hydrogen removed and a $-CH_3$ put in its place. Again, the carbon bonds have a tetrahedral geometry as shown in Figure 11.10. Ethane is a minor component of natural gas.

Propane, the third alkane, has the molecular formula of C_3H_8 and the structural formula $CH_3-CH_2-CH_3$. Again, we can produce this molecule by removing a hydrogen atom from the preceding compound (ethane) and substituting a $-CH_3$ in its place (see Figure 11.11). Since all six hydrogen atoms of ethane are equivalent, it makes no difference which one is replaced. Propane is used extensively as an industrial fuel, as well as for home heating and cooking (see Figure 11.11c).



© Jim Pickrell/Stock Boston Inc.

Figure 11.8 This golf club has the strength and light weight of aluminum yet is made from graphite (carbon) fibers reinforced with plastic. What other sports equipment is made from graphite fibers?

Figure 11.9 Structural representations of methane, CH_4 .

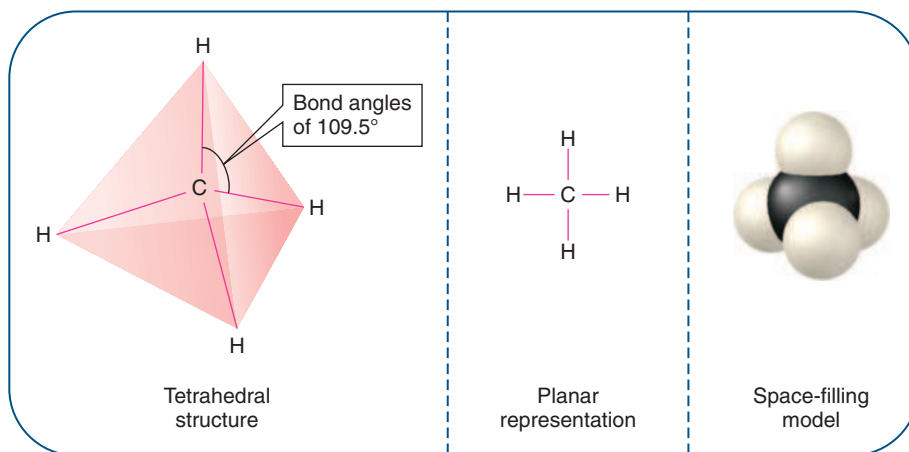
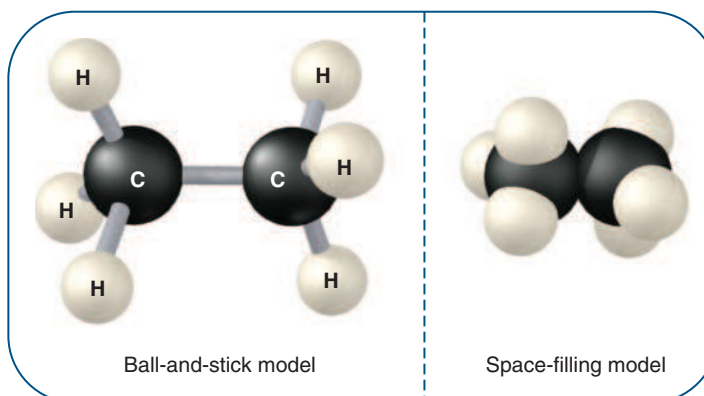


Figure 11.10 Perspective models of the ethane molecule, CH_3CH_3 .



The fourth member of the series, butane, with molecular formula C_4H_{10} , can also be produced by removing a hydrogen atom (this time from propane) and adding a —CH_3 . However, all the hydrogen atoms of propane are not geometrically equivalent, and more than one position is available for substitution. Replacing a hydrogen atom on one of the

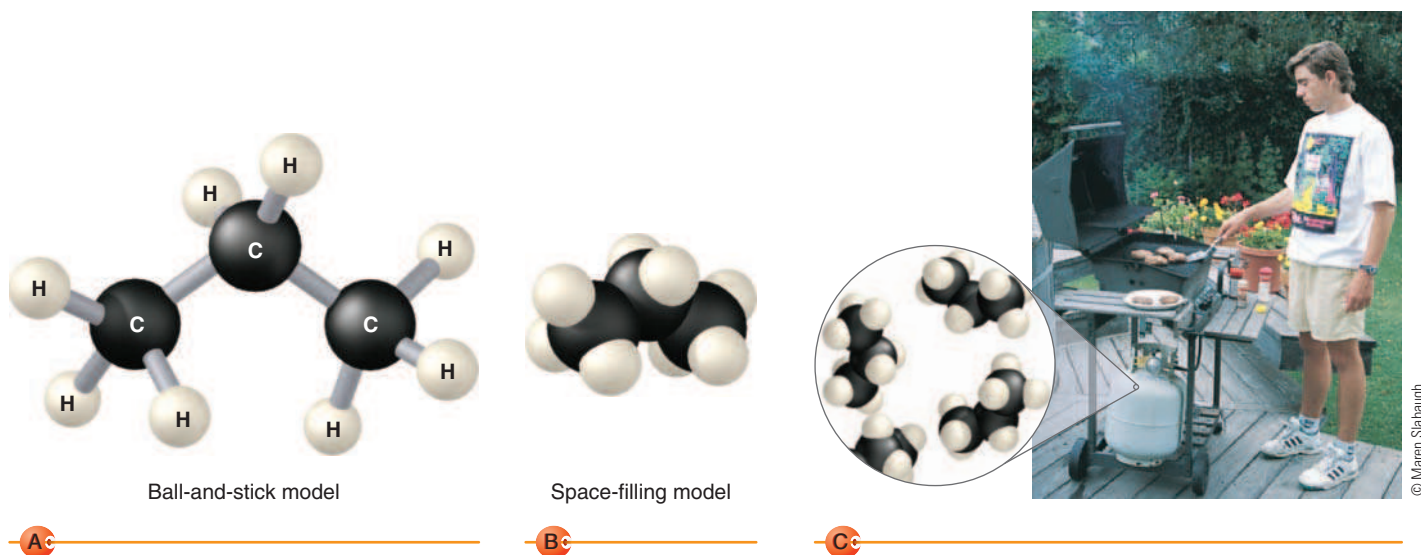


Figure 11.11 Perspective models **A** and **B** of the propane molecule $\text{CH}_3\text{CH}_2\text{CH}_3$. Propane is a common fuel for gas grills **C**.

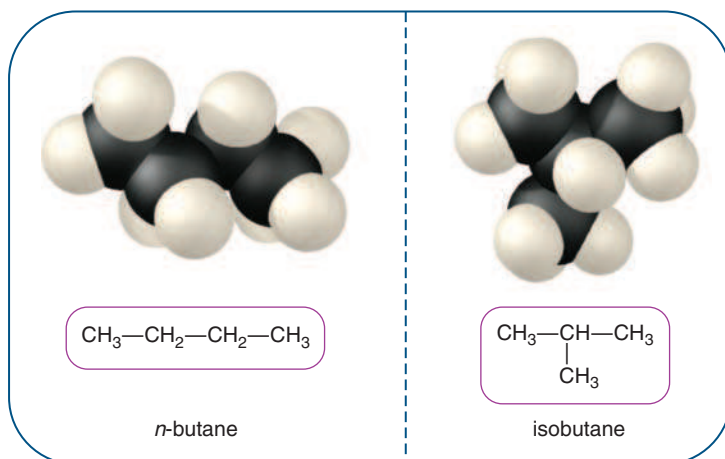
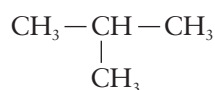


Figure 11.12 Space-filling models of the isomeric butanes.

end carbons of propane with —CH_3 produces a butane molecule that has the structural formula $\text{CH}_3\text{—CH}_2\text{—CH}_2\text{—CH}_3$. If, however, substitution is made on the central carbon atom of propane, the butane produced is



Notice that both butanes have the same molecular formula, C_4H_{10} . These two possible butanes are structural isomers because they have the same molecular formulas, but they have different atom-to-atom bonding sequences. The straight-chain isomer is called a **normal alkane** and the other is a **branched alkane** (► Figure 11.12).

The number of possible structural isomers increases dramatically with the number of carbon atoms in an alkane, as shown in ► Table 11.3.

► Learning Check 11.5

- Determine the molecular formula of the alkane containing eight carbon atoms.
- Draw the condensed structural formula of the normal isomer of the compound in part a.

normal alkane Any alkane in which all the carbon atoms are aligned in a continuous chain.

branched alkane An alkane in which at least one carbon atom is not part of a continuous chain.

Table 11.3 Molecular Formulas and Possible Structural Isomers of Alkanes

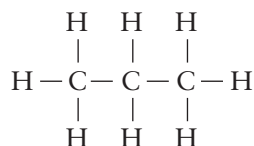
Molecular Formula	Number of Possible Structural Isomers
C_4H_{10}	2
C_5H_{12}	3
C_6H_{14}	5
$\text{C}_{10}\text{H}_{22}$	75
$\text{C}_{20}\text{H}_{42}$	366,319
$\text{C}_{30}\text{H}_{62}$	4,111,846,763

11.6 Conformations of Alkanes

Learning Objective

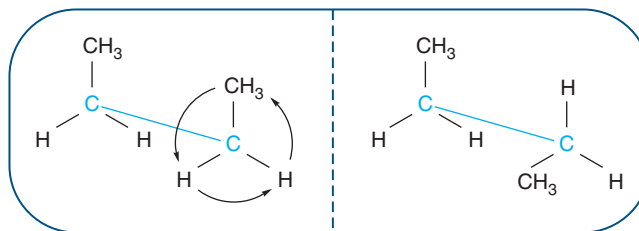
- Use structural formulas to determine whether compounds are structural isomers.

Remember that planar representations such as $\text{CH}_3\text{—CH}_2\text{—CH}_3$ or



are given with no attempt to accurately portray correct bond angles or molecular geometries. Structural formulas are usually written horizontally simply because it is convenient. It is also important to know that actual organic molecules are in constant motion—twisting, turning, vibrating, and bending. Groups connected by a single bond are capable of rotating about that bond much like a wheel rotates around an axle (see ► Figure 11.13). As a result of such rotation about single bonds, a molecule can exist in many different orientations,

Figure 11.13 Rotation about single bonds (n-butane molecule).

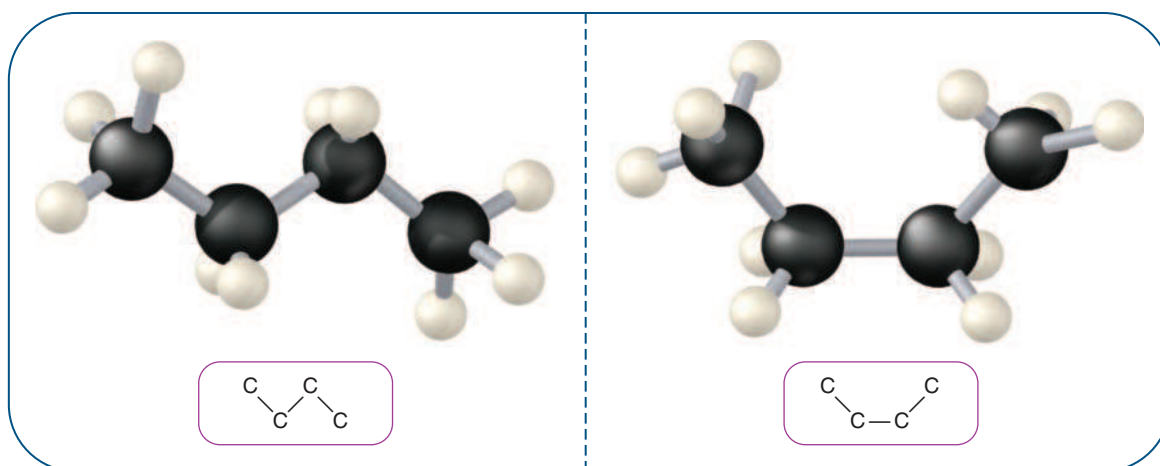
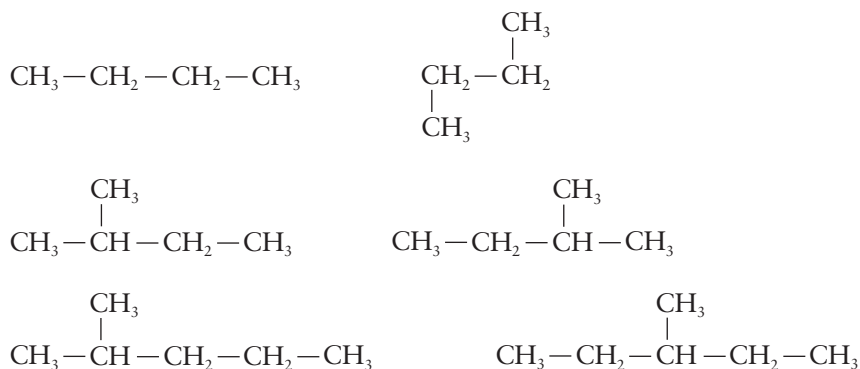


conformations The different arrangements of atoms in space achieved by rotation about single bonds.

called **conformations**. In a sample of butane containing billions of identical molecules, there are countless conformations present at any instant, and each conformation is rapidly changing into another. Two of the possible conformations of butane are shown in **Active Figure 11.14**. We must be sure to recognize that these different conformations do not represent different structural isomers. In each case, the four carbon atoms are bonded in a continuous (unbranched) chain. Since the order of bonding is not changed, the conformations correspond to the same molecule. Two structures would be structural isomers only if bonds had to be broken and remade to convert one into the other.

Example 11.3

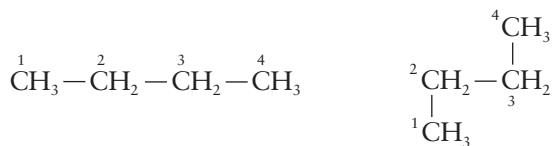
Which of the following pairs are structural isomers, and which are simply different representations of the same molecule?



Active Figure 11.14 Perspective models and carbon skeletons of two conformations of n-butane. Go to www.cengage.com/chemistry/seager or OWL to explore an interactive version of this figure.

Solution

- a. Same molecule: In both molecules, the four carbons are bonded in a continuous chain.



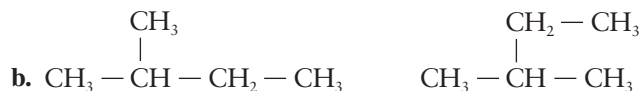
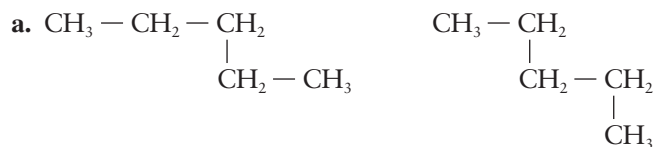
- b. Same molecule: In both molecules, there is a continuous chain of four carbons with a branch at position 2. The molecule has simply been turned around.



- c. Structural isomers: Both molecules have a continuous chain of five carbons, but the branch is located at different positions.



► **Learning Check 11.6** Which of the following pairs represent structural isomers, and which are simply the same compound?



11.7 Alkane Nomenclature

Learning Objective

8. Assign IUPAC names and draw structural formulas for alkanes.

When only a relatively few organic compounds were known, chemists gave them what are today called trivial or common names, such as methane, ethane, propane, and butane. The names for the larger alkanes were derived from the Greek prefixes that indicate the number of carbon atoms in the molecule. Thus, *pentane* contains five carbons, *hexane* has six, *heptane* has seven, and so forth, as shown in ► Table 11.4.

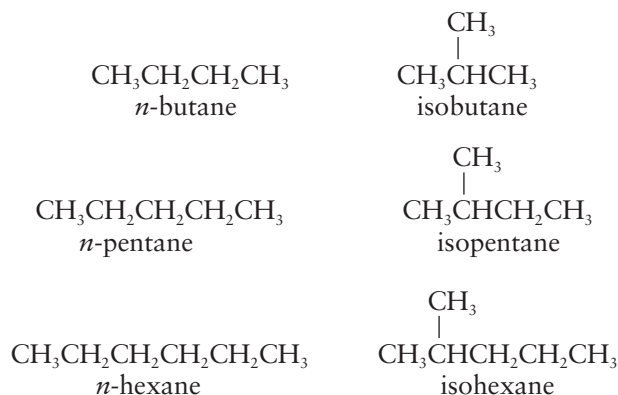
As more compounds and isomers were discovered, however, it became increasingly difficult to devise unique names and much more difficult to commit them to memory. Obviously, a systematic method was needed. Such a method is now in use, but so are several common methods.

The names for the two isomeric butanes (*n*-butane and isobutane) illustrate the important features of the common nomenclature system used for alkanes. The stem *but*- indicates

Table 11.4 Names of Alkanes

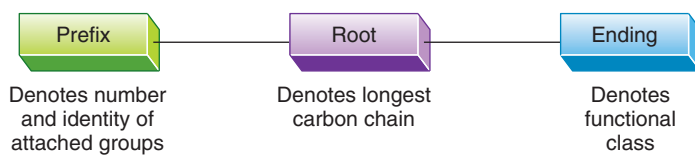
Number of Carbon Atoms	Name	Molecular Formula	Structure of Normal Isomer
1	methane	CH ₄	CH ₄
2	ethane	C ₂ H ₆	CH ₃ CH ₃
3	propane	C ₃ H ₈	CH ₃ CH ₂ CH ₃
4	butane	C ₄ H ₁₀	CH ₃ CH ₂ CH ₂ CH ₃
5	pentane	C ₅ H ₁₂	CH ₃ CH ₂ CH ₂ CH ₂ CH ₃
6	hexane	C ₆ H ₁₄	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
7	heptane	C ₇ H ₁₆	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
8	octane	C ₈ H ₁₈	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
9	nonane	C ₉ H ₂₀	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
10	decane	C ₁₀ H ₂₂	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃

that four carbons are present in the molecule. The *-ane* ending signifies the alkane family. The prefix *n-* indicates that all carbons form an unbranched chain. The prefix *iso-* refers to compounds in which all carbons except one are in a continuous chain and in which that one carbon is branched from a next-to-the-end carbon, as shown:



This common naming system has limitations. Pentane has three isomers, and hexane has five. The more complicated the compound, the greater the number of isomers, and the greater the number of special prefixes needed to name all the isomers. It would be extremely difficult and time-consuming to try to identify each of the 75 isomeric alkanes containing 10 carbon atoms by a unique prefix or name.

To devise a system of nomenclature that could be used for even the most complicated compounds, committees of chemists have met periodically since 1892. The system resulting from these meetings is called the IUPAC (International Union of Pure and Applied Chemistry) system. This system is much the same for all classes of organic compounds. The IUPAC name for an organic compound consists of three component parts:



The *root* of the IUPAC name specifies the longest continuous chain of carbon atoms in the compound. The roots for the first 10 normal hydrocarbons are based on the names given in Table 11.4: C₁ *meth*-, C₂ *eth*-, C₃ *prop*-, C₄ *but*-, C₅ *pent*-, C₆ *hex*-, C₇ *hept*-, C₈ *oct*-, C₉ *non*-, C₁₀ *dec*-.

The *ending* of an IUPAC name specifies the functional class or the major functional group of the compound. The ending *-ane* specifies an alkane. Each of the other functional classes has a characteristic ending; for example, *-ene* is the ending for alkenes, and the *-ol* ending designates alcohols.

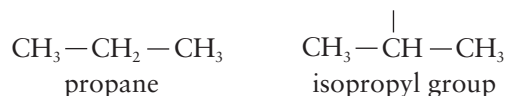
Prefixes are used to specify the identity, number, and location of atoms or groups of atoms that are attached to the longest carbon chain. Table 11.5 lists several common carbon-containing groups referred to as **alkyl groups**. Each alkyl group is a collection of atoms that can be thought of as an alkane minus one hydrogen atom. Alkyl groups are named simply by dropping *-ane* from the name of the corresponding alkane and replacing it with *-yl*. For example, CH₃— is called a methyl group and CH₃—CH₂— an ethyl group:



Two different alkyl groups can be derived from propane, depending on which hydrogen is removed. Removal of a hydrogen from an end carbon results in a propyl group:



Removal of a hydrogen from the center carbon results in an isopropyl group:



alkyl group A group differing by one hydrogen from an alkane.

Table 11.5 Common Alkyl Groups

Parent Alkane	Structure of Parent Alkane	Structure of Alkyl Group	Name of Alkyl Group
methane	CH ₄	CH ₃ —	methyl
ethane	CH ₃ CH ₃	CH ₃ CH ₂ —	ethyl
propane	CH ₃ CH ₂ CH ₃	CH ₃ CH ₂ CH ₂ —	propyl
		$\begin{array}{c} \\ \text{CH}_3\text{CHCH}_3 \end{array}$	isopropyl
<i>n</i> -butane	CH ₃ CH ₂ CH ₂ CH ₃	CH ₃ CH ₂ CH ₂ CH ₂ —	butyl
		$\begin{array}{c} \\ \text{CH}_3\text{CH}_2\text{CHCH}_3 \end{array}$	<i>sec</i> -butyl (secondary-butyl)
isobutane	$\begin{array}{c} \text{CH}_3 \\ \\ \text{CH}_3\text{CHCH}_3 \end{array}$	$\begin{array}{c} \text{CH}_3 \\ \\ \text{CH}_3\text{CHCH}_2— \end{array}$	isobutyl
		$\begin{array}{c} \text{CH}_3 \\ \\ \text{CH}_3\text{CCH}_3 \\ \end{array}$	<i>t</i> -butyl (tertiary-butyl)

Table 11.6 Common Nonalkyl Groups

Group	Name
—F	fluoro
—Cl	chloro
—Br	bromo
—I	iodo
—NO ₂	nitro
—NH ₂	amino

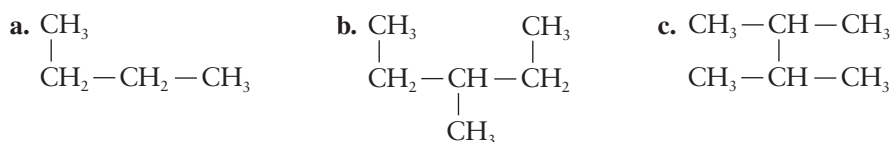
An isopropyl group also can be represented by $(\text{CH}_3)_2\text{CH—}$. As shown in Table 11.5, four butyl groups can be derived from butane, two from the straight-chain, or normal, butane, and two from the branched-chain isobutane. A number of nonalkyl groups are also commonly used in naming organic compounds (see Table 11.6).

The following steps are useful when the IUPAC name of an alkane is written on the basis of its structural formula:

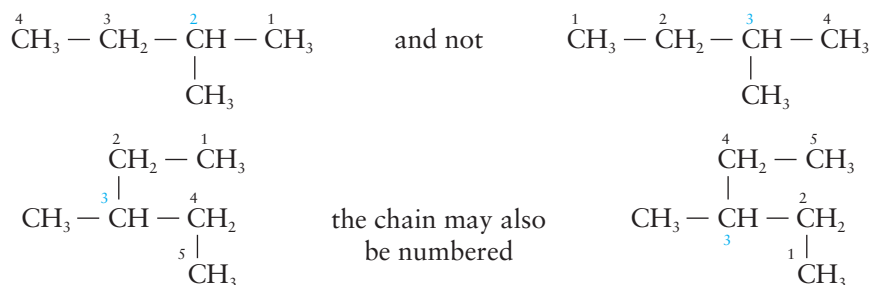
Step 1. Name the longest chain. The longest continuous carbon-atom chain is chosen as the basis for the name. The names are those given in Table 11.4.

Example	Longest Chain	Comments
$\begin{array}{c} \text{CH}_3 - \text{CH}_2 - \text{CH} - \text{CH}_3 \\ \\ \text{CH}_3 \end{array}$	$\begin{array}{c} \text{C} - \text{C} - \text{C} - \text{C} \\ \\ \text{C} \end{array}$	This compound is a butane.
$\begin{array}{c} \text{CH}_2 - \text{CH}_3 \\ \\ \text{CH}_3 - \text{CH} - \text{CH}_2 \\ \\ \text{CH}_3 \end{array}$	$\begin{array}{c} \text{C} - \text{C} \\ \\ \text{C} - \text{C} - \text{C} \\ \\ \text{C} \end{array}$	This compound is a pentane.
$\begin{array}{c} \text{CH}_2 - \text{CH}_3 \\ \\ \text{CH}_3 - \text{CH}_2 - \text{C} - \text{CH}_3 \\ \\ \text{CH}_2 \\ \\ \text{CH}_3 - \text{CH}_2 \end{array}$	$\begin{array}{c} \text{C} - \text{C} \\ \\ \text{C} - \text{C} - \text{C} - \text{C} \\ \\ \text{C} \\ \\ \text{C} - \text{C} \\ \\ \text{C} - \text{C} \\ \\ \text{C} - \text{C} \end{array}$	Often there is more than one way to designate the longest chain. Either way, in this case, the compound is a hexane.

Learning Check 11.7 Identify the longest carbon chain in the following:



Step 2. Number the longest chain. The carbon atoms in the longest chain are numbered consecutively from the end that will give the lowest possible number to any carbon to which a group is attached.



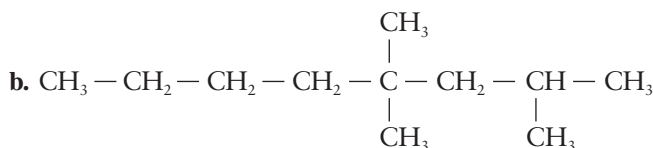
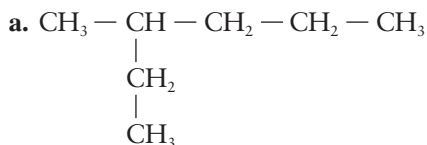
$$\begin{array}{c} 1 \qquad 2 \qquad 3 \qquad 4 \qquad 5 \qquad 6 \\ \text{CH}_3 - \text{CH} - \text{CH}_2 - \text{CH} - \text{CH}_2 - \text{CH}_3 \\ | \qquad \qquad | \\ \text{CH}_3 \qquad \text{CH}_2 - \text{CH}_3 \end{array} \quad \text{and not} \quad \begin{array}{c} 6 \qquad 5 \qquad 4 \qquad 3 \qquad 2 \qquad 1 \\ \text{CH}_3 - \text{CH} - \text{CH}_2 - \text{CH} - \text{CH}_2 - \text{CH}_3 \\ | \qquad \qquad | \\ \text{CH}_3 \qquad \text{CH}_2 - \text{CH}_3 \end{array}$$

Groups are located at positions 3 and 5

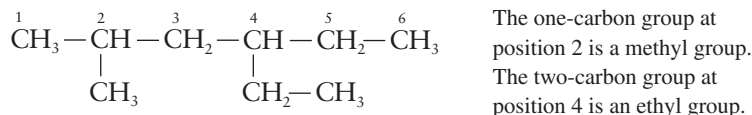


Groups are located at
positions 2,4,4

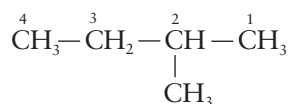
► **Learning Check 11.8** Decide how to correctly number the longest chain in the following according to IUPAC rules:


$$\begin{array}{ccccccc} & & 3 & & 2 & & 1 \\ \text{CH}_3 & - & \text{CH}_2 & - & \text{CH} & - & \text{CH}_3 \\ & & & & | & & \\ & & & & \text{CH}_3 & & \end{array}$$

The attached group is located on carbon 2 of the chain, and it is a methyl group.


$$\begin{array}{ccccccc}
 & & & & \text{CH}_3 & & \\
 & & & & | & & \\
 & & & & \text{CH} - \text{CH}_3 & & \\
 & & & | & & & \\
 \text{---} & | & | & | & | & | & \text{---} \\
 \text{CH}_3 & \text{CH}_2 & & \text{CH} - \text{CH}_3 & & \text{CH}_2 - \text{CH}_3 & \\
 & | & & | & & & \\
 & \text{CH}_2 & & \text{CH}_2 & & & \\
 & | & & | & & & \\
 & \text{CH}_3 & & \text{CH}_3 & & &
 \end{array}$$

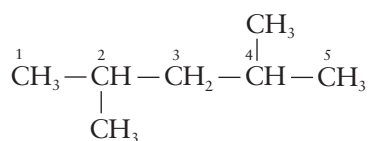
Step 4. Combine the longest chain and the branches into the name. The position and the name of the attached alkyl group are added to the name of the longest chain and written as one word:



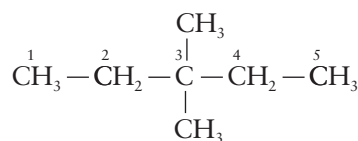
2-methylbutane

Additional steps are needed when more than one alkyl group is attached to the longest chain.

Step 5. Indicate the number and position of attached alkyl groups. If two or more of the same alkyl group occur as branches, the number of them is indicated by the prefixes *di-*, *tri-*, *tetra-*, *penta-*, etc., and the location of each is again indicated by a number. These position numbers, separated by commas, are put just before the name of the group, with hyphens before and after the numbers when necessary:

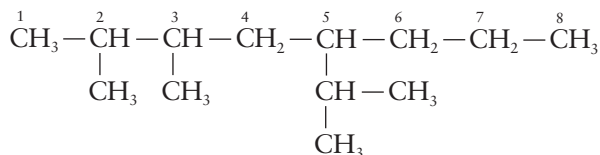


2,4-dimethylpentane



3,3-dimethylpentane

If two or more *different* alkyl groups are present, their names are alphabetized and added to the name of the basic alkane, again as one word. For purposes of alphabetizing, the prefixes *di-*, *tri-*, and so on are ignored, as are the italicized prefixes secondary (*sec*) and tertiary (*t*). The prefix *iso-* is an exception and is used for alphabetizing:



5-isopropyl-2,3-dimethyloctane

Learning Check 11.10 Give the correct IUPAC name to each of the following:

- $$\begin{array}{ccccccc} & & & & \text{CH}_3 & & \\ & & & & | & & \\ \text{CH}_3 & - & \text{CH}_2 & - & \text{CH}_2 & - & \text{CH}_2 & - & \text{CH} & - & \text{CH}_3 \end{array}$$
- $$\begin{array}{ccccccc} & & & & \text{CH}_2 & - & \text{CH}_3 & & \\ & & & & | & & & & \\ \text{CH}_3 & - & \text{CH}_2 & - & \text{CH}_2 & - & \text{CH} & - & \text{CH}_2 & - & \text{CH} & - & \text{CH}_3 \\ & & & & & & & & | & & \\ & & & & & & & & \text{CH}_2 & - & \text{CH}_3 \end{array}$$
- $$\begin{array}{ccccccc} & & & & \text{CH}_3 & & \\ & & & & | & & \\ \text{CH}_3 & - & \text{CH}_2 & - & \text{CH}_2 & - & \text{CH} & - & \text{CH} & - & \text{CH} & - & \text{CH}_3 \\ & & & & & & | & & | & & \\ & & & & & & \text{CH} & - & \text{CH}_3 & & \text{CH}_3 \\ & & & & & & | & & \\ & & & & & & \text{CH}_3 & & \end{array}$$

Naming compounds is a very important skill, as is the reverse process of using IUPAC nomenclature to specify a structural formula. The two processes are very similar.

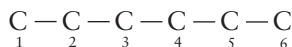
To obtain a formula from a name, determine the longest chain, number the chain, and add any attached groups.

Example 11.4

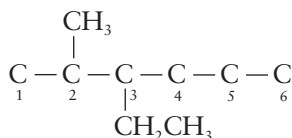
Draw a condensed structural formula for 3-ethyl-2-methylhexane.

Solution

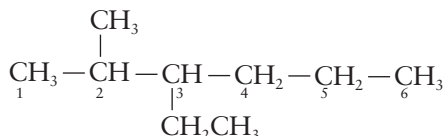
Use the last part of the name to determine the longest chain. Draw a chain of six carbons. Then, number the carbon atoms.



Attach a methyl group at position 2 and an ethyl group at position 3.



Complete the structure by adding enough hydrogen atoms so that each carbon has four bonds.



Learning Check 11.11 Draw a condensed structural formula for each of the following compounds:

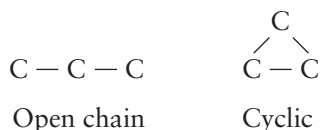
- 2,2,4-trimethylpentane
- 3-isopropylhexane
- 3-ethyl-2,4-dimethylheptane

11.8 Cycloalkanes

Learning Objective

9. Assign IUPAC names and draw structural formulas for cycloalkanes.

From what we have said so far, the formula C_3H_6 cannot represent an alkane. Not enough hydrogens are present to allow each carbon to form four bonds, unless there are multiple bonds. For example, the structural formula $\text{CH}_3-\text{CH}=\text{CH}_2$ fits the molecular formula but cannot represent an alkane because of the double bond. The C_3H_6 formula does become acceptable for an alkane if the carbon atoms form a ring, or cyclic, structure rather than the open-chain structure shown:



The resulting saturated cyclic compound, called cyclopropane, has the structural formula

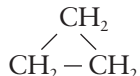






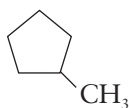
Table 11.7 Structural Formulas and Symbols for Common Cycloalkanes

Name	Structural Formula	Condensed Formula
cyclopropane	$ \begin{array}{c} \text{CH}_2 \\ \diagup \quad \diagdown \\ \text{H}_2\text{C} - \text{CH}_2 \end{array} $	
cyclobutane	$ \begin{array}{c} \text{H}_2\text{C} - \text{CH}_2 \\ \quad \\ \text{H}_2\text{C} - \text{CH}_2 \end{array} $	
cyclopentane	$ \begin{array}{c} \text{H}_2 \\ \diagup \quad \diagdown \\ \text{C} \\ \diagdown \quad \diagup \\ \text{H}_2\text{C} \quad \text{CH}_2 \\ \quad \\ \text{H}_2\text{C} - \text{CH}_2 \end{array} $	
cyclohexane	$ \begin{array}{c} \text{H}_2 \\ \diagup \quad \diagdown \\ \text{C} \\ \diagdown \quad \diagup \\ \text{H}_2\text{C} \quad \text{CH}_2 \\ \quad \\ \text{H}_2\text{C} \quad \text{CH}_2 \\ \quad \\ \text{C} \\ \diagup \quad \diagdown \\ \text{H}_2 \end{array} $	

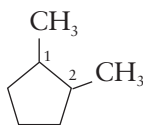
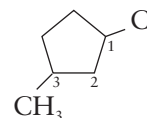
cycloalkane An alkane in which carbon atoms form a ring.

Alkanes containing rings of carbon atoms are called **cycloalkanes**. Like the other alkanes, cycloalkanes are not found in human cells. However, several important molecules in human cells do contain rings of five or six atoms, and the study of cycloalkanes will help you better understand the chemical behavior of these complex molecules.

According to IUPAC rules, cycloalkanes are named by placing the prefix *cyclo-* before the name of the corresponding alkane with the same number of carbon atoms. Chemists often abbreviate the structural formulas for cycloalkanes and draw them as geometric figures (triangles, squares, etc.) in which each corner represents a carbon atom. The hydrogens are omitted (see Table 11.7). It is important to remember that each carbon atom still possesses four bonds, and that hydrogen is assumed to be bonded to the carbon atoms unless something else is indicated. When substituted cycloalkanes (those with attached groups) are named, the position of a single attached group does not need to be specified in the name because all positions in the ring are equivalent. However, when two or more groups are attached, their positions of attachment are indicated by numbers, just as they were for alkanes. The ring numbering begins with the carbon attached to the first group alphabetically and proceeds around the ring in the direction that will give the lowest numbers for the locations of the other attached groups.

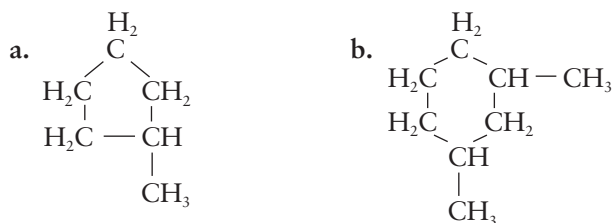


methylcyclopentane

1,2-dimethylcyclopentane
not 1,5-dimethylcyclopentane1-chloro-3-methylcyclopentane
not 1-chloro-4-methylcyclopentane
not 3-chloro-1-methylcyclopentane

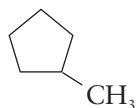
Example 11.5

Represent each of the following cycloalkanes by a geometric figure, and name each compound:

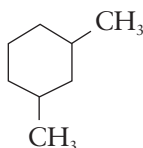


Solution

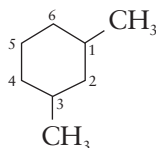
- a. A pentagon represents a five-membered ring, which is called cyclopentane. This compound has a methyl group attached, so the name is methylcyclopentane. The position of a single alkyl group is not indicated by a number because the positions of all carbons in the ring are equivalent.



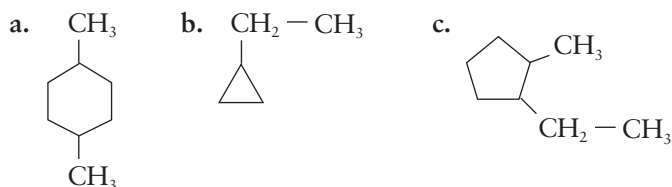
- b. A hexagon represents a six-carbon ring, which is called cyclohexane. Two methyl groups are attached; thus we have a dimethylcyclohexane.



However, the positions of the two alkyl groups must be indicated. The ring is numbered beginning with a carbon to which a methyl group is attached, counting in the direction giving the lowest numbers. The correct name is 1,3-dimethylcyclohexane. Notice that a reverse numbering beginning at the same carbon would have given 1,5-dimethylcyclohexane. The number 3 in the correct name is lower than the 5 in the incorrect name.



► **Learning Check 11.12** Give each of the following compounds the correct IUPAC name:



11.9 The Shape of Cycloalkanes

Learning Objective

10. Name and draw structural formulas for geometric isomers of cycloalkanes.

Recall from Section 11.5 that a tetrahedral orientation of bonds with bond angles of 109.5° is characteristic of carbon atoms that form four single bonds. A tetrahedral arrangement is the most stable because it results in the least crowding of the atoms. In certain cycloalkanes, however, a tetrahedral arrangement for all carbon-to-carbon bonds

Figure 11.15 Ball-and-stick models for common cycloalkanes.

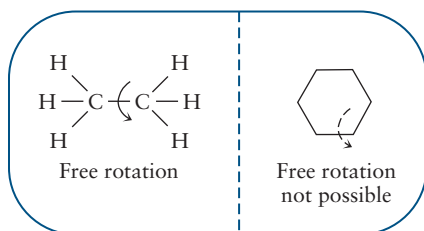
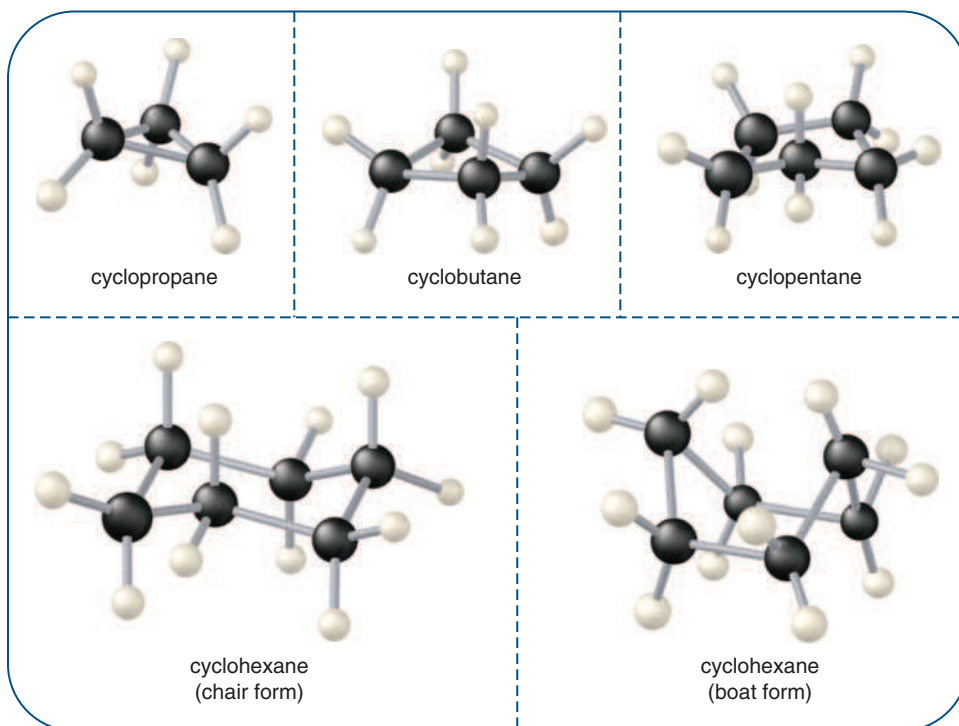


Figure 11.16 Rotation about C—C single bonds occurs in open-chain compounds but not within rings.

is not possible. For example, the bond angles between adjacent carbon–carbon bonds in planar cyclopropane molecules must be 60° (see [Figure 11.15](#)). In cyclobutane, they are close to 90° . As a result, cyclopropane and cyclobutane rings are much less stable than compounds with bond angles of about 109° . Both cyclobutane and cyclopentane bend slightly from a planar structure to reduce the crowding of hydrogen atoms ([Figure 11.15](#)). In larger cycloalkanes, the bonds to carbon atoms can be tetrahedrally arranged only when the carbon atoms do not lie in the same plane. For example, cyclohexane can assume several nonplanar shapes. The chair and boat forms, where all the bond angles are 109.5° , are shown in [Figure 11.15](#).

The free rotation that can take place around C—C single bonds in alkanes ([Section 11.6](#)) is not possible for the C—C bonds of cycloalkanes. The ring structure allows bending or puckering but prevents free rotation (see [Figure 11.16](#)). Any rotation of one carbon atom 180° relative to an adjacent carbon atom in a cycloalkane would require a single carbon–carbon bond to be broken somewhere in the ring. The breaking of such bonds would require a large amount of energy.

The lack of free rotation around C—C bonds in disubstituted cycloalkanes leads to an extremely important kind of isomerism called *stereoisomerism*. Two different compounds that have the same molecular formula and the same structural formula but different spatial arrangements of atoms are called **stereoisomers**. For example, consider a molecule of 1,2-dimethylcyclopentane. The cyclopentane ring is drawn in [Figure 11.17](#) as a planar pentagon with the heavy lines indicating that two of the carbons are in front as one views the structure. The groups attached to the ring project above or below the plane of the ring. Two stereoisomers are possible: Either both groups may project in the same direction from the plane, or they may project in opposite directions from the plane of the ring. Since the methyl groups cannot rotate from one side of the ring to the other, molecules of the two compounds represented in [Figure 11.17](#) are distinct.

These two compounds have physical and chemical properties that are quite different and therefore can be separated from each other. Stereoisomers of this type, in which the spatial arrangement or geometry of their groups is maintained by rings, are called **geometric isomers** or *cis-trans* isomers. The prefix *cis-* denotes the isomer in which both groups are on the same side of the ring, and *trans-* denotes the isomer in which they are on opposite sides. To exist as geometric isomers, a disubstituted cycloalkane must be bound

stereoisomers Compounds with the same structural formula but different spatial arrangements of atoms.

geometric isomers Molecules with restricted rotation around C—C bonds that differ in the three-dimensional arrangements of their atoms in space and not in the order of linkage of atoms.

cis- On the same side (as applied to geometric isomers).

trans- On opposite sides (as applied to geometric isomers).

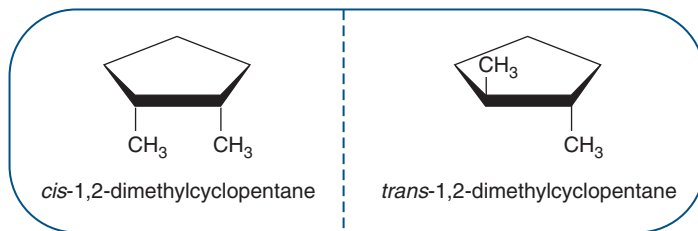
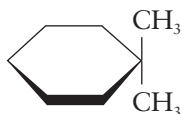


Figure 11.17 Two geometric isomers of 1,2-dimethylcyclopentane.

to groups at two different carbons of the ring. For example, there are no geometric isomers of 1,1-dimethylcyclohexane:

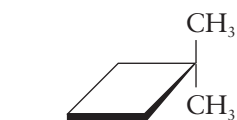


Example 11.6

Name and draw structural formulas for all the isomers of dimethylcyclobutane. Indicate which ones are geometric isomers.

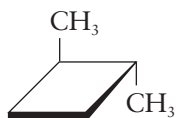
Solution

There are three possible locations for the two methyl groups: positions 1,1, positions 1,2, and positions 1,3.



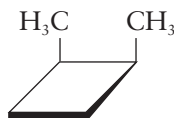
1,1-dimethylcyclobutane

Geometric isomerism is not possible in this case with the two groups bound to the same carbon of the ring



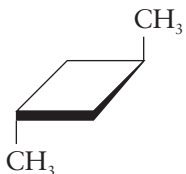
trans-1,2-dimethylcyclobutane

Two groups on opposite sides of the planar ring



cis-1,2-dimethylcyclobutane

Two groups on the same side of the ring



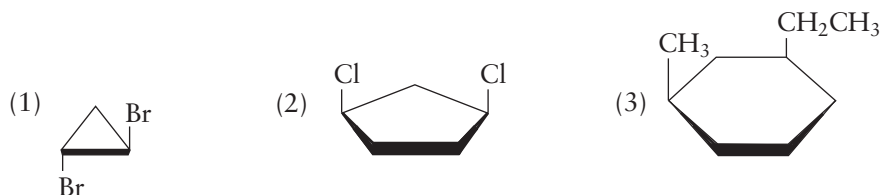
trans-1,3-dimethylcyclobutane



cis-1,3-dimethylcyclobutane

Learning Check 11.13

a. Identify each of the following cycloalkanes as a *cis*- or *trans*- compound:



b. Draw the structural formula for *cis*-1,2-dichlorocyclobutane.



Petroleum, the most important of the fossil fuels used today, is sometimes called “black gold” in recognition of its importance in the 20th century. At times, the need for petroleum to keep society fueled has seemed second only to our need for food, shelter, and clothing.

It is generally believed that this complex mixture of hydrocarbons was formed over eons through the gradual decay of ocean-dwelling microscopic animals. The resulting crude oil, a viscous black liquid, collects in vast underground pockets in sedimentary rock. It is brought to the surface via drilling and pumping.

Useful products are obtained from crude oil by heating it to high temperatures to produce various fractions according to boiling point (see table). Most petroleum products are eventually burned as a fuel, but about 2% are used to synthesize organic compounds. This seem-

ingly small amount is quite large in actual tonnage because of the huge volume of petroleum that is refined annually. In fact, more than half of all industrial synthetic organic compounds are made from this source. These industrial chemicals are eventually converted into dyes, drugs, plastics, artificial fibers, detergents, insecticides, and other materials deemed indispensable by many in industrialized nations.

Fraction	Boiling Point Range (°C)	Molecular Size Range	Typical Uses
Gas	−164–30	C ₁ –C ₄	Heating, cooking
Gasoline	30–200	C ₅ –C ₁₂	Motor fuel
Kerosene	175–275	C ₁₂ –C ₁₆	Fuel for stoves and diesel and jet engines
Heating oil	Up to 375	C ₁₅ –C ₁₈	Furnace oil
Lubricating oils	350 and up	C ₁₆ –C ₂₀	Lubrication, mineral oil
Greases	Semisolid	C ₁₈ –up	Lubrication, petroleum jelly
Paraffin (wax)	Melts at 52–57	C ₂₀ –up	Candles, toiletries
Pitch and tar	Residue in boiler	High	Roofing, asphalt paving



Asphalt for paving roads is a petroleum product.

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11.10 Physical Properties of Alkanes

Learning Objective

11. Describe the key physical properties of alkanes.

Since alkanes are composed of nonpolar carbon–carbon and carbon–hydrogen bonds, alkanes are nonpolar molecules. Alkanes have lower melting and boiling points than other organic compounds of comparable molecular weight (see ▶ Table 11.8). This is because their nonpolar molecules exert very weak attractions for each other. Alkanes are odorless compounds.

The normal, or straight-chain, alkanes make up what is called a **homologous series**. This term describes any series of compounds in which each member differs from a previous member only by having an additional —CH₂— unit. The physical and chemical properties of compounds making up a homologous series are usually closely related and vary in a systematic and predictable way. For example, the boiling points of normal alkanes increase smoothly as the length of the carbon chain increases (see ▶ Figure 11.18). This pattern results from increasing dispersion forces as molecular weight increases. At ordinary temperatures and pressures, normal alkanes with 1 to 4 carbon atoms are gases, those with 5 to 20 carbon atoms are liquids, and those with more than 20 carbon atoms are waxy solids.

Because they are nonpolar, alkanes and other hydrocarbons are insoluble in water, which is a highly polar solvent. They are also less dense than water and thus float on it. These two properties of hydrocarbons are partly responsible for the well-known serious effects of oil spills from ships (see ▶ Figure 11.19).

homologous series Compounds of the same functional class that differ by a —CH₂— group.

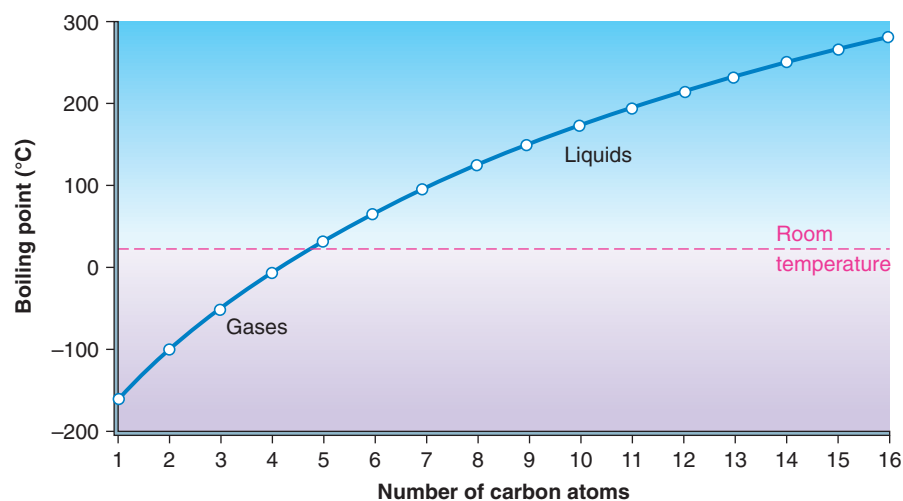
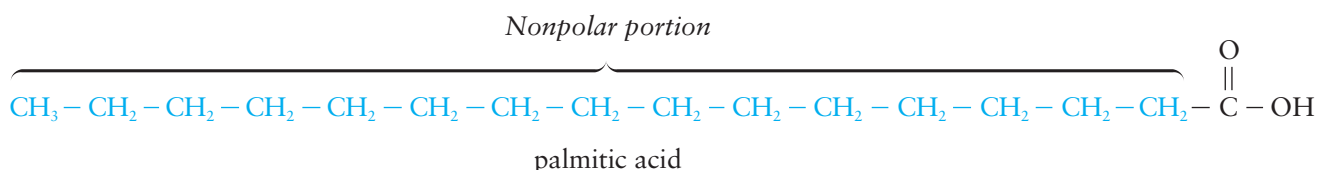
Table 11.8 Physical Properties of Some Normal Alkanes

Number of Carbon Atoms	IUPAC Name	Condensed Structural Formula	Melting Point (°C)	Boiling Point (°C)	Density (g/mL)
1	methane	CH ₄	-182.5	-164.0	0.55
2	ethane	CH ₃ CH ₃	-183.2	-88.6	0.57
3	propane	CH ₃ CH ₂ CH ₃	-189.7	-42.1	0.58
4	butane	CH ₃ CH ₂ CH ₂ CH ₃	-133.4	-0.5	0.60
5	pentane	CH ₃ CH ₂ CH ₂ CH ₂ CH ₃	-129.7	36.1	0.63
6	hexane	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-95.3	68.9	0.66
7	heptane	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-90.6	98.4	0.68
8	octane	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-56.8	125.7	0.70
9	nonane	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-53.5	150.8	0.72
10	decane	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃	-29.7	174.1	0.73

Liquid alkanes of higher molecular weight behave as emollients (skin softeners) when applied to the skin. An alkane mixture known as mineral oil is sometimes used to replace natural skin oils washed away by frequent bathing or swimming. Petroleum jelly (Vaseline is a well-known brand name) is a semisolid mixture of alkanes that is used as both an emollient and a protective film. Water and water solutions such as urine don't dissolve or penetrate the film, and the underlying skin is protected. Many cases of diaper rash have been prevented or treated this way.

The word **hydrophobic** (literally “water fearing”) is often used to refer to molecules or parts of molecules that are insoluble in water. Many biomolecules, the large organic molecules associated with living organisms, contain nonpolar (hydrophobic) parts. Thus, such molecules are not water-soluble. Palmitic acid, for example, contains a large nonpolar hydrophobic portion and is insoluble in water.

hydrophobic Molecules or parts of molecules that repel (are insoluble in) water.

**Figure 11.18** Normal alkane boiling points depend on chain length.**Figure 11.19** Oil spills can have serious and long-lasting effects on the environment because of the insolubility of hydrocarbons in water.



Healthy, normal human skin is kept moist by natural body oils contained in sebum, a secretion of the skin's sebaceous glands. The oily sebum helps the epidermis, or outer layer of the skin, retain the 10–30% of water it normally contains. However, some people suffer from skin that is naturally dry, or dry because of aging or contact with materials like paint thinner that dissolve and remove the sebum. Such individuals may find some relief by using OTC products called *moisturizers*.

Two types of skin moisturizers are commonly available. One type, which behaves like natural sebum, contains nonpolar oily substances that form a barrier and prevent water from passing through and evaporating from the skin. These barrier-forming products often contain combinations of materials from various sources, including mineral oil and petroleum jelly from petroleum, vegetable oils such as apricot oil, sesame seed oil, palm kernel oil, olive oil, and safflower oil, and lanolin, an animal fat from sheep oil glands and wool. While they effectively keep water from leaving the skin, these products are somewhat messy to use and leave the skin feeling greasy.

A second, more popular type of moisturizer works by attracting water from the air and skin. These products form a water-rich layer that adheres to the skin without giving it a greasy feel. The substances that attract water are called humectants, and are compounds capable of forming hydrogen bonds with water. Some examples of humectants used in products of this type are glycerol, urea, lactic acid, and propylene glycol.

If you are shopping for a moisturizer, remember that the main characteristic you should look for in a product is the ability to form a

barrier to prevent water evaporation, or the ability to act as a humectant. Some expensive products advertise that in addition to moisturizing, they also beautify the skin and even reverse aging because they contain proteins such as collagen and elastin, vitamins, hormones, or even DNA. It is unlikely that such substances can pass through the epidermis of the skin in sufficient amounts to provide the advertised benefits.



Alkane products marketed to soften and protect the skin.

11.11 Alkane Reactions

Learning Objective

12. Write alkane combustion reactions.

The alkanes are the least reactive of all organic compounds. In general, they do not react with strong acids (such as sulfuric acid), strong bases (such as sodium hydroxide), most oxidizing agents (such as potassium dichromate), and most reducing agents (such as sodium metal). This unreactive nature is reflected in the name *paraffins*, sometimes used to identify alkanes (from Latin words that mean “little affinity”). Paraffin wax, sometimes used to seal jars of homemade preserves, is a mixture of solid alkanes. Paraffin wax is also used in the preparation of milk cartons and wax paper. The inertness of the compounds in the wax makes it ideal for these uses. Alkanes do undergo reactions with halogens such as chlorine and bromine, but these are not important for our purposes.

Perhaps the most significant reaction of alkanes is the rapid oxidation called combustion. In the presence of ample oxygen, alkanes burn to form carbon dioxide and water, liberating large quantities of heat:



It is this type of reaction that accounts for the wide use of hydrocarbons as fuels.

Natural gas contains methane (80–95%), some ethane, and small amounts of other hydrocarbons. Propane and butane are extracted from natural gas and sold in pressurized metal containers (bottled gas). In this form, they are used for heating and cooking in campers, trailers, boats, and rural homes.

Ice Storms and Deadly Carbon Monoxide



Every winter a killer waits to strike. The news headlines tell of a devastating winter storm as the country is pummeled with ice and snow. The storm is blamed for a number of deaths, many resulting from carbon monoxide poisoning after residents pulled power generators inside their homes to keep warm.

The incomplete combustion of alkanes and other carbon-containing fuels may produce carbon monoxide, CO. It is a common component of furnace, stove, and automobile exhaust. Normally, these fumes are emitted into the air and dispersed. However, whenever CO gas is released into closed spaces, lethal concentrations may develop. All of the following are known to have caused CO poisoning deaths: poor ventilation of a furnace or stove exhaust, running the engine of an automobile for a period of time in a closed garage, idling the engine of a snowbound car to keep the heater working, and using a barbecue grill as a source of heat in a house or camper (burning charcoal generates a large amount of carbon monoxide). Recently, the exhaust systems of snow-blocked vehicles have been identified as another cause of carbon monoxide poisoning. The problem is created when the end of a vehicle's exhaust pipe is obstructed or plugged by snow. CO-containing exhaust can leak through cracks in the exhaust system, penetrate the floorboard, and enter the passenger compartment of the vehicle. An estimated 1000 Americans die each year from unintentional CO poisoning, and as many as 10,000 require medical treatment.

Carbon monoxide is dangerous because of its ability to bind strongly with hemoglobin in red blood cells. When this occurs, the ability of the blood to transport oxygen (O₂) is reduced because the CO molecules occupy sites on hemoglobin molecules that are normally occupied by O₂ molecules.

Because it is colorless and odorless, carbon monoxide gas is a silent and stealthy potential killer. The following symptoms of CO poisoning are given in order of increasing severity and seriousness:

1. Slight headache, dizziness, drowsiness
2. Headache, throbbing temples
3. Weakness, mental confusion, nausea
4. Rapid pulse and respiration, fainting
5. Possibly fatal coma

Gasoline is a mixture of hydrocarbons (primarily alkanes) that contain 5 to 12 carbon atoms per molecule. Diesel fuel is a similar mixture, except the molecules contain 12 to 16 carbon atoms. The hot CO₂ and water vapor generated during combustion in an internal combustion engine have a much greater volume than the air and fuel mixture. It is this sudden increase in gaseous volume and pressure that pushes the pistons and delivers power to the crankshaft.

If there is not enough oxygen available, incomplete combustion of hydrocarbons occurs, and some carbon monoxide (CO) or even carbon (see Figure 11.20) may be produced (Reactions 11.5 and 11.6):



These reactions are usually undesirable because CO is toxic, and carbon deposits hinder engine performance. Occasionally, however, incomplete combustion is deliberately caused; specific carbon blacks (particularly lampblack, a pigment for ink) are produced by the incomplete combustion of natural gas.

All except the most severe cases of CO poisoning are reversible. The most important first aid treatment is to get the victim fresh air. Any person who feels ill and suspects CO poisoning might be the cause should immediately evacuate the building, get fresh air, and summon medical assistance. Victims are often treated with 100% oxygen delivered from a mask.

Consumer safety experts stress that combustion appliances should be inspected regularly for leaks or other malfunctions. Also, it is recommended that drivers should inspect exhaust pipes and clear any obstructions before starting vehicles that have been parked in snow. The Consumer Product Safety Commission recommends that CO detectors be installed near bedrooms in residences. These detectors sound an alarm before the gas reaches potentially lethal concentrations in the air.



Carbon monoxide detectors are relatively inexpensive and save lives.

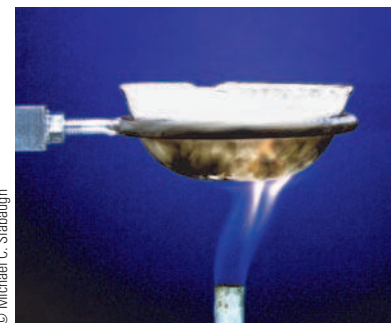


Figure 11.20 A luminous yellow flame from a laboratory burner produces a deposit of carbon when insufficient air (O₂) is mixed with the gaseous fuel.

Concept Summary

Carbon: The Element of Organic Compounds. Organic compounds contain carbon, and organic chemistry is the study of those compounds. Inorganic chemistry is the study of the elements and all noncarbon compounds. **Objective 2, Exercise 11.4** Carbon compounds are of tremendous everyday importance to life on Earth and are the basis of all life processes.

Objective 1 Exercise 11.2

Organic and Inorganic Compounds Compared. The properties of organic and inorganic compounds often differ, largely as a result of bonding differences. Organic compounds contain primarily covalent bonds, whereas ionic bonding is more prevalent in inorganic compounds.

Objective 3 Exercise 11.8

Bonding Characteristics and Isomerism. Large numbers of organic compounds are possible because carbon atoms link to form chains and networks. An additional reason for the existence of so many organic compounds is the phenomenon of isomerism. Isomers are compounds that have the same molecular formula but different arrangements of atoms.

Objective 4, Exercise 11.20

Functional Groups: The Organization of Organic Chemistry.

All organic compounds are grouped into classes based on characteristic features called functional groups. Compounds with their functional groups are represented by two types of structural formulas. Expanded structural formulas show all covalent bonds, whereas condensed structural formulas show no covalent bonds or only selected bonds.

Objective 5 Exercise 11.24

Alkane Structures. Alkanes are hydrocarbons that contain only single covalent bonds and can be represented by the formula C_nH_{2n+2} . Alkanes possess a three-dimensional geometry in which each carbon is surrounded by four bonds directed to the corners of a tetrahedron. Methane, the simplest alkane, is an important fuel (natural gas) and a chemical feedstock for the preparation of other organic compounds. The number of structural isomers possible for an alkane increases dramatically with the number of carbon atoms present in the molecule.

The straight-chain isomer is called a normal alkane; others are called branched isomers.

Objective 6 Exercise 11.28

Conformations of Alkanes. Rotation about the single bonds between carbon atoms allows alkanes to exist in many different conformations. When an alkane is drawn using only two dimensions, the structure can be represented in a variety of ways as long as the order of bonding is not changed.

Objective 7 Exercise 11.30

Alkane Nomenclature. Some simple alkanes are known by common names. More complex compounds are usually named using the IUPAC system. The characteristic IUPAC ending for alkanes is *-ane*.

Objective 8 Exercise 11.34

Cycloalkanes. These are alkanes in which the carbon atoms form a ring. The prefix *cyclo-* is used in the names of these compounds to indicate their cyclic nature.

Objective 9 Exercise 11.44

The Shape of Cycloalkanes. The carbon atom rings of cycloalkanes are usually shown as planar, although only cyclopropane is planar. Because rotation about the single bonds in the ring is restricted, certain disubstituted cycloalkanes can exist as geometric (*cis-trans*) isomers.

Objective 10 Exercise 11.54

Physical Properties of Alkanes. The physical properties of alkanes are typical of all hydrocarbons: nonpolar, insoluble in water, less dense than water, and increasing melting and boiling points with increasing molecular weight.

Objective 11 Exercise 11.56

Alkane Reactions. Alkanes are relatively unreactive and remain unchanged by most reagents. The reaction that is most significant is combustion.

Objective 12 Exercise 11.60

Key Terms and Concepts

Alkane (11.5)

Alkyl group (11.7)

Branched alkane (11.5)

cis- (11.9)

Condensed structural formula (11.4)

Conformations (11.6)

Cycloalkane (11.8)

Expanded structural formula (11.4)

Functional group (11.4)

Geometric isomers (11.9)

Homologous series (11.10)

Hybrid orbital (11.3)

Hydrocarbon (11.5)

Hydrophobic (11.10)

Inorganic chemistry (11.1)

Isomerism (11.3)

Normal alkane (11.5)

Organic chemistry (11.1)

Organic compound (11.1)

Saturated hydrocarbon (11.5)

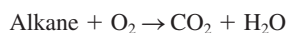
Stereoisomers (11.9)

Structural isomers (11.3)

trans- (11.9)

Key Reactions


1. Complete combustion of alkanes (Section 11.11):



2. Incomplete combustion of alkanes (Section 11.11):



Exercises

 Interactive versions of these problems are assignable in OWL.

Even-numbered exercises are answered in Appendix B.

Blue-numbered exercises are more challenging.

Carbon: The Element of Organic Compounds (Section 11.1)

- 11.1 Why were the compounds of carbon originally called organic compounds?
- 11.2 Name at least six items you recognize to be composed of organic compounds.
- 11.3 Describe what Wöhler did that made the vital force theory highly questionable.
- 11.4 What is the unique structural feature shared by all organic compounds?
- 11.5 Classify each of the following compounds as organic or inorganic:
- | | |
|---------------------|-------------------------------------|
| a. KBr | d. LiOH |
| b. H ₂ O | e. CH ₃ —NH ₂ |
| c. H—C≡C—H | |

Organic and Inorganic Compounds Compared (Section 11.2)

- 11.6 What kind of bond between atoms is most prevalent among organic compounds?
- 11.7 Are the majority of all compounds that are insoluble in water organic or inorganic? Why?
- 11.8 Indicate for each of the following characteristics whether it more likely describes an inorganic or an organic compound. Give one reason for your answer.
- This compound is a liquid that readily burns.
 - A white solid upon heating is found to melt at 735°C.
 - A liquid added to water floats on the surface and does not dissolve.
 - This compound exists as a gas at room temperature and ignites easily.
 - A solid substance melts at 65°C.
- 11.9 Devise a test, based on the general properties in Table 11.1, that you could use to quickly distinguish between the substances in each of the following pairs:
- Gasoline (liquid, organic) and water (liquid, inorganic)

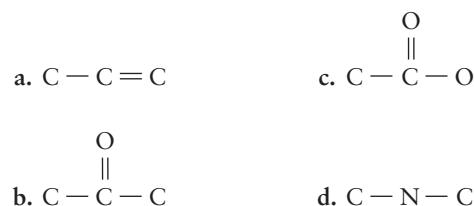
- Naphthalene (solid, organic) and sodium chloride (solid, inorganic)
- Methane (gaseous, organic) and hydrogen chloride (gaseous, inorganic)

11.10 Explain why organic compounds are nonconductors of electricity.

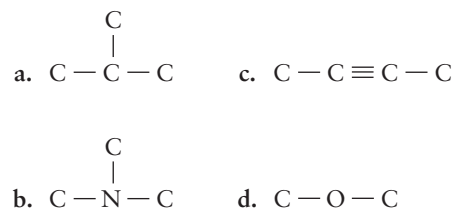
11.11 Explain why the rate of chemical reactions is generally slow for organic compounds and usually fast for inorganic compounds.

Bonding Characteristics and Isomerism (Section 11.3)

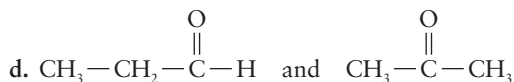
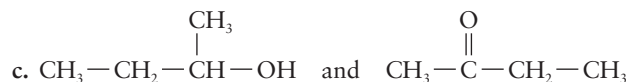
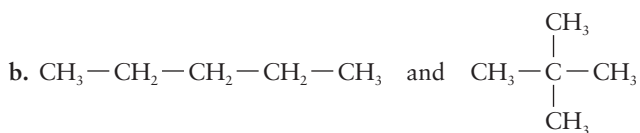
- 11.12 Give two reasons for the existence of the tremendous number of organic compounds.
- 11.13 How many of carbon's electrons are unpaired and available for bonding according to an sp^3 hybridization model?
- 11.14 Describe what atomic orbitals overlap to produce a carbon–hydrogen bond in CH₄.
- 11.15 What molecular geometry exists when a central carbon atom bonds to four other atoms?
- 11.16 Compare the shapes of unhybridized p and hybridized sp^3 orbitals.
- 11.17 Use Example 11.1 and Tables 11.2 and 11.6 to determine the number of covalent bonds formed by atoms of the following elements: carbon, hydrogen, oxygen, nitrogen, and bromine.
- 11.18 Complete the following structures by adding hydrogen atoms where needed:



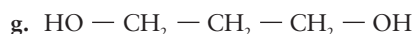
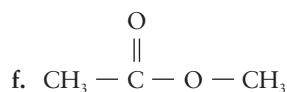
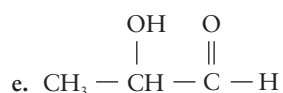
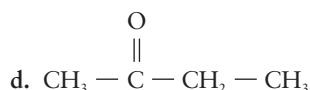
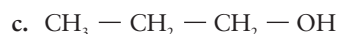
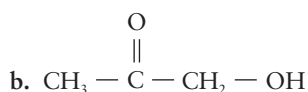
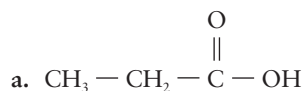
11.19 Complete the following structures by adding hydrogen atoms where needed.



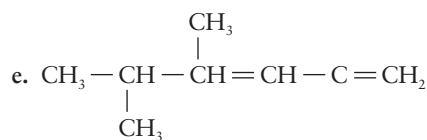
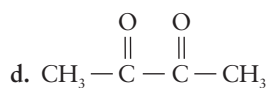
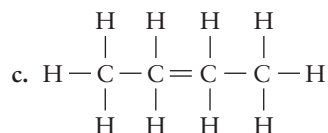
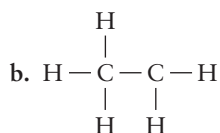
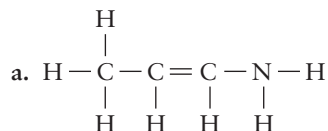
11.20 Which of the following pairs of compounds are structural isomers?



11.21 Group all the following compounds together that represent structural isomers of each other:

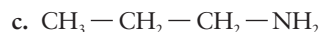
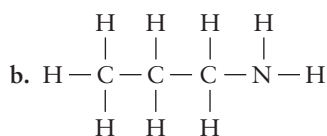


11.22 On the basis of the number of covalent bonds possible for each atom, determine which of the following structural formulas are correct. Explain what is wrong with the incorrect structures.

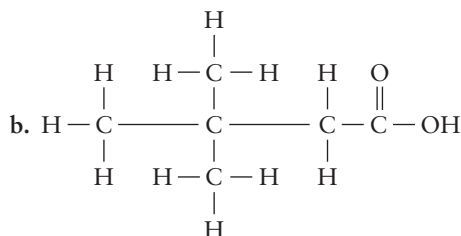
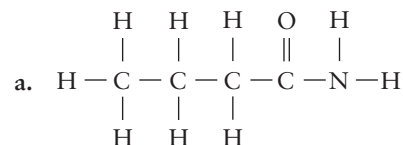


Functional Groups: The Organization of Organic Chemistry (Section 11.4)

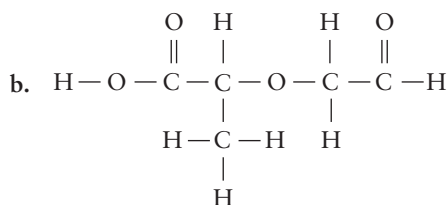
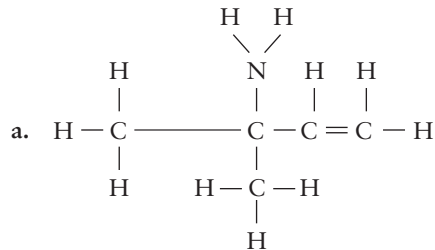
11.23 Identify each of the following as a condensed structural formula, expanded structural formula, or molecular formula:



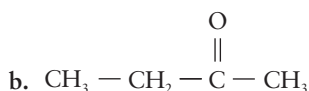
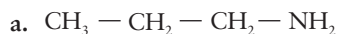
11.24 Write a condensed structural formula for the following compounds:



11.25 Write a condensed structural formula for the following compounds:



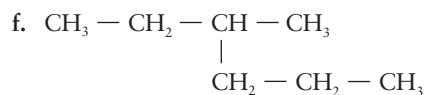
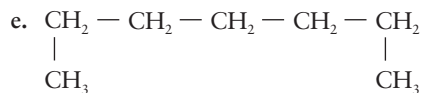
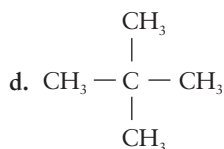
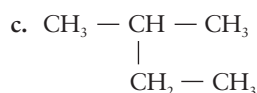
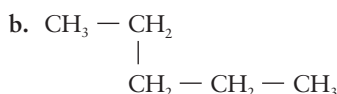
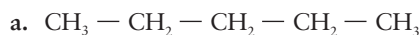
11.26 Write an expanded structural formula for the following:



Alkane Structures (Section 11.5)

11.27 The name of the normal alkane containing 9 carbon atoms is nonane. What are the molecular and condensed structural formulas for nonane?

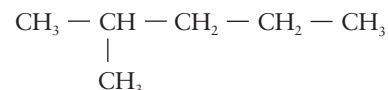
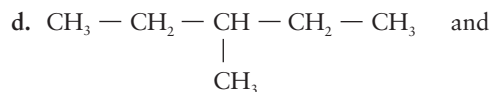
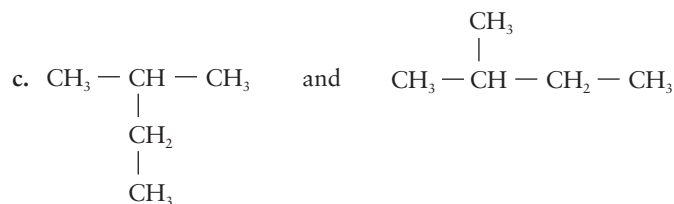
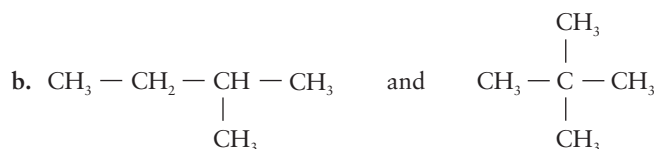
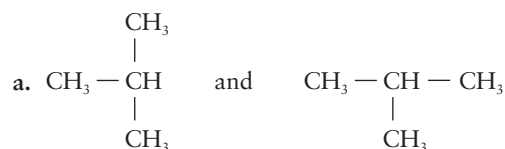
11.28 Classify each of the following compounds as a normal alkane or a branched alkane:



Conformations of Alkanes (Section 11.6)

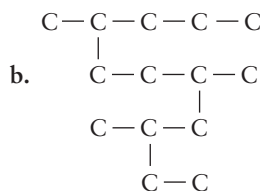
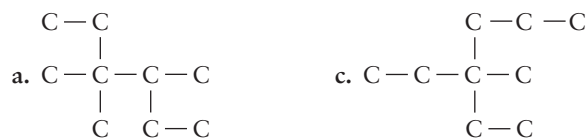
11.29 Why are different conformations of an alkane not considered structural isomers?

11.30 Which of the following pairs represent structural isomers, and which are simply the same compound?

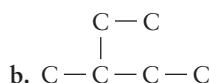
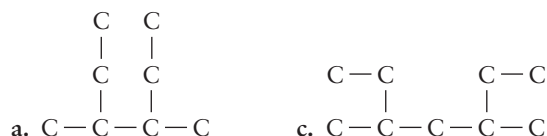


Alkane Nomenclature (Section 11.7)

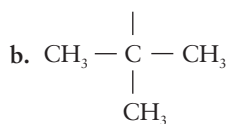
11.31 For each of the following carbon skeletons, give the number of carbon atoms in the longest continuous chain:



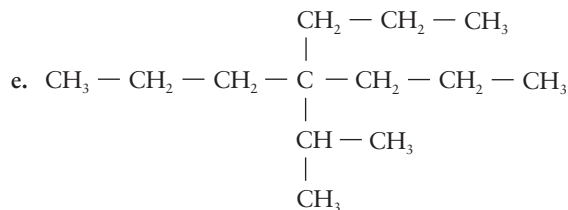
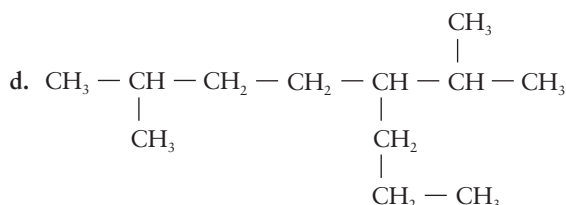
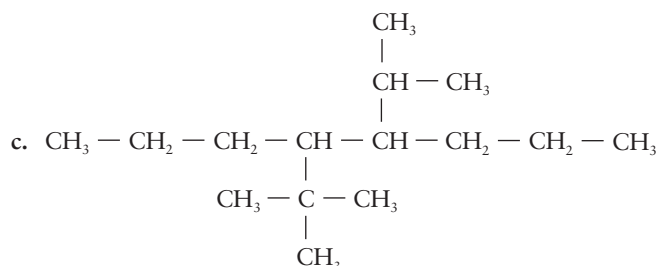
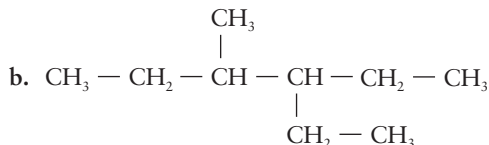
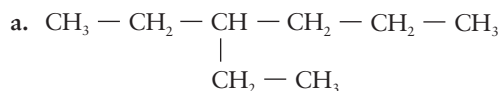
11.32 For each of the following carbon skeletons, give the number of carbon atoms in the longest continuous chain:



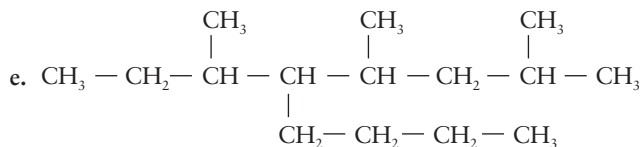
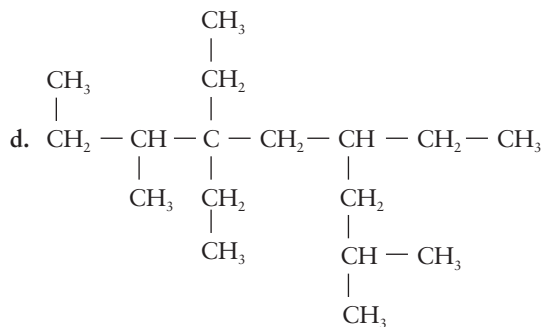
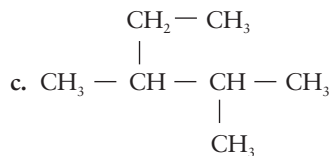
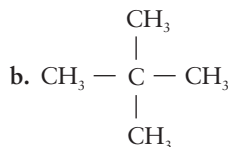
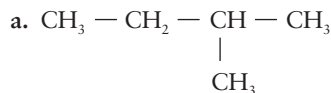
11.33 Identify the following alkyl groups:



11.34 Give the correct IUPAC name for each of the following alkanes:



11.35 Give the correct IUPAC name for each of the following alkanes:



11.36 Draw a condensed structural formula for each of the following compounds:

- 3-ethylpentane
- 2,2-dimethylbutane
- 4-ethyl-3,3-dimethyl-5-propyldecane
- 5-*sec*-butyldecane

11.37 Draw a condensed structural formula for each of the following compounds:

- 2,2,4-trimethylpentane
- 4-isopropyloctane
- 3,3-diethylhexane
- 5-*t*-butyl-2-methylnonane

11.38 Draw the condensed structural formula for each of the three structural isomers of C_5H_{12} , and give the correct IUPAC names.

11.39 Isooctane is 2,2,4-trimethylpentane. Draw structural formulas for and name a branched heptane, hexane, pentane, and butane that are structural isomers of isooctane.

11.40 Draw structural formulas for the compounds and give correct IUPAC names for the five structural isomers of C_6H_{14} .

11.41 The following names are incorrect, according to IUPAC rules. Draw the structural formulas and tell why each name is incorrect. Write the correct name for each compound.

- 2,2-methylbutane
- 4-ethyl-5-methylheptane
- 2-ethyl-1,5-dimethylhexane

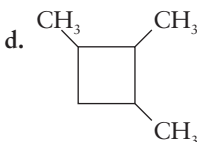
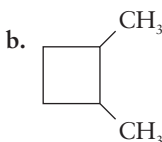
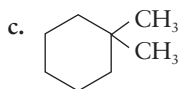
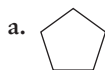
11.42 The following names are incorrect, according to IUPAC rules. Draw the structural formulas and tell why each name is incorrect. Write the correct name for each compound.

- 1,2-dimethylpropane
- 3,4-dimethylpentane
- 2-ethyl-4-methylpentane
- 2-bromo-3-ethylbutane

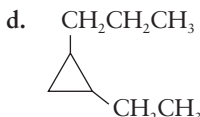
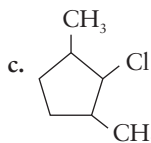
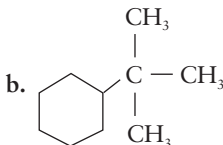
Cycloalkanes (Section 11.8)

11.43 The general formula for alkanes is C_nH_{2n+2} . Write a general formula for cycloalkanes.

11.44 Write the correct IUPAC name for each of the following:



11.45 Write the correct IUPAC name for each of the following:



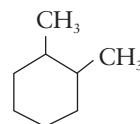
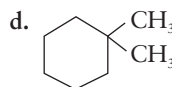
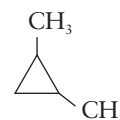
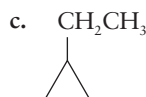
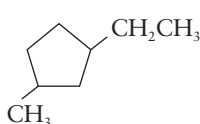
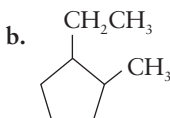
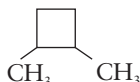
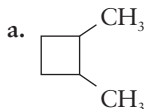
11.46 Draw the structural formulas corresponding to each of the following IUPAC names:

- isopropylcyclopentane
- 1,1-dimethylcyclobutane
- 1-isobutyl-3-isopropylcyclohexane

11.47 Draw the structural formulas corresponding to each of the following IUPAC names:

- 1,2-diethylcyclopentane
- 1,2,4-trimethylcyclohexane
- propylcyclobutane

11.48 Which of the following pairs of cycloalkanes represent structural isomers?



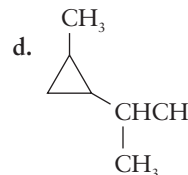
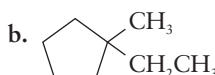
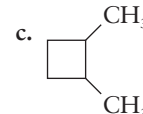
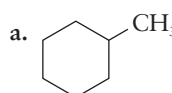
11.49 Draw structural formulas for the five structural isomers of C_5H_{10} that are cycloalkanes.

The Shape of Cycloalkanes (Section 11.9)

11.50 Why does cyclohexane assume a chair form rather than a planar hexagon?

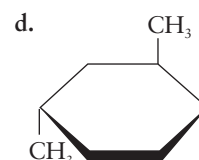
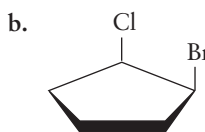
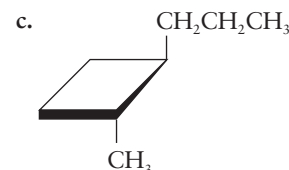
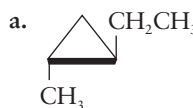
11.51 Explain the difference between geometric and structural isomers.

11.52 Which of the following cycloalkanes could show geometric isomerism? For each that could, draw structural formulas, and name both the *cis*- and the *trans*- isomers.



11.53 Draw structural formulas for *cis*- and *trans*-1,3-dibromocyclobutane.

11.54 Using the prefix *cis*- or *trans*-, name each of the following:



11.55 For each of the following molecular formulas, give the structural formulas requested. In most cases, there are several possible structures.

- C_6H_{14} , a normal alkane
- C_6H_{14} , a branched alkane
- C_5H_{12} , a pair of conformations
- C_5H_{12} , a pair of structural isomers
- C_5H_{10} , a cyclic hydrocarbon
- C_6H_{12} , two cycloalkane geometric isomers
- C_6H_{12} , a cycloalkane that has no geometric isomers

Physical Properties of Alkanes (Section 11.10)

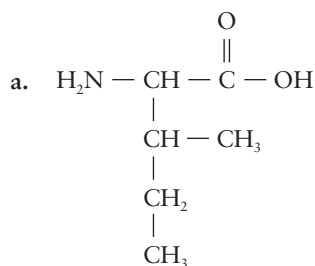
11.56 The compound decane is a straight-chain alkane. Predict the following:

- Is decane a solid, liquid, or gas at room temperature?
- Is it soluble in water?
- Is it soluble in hexane?
- Is it more or less dense than water?

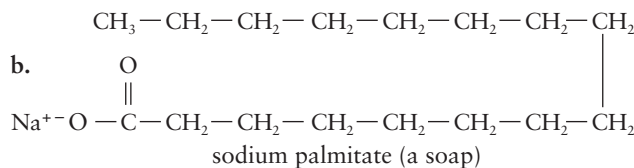
11.57 Explain why alkanes of low molecular weight have lower melting and boiling points than water.

11.58 Suppose you have a sample of 2-methylhexane and a sample of 2-methylheptane. Which sample would you expect to have the higher melting point? Boiling point?

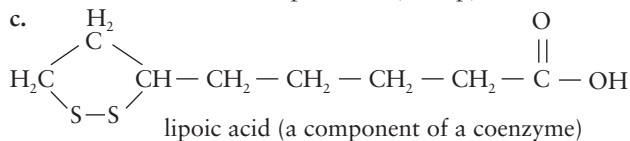
11.59 Identify (circle) the alkane-like portions of the following molecules:



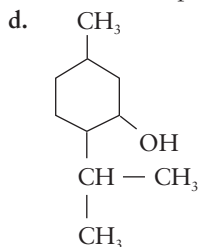
isoleucine (an amino acid)



sodium palmitate (a soap)



lipoic acid (a component of a coenzyme)

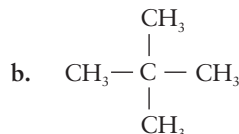


menthol (a flavoring)

Alkane Reactions (Section 11.11)

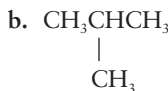
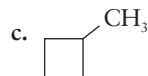
11.60 Write a balanced equation to represent the complete combustion of each of the following:

a. Butane



11.61 Write a balanced equation to represent the complete combustion of each of the following:

a. Propane



11.62 Write a balanced equation for the incomplete combustion of hexane, assuming the formation of carbon monoxide and water as the products.

11.63 Why is it dangerous to relight a furnace when a foul odor is present?

Additional Exercises

11.64 Using the concept of dispersion forces, explain why most cycloalkanes have higher boiling points than normal alkanes with the same number of carbon atoms.

11.65 Draw a Lewis electron dot formula for ethane ($\text{CH}_3 - \text{CH}_3$). Explain why ethane molecules do not hydrogen-bond.

11.66 Your sometimes inept lab technician performed experiments to determine the vapor pressure of pentane, hexane, and heptane at 20°C . He gave you back the numbers of 113.9, 37.2, and 414.5 torr without identifying the compounds. Which vapor pressure value goes with which compound?

11.67 How many grams of water will be produced by the complete combustion of 10.0 g of methane (CH_4)?

11.68 How many liters of air at STP are needed to completely combust 1.00 g of methane (CH_4)? Air is composed of about 21% v/v oxygen (O_2).

Allied Health Exam Connection

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11.69 Which of the following are examples of organic compounds?

- CaCO_3
- NH_3
- NaCl
- $\text{C}_6\text{H}_{12}\text{O}_6$

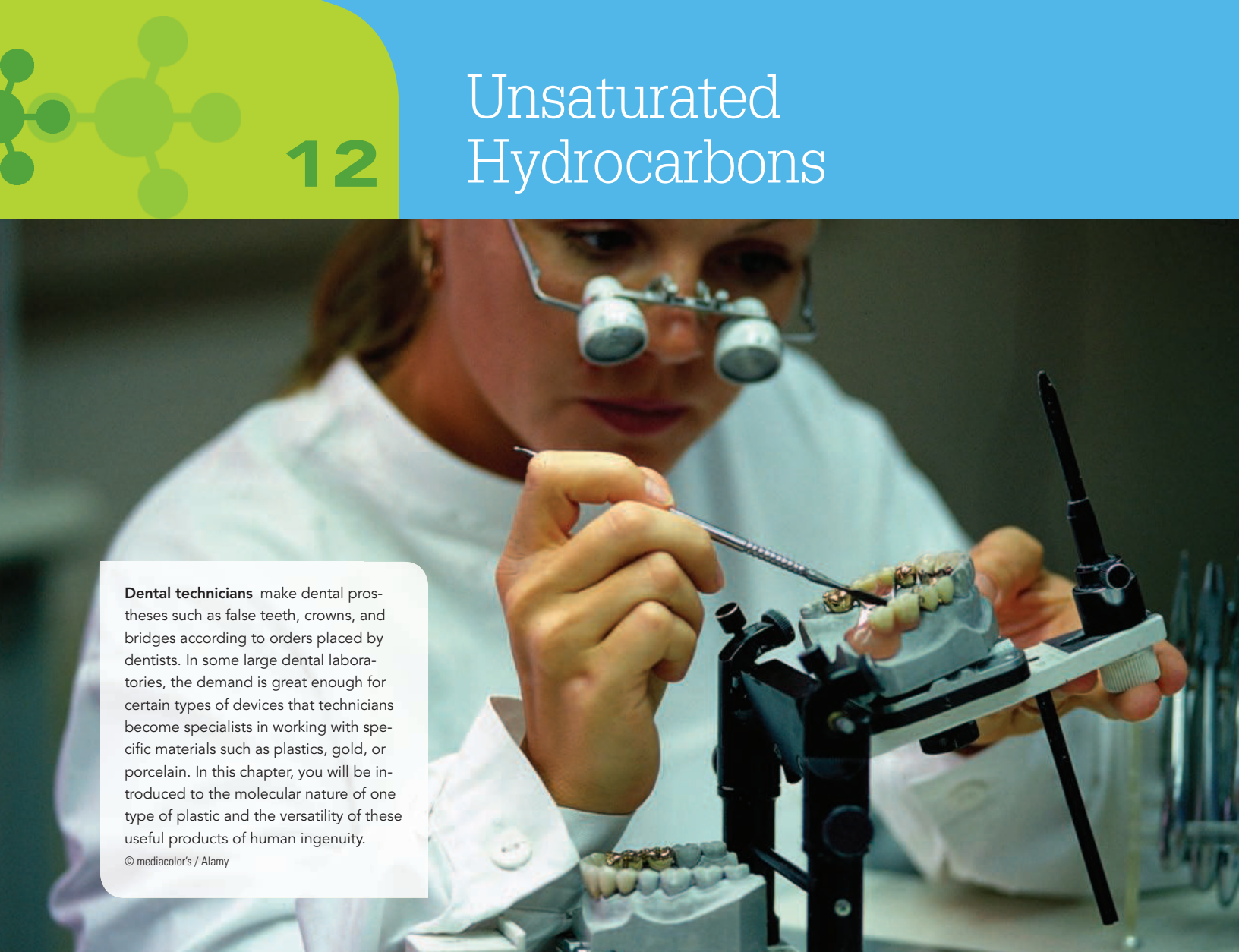
11.70 Use the generic formula for alkanes ($\text{C}_n\text{H}_{2n+2}$) to derive molecular and condensed structural formulas for:

- Propane, 3 carbon atoms
- Octane, 8 carbon atoms
- Butane, 4 carbon atoms

- 11.71** Organic compounds are the basis for life as we know it because:
- Carbon-to-carbon bonds are strong
 - Carbon can form long chains
 - Carbon chains can include other elements to give rise to different functional groups
 - All of the above are correct
- 11.72** Which of the following is an example of an alkane?
- C_2H_6
 - C_2H_4
 - CH_3OH
 - $C_6H_{12}O_6$
- 11.73** The name of the compound $CH_3-CH_2-CH_2-CH_3$ is most likely:
- cyclobutane
 - butane
 - butene
 - butyne
- 11.74** Which of the following are the general products of a combustion reaction?
- $C_{(s)}$, O_2 , and H_2
 - $C_{(s)}$, H_2O , and O_2
 - CO_2 and H_2
 - CO_2 and H_2O
- 11.75** Which is mostly methane?
- oil
 - natural gas
 - coal waste
 - propane
- 11.76** The deadly property of carbon monoxide, if inhaled, is due to its:
- high affinity for O_2
 - low affinity for hemoglobin
 - high affinity for hemoglobin
 - conversion to cyanide gas

Chemistry for Thought

- 11.77** Would you expect a molecule of urea produced in the body to have any different physical or chemical properties from a molecule of urea prepared in a laboratory?
- 11.78** Why might the study of organic compounds be important to someone interested in the health or life sciences?
- 11.79** Why do very few aqueous solutions of organic compounds conduct electricity?
- 11.80** The ski wax being examined in Figure 11.2 has a relatively low melting point. What does that fact reveal about the forces between molecules?
- 11.81** Charcoal briquettes sometimes burn with incomplete combustion when the air supply is limited. Why would it be hazardous to place a charcoal grill inside a home or a camper in an attempt to keep warm?
- 11.82** If carbon did not form hybridized orbitals, what would you expect to be the formula of the simplest compound of carbon and hydrogen?
- 11.83** What types of sports equipment are made from graphite fibers besides that shown in Figure 11.8?
- 11.84** A semi truck loaded with cyclohexane overturns during a rainstorm, spilling its contents over the road embankment. If the rain continues, what will be the fate of the cyclohexane?
- 11.85** On the way home from school, you drove through a construction zone, resulting in several tar deposits on the car's fender. What substances commonly found in the kitchen might help in removing the tar deposits?
- 11.86** Oil spills along coastal shores can be disastrous to the environment. What physical and chemical properties of alkanes contribute to the consequences of an oil spill?
- 11.87** Why might some farmers hesitate to grow and sell organic produce?




Dental technicians make dental prostheses such as false teeth, crowns, and bridges according to orders placed by dentists. In some large dental laboratories, the demand is great enough for certain types of devices that technicians become specialists in working with specific materials such as plastics, gold, or porcelain. In this chapter, you will be introduced to the molecular nature of one type of plastic and the versatility of these useful products of human ingenuity.

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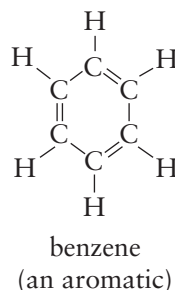
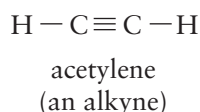
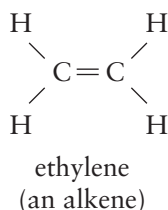
Learning Objectives

When you have completed your study of this chapter, you should be able to:

- 1 Classify unsaturated hydrocarbons as alkenes, alkynes, or aromatics. (Section 12.1)
- 2 Write the IUPAC names of alkenes from their molecular structures. (Section 12.1)
- 3 Predict the existence of geometric (cis-trans) isomers from formulas of compounds. (Section 12.2)
- 4 Write the names and structural formulas for geometric isomers. (Section 12.2)
- 5 Write equations for addition reactions of alkenes, and use Markovnikov's rule to predict the major products of certain reactions. (Section 12.3)
- 6 Write equations for addition polymerization, and list uses for addition polymers. (Section 12.4)
- 7 Write the IUPAC names of alkynes from their molecular structures. (Section 12.5)
- 8 Classify organic compounds as aliphatic or aromatic. (Section 12.6)
- 9 Name and draw structural formulas for aromatic compounds. (Section 12.7)
- 10 Recognize uses for specific aromatic compounds. (Section 12.8)

 **OWL** Online homework for this chapter may be assigned in OWL.

Unsaturated hydrocarbons contain one or more double or triple bonds between carbon atoms and belong to one of three classes: alkenes, alkynes, or aromatic hydrocarbons. **Alkenes** contain one or more double bonds, **alkynes** contain one or more triple bonds, and **aromatic hydrocarbons** contain three double bonds alternating with three single bonds in a six-carbon ring. Ethylene (the simplest alkene), acetylene (the simplest alkyne), and benzene (the simplest aromatic) are represented by the following structural formulas:



Alkenes and alkynes are called *unsaturated* because more hydrogen atoms can be added in somewhat the same sense that more solute can be added to an unsaturated solution. Benzene and other aromatic hydrocarbons also react to add hydrogen atoms; in general, however, they have chemical properties very different from those of alkenes and alkynes.

12.1 The Nomenclature of Alkenes

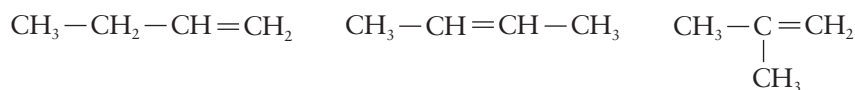
Learning Objectives

1. Classify unsaturated hydrocarbons as alkenes, alkynes, or aromatics.
2. Write the IUPAC names of alkenes from their molecular structure.

The general formula for alkenes is C_nH_{2n} (the same as that for cycloalkanes). The simplest members are well known by their common names, ethylene and propylene:



Three structural isomers have the formula C_4H_8 :



The number of structural isomers increases rapidly as the number of carbons increases because, besides variations in chain length and branching, variations occur in the position of the double bond. IUPAC nomenclature is extremely useful in differentiating among these many alkene compounds.

The IUPAC rules for naming alkenes are similar to those used for the alkanes, with a few additions to indicate the presence and location of double bonds.

- Step 1.** Name the longest chain that contains the double bond. The characteristic name ending is *-ene*.
- Step 2.** Number the longest chain of carbon atoms so that the carbon atoms joined by the double bond have numbers as low as possible.
- Step 3.** Locate the double bond by the lower-numbered carbon atom bound by the double bond.
- Step 4.** Locate and name attached groups.
- Step 5.** Combine the names for the attached groups and the longest chain into the name.

unsaturated hydrocarbon A hydrocarbon containing one or more multiple bonds.

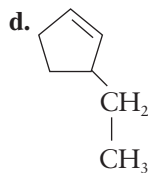
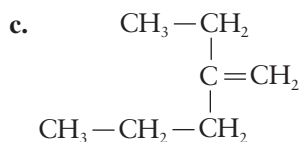
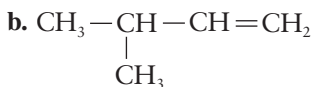
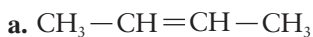
alkene A hydrocarbon containing one or more double bonds.

alkyne A hydrocarbon containing one or more triple bonds.

aromatic hydrocarbons Any organic compound that contains the characteristic benzene ring or similar feature.

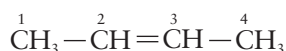
Example 12.1

Name the following alkenes:



Solution

- a. The longest chain containing a double bond has four carbon atoms. The four-carbon alkane is butane. Thus, the compound is a butene:



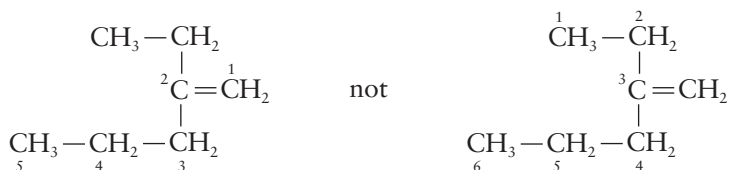
The chain can be numbered from either end because the double bond will be between carbons 2 and 3 either way. The position of the double bond is indicated by the lower-numbered carbon atom that is double-bonded, carbon 2 in this case. The name is 2-butene.

- b. To give lower numbers to the carbons bound by the double bond, the chain is numbered from the right:



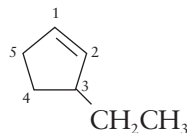
Thus, the compound is a 1-butene with an attached methyl group on carbon 3. Therefore, the name is 3-methyl-1-butene.

- c. Care must be taken to select the longest chain containing the double bond. This compound is named as a pentene and not as a hexene because the double bond is not contained in the six-carbon chain:



The compound is a 1-pentene with an ethyl group at position 2. Therefore, the name is 2-ethyl-1-pentene.

- d. In cyclic alkenes, the ring is numbered so as to give the lowest possible numbers to the double-bonded carbons (they become carbons 1 and 2). The numbering direction around the ring is chosen so that attached groups are located on the lowest-numbered carbon atoms possible. Thus, the name is 3-ethylcyclopentene:



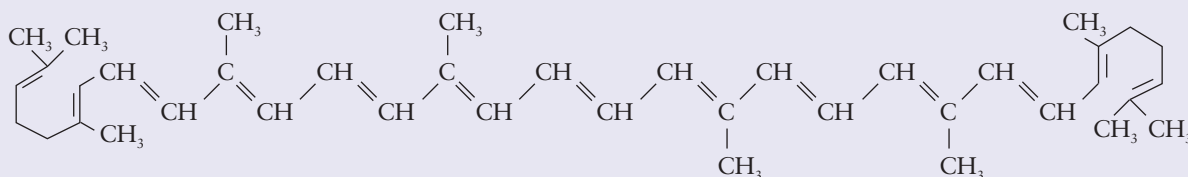
Notice that it is not called 3-ethyl-1-cyclopentene because the double bond is always between carbons 1 and 2, and therefore its position need not be indicated.

Watermelon: A Source of Lycopene



Summer conjures up images of picnic tables and cool, fresh slices of watermelon. What most watermelon eaters don't realize is that in addition to enjoying a cool summer snack, they are also benefitting

from a great source of lycopene, a red substance with the following highly unsaturated molecular structure:



Lycopene is known to help prevent certain types of cancer as well as heart disease, and watermelon is one of a small number of foods that contain this useful compound in large quantities. Other good food sources of lycopene are tomatoes, guava, and pink grapefruit.

The anti-cancer characteristics of lycopene are attributed to its antioxidant properties. It slows or prevents highly reactive oxygen-containing molecules from oxidizing cell components and causing the cells to malfunction. Research results indicate that tomatoes in the diet—especially cooked tomatoes, which contain concentrated amounts of lycopene—reduce the incidence of prostate cancer. In a study conducted at Harvard University, the incidence of prostate cancer was one-third lower in men who ate a lycopene-rich diet compared to a group who ate a low-lycopene diet.

It was long thought that heat-processed tomatoes represented the best source of lycopene in the diet. This was based on the large amounts of lycopene found in small servings of tomato juice. However, recent studies have found that red seedless watermelons contain as much lycopene as cooked tomatoes and in some cases more, depending on the variety of melon and the growing conditions.

Another factor to consider when looking for a good lycopene source is the ability of the body to digest and use the compound (bioavailability) when a lycopene-containing food is eaten. For example, the lycopene in cooked tomatoes is absorbed more readily during digestion than the lycopene in raw tomatoes. Also, it has been found that the absorption rate of lycopene goes up if the food containing it is eaten with food that contains fat. This characteristic is related to the nonpolar nature of lycopene molecules and their resulting significant solubility in fats. The level of lycopene in body fat is used as an indication of how much lycopene has been consumed in the diet. Increased lycopene levels in fat tissue have been linked to reduced risk of heart attack. The bioavailability of lycopene from raw watermelon has been found to be equal to that of lycopene obtained from

cooked tomatoes. Studies are now being conducted to determine if the bioavailability of watermelon lycopene increases if the watermelon is eaten in a diet that also includes fat-containing foods.

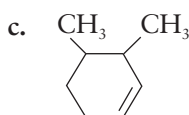
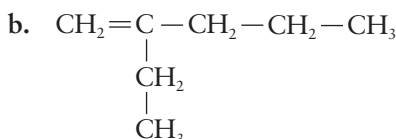
In addition to its ability to satisfy a sweet tooth and serve as a good source of lycopene, watermelon is also an excellent source of the vitamins A, B₆, C, and thiamin. It is also fat free and low in calories, so why not include a big juicy slice of it as one of the five servings of fruits and vegetables recommended by the American Institute of Cancer Research that will reduce your risk of cancer? Enjoy yourself and try not to drip juice on your clothes.



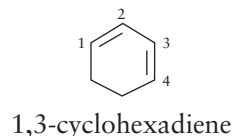
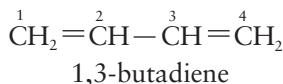
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Watermelon is an excellent source of lycopene.

Learning Check 12.1 Give the IUPAC name for each of the following:

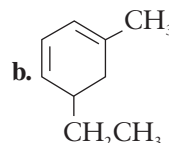
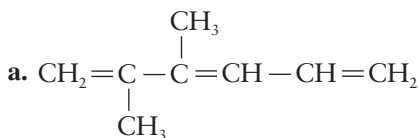


Some compounds contain more than one double bond per molecule. Molecules of this type are important components of natural and synthetic rubber and other useful materials. The nomenclature of these compounds is the same as for the alkenes with one double bond, except that the endings *-diene*, *-triene*, and the like are used to denote the number of double bonds. Also, the locations of all the multiple bonds must be indicated in all molecules, including those with rings:



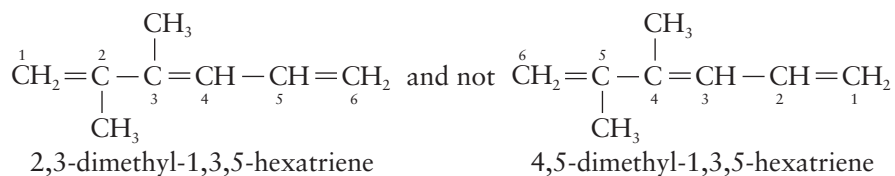
Example 12.2

Name the following compounds:

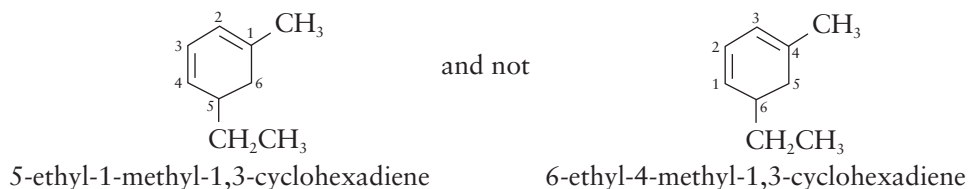


Solution

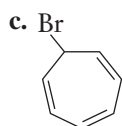
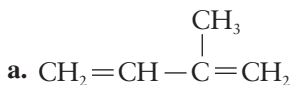
- a. This compound is a methyl-substituted hexatriene. The chain is numbered from the end nearest the branch because the direction of numbering, again, makes no difference in locating the double bonds correctly. The name is 2,3-dimethyl-1,3,5-hexatriene:



- b. This compound is a substituted cyclohexadiene. The ring is numbered as shown. The name is 5-ethyl-1-methyl-1,3-cyclohexadiene:



Learning Check 12.2 Give the IUPAC name for each of the following:



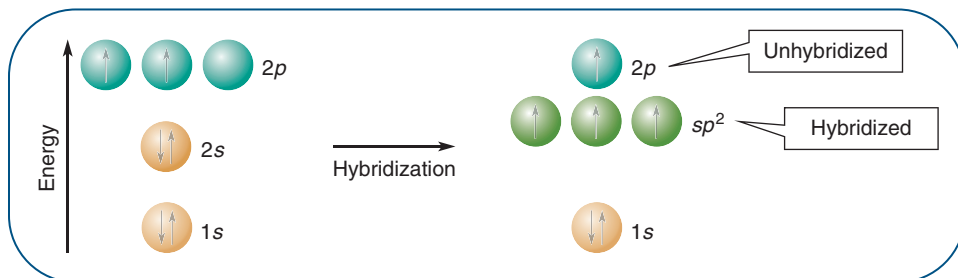


Figure 12.1 A representation of sp^2 hybridization of carbon. During hybridization, two of the $2p$ orbitals mix with the single $2s$ orbital to produce three sp^2 hybrid orbitals. One $2p$ orbital is not hybridized and remains unchanged.

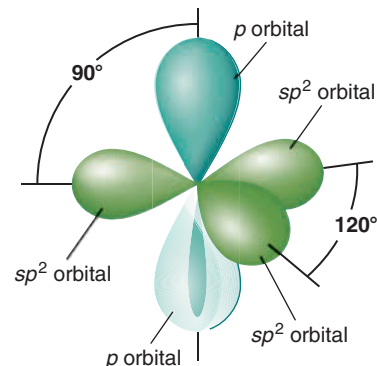


Figure 12.2 The unhybridized p orbital is perpendicular to the plane of the three sp^2 hybridized orbitals.

12.2 The Geometry of Alkenes

Learning Objectives

- Predict the existence of geometric (cis-trans) isomers from the formulas of compounds.
- Write the names and structural formulas for geometric isomers.

The hybridization of atomic orbitals discussed in Section 11.3 to explain the bonding characteristics of carbon atoms bonded to four other atoms can also be used to describe alkenes, compounds in which some carbon atoms are bonded to only three atoms. This hybridization involves mixing a $2s$ orbital and two $2p$ orbitals of a carbon atom to form three hybrid sp^2 orbitals (see Figure 12.1).

The three sp^2 hybrid orbitals lie in the same plane and are separated by angles of 120° . The unhybridized $2p$ orbital of the carbon atom is located perpendicular to the plane of the sp^2 hybrid orbitals (see Figure 12.2).

The bonding between the carbon atoms of ethylene results partially from the overlap of one sp^2 hybrid orbital of each carbon to form a sigma (σ) bond. The second carbon-carbon bond is formed when the unhybridized $2p$ orbitals of the carbons overlap sideways to form what is called a pi (π) bond. The remaining two sp^2 hybrid orbitals of each carbon overlap with the $1s$ orbitals of the hydrogen atoms to form sigma bonds. Thus, each ethylene molecule contains five sigma bonds (one carbon-carbon and four carbon-hydrogen) and one pi bond (carbon-carbon), as shown in Figure 12.3.

Experimental data support this hybridization model. Ethylene has been found to be a planar molecule with bond angles close to 120° between the atoms (see Figure 12.4).

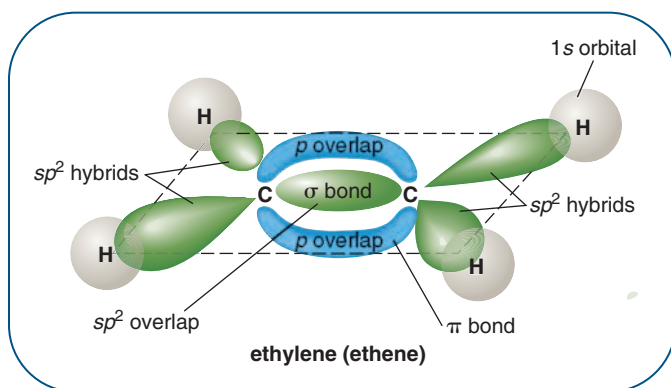


Figure 12.3 The bonding in ethylene is explained by combining two sp^2 hybridized carbon atoms. The C—H sigma bonds all lie in the same plane. The unhybridized p orbitals of the carbon atoms overlap above and below the molecular plane to form the pi bond.

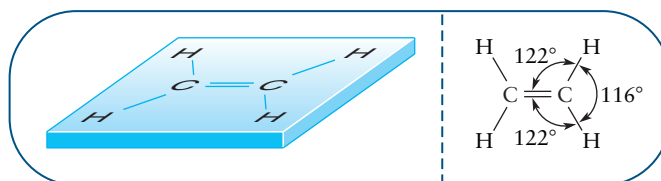


Figure 12.4 All six atoms of ethylene lie in the same plane.