CHAPTER 23

Biological Chemistry

Chemical reactions occur in all living organisms.

Honeybee Pollinating a Poppy

Carbohydrates and Lipids

SECTION 1

OBJECTIVES

- Describe the structural characteristics of simple carbohydrates and complex carbohydrates.
- Explain the role of carbohydrates in living systems.
- Describe the structural characteristics of lipid molecules.
- Identify the functions of lipids in living cells.

Biochemistry is the study of the chemicals and reactions that occur in living things. Biochemical compounds are often large and complex organic molecules, but their chemistry is similar to that of the smaller organic molecules you studied in Chapter 22. Now you will study many important biochemical molecules and learn why they are needed to stay healthy. Two of the most common types of molecules that you may know about are *carbohydrates* and *lipids*. These molecules are important parts of the food that you eat and provide most of the energy that your body needs.

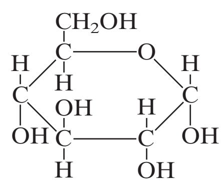
Carbohydrates

Sugars, starches, and *cellulose* belong to the large group of biochemical molecules called carbohydrates. **Carbohydrates** *are molecules that are composed of carbon, hydrogen, and oxygen atoms in a 1:2:1 ratio, and provide nutrients to the cells of living things.* They are produced by plants through a process called *photosynthesis.* Cellulose provides structure and support for plants and starch stores energy in plants. Because animals cannot make all of their own carbohydrates, they must get them from food. Carbohydrates provide nearly all of the energy that is available in most plant-derived food.

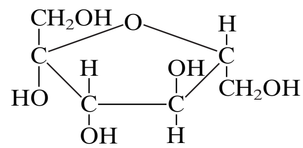
Monosaccharides

A **monosaccharide** *is a simple sugar that is the basic subunit of a carbohydrate.* A single monosaccharide molecule contains three to seven carbon atoms. Monosaccharide compounds are typically sweet-tasting, white solids at room temperature. Because they have polar, hydroxyl (−OH) groups in their molecular structures, they are very soluble in water. The most common monosaccharides are *glucose* (also called *dextrose*) and *fructose.* Although both of these monosaccharides have the formula $C_6(H_2O)_6$, their structural formulas differ. As **Figure 1** shows, glucose in a water solution forms a ring made up of five carbon atoms and one oxygen atom, and fructose in a water solution forms a ring made up of four carbon atoms and one oxygen atom. Notice that both compounds have five −OH groups in their structures.

FIGURE 1 Glucose and fructose both have 6 C, 12 H, and 6 O atoms. The arrangement of the C, H, and O atoms determines the shape and properties of each sugar.



glucose



fructose

Glucose is the most abundant monosaccharide in nature. It is also the most important monosaccharide nutritionally for animals because glucose provides energy for cellular activities. The carbohydrates we eat are broken down into glucose, which may be used immediately by cells or stored in the liver as glycogen for later use. Glucose is also found in some fruits, corn, and the sap of plants.

Fructose, also called *fruit sugar,* is found in most fruits and in honey. The sweetest naturally occurring sugar, fructose is sweeter than table sugar. Because of its sweetness, fructose is sometimes used as a low-calorie sweetener because less fructose is needed to produce the same sweetness that table sugar does.

FIGURE 2 Most of the sugar produced throughout the world comes from sugar beets such as those shown here, or from sugar cane.

Disaccharides

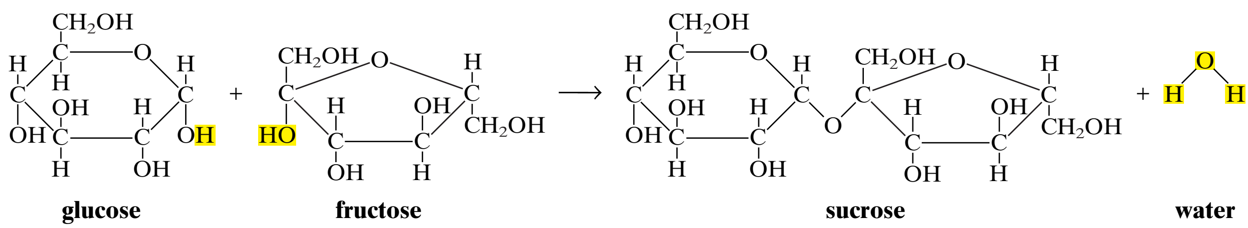
Generally, when someone asks for "sugar," the person is asking for the disaccharide *sucrose*, $C_{12}H_{22}O_{11}$. *A* **disaccharide** *is a sugar that consists of two monosaccharide units that are joined together.* Like monosaccharides, disaccharides have polar hydroxy groups in their molecular structures and therefore are water soluble. A molecule of sucrose forms when a glucose molecule bonds to a fructose molecule. Commercially available sugar comes from sugar cane or sugar beets, such as those shown in **Figure 2.** Another important disaccharide is *lactose*, or milk sugar. Lactose is made up of a sugar called *galactose* and glucose. Human milk is 7 to 8% lactose, but cow's milk is only 4% to 5% lactose. Infant formula may be enriched with lactose to simulate human milk.

Carbohydrate Reactions

Carbohydrates undergo two important kinds of reactions: condensation reactions and hydrolysis reactions. *A* **condensation reaction** *is a reaction in which two molecules or parts of the same molecule combine.* **Figure 3** shows a condensation reaction in which a molecule of glucose combines with a molecule of fructose to yield a molecule of sucrose. Note that in this reaction a molecule of water is also formed.

Disaccharides and longer-chain polysaccharides can be broken down into smaller sugar units by hydrolysis. **Hydrolysis** *is a chemical reaction between water and another substance to form two or more new substances.* Sucrose will undergo a hydrolysis reaction with water, forming glucose and fructose. This hydrolysis, or "water-splitting," reaction occurs in many common processes, such as in the making of jams and jellies.

FIGURE 3 The disaccharide sucrose is formed by a condensation reaction between glucose and fructose.



Cooking sucrose with high acid foods, such as berries and fruits, causes it to break down into a mixture of equal parts of glucose and fructose. This new mixture provides the sweet taste in jams and jellies, which is sweeter than the starting sugar. When lactose is broken down, glucose and galactose are formed. Some people do not produce the enzyme needed to break down the milk sugar in dairy products. This condition is called *lactose intolerance*. People who have this can become ill when they drink milk or eat foods that have milk in them.

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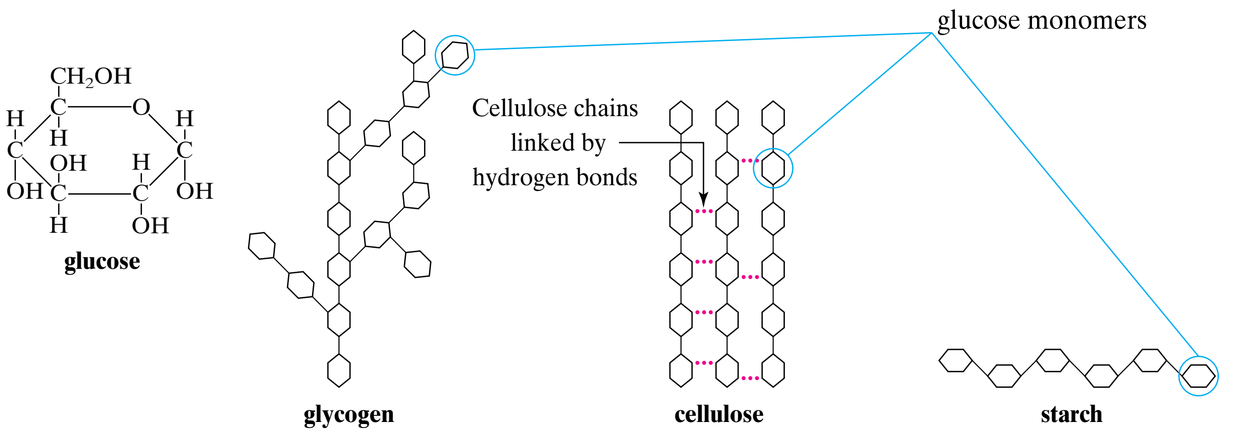
Polysaccharides

When many monosaccharides or disaccharides combine in a series of condensation reactions, they form a polysaccharide. *A* **polysaccharide** *is a carbohydrate made up of long chains of simple sugars.* Cellulose, starch, and glycogen are *polymers* of glucose, or polysaccharides, that contain many glucose monomer units.

As shown in **Figure 4,** the glucose molecules in cellulose chains are arranged in such a way that hydrogen bonds link the hydroxy groups of adjacent glucose molecules to form insoluble fibrous sheets. These sheets of cellulose are the basic component of plant cell walls. More than 50% of the total organic matter in the world is cellulose. People cannot digest cellulose, but when we eat fiber, which is cellulose, it speeds the movement of food through the digestive tract. Microorganisms that can digest cellulose are present in the digestive tracts of some animals. Cows and other ruminants have a special stomach chamber that holds the plants they eat for long periods of time, during which these micro-organisms can break down the cellulose into glucose.

Starch is the storage form of glucose in plants. Starch from foods such as potatoes and cereal grains makes up about two-thirds of the food eaten by people throughout the world. Starch in food is broken down into glucose during digestion. Glucose is broken down further in metabolic reactions that will be discussed later in this chapter.

FIGURE 4 Glucose is the monosaccharide subunit for glycogen, cellulose, and starch. Notice that these three polymers differ in their arrangement of glucose monomers.



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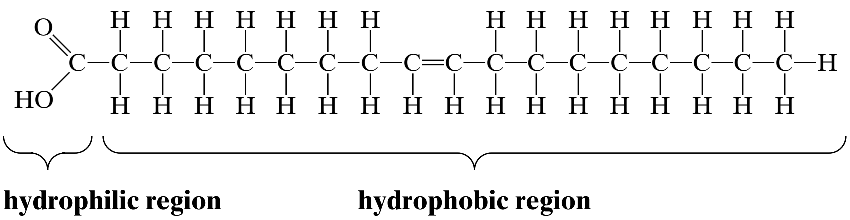
Lipids

Lipids are another important class of nutrients in our diet. They are found in dairy products, grains, meats, and oils. *A* **lipid** *is a type of biochemical that does not dissolve in water, has a high percentage of C and H atoms, and is soluble in nonpolar solvents.* As a class, lipids are not nearly as similar to each other as carbohydrates are. Long-chain fatty acids, phospholipids, steroids, and cholesterol are all lipids.

Fatty Acids and Triglycerides

Fatty acids *consist of a long, nonpolar hydrocarbon "tail" and a polar carboxylic acid functional group at the "head."* Fatty acids are the simplest lipid molecules. They have *hydrophilic* polar heads, but their hydrocarbon chains make them insoluble in water. Fatty acids can also be saturated or unsaturated. *Saturated fatty acids* have no carbon-carbon double bonds, while *unsaturated fatty acids* have one or more double bonds in the hydrocarbon chain. The lipid shown below is oleic acid, which is unsaturated.



Triglycerides are the major component of the fats and oils in your diet. They are formed by condensation reactions in which three fatty acid molecules bond to one glycerol (a type of alcohol) molecule. Fats, such as butter and lard, come from animals, while oils come from plant sources, such as coconuts, peanuts, corn, and olives, as shown in **Figure 5.** Because they have a large amount of saturated fatty acids, fats are solids at room temperature. Oils have more unsaturated fatty acids than fats, and are liquids. Like other animals, humans make fat, which is stored in *adipose* tissue until it is needed as an energy source. Fat has about twice as much energy per gram as carbohydrates or proteins do. Thus, fat is an efficient form of energy storage.

FIGURE 5 Fats, such as lard and butter, are obtained from animals. Oils are found in many different plants.

Fats have another important commercial value based on their ability to react with sodium hydroxide, NaOH, commonly known as *lye*. *When a fat combines with NaOH, an acid-base reaction called* **saponification** *occurs, and a salt and water form.* This salt is made from molecules that have long carboxylic acid chains and is called *soap*. A molecule of soap has a charged ionic head and a nonpolar hydrocarbon tail. This structure allows the ionic head of a soap molecule to dissolve in water and the nonpolar tail to dissolve in nonpolar greases. This property gives the soap its cleaning ability. The chemistry of this reaction is also used as a way of classifying lipids. Lipids that react with a base to form soap are called *saponifiable lipids*, which include fats, oils, and fatty acids.

Other Important Lipids

Compound saponifiable lipids play an important role in biochemical processes. These lipids are structurally similar to triglycerides in that at least one fatty acid is bonded to the central glycerol or glycerol-like unit. These molecules may also have phosphate groups, sugar units, or nitrogen-containing groups. Phospholipids, shown in **Figure 6,** are compound saponifiable lipids and are the main structural component of cell membranes. Phospholipids are arranged in a *bilayer,* or double layer, at the surface of the cell. As **Figure 6** shows, the hydrophilic heads of the phospholipids are on the outside surfaces of the bilayer. The heads are in contact with water-containing solutions inside of the cell and surrounding the cell. The hydrophobic tails point toward the interior of the membrane, away from water-containing solutions. The cell membrane forms a boundary between the cell and its external environment. Only certain substances may pass through the cell membrane. This enables the cell to maintain a stable internal environment.

Nonsaponifiable lipids are nonpolar compounds that do not form soap. They include *steroids,* many *vitamins,* and *bile acids. Cholesterol* is a steroid present in animal cell membranes and is a precursor of many hormones.

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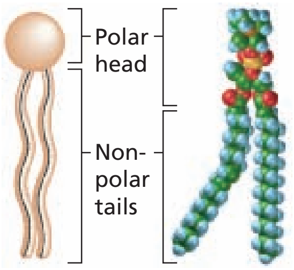
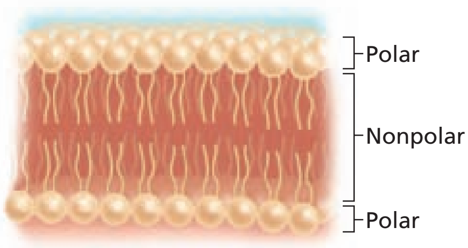
FIGURE 6



A phospholipid's "head" is polar, and its two fatty "tails" are nonpolar.



Phospholipids are arranged in a bilayer, with the hydrophilic heads pointing outward and the hydrophobic tails pointing inward.

SECTION REVIEW

1. Describe two functions of carbohydrates in living systems.
2. Carbohydrates make up about 2% of the mass of the human body, yet we need about 1 tsp of glucose every 15 min to maintain energy for our cells. Where does all of this glucose come from?
3. What is the difference between saponifiable and nonsaponifiable lipids?

Critical Thinking

4. **ANALYZING RELATIONSHIPS** Glucose is soluble in water. Why is cellulose, which is made up of glucose, insoluble in water?
5. **EVALUATING IDEAS** Carbohydrates make up about 90% of the mass of cotton. Why don't humans include cotton in their diet?

SECTION 2

Amino Acids and Proteins

Objectives

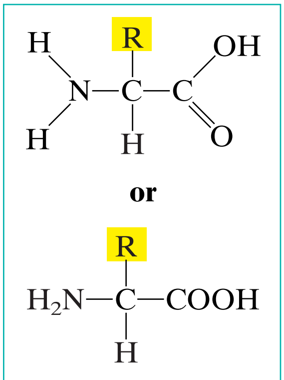
- Describe the basic structure of amino acids and the formation of polypeptides.
- Determine the significance of amino acid side chains to the three-dimensional structure of a protein and the function of a protein.
- Describe the functions of proteins in cells.
- Identify the effects of enzymes on biological molecules.

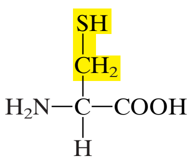
Amino acid molecules are the basic building blocks of proteins. Although only 20 types of amino acids are found in human proteins, more than 700 types of amino acids occur in nature. The human body can synthesize only 11 of the 20 amino acids as needed. The other nine, called the *essential amino acids,* have to be supplied by the food that we eat.

Amino Acids

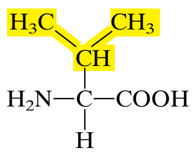
Amino acids *are organic molecules that contain two functional groups: a basic* $-NH_2$ *amino group and an acidic* $-COOH$ *carboxylic acid group.* All of the 20 amino acids have the general structure shown in **Figure 7.** The "R" represents a side chain that is different for each amino acid. The R-groups of the amino acids present in a protein determine the protein's biological activity. The structures of four amino acids—cysteine, valine, glutamic acid, and histidine—are shown below.

FIGURE 7 Amino acids have the same basic structure. The *R* represents a side chain.

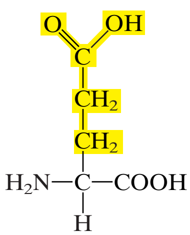




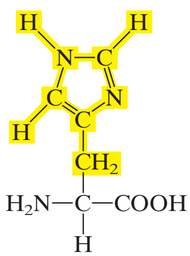
cysteine



valine



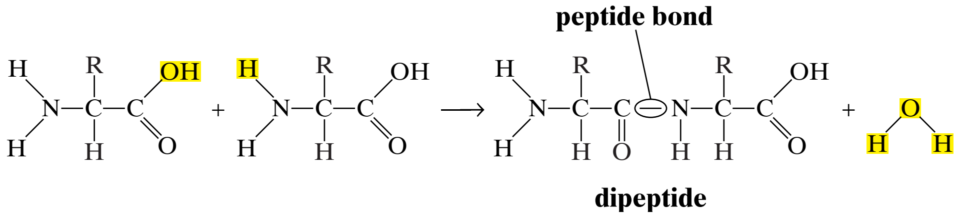
glutamic acid



histidine

Amino Acid Reactions

Two amino acids can react with each other in an acid-base reaction similar to those discussed in Chapter 14. The basic amino group of one amino acid reacts with the acidic carboxylic acid group of another amino acid, forming a *peptide,* and a molecule of water is lost. This reaction, shown below, is classified as a condensation reaction because the two amino acid molecules join together, and water is formed. The bond formed is called a *peptide bond,* and the product is a *dipeptide* because it is made up of two amino acid units. Longer chains are called *polypeptides,* and chains of 50 or more amino acids are called *proteins.*



Peptide bonds can be broken by enzymes called *proteases.* These enzymes are found in cells and tissues where they aid in the digestion of proteins from food, or where they degrade unneeded or damaged proteins.

Proteins

Proteins are found in all living cells and are the most complex and varied class of biochemical molecules. *A* **protein** *is an organic biological polymer that is made up of polypeptide chains of 50 or more amino acids and is an important building block of all cells.* The name *protein* comes from the Greek *proteios,* which means "of first importance." This name was chosen to show the importance of proteins in living things.

Proteins are the second most common molecules found in the human body (after water) and make up about 10% to 20% of the mass of a cell. Made up of specific sequences of amino acids, proteins have molecular masses that range from 6000 to more than 9 million atomic mass units. About 9000 different protein molecules are found in cells in the human body. Nitrogen accounts for about 15% of the mass of a protein molecule, which makes the structure of a protein quite different from that of a carbohydrate or lipid. Most proteins also contain sulfur, and some contain phosphorus or other elements, such as iron, zinc, and copper.

The importance of proteins in living things comes from their many different functions. Besides being the body's main food source for nitrogen and sulfur, proteins have many important catalytic, structural, regulatory, and antibody defense functions. Some different kinds of proteins are *keratin,* which is the main component of hair and fingernails; *enzymes,* which catalyze biochemical reactions; *hemoglobin*, which carries oxygen in the blood; *insulin*, which regulates glucose levels; and *antibodies*, which protect the body from foreign substances.

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FIGURE 8 A scanning electron micrograph showing crystals of the amino acid glycine, one of the building blocks of proteins.

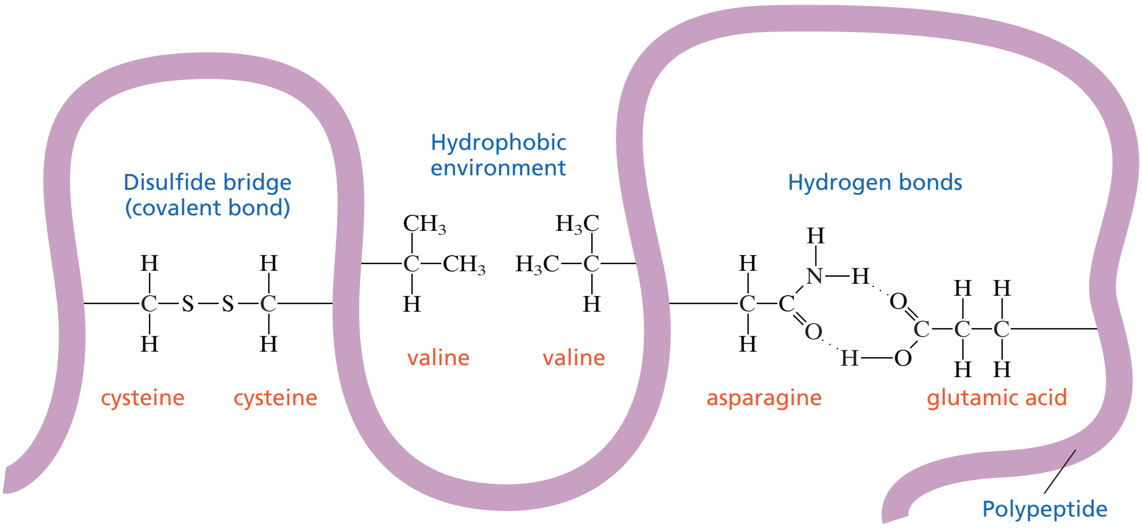
Arrangement of Amino Acids in Peptides and Proteins

Each peptide, polypeptide, or protein is made up of a special sequence of amino acids. A simple set of three-letter abbreviations is used to represent each amino acid in these kinds of molecules. For example, the dipeptide from glycine, shown in **Figure 8,** and glutamic acid would be written as Gly–Glu. The dipeptide Glu–Gly is an isomer of Gly–Glu. Both have the same numbers of C, H, O, and N atoms but in a different order. For the tripeptide Val–Asp–His, made up of valine, asparagine, and histidine, there are five isomers. There are 120 possible isomers for a pentapeptide of five different amino acids. Even though there are only 20 types of amino acids in proteins found in the human body, an incredibly large number of polypeptide and protein molecules are possible. Even for a small protein made up of 100 amino acids, the number of possible combinations of the 20 amino acids is 20^{100}! Polypeptide and protein function depend not only on the kinds and number of amino acids but also on their order. Later, you will see that even the difference of only one amino acid in a polypeptide or protein chain can cause a big change in a protein's activity in a cell.

Amino Acid Side-Chain Reactions

The properties of amino acids—and ultimately polypeptides and proteins—depend on the properties of the side chains present. For example, the side chain of *glutamic acid* is acidic, and the side chain of *histidine* is basic. The side chains of *asparagine* and several other amino acids are polar. In addition, both glutamic acid and asparagine can form hydrogen bonds, shown in **Figure 9.** Some amino acid side chains can form ionic or covalent bonds with other side chains. *Cysteine* is a unique amino acid, because the $-$SH group in cysteine can form a covalent bond with another cysteine side chain. **Figure 9** shows that two cysteine units—at different points on a protein molecule—can bond to form a *disulfide bridge*. Such bonding can link two separate polypeptides or can cause one long protein to bond onto itself to form a loop. In fact, curly hair is a result of the presence of many disulfide bridges in hair protein.

FIGURE 9 Three kinds of interactions between side chains on a polypeptide molecule are shown here. These interactions help determine the shape of a protein.



Shape and Structure of Protein Molecules

The interaction of amino acid side chains determines the shape and structure of proteins, which in turn are important to the proteins' biological functions. In a polypeptide chain or protein, the sequence of the amino acids is called the *primary* (1°) *structure*. The *secondary* (2°) *structure* describes how the chain is coiled or otherwise arranged in space. For example, the alpha (α) helix is a secondary structure that resembles a coiled spring. Another type of secondary structure is the beta (β) pleated sheet, which has accordion-like folds. Both of these secondary structures form because hydrogen bonding occurs between a hydrogen atom attached to the nitrogen atom in one peptide bond and the oxygen atom of another peptide bond farther down the backbone of the protein.

In a protein, the amino acid side chains project out in such a way that they often interact with other side chains located at various positions along the protein backbone. These interactions give the protein its characteristic three-dimensional shape, which is called its *tertiary* (3°) *structure*. The side-chain interactions can include hydrogen bonding, salt bridges, and cysteine-cysteine disulfide bonds. Hydrophobic interactions that occur between nonpolar side chains also contribute to a protein's tertiary structure. Because nonpolar side groups are repulsed by the water found in cells and body fluids, these groups tend to be found in the interior of the protein, where contact with water is minimal. Polar and ionic side chains tend to be on the protein surface, where they are in contact with water. In some proteins, different polypeptides, each of which has its own 3° structure, come together. In the case of hemoglobin, four different polypeptides make up the *quaternary* (4°) *structure*. The four structural levels of proteins are shown in **Figure 10.**

FIGURE 10

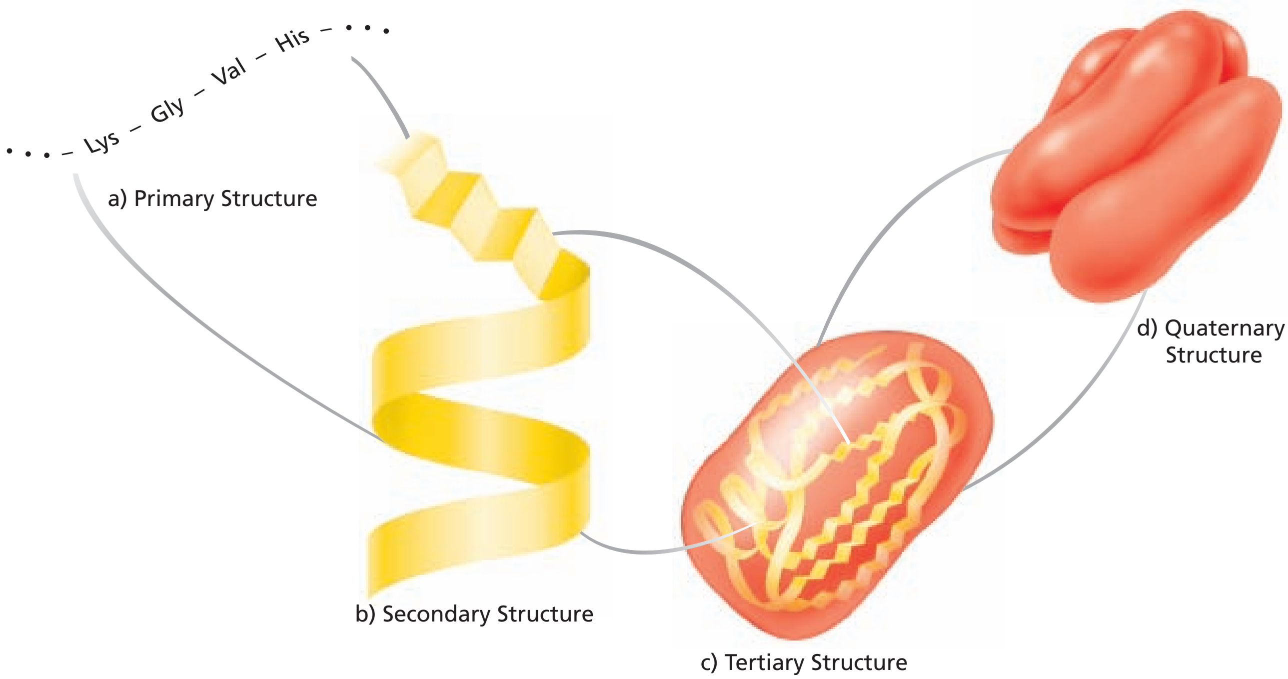


TABLE 1 *Biological Functions of Proteins*

Type of Protein	Function	Examples
Storage	storage of amino acids	*Casein* protein in milk supplies amino acids for baby mammals. Egg white protein, or *ovalbumin,* is a source of amino acids for developing embryos. Plants store proteins in seeds.
Transport	transport of substances	Proteins transport molecules across cell membranes. *Hemoglobin* in blood transports oxygen.
Structural	support	Spiders produce silk fibers, which are proteins, to make webs. *Collagen* and *elastin* give connective tissues strength and flexibility. *Keratin* is found in hair, feathers, horns, hooves, and nails.
Contractile	movement	*Actin* and *myosin* fibers cause movement in muscles. Contractile fibers in cilia and flagella help propel single-celled organisms.
Enzymatic	catalysis of chemical reactions	Enzymes break down large molecules in food within the digestive system.
Hormonal	coordination of processes in an organism	Pancreatic insulin helps regulate blood-sugar levels.
Receptor	response of cell to chemical stimuli	Nerve cell membranes have chemical receptors that detect chemical signals released by other nerve cells.
Defensive	protection against disease	Antibodies attack pathogenic viruses and bacteria.

Biological Functions of Proteins

From **Table 1,** you can see that almost everything that occurs in a living organism depends on one or more proteins. Scientists have discovered that the specific function of a protein is related to the protein's shape. The shape of a protein can generally be described as fibrous or globular. *Fibrous proteins* are insoluble in water and are long, thin, and physically strong. *Globular proteins* are generally soluble in water and are twisted and folded into a globe-like shape.

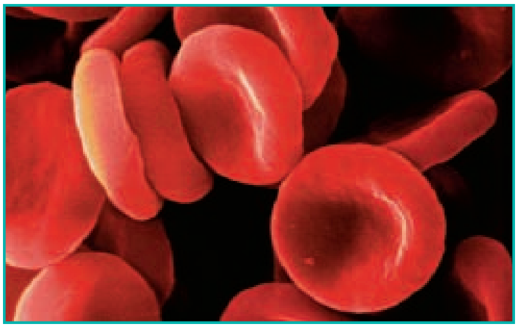
Fibrous proteins give strength and protection to structures in living things. *Keratin* is a fibrous protein whose secondary structure is almost entirely alpha helical in shape. The keratin in nails and hooves is much stiffer than the keratin in fur or wool because of the large number of side-chain interactions that occur between the nail and hoof proteins. *Collagen,* found in bone and tendons, is a triple helix of three intertwined alpha helices, which gives these tissues their strength. *Fibrin* found in silk has a beta-pleated sheet structure. *Elastins* in blood tissue, *fibrins* in blood clots, and *myosins* found in muscle tissue are other kinds of fibrous proteins.

Globular proteins regulate body functions, catalyze reactions, and transport substances. The regulatory hormone *insulin* is a small protein of 51 amino acids in two polypeptide chains. *Myoglobin* transports oxygen in the muscles, and *hemoglobin* transports oxygen in the blood. *Casein,* found in milk and used for food, is also a globular protein. It contains phosphorus, which is needed for bone growth.

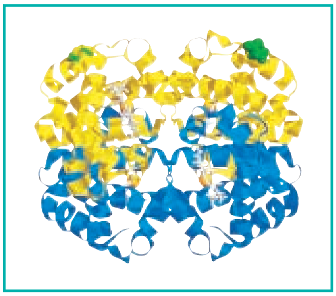
Amino Acid Substitution

A protein's amino acid sequence determines its three-dimensional structure, which in turn determines its function. Even a single substitution of one amino acid for another can change the shape and function of a protein. For example, the genetic disease sickle cell anemia can happen when one amino acid—glutamic acid—is replaced by a molecule of valine. This change in only 1 of 146 amino acids in one of the two protein chains in the hemoglobin molecule causes a major change in the shape of the molecule. This change in the shape of the hemoglobin molecule causes the red blood cells to sickle when oxygen levels are relatively low (as is the case in most body tissues). The sickled cells tend to clog small blood vessels, which prevents the transport of enough oxygen to tissue cells. As a result, people who have sickle cell anemia suffer from shortness of breath. **Figure 11** shows the shape of normal red blood cells and sickled cells. The sickle cell gene is more common in some groups of people than it is in others. In areas where the disease malaria is common, scientists have discovered that sickled cells are more resistant to malarial infection than other cells are. So, people who have sickle cell anemia are more resistant to malaria than other people are.

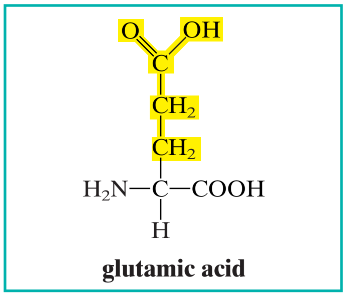
FIGURE 11



(a) The round, flat shape of healthy red blood cells shows they have normal hemoglobin molecules.

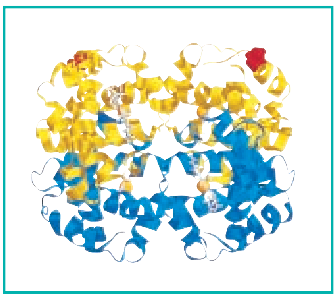


(b) Hemoglobin consists of four polypeptide chains; the area where the change in sickle cell hemoglobin occurs is shown in green.

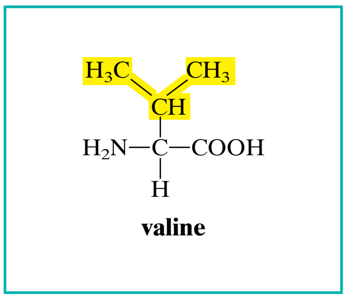


(c) Each of the chains is a polymer of 141 or 146 amino acid units, such as the glutamic acid monomer shown here.

(d) Because of their shape, sickle cells clog small blood vessels.



(e) A genetic mutation causes one glutamic acid to be replaced by valine in the hemoglobin molecules, as shown in red.



(f) The sickle shape of the cell comes from the different shape of the hemoglobin caused by the substitution of one valine for one glutamic acid in the 146 amino acids.

HISTORICAL CHEMISTRY

Dr. Charles Drew and Blood Transfusions

Prior to the 1900s, severe bleeding often resulted in death. But today blood is stored at blood banks, where people "deposit" blood so that they or others can "withdraw" it when needed. Charles Drew was a pioneer in the work of blood transfusions, especially in the use of plasma and the development of blood banks.

The Need for Blood

While in medical school, Drew realized that many lives could be saved if blood could be stored for transfusions. Before 1937, most patients needing blood received it directly from a donor at the time it was needed. In 1938, Drew and physician John Scudder from Britain studied the chemistry of blood to try to find a way to preserve blood. Drew recognized that extracting blood plasma could help solve the problems of storing blood.

There are two main components of blood: blood cells and plasma. Red blood cells, white blood cells, and platelets make up 45% of the volume of blood, while plasma makes up 55% of blood. Plasma is amber colored and is about 90% water. It has more than 100 solutes, including nutrients, antibodies, hormones, and proteins. Drew discovered that plasma could be used as an emergency substitute for blood without testing for blood type, because the ABO blood types are removed with the red blood cells. He also found that plasma could be dehydrated and stored.

▲ *Charles Drew was a pioneer in the development of blood banks.*

During World War II, Drew was the medical supervisor for the "Blood for Britain" program. He also coordinated the American blood storage program by setting up collection centers, standardizing blood bank equipment, and establishing record-keeping procedures and criteria to ensure the safety of blood products. When the United States entered the war, the blood supply was ready, and stored blood and plasma saved several thousand lives.

A Safer Blood Supply

With the emergence of HIV and AIDS in the 1980s, the methods of blood transfusion needed to be modified. Some transfused blood was found to be contaminated with HIV. Before modifications were put into place, as many as 50% of hemophiliacs in the United States could have been infected with HIV because of their frequent transfusions. Although current screening procedures to detect HIV, hepatitis, and other diseases have almost completely removed the risk of transfusing infected blood, some people choose to bank their own blood before undergoing surgery.

Blood Substitutes

The AIDS epidemic has sped up attempts to find an artificial blood substitute. Perfluorocarbons (PFCs), which are made by replacing the hydrogen atoms of hydrocarbons with fluorine atoms, are possible substitutes. Oxygen binds to PFCs 10 to 20 times more readily than it binds to plasma or water, and products using PFCs have a shelf life of up to two years. Another substitute is a hemoglobin-based oxygen carrier (HBOC). HBOCs may be able to transport oxygen until a body is able to replenish its red blood cells. If these products are successful, they could provide a supply of blood substitutes that do not require typing and do not depend on donors.

Questions

1. What was Dr. Drew's contribution to medicine?
2. Why are scientists searching for blood substitutes?

Proteins as Enzymes

An **enzyme** *is a protein that catalyzes a biochemical reaction.* Enzymes make up the largest and most highly specialized class of proteins. All living things have enzymes; as many as 3000 enzymes can be found in a single cell. Most enzymes are water-soluble, globular proteins. The amino acid side chains and the three-dimensional shape of enzymes play a very important role in the enzymatic activity. You will remember from Chapter 17 that catalysts speed up a reaction but do not change as a result of the reaction. An enzyme also does not change the amount of product that is formed in a reaction; it only decreases the time it takes to form the product. Enzymes catalyze both decomposition and synthesis reactions.

Enzymes are also very efficient. For example, the enzyme *carbonic anhydrase* acts on carbonic acid to break it down into carbon dioxide and water. A single molecule of carbonic anhydrase can break down 36 million carbonic acid molecules in 1 minute! Carbon dioxide is carried from all the parts of the body to the lungs as carbonic acid in the blood. In the lungs, carbonic acid is broken down by carbonic anhydrase into CO_2 and water vapor and then is removed from the lungs as a person exhales.

Enzyme Specificity

In addition to being highly efficient, enzymes are very specific. Often, they catalyze just a single reaction as carbonic anhydrase does. In another example, the enzyme peroxidase is responsible only for the decomposition of hydrogen peroxide to water and oxygen. Have you ever put hydrogen peroxide on a minor cut? In the bottle, hydrogen peroxide is relatively stable, but as soon as it comes into contact with your wound, the peroxidase enzyme in the blood catalyzes its decomposition. You see bubbling, which is the result of the gaseous oxygen from the decomposing hydrogen peroxide. As you can see in **Figure 12,** enzymes act by binding to a specific *substrate* molecule. For example, hydrogen peroxide is the substrate in the

FIGURE 12

(a) The enzyme reacts with the substrate in a fast, reversible reaction.

(b) The substrate-enzyme complex can either revert to the reactants or . . .

(c) . . . proceed to the products.

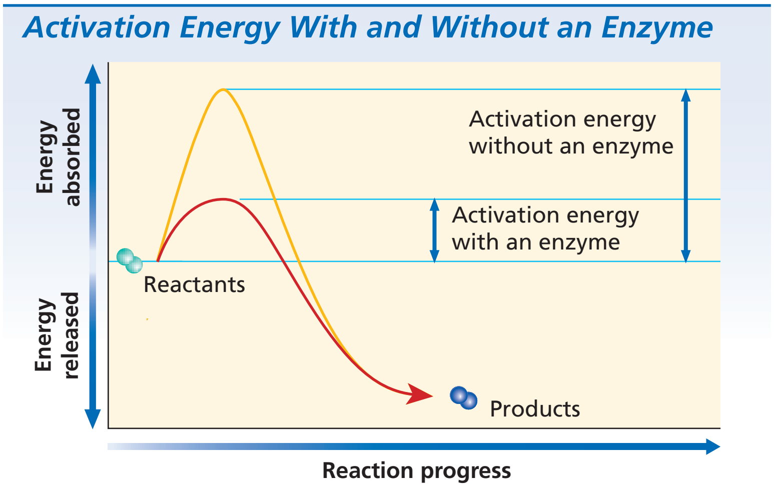


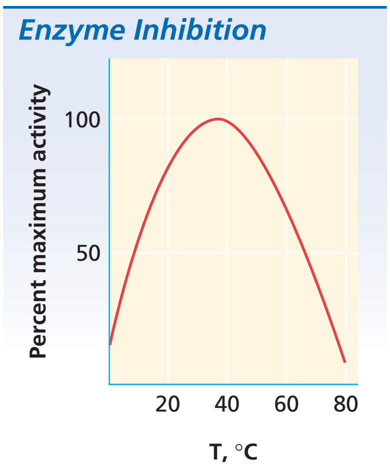
FIGURE 13 Enzymes decrease the activation energy of a chemical reaction. However, the energy change from reactants to products is the same for both the catalyzed and the non-catalyzed reaction.

reaction just discussed. The shape of the enzyme is such that a molecule of hydrogen peroxide can fit into the enzyme at a specific part of the enzyme molecule, called the *active site*. The resulting compound is called the *enzyme-substrate complex*. In the enzyme-substrate complex, hydrogen-oxygen bonds break, and oxygen-oxygen bonds form. Then, the enzyme releases the products and is ready to react with another substrate molecule. This model of enzyme action is called the *lock and key model*.

Enzymes and Reaction Rates

The presence of an enzyme in a chemical reaction can increase the rate of a reaction by a factor of up to 10^{20}. In Chapter 17, you saw that a reaction can occur when two atoms or molecules collide. But only collisions that have enough energy to overcome the *activation energy* and have the proper orientation change reactants into products. As you can see from the graph in **Figure 13,** an enzyme that catalyzes a chemical reaction causes an increase in the rate of the reaction by reducing the activation energy. The enzyme lowers the activation energy by forming the enzyme-substrate complex, which makes breaking bonds in the reactants and forming new bonds in the products easier. Even though the reaction is faster, the net amount of energy required for the reaction or released by the reaction is not changed by the action of an enzyme.

FIGURE 14 Most enzymes have maximum activity within a narrow temperature range. Denaturation in many occurs at temperatures above 50°C and causes a decrease in activity.



Temperature and Enzyme Activity

Proteins, including enzymes, are also affected by changes in temperature. The graph in **Figure 14** shows the relatively narrow range of temperatures within which enzymes typically have maximum activity. Enzymes in the human body work optimally at the normal body temperature of 37°C (98.6°F). At temperatures above 50°C to 60°C enzymes typically show a decline in activity. High heat can denature, or alter, the shape of a protein, which in turn alters the protein's function. **Denaturation** *is a change in a protein's characteristic three-dimensional shape due to changes of its secondary, tertiary, and quaternary structure.* If you have ever cooked an egg, you have caused protein denaturation. The white of an egg is a solution of the protein albumin. When the egg is placed in a hot frying pan, a dramatic change takes place and the semitransparent solution turns into a white solid. Because the primary structure is retained in denaturation, the nutritional value of the egg white is not affected. However, the process is not reversable. Cooling a fried egg does not reverse the denaturation. When food is cooked, the three-dimensional structure of the protein is altered, making the food easier to digest.

pH and Enzyme Activity

Enzymes also typically have maximum activity within a relatively narrow range of pH. The optimal pH for normal cell enzyme functions is almost neutral, about 7.3 to 7.4. Changes in pH can cause changes in protein structure and shape. For example, adding acid or base can interfere with the side-chain interactions and thus change the shape of a protein. Most enzymes become *inactivated,* or no longer work, because of denaturation when the pH changes. When milk sours (because lactic acid has formed), it curdles, and curds of the protein casein form. Yogurt is made by growing acid-producing bacteria in milk, which causes the casein to denature, giving yogurt its consistency.

The digestion of dietary protein by enzymes begins in the stomach. When food is swallowed, the stomach lining produces HCl and *pre-enzymes,* inactive forms of protein-digesting enzymes. These pre-enzymes travel from the stomach lining into the stomach before they become activated by the stomach's low pH of 1.5 to 2.0. This process is important because it prevents the active form of the enzymes from digesting the stomach lining. A layer of mucus protects the lining of the stomach from the enzymes it contains. Once activated, the enzymes catalyze the breakdown of the proteins in food into shorter polypeptide segments. *Pepsin* is a stomach enzyme found in adults. The partially digested protein in food travels into the small intestine, where the pH is 7 to 8. Under these conditions, the enzyme *trypsin* becomes active. It catalyzes the hydrolysis of the polypeptide segments into amino acids, which are absorbed through the intestinal wall and enter the bloodstream. The body uses these newly acquired amino acids to make other amino acids and new protein molecules. **Figure 15** shows how the protein in raw fish looks before and after it is soaked in acidic lime juice. Because the acidic lime juice denatures protein in the fish, the acid-treated fish looks very different.

FIGURE 15 The fish treated with lime has turned white because the acidic lime juice denatures the protein in raw fish.

SECTION REVIEW

1. Which elements do amino acids and proteins have in common with carbohydrates and lipids?
2. What is the difference between an amino acid and a protein?
3. Explain the difference between fibrous proteins and globular proteins.
4. Why are only small amounts of enzymes found in the body?

Critical Thinking

5. **RELATING IDEAS** Explain how the ball-like structure of globular proteins allows them to be water soluble.
6. **INFERRING CONCLUSIONS** If an essential amino acid is in short supply in the diet, it can become a limiting reactant in building any protein that contains the amino acid. Explain why, under these conditions, the only way that the cell could make that protein would be to degrade one of its proteins that contain the limiting amino acid.

SECTION 3

Metabolism

OBJECTIVES

- Describe the role of ATP in cells.
- Explain how energy is released by metabolic reactions.
- Summarize the relationship between anabolism and catabolism.

Metabolism *is the sum of all the chemical processes that occur in an organism.* Complex molecules are broken down into smaller ones through *catabolism,* and simple molecules are used to build bigger ones through a process called *anabolism.* A *metabolic pathway* is a series of linked chemical reactions that occur within a cell and result in a specific product or products. The major metabolic pathways for most organisms are similar. So, one can study the basic metabolic pathways in simple organisms to get information about the reactions in organisms that are more complex, including humans.

ATP: Energy for the Cell

Just as it takes energy to run a chemical factory, cells require energy to make the proteins, carbohydrates, lipids, and nucleic acids that are necessary for life. In addition, the body needs energy as heat to keep warm, mechanical energy to move muscles and pump blood, and electrical energy to move ions across cell membranes. The original source for almost all of the energy needed by living systems is the sun. **Autotrophs,** such as plants, algae, and photosynthetic bacteria, can use sunlight, water, and CO_2 to make carbon-containing biomolecules, including carbohydrates. This process is called *photosynthesis* and occurs in the cells of plants and algae, such as those shown in **Figure 16,** within structures called *chloroplasts*. Chloroplasts contain *chlorophyll,* an organic molecule that absorbs solar energy. This energy is captured immediately in two compounds, one of which is **adenosine triphosphate (ATP). ATP** *is an energy storage molecule that plant cells use to make carbohydrates.* The other compound, known as *NADPH,* is also used in carbohydrate-forming reactions.

Unlike plants, animals cannot use the sun's energy to convert CO_2 into food. Animals must get the energy that they need to sustain life by consuming plants and other animals. *Living things, including most microorganisms, which depend on plants or other animals for food, are called* **heterotrophs.** Heterotrophs use the energy released when complex molecules react to form simpler products to drive chemical reactions in cells. The carbohydrates, lipids, and amino acids that heterotrophs consume undergo several energy-yielding reactions as they break down into simpler molecules. Some of this energy is stored in the ATP molecules, which cells use to drive a wide variety of metabolic reactions.

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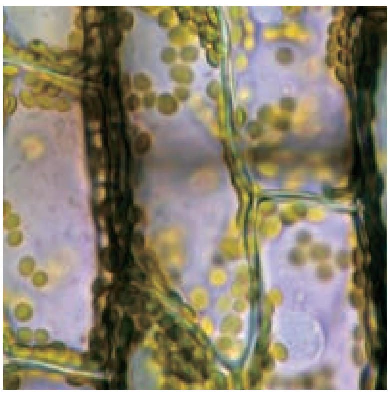
Topic: Exothermic and Endothermic Reactions
Code: HC60555

FIGURE 16 The cells of algae and green plants contain chlorophyll, the green pigment that absorbs light energy from the sun.



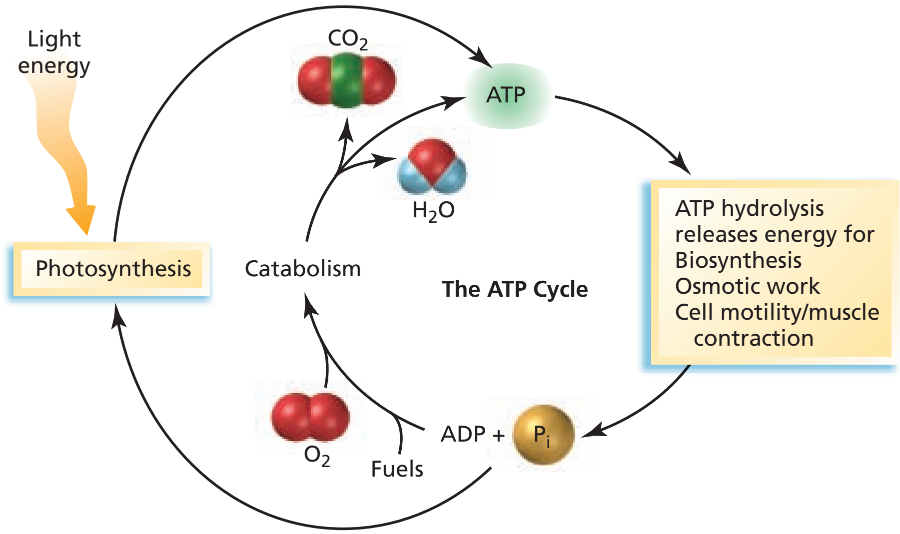


FIGURE 17 ATP is formed by photosynthesis or catabolism. Cell activities that require energy are powered by ATP hydrolysis. "P_i" represents a phosphate.

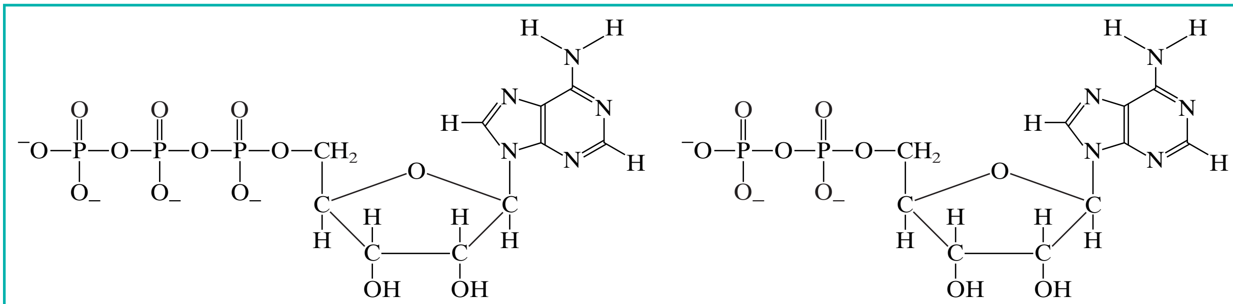
Energy Activities

The cycle between ATP and *ADP, adenosine diphosphate,* is the primary energy exchange mechanism in the body. **Figure 17** provides an overview of the ATP cycle in cells. In this energy cycle, ATP is the molecule that serves to carry energy from energy-storing molecules, carbohydrates, lipids, and proteins to specific energy-requiring processes in cells. When ATP is hydrolyzed to ADP, energy is released to power the cell's activities. The molecular structures of ATP and ADP are closely related, as shown in **Figure 18.**

The difference in the number of phosphate groups between ATP and ADP molecules is the key to the energy exchange in metabolic reactions. The chemical equation below shows the hydrolysis reaction by which ATP is converted into ADP and a phosphate group (represented by the gold-colored ball in **Figure 17**). The free energy for this reaction is –31 kJ, which is the amount of energy available to do work.

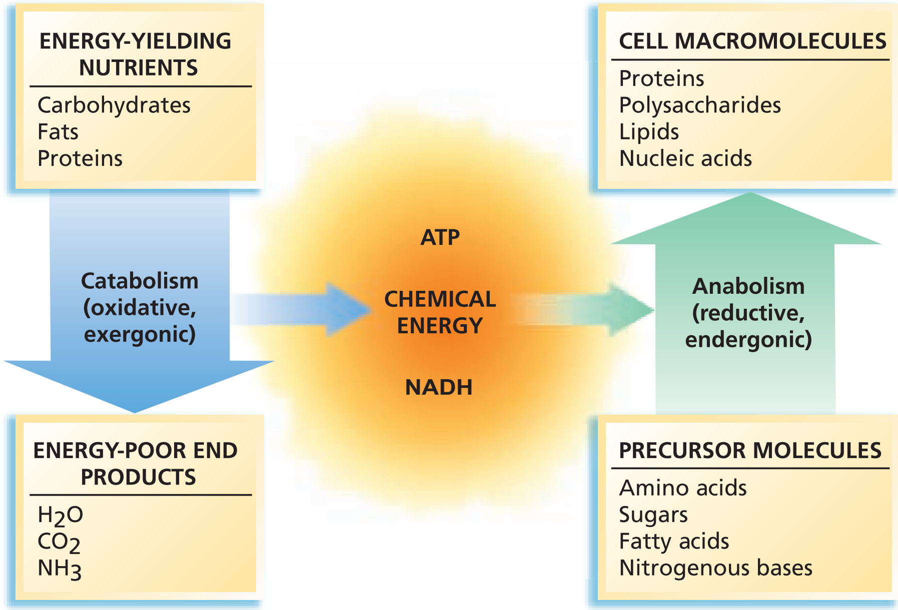
$$ATP^{4-}(aq) + H_2O(l) \longrightarrow ADP^{3-}(aq) + H_2PO_4^-(aq) \quad \Delta G = -31 \text{ kJ}$$

FIGURE 18 The hydrolysis of ATP produces ADP and releases energy.



Adenosine triphosphate (ATP) **Adenosine diphosphate (ADP)**

FIGURE 19 Catabolic pathways release free energy in the form of ATP and NADH. Anabolic pathways consume energy released by catabolic pathways.



Catabolism

The energy that your body needs to maintain its temperature and drive its biochemical reactions is provided through *catabolic* processes. **Figure 19** illustrates the relationship between the pathways of catabolism and anabolism. **Catabolism** *is the part of metabolism in which complex compounds break down into simpler ones and is accompanied by the release of energy.* First, enzymes break down the complex compounds in food—carbohydrates, fats, and proteins—into simpler molecules.

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Topic: Food as Fuel
Code: HC60591

Carbohydrate digestion begins in the mouth, where the enzyme *amylase* in saliva begins to break down polysaccharides. The food then passes through the esophagus, then the stomach, and into the small intestine. Here, additional enzymes are secreted to complete the hydrolysis of carbohydrates to form glucose and other monosaccharides.

Digestion of fats occurs only in the small intestine. Protein digestion begins in the stomach and is completed in the small intestine. During the digestion of both fats and proteins, complex molecules hydrolyze into simpler ones. Fats are broken down into fatty acids and glycerol. Proteins are broken down into amino acids.

These products are absorbed across the wall of the small intestine into the blood and are transported to cells. Once in the cells, each glucose molecule is broken down through glycolysis into two molecules of pyruvate, which enter the mitochondria and feed into a complex series of reactions called the *citric acid cycle,* or *Krebs cycle*. The citric acid cycle produces carbon dioxide and other molecules, such as NADH and ATP. This NADH and ATP then move through another set of reactions to produce more ATP and water.

The amount of energy released in catabolism depends on the amount of ATP that is made as the products of digestion are oxidized. The catabolism of 1 glucose molecule generally may produce up to a total of 36 ATP molecules. This accounts for about 40% of the energy released by the complete oxidation of glucose. Most of the energy not converted to ATP is lost by the body as energy in the form of heat. **Table 2** shows how much ATP is needed for some daily activities.

TABLE 2 Approximate "Cost" of Daily Activities

Activity	Energy required (kJ)	ATP required (mol)
Running	1120	56
Swimming	840	42
Bicycling	1400	70
Walking	560	28

Anabolism

Cells use the simple molecules that result from the breakdown of food to make larger, more complex molecules. Energy released during catabolism powers the synthesis of new molecules as well as the active transport of ions and molecules across cell membranes. *Anabolic* processes are the energy-consuming pathways by which cells produce the molecules that they need for sustaining life and for growth and repair. *The conversion of small biomolecules into larger ones is called* **anabolism.**

In an anabolic pathway, small precursor molecules are converted into complex molecules, including lipids, polysaccharides, proteins, and nucleic acids. Energy from ATP and NADH is necessary for these biosynthesis reactions to occur. **Figure 19** illustrates that catabolism and anabolism occur simultaneously and that ATP and NADH serve as chemical "links" between the two processes.

One important anabolic pathway that is common to animals, plants, fungi, and microorganisms is *gluconeogenesis*. As the name implies, glucose is synthesized in this pathway from non-carbohydrate substances, including lactate, pyruvate, glycerol, and most of the amino acids. In mammals, glucose from the blood is the fuel source for the brain and nervous system as well as for the kidney medulla, red blood cells, and embryonic tissues.

SECTION REVIEW

1. List four ways in which the body uses energy.
2. What is the total energy (in kilojoules) stored in the 36 ATP molecules that are made from the metabolism of 1 molecule of glucose in skeletal tissue?
3. The teeth break large pieces of food into smaller ones. However, this process is not considered part of digestion. Explain why.
4. How does the digestion of fat in a strip of ham differ from the digestion of starch in a piece of toast?
5. Why are diets that are severely restrictive in carbohydrate intake potentially dangerous to a person's health?

Critical Thinking

6. **RELATING CONCEPTS** When a molecule of glucose is oxidized, only about 40% of the energy is captured in ATP molecules, and the rest is lost as energy in the form of heat. What do you think would happen if this energy remained in the body as heat?

SECTION 4

Nucleic Acids

Objectives

- Describe the structure of the nucleic acids DNA and RNA.
- Explain the functions of DNA and RNA in the cell.
- Describe applications of modern gene technology.

Nucleic acids contain all of the genetic information of an organism. They are the means by which a living organism stores and conveys instructional information for all of its activities. They are also the means by which an organism can reproduce. The two nucleic acids found in organisms are *deoxyribonucleic acid (DNA)* and *ribonucleic acid (RNA)*.

Nucleic Acid Structure

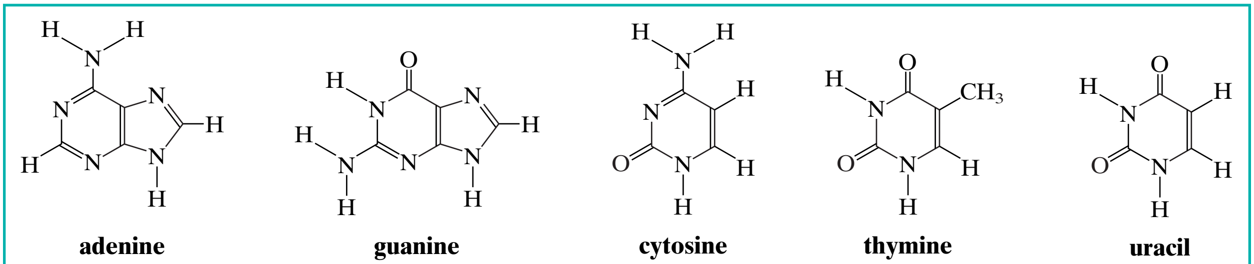
A **nucleic acid** *is an organic compound, either RNA or DNA, whose molecules carry genetic information and is made up of one or two chains of monomer units called* nucleotides. However, unlike the monomer units in polysaccharides and polypeptides, each nucleotide monomer can be further hydrolyzed into three different molecules. A nucleotide molecule is composed of a five-carbon sugar unit that is bonded to both a phosphate group and a cyclic organic base containing nitrogen.

The sugar unit in DNA is deoxyribose, and the sugar unit in RNA is ribose. The diagram below shows the sugar-phosphate arrangement in three nucleotides.

```
—phosphate—sugar—phosphate—sugar—phosphate—sugar—
             |                |                |
            base             base             base
```

The five nitrogenous bases found in nucleic acids are shown in **Figure 20.** *Adenine* (A), *guanine* (G), and *cytosine* (C) are found in both DNA and RNA. *Thymine* (T) is found only in DNA, and *uracil* (U) is found only in RNA.

FIGURE 20 There are five common nitrogenous bases. Thymine (T), cytosine (C), and uracil (U) have a single six-member ring. Adenine (A) and guanine (G) have a six-member ring connected to a five-member ring.



DNA: Deoxyribonucleic Acid

Every single instruction for all of the traits that you have inherited and all of the life processes that occur in your cells is contained in your DNA. It is no wonder then that DNA molecules are the largest molecules found in cells. Living organisms vary widely in the size and number of DNA molecules in their cells. Some bacterial cells contain only 1 DNA molecule, while human cells contain 46 relatively large DNA molecules. The DNA in each human cell is about 2 m long. This DNA is divided and packed into the cell's 46 *chromosomes.* An average cell is only about 6 μm in diameter and contains many organelles and structures. To fit in a cell, DNA must undergo extensive twisting, coiling, folding, and wrapping.

The Swedish scientist Friedrich Miescher first extracted DNA from cells in 1868, but its three-dimensional structure was not discovered until 1953. Using the X-ray data of Maurice Wilkins and Rosalind Franklin, James Watson of the United States and Francis Crick of England proposed that DNA was a double helix. In this structure, which has been confirmed by numerous methods, two strands of the sugar-phosphate backbone are wound around each other, and the nitrogenous bases point inward, as shown in **Figure 21.** The sequence of these nitrogenous bases along the phosphate-sugar backbone in DNA forms the code responsible for transferring genetic information. The three-dimensional DNA molecule is similar to a twisted ladder. The sides of the ladder are the sugar-phosphate backbone, and the rungs are base pairs of A—T (adenine-thymine) or G—C (guanine-cytosine) bases extending between the two backbones. Hydrogen bonding between these pairs of nitrogenous bases holds the rungs of the ladder together.

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Topic: DNA
Code: HC60418
Topic: DNA Replication
Code: HC60420

FIGURE 21 Hydrogen bonding between base pairs makes the three-dimensional structure of DNA stable. Base pairing occurs between adenine and thymine or guanine and cytosine, keeping the distance between the strands constant.

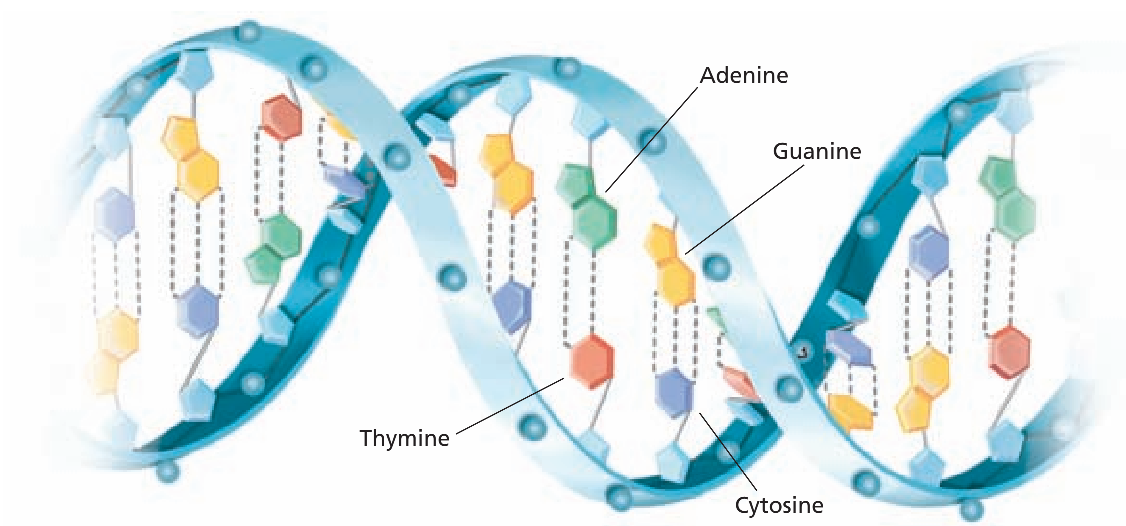


FIGURE 22 The double helix of DNA can be seen by using scanning tunneling microscopy (STM).

Nitrogenous Base Pairs

In the DNA double helix, base pairing exists only between A—T and between C—G, as you saw in **Figure 21.** The reason is that the pairing between one single-ringed base and one double-ringed base provides the correct orientation for the hydrogen bonds to form between the two sides of the DNA ladder. One thymine and one adenine form a link between the two strands of a DNA molecule that is exactly the same size as the link formed by one cytosine and one guanine.

The double-helix configuration of DNA, shown in **Figure 22,** can be seen by using a scanning tunneling microscope (STM). The discovery of the relative quantities of the nitrogenous bases A, T, G, and C present in DNA was the key to determining the three-dimensional molecular structure. Analysis of DNA from different organisms reveals that the amounts of A and T are the same and that the amounts of G and C are the same for all members of the same species. In humans, DNA is about 30% A, 30% T, 20% G, and 20% C.

The interaction between base pairs accounts for the ability of DNA to replicate, as you will see in the next section. Just as combinations of the 26 letters of the alphabet form words that tell a story in a novel, combinations of the four-letter alphabet of A, T, G, and C form the *genes* that define our heredity. Each gene is a section of DNA that contains a specific sequence of four bases (A, G, T, and C) and encodes instructions for protein synthesis. Researchers using the Human Genome Project data predict that the human body contains about 20 000 to 25 000 genes. Laboratory verification of this count will require many more years.

extension

Historical Chemistry

Go to **go.hrw.com** for full-length article on the discovery of the structure of DNA.

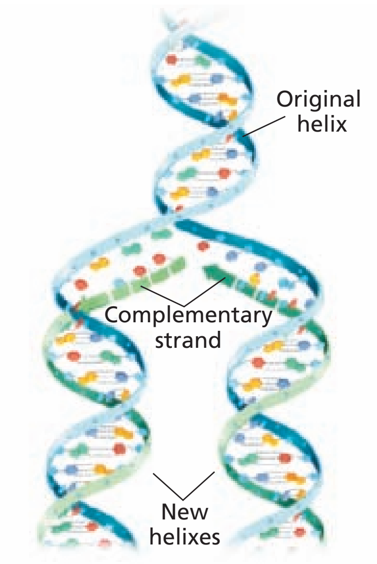
Keyword: HC6BIOX

DNA Replication

Like the two sides of a zipper, the two strands of the double helix of DNA are not identical. Instead, the two strands are complements of each other. Thus, a base on one strand is paired through hydrogen bonding to its complementary base on the other strand. For example, if one strand sequence is AGCTC, the complementary strand sequence will be TCGAG.

Each time a cell divides, an exact copy of the DNA of the parent cell is reproduced for the daughter cells. *The process by which an identical copy of the original DNA is formed is called* **DNA replication.** As replication begins, a portion of the two strands of the original DNA unzips, as shown in **Figure 23.** Each strand can then act as a template for the synthesis of a new, complementary strand. The result is two new DNA molecules, which have the same base pair sequence as the original double helix.

FIGURE 23 DNA replicates when its double helix unwinds and becomes single stranded. The single strands are used as a template for the formation of new complementary strands.



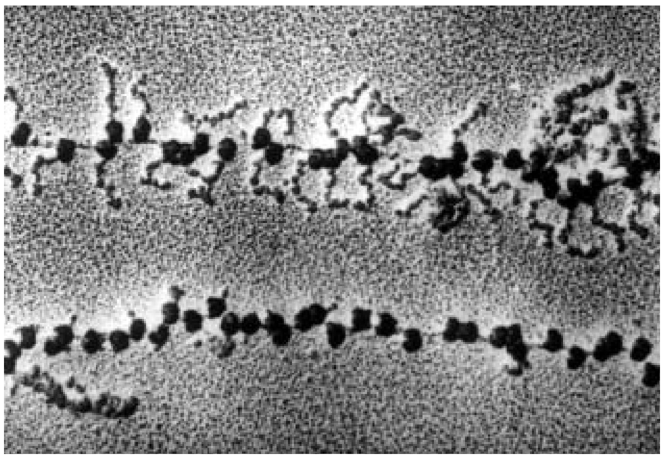
RNA: Ribonucleic Acid

RNA molecules are responsible for the synthesis of proteins, which in turn control the operation and function of the cell. RNA differs from DNA in four basic ways: (1) the sugar unit in the backbone of RNA is ribose rather than deoxyribose, (2) RNA contains the base uracil, U, instead of thymine, which occurs in DNA, (3) RNA is a single-stranded molecule rather than a double-stranded helix like DNA, and (4) RNA molecules typically consist of 75 to a few thousand nucleotide units rather than the millions that exist in DNA. Even though RNA is much smaller than DNA, RNA is still large enough to twist, coil, bend, and fold back onto itself. In fact, it is not uncommon for up to 50% of an RNA molecule to have a double-helix structure. The reason is that the base sequences along the helical regions of the RNA strand are complementary, which makes hydrogen bonding between bases possible.

Synthesis of RNA

RNA is synthesized in the nucleus of eukaryotic cells, where DNA and protein molecules actually help synthesize specific RNA molecules. RNA can also be seen by STM, as shown in **Figure 24.** As RNA is synthesized, some information contained in the DNA is transferred to the RNA molecules. Like the genetic information of DNA, the genetic information of RNA is carried in its nucleotide sequence. One type of RNA molecule is called *messenger RNA* (mRNA) because it carries the instructions for making proteins out into the cytosol, where proteins are produced on *ribosomes. A* **ribosome** *is a cell organelle that is composed of RNA and protein. Ribosomes are the main site of protein production in cells.* The DNA template is also used to make two other types of RNA molecules: *ribosomal RNA* (rRNA) and *transfer RNA* (tRNA). Both of these types of RNA also leave the nucleus and come together in the ribosome where they help synthesize proteins. Ribosomal RNA becomes part of the structure of the ribosome, and tRNA is used to transfer amino acids into the ribosome. Only mRNA carries the coded genetic information that is translated into proteins.

FIGURE 24 Scanning tunneling micrograph of RNA strands being transcribed in a cell.



DNA supplies all of the information necessary for RNA to be used to make the proteins needed by the body. The portion of DNA that holds the specific genetic code for a single, specific mRNA molecule is a gene. As you learned previously, each gene is a section of the DNA chain that contains a specific sequence of the bases A, G, T, and C. A gene has the information necessary in this sequence to direct RNA to produce several proteins that have specific functions.

CAREERS in Chemistry

Forensic Chemist

A forensic chemist applies scientific methodology to physical evidence. Forensic chemists focus on analyzing evidence that has been collected at a crime scene. They then report any conclusions that they can draw from the analysis. Understanding evidence requires knowledge of biology, materials science, and genetics in addition to chemistry. Because forensic chemists are often asked to testify in court, they need to be comfortable speaking in public and able to give clear and concise explanations to those who do not have a background in science.

RNA and Protein Synthesis

RNA is made from DNA in a process that is similar to how DNA replicates itself. At a gene, a portion of DNA unwinds and RNA is assembled using the same complementary base pairs as DNA except that uracil replaces the thymine. When a signal to stop is received, the RNA is released. As in DNA replication, the RNA sequence that forms has the complementary base pairs of the DNA gene. The DNA sequence below would form the complementary RNA sequence shown.

DNA strand: C C C C A C C C T A C G G T G
RNA strand: G G G G U G G G A U G C C A C

A sequence of three bases in mRNA codes for a specific amino acid. Thus, the sequence CAG codes for glutamic acid, and GUC codes for valine. There are 64 (4^3) unique combinations of three-base sequences made from four bases. Because only 20 amino acids require codes, some of the amino acids have more than one code. For example, leucine is coded by six three-base sequences: UUA, UUG, CUU, CUC, CUA, and CUG. The genetic code is universal, meaning that the same three-base sequence always codes for the same amino acid regardless of whether the organism is a bacterium or a human. The "stop" signal in the gene is also a three-base code: UAG, UAA, or UGA.

Technology and Genetic Engineering

The discovery of DNA's function in life has provided new options for the production of food, medical diagnosis and treatments, and increased understanding of genetic disorders. Scientists in the field of genetic engineering study how manipulation of an organism's genetic material can modify the proteins that are produced and the changes that result in the organism. Although the selective breeding of plants and animals has been practiced for hundreds of years, today genetic engineering refers to recombinant DNA technology that is used for cloning and the creation of new forms of life. Because the technique is so powerful, it is controversial and must be used responsibly.

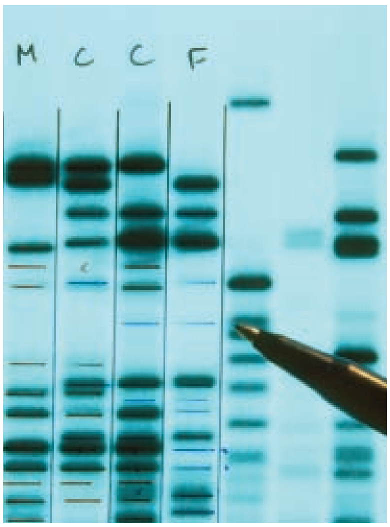


FIGURE 25 A DNA autoradiograph shows the pattern of DNA fragments of an organism after they have been separated from each other by electrophoresis.

DNA Fingerprinting

One of the most visible uses of molecular technology is DNA fingerprinting. DNA is unique to an individual except for identical twins. This technology is used in criminal investigations and victim identification. Often there are only very small samples available, such as a single drop of blood or one strand of hair. The technique of the *polymerase chain reaction* (PCR) may be used to copy a DNA sample to supply sufficient DNA for identification. The processes of electrophoresis and autoradiography may be used to compare DNA from a sample with the DNA of a known individual to confirm identification, as **Figure 25** shows. DNA technology can also be used to test paternity or to trace heredity.

Cloning

One meaning of the word **cloning** *is the process of making an exact copy of an organism*. One example of a natural occurrence of cloning, the formation of identical twins, is the result of a chance splitting of the embryonic cells very early in the growth of a zygote. Artificial cloning, using stem cells from animals or meristem cells from plants, can produce identical replicas of the parent cells or, under specialized conditions, a complete organism that is identical to the original organism. The orchid shown in **Figure 26** is a clone of its parent plant. Cloning of plants may hold promise for increasing the yields of crops. Recently, scientists at Pennsylvania State University cloned cocoa plants from cocoa flowers. When cocoa trees are planted from seed, as many as 50% do not mature with the desired characteristics. By planting young trees that are clones of plants with desirable characteristics, farmers may be able to increase their cocoa production.

The first animal to be cloned, a sheep named Dolly, was born in 1996 in Scotland. Dolly was euthanized in 2003 because of lung disease. She had also been diagnosed with arthritis. Both diseases are normally found in sheep older than Dolly was.

FIGURE 26 Growers can produce many orchids by artificial cloning of the meristem tissue of a single orchid plant.

Recombinant DNA Technology

Recombinant DNA technology has been used to insert DNA from one organism into another. One technique involves splicing a gene from one organism's DNA into a molecule of DNA from another organism. *Escherichia coli,* a bacterium found in animal intestinal tracts, are often used by biologists as cellular factories for the production or manufacture of DNA fragments cloned from other organisms. In some instances, *E. coli* can even be used to produce protein from DNA cloned from other organisms.

One of the first applications of genetic engineering was the synthesis of human insulin. Previously, most diabetics had to use either pig or cow insulin. But insulin from animals is not exactly the same as human insulin. Today, most insulin used is produced in bacteria and harvested. Human growth hormone is also commercially produced by using recombinant DNA technology.

SECTION REVIEW

1. What sugar is present in DNA? What sugar is present in RNA?
2. Explain why the two strands of the DNA double helix are said to be complementary instead of identical.
3. Describe how DNA uses the genetic code to control the synthesis of proteins.
4. Why is a very small trace of blood enough for DNA fingerprinting?

Critical Thinking

5. **INTERPRET AND APPLY** Is it possible to specify 20 amino acids by using only two base pairs instead of three for coding?
6. **DRAWING CONCLUSIONS** Why is the arrangement of base pairs that is found in DNA ideal for holding the double helix of DNA together?

CHAPTER HIGHLIGHTS

Carbohydrates and Lipids

Vocabulary
carbohydrate
monosaccharide
disaccharide
condensation reaction
hydrolysis
polysaccharide
lipid
fatty acid
saponification

- Carbohydrates are nutrients that are produced by plants and are made up of carbon, oxygen, and hydrogen.
- Monosaccharides are the simplest carbohydrates. Carbohydrates made of two monosaccharides are called *disaccharides,* and carbohydrates made of more than two monosaccharides are called *polysaccharides.*
- Carbohydrates undergo condensation reactions and hydrolysis reactions.
- Lipids are a varied group of biochemical molecules that have a high percentage of C and H atoms.

Amino Acids and Proteins

Vocabulary
amino acid
protein
enzyme
denaturation

- Amino acid molecules are the basic building blocks of proteins.
- Proteins are biopolymers, each of which has a unique sequence of the acid monomer molecules.
- The specific function of a protein is related to the shape of the protein.
- Side-chain interactions between amino acids result in secondary, tertiary, and quaternary protein structures.

Metabolism

Vocabulary
metabolism
autotroph
adenosine triphosphate (ATP)
heterotroph
adenosine diphosphate (ADP)
catabolism
anabolism

- ATP is a high-energy storage compound that the body uses to store and provide energy for life.
- The metabolic pathways involve both the conversion of ATP to ADP and the conversion of ADP to ATP.
- Metabolic pathways are classified as two types: catabolism and anabolism.
- Catabolism includes reactions in which large molecules are changed into simpler molecules. These reactions release energy.
- Anabolic processes are energy-consuming pathways by which cells produce the molecules needed for growth and repair.

Nucleic Acids

Vocabulary
nucleic acid
DNA replication
cloning

- Deoxyribonucleic acid (DNA) and ribonucleic acid (RNA) are nucleic acids, the compounds by which living organisms can reproduce themselves.
- Nucleic acids are polymers of monomer units called *nucleotides*.
- The two strands of the double helix of DNA are complementary to each other, not identical. These strands are held together by hydrogen bonding of the base pairs.
- RNA is used as a template to produce proteins in the cell.

CHAPTER REVIEW

Carbohydrates and Lipids

SECTION 1 REVIEW

1. Describe the general chemical formula of carbohydrates.
2. Name two examples from each of the following classes of carbohydrates: monosaccharides, disaccharides, and polysaccharides.
3. What different roles do the polysaccharides starch and cellulose play in plant systems?
4. What word is used to describe fatty acids that contain at least one double bond?
5. Why are some triglycerides liquid, while others are solid?
6. What reagents are used to make soaps?

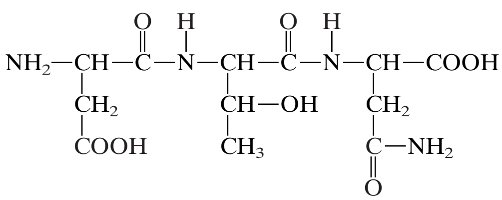
PRACTICE PROBLEMS

7. Draw the structural formula for glucose.
8. Using structural formulas, write the equation showing the formation of maltose, which is the disaccharide made of two glucose units.
9. Write the equation representing the formation of a soap molecule from stearic acid, $C_{17}H_{35}COOH$, and sodium hydroxide.

Amino Acids and Proteins

SECTION 2 REVIEW

10. Describe the structure of an amino acid. Then, explain how amino acids become linked together to form a protein.
11. Circle and identify the carboxylic acid groups and the amino groups in the following molecule:



12. Can two types of enzymes contain the same number and kinds of amino acids? Explain.
13. What happens when a protein is denatured?
14. Explain the cause of the genetic disease sickle cell anemia.
15. Why is the water solubility of fibrous proteins so different from that of globular proteins?

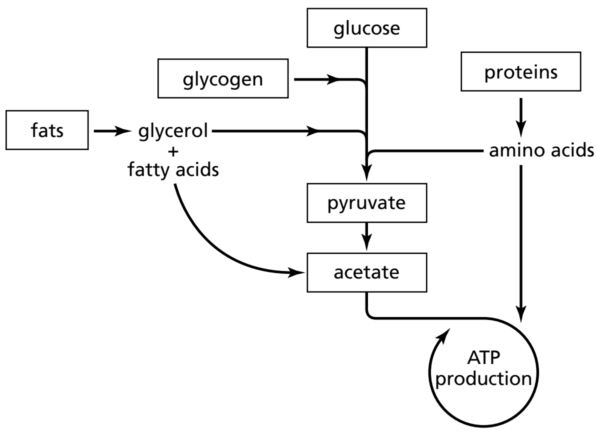
PRACTICE PROBLEMS

16. Draw the structures of two dipeptides made up of glycine and valine.
17. How many different tripeptides can be formed from two molecules of glycine and one molecule of cysteine? Write all of the structures by using the three-letter codes Gly and Cys.

Metabolism

SECTION 3 REVIEW

18. What chemical gains the metabolic energy that is released as glucose is broken down in the body?
19. What does *ATP* stand for? What is the role of ATP in living things?
20. Describe the steps that occur in the digestion of fats.
21. Review the following diagram of catabolism.



According to the diagram, what could happen in the cell when glucose and glycogen reserves are nearly gone?

PRACTICE PROBLEMS

22. Draw the structure of ATP. Circle the bond that breaks when ADP forms.

Nucleic Acids

SECTION 4 REVIEW

23. What are the three components of a nucleotide?

24. How are the two polynucleotide chains of DNA held together?

25. Describe in general terms the process of DNA replication.

26. What are the main differences between DNA and RNA?

27. Describe the similarities and differences between the three kinds of RNA molecules.

28. What is a ribosome? What is the function of a ribosome in a cell?

PRACTICE PROBLEMS

29. The following sequence of bases might be found on the gene that codes for oxytocin, the human pituitary hormone:
TACACAATGTAAGTTTTGACGGGGGAC-
CCTATC

a. What is the sequence of bases on the complementary strand of DNA?

b. What is the sequence of bases that would occur on a strand of mRNA that was transcribed from the oxytocin DNA sequence?

MIXED REVIEW

30. Name the four main elements that make up compounds found in living organisms.

31. In each of the following groups, one of the items does not belong in the group. Identify the odd item in the group and explain why it is different. Explain how the other items are related.

a. glycogen, starch, fructose, and cellulose

b. amino acids, dipeptides, polypeptides, and proteins

c. fats, oils, and fatty acids

d. cytosine, adenine, and guanine

32. What is the human body's *storage* form of each of the following?

a. glucose

b. lipids

c. protein

33. Is each of the following statements about proteins and triglycerides true or false?

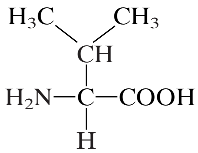
a. Both contain the amide functional group.

b. Both are a part of a major class of biochemical molecules.

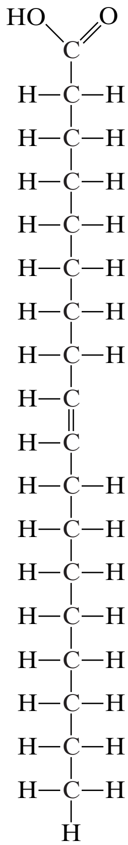
c. Both hydrolyze in order to enter the metabolic pathway in humans.

34. Circle the hydrophobic part in each of the figures shown below.

a.



b.



35. Both celery and potato chips are composed of molecules that are polymers of glucose. Explain why celery is a good snack for people on a diet while potato chips are not.
36. Carbohydrates, fats, and proteins can provide energy for an organism.
    a. Which class of substances most rapidly provides energy?
    b. Which class can be used as a building material in the human body?
    c. Which is the most efficient as an energy storage system?
37. Describe the basic structure of the cell membrane. What is the cell membrane's main function?

CRITICAL THINKING

38. **Interpreting Concepts** A diet that consists primarily of corn can result in a protein-deficiency disease called *kwashiorkor*. What does this information indicate about the protein content of corn?
39. **Inferring Relationships** Explain how a similar reaction forms three kinds of biological polymers: polysaccharides, polypeptides, and nucleic acids.
40. **Evaluating Ideas** Some diets recommend severely restricting or eliminating the intake of carbohydrates. Why is it not a good idea to eliminate all carbohydrates from the diet?
41. **Using Analogies** Explain why the model of enzyme action is called the "lock and key" model.

RESEARCH & WRITING

42. Conduct library research about how Olestra® decreases fat and caloric content of potato chips. What are the advantages and disadvantages of Olestra in food products?
43. Write a summary discussing what you have learned about the four major classes of organic compounds found in living things—carbohydrates, lipids, proteins, and nucleic acids. Include a description of how these organic molecules are used by the body.

ALTERNATIVE ASSESSMENT

44. Amylase, the enzyme present in the mouth, catalyzes the digestion of starch. The pH of the mouth is almost neutral.
    a. Do you think that amylase is active in the stomach after you swallow the food? Why or why not?
    b. Design an experiment you could perform to test your answer to item a. Note: A common test for the presence of starch is the addition of tincture of iodine, which will produce a blue color if starch is present.

Math Tutor INTERPRETATION OF THE GENETIC CODE

In protein synthesis, the DNA sequence of bases is transcribed onto messenger RNA (mRNA). The mRNA base sequence is the complement of the DNA sequence except that uracil takes the place of thymine as the complement of adenine.

Problem-Solving TIPS

- Find the first base of the mRNA triplet along the left side of the table.
- Follow that row to the right until you are beneath the second base of the triplet.
- Move up or down in the square that corresponds to the second base until you are even, on the right side of the chart, with the third base of the triplet.

The Genetic Code (mRNA)

First base	Second base: U	Second base: C	Second base: A	Second base: G	Third base
U	UUU Phenylalanine	UCU Serine	UAU Tyrosine	UGU Cysteine	U
U	UUC Phenylalanine	UCC Serine	UAC Tyrosine	UGC Cysteine	C
U	UUA Leucine	UCA Serine	UAA Stop	UGA—Stop	A
U	UUG Leucine	UCG Serine	UAG Stop	UGG—Tryptophan	G
C	CUU Leucine	CCU Proline	CAU Histidine	CGU Arginine	U
C	CUC Leucine	CCC Proline	CAC Histidine	CGC Arginine	C
C	CUA Leucine	CCA Proline	CAA Glutamine	CGA Arginine	A
C	CUG Leucine	CCG Proline	CAG Glutamine	CGG Arginine	G
A	AUU Isoleucine	ACU Threonine	AAU Asparagine	AGU Serine	U
A	AUC Isoleucine	ACC Threonine	AAC Asparagine	AGC Serine	C
A	AUA Isoleucine	ACA Threonine	AAA Lysine	AGA Arginine	A
A	AUG—Start	ACG Threonine	AAG Lysine	AGG Arginine	G
G	GUU Valine	GCU Alanine	GAU Aspartic acid	GGU Glycine	U
G	GUC Valine	GCC Alanine	GAC Aspartic acid	GGC Glycine	C
G	GUA Valine	GCA Alanine	GAA Glutamic acid	GGA Glycine	A
G	GUG Valine	GCG Alanine	GAG Glutamic acid	GGG Glycine	G

SAMPLE

What sequence of amino acids will be incorporated into protein as a result of the mRNA sequence UUACCCGAGAAGUCC?

Divide the sequence into groups of three to clearly see the separate codons.

UUACCCGAGAAGUCC = UUA ∣ CCC ∣ GAG ∣ AAG ∣ UCC

Now, use the table to determine the match between codons and amino acids.

UUA	CCC	GAG	AAG	UCC
leucine	proline	glutamic acid	lysine	serine

PRACTICE PROBLEMS

1. What amino acid sequence will be added to a protein as a result of the mRNA sequence UUACACGACUAUAAUUGG?

2. What amino acid sequence will be added to a protein as a result of the mRNA sequence CUAACCGGGUGAGCUUCU?

Standardized Test Prep

Answer the following items on a separate piece of paper.

MULTIPLE CHOICE

1. Which of the following statements about enzymes is true?
   A. Enzymes can be biological catalysts.
   B. Enzymes increase the speed of a chemical reaction.
   C. Enzymes are highly specific.
   D. All of the above

2. Which of the following statements about denaturing is true?
   A. Denaturing occurs when a protein unfolds.
   B. Denaturing occurs when a carbohydrate is heated.
   C. Denaturing does not affect the tertiary structure of an enzyme.
   D. Denaturing increases the rate of a chemical reaction.

3. The process in which molecules in a cell break down to produce smaller molecules and energy is called
   A. glycogenesis.
   B. biosynthesis.
   C. catabolism.
   D. metabolism.

4. Which of the following is partially digested by saliva in the mouth?
   A. glucose
   B. starch
   C. fat
   D. protein

5. In the human body, the storage form of
   A. glucose is maltose.
   B. triglycerides is protein.
   C. carbohydrates is glycogen.
   D. nucleic acids is amino acids.

6. The purpose of insulin is to
   A. regulate glucose levels in the body.
   B. catalyze the oxidation of fatty acids.
   C. stimulate RNA production.
   D. initiate DNA replication.

7. Which of the following statements about fats is true?
   A. Fats serve as a reserve supply of energy.
   B. Fats are stored in the adipose tissue.
   C. Fats act as insulators.
   D. All of the above

8. When carbohydrates are unavailable or unable to provide the energy needs of the body,
   A. glucose is converted to glycogen.
   B. proteins or fats are used for energy.
   C. amino acids form proteins.
   D. All of the above

9. Which of the following statements is true?
   A. RNA contains the base uracil rather than thymine, which occurs in DNA.
   B. Both RNA and DNA are double-stranded helixes.
   C. The ribose sugar unit is in the backbone of DNA.
   D. None of the above

SHORT ANSWER

10. Draw a simple dipeptide, and label the functional groups and peptide linkage.

11. Describe the shape of the DNA molecule, and discuss how the DNA molecule is held in this shape.

EXTENDED RESPONSE

12. The body has numerous energy reserves. What are they, and where are they found in the body? Which of these reserves provides the greatest source of quick energy?

13. Explain how it is possible to denature a protein without breaking the polypeptide chain.

Test TIP If a question or an answer choice contains an unfamiliar term, try to break the word into parts to determine its meaning.

SEE PRE-LAB
EXTRACTION AND FILTRATION

Casein Glue

OBJECTIVES

- ***Recognize*** **the structure of a protein.**
- ***Predict*** **and** ***observe*** **the result of acidifying milk.**
- ***Prepare*** **and** ***test*** **a casein glue.**
- ***Deduce*** **the charge distribution in proteins as determined by pH.**

MATERIALS

- **100 mL graduated cylinder**
- **250 mL beaker**
- **250 mL Erlenmeyer flask**
- **funnel**
- **glass stirring rod**
- **hot plate**
- **medicine dropper**
- **baking soda, $NaHCO_3$**
- **nonfat milk**
- **paper**
- **paper towel**
- **thermometer**
- **white vinegar**
- **wooden splints, 2**

BACKGROUND

Cow's milk contains averages of 4.4% fat, 3.8% protein, and 4.9% lactose. At the normal pH of milk, 6.3 to 6.6, the protein remains dispersed because it has a net negative charge due to the dissociation of the carboxylic acid group, as shown in **Figure A** below. As the pH is lowered by the addition of an acid, the protein acquires a net charge of zero, as shown in **Figure B.** After the protein loses its negative charge, it can no longer remain in solution, and it coagulates into an insoluble mass. The precipitated protein is known as casein and has a molecular mass between 75 000 and 375 000 amu. The pH at which the net charge on a protein becomes zero is called the isoelectric pH. For casein, the isoelectric pH is 4.6.

$$H_2N-\boxed{\text{protein}}-COO^-$$

FIGURE A

$$^{+}H_3N-\boxed{\text{protein}}-COO^-$$

FIGURE B

In this experiment, you will coagulate the protein in milk by adding acetic acid. The casein can then be separated from the remaining solution by filtration. This process is known as separating the curds from the whey. The excess acid in the curds can be neutralized by the addition of sodium hydrogen carbonate, $NaHCO_3$. The product of this reaction is casein glue. Do not eat or drink any materials or products of this lab.

SAFETY

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

PREPARATION

1. Prepare your notebook for recording observations at each step of the procedure.
2. Predict the characteristics of the product that will be formed when the acetic acid is added to the milk. Record your predictions in your notebook.

PROCEDURE

1. Pour 125 mL of nonfat milk into a 250 mL beaker. Add 20 mL of 4% acetic acid (white vinegar).
2. Place the mixture on a hot plate and heat it to 60°C. Record your observations in your lab notebook, and compare them with the predictions you made in Preparation step 2.
3. Filter the mixture through a folded piece of paper towel into an Erlenmeyer flask, as shown in **Figure C.**
4. Discard the filtrate, which contains the whey. Scrape the curds from the paper towel back into the 250 mL beaker.
5. Add 1.2 g of $NaHCO_3$ to the beaker and stir. Slowly add drops of water, stirring intermittently, until the consistency of white glue is obtained.
6. Use your glue to fasten together two pieces of paper. Also fasten together two wooden splints. Allow the splints to dry overnight, and then test the joint for strength.

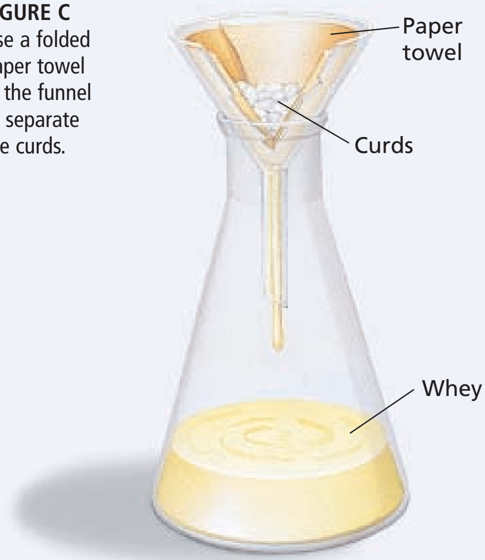
CLEANUP AND DISPOSAL

7. Clean all apparatus and your lab station. Return equipment to its proper place. Dispose of chemicals and solutions in the containers designated by your teacher. Do not pour any chemicals down the drain or in the trash unless your teacher directs you to do so. Wash your hands thoroughly before you leave the lab and after all work is finished.

ANALYSIS AND INTERPRETATION

1. **Organizing Ideas:** Write the net ionic equation for the reaction between the excess acetic acid and the sodium hydrogen carbonate. Include the physical states of the reactants and products.

FIGURE C
Use a folded paper towel in the funnel to separate the curds.



2. **Evaluating Methods:** In this experiment, what happened to the lactose and fat portions of the milk?

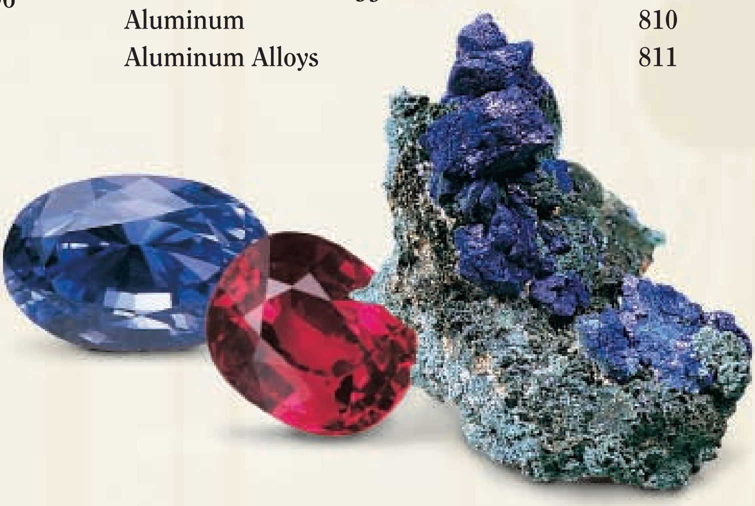
CONCLUSIONS

1. **Inferring Conclusions: Figure A** shows that the net charge on a protein is negative at pH values higher than its isoelectric pH because the carboxyl group is ionized. **Figure B** shows that at the isoelectric pH, the net charge is zero. Predict the net charge on a protein at pH values lower than the isoelectric point, and draw a diagram to represent the protein.

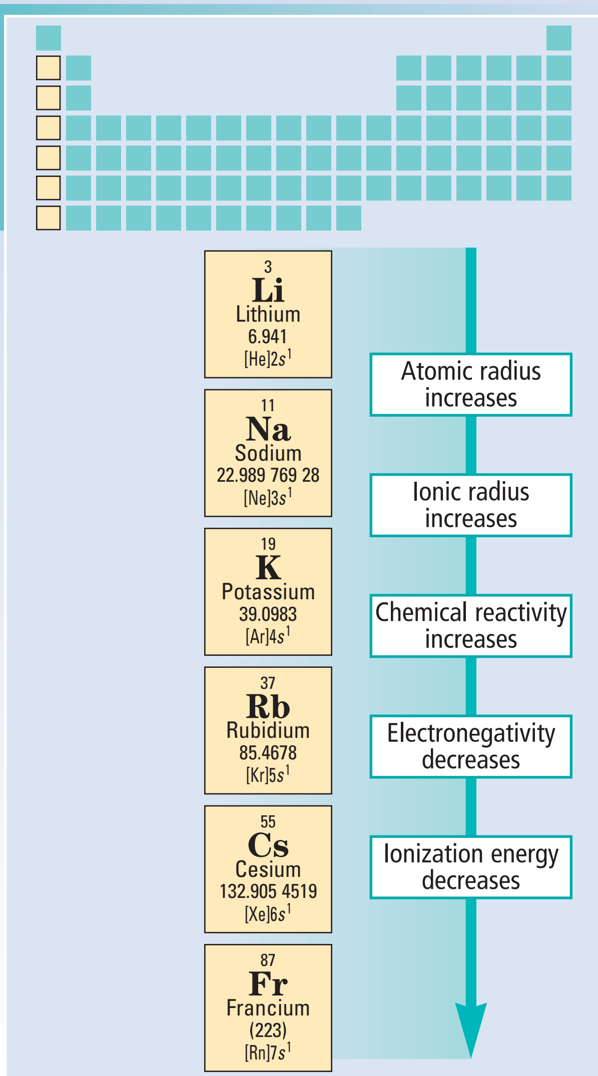
EXTENSIONS

1. **Relating Ideas: Figure B** represents a protein as a dipolar ion, or zwitterion. The charges in a zwitterion suggest that the carboxyl group donates a hydrogen ion to the amine group. Is there any other way to represent the protein in **Figure B** so that it still has a net charge of zero?
2. **Designing Experiments:** Design a strength-testing device for the glue joint between the two wooden splints. If your teacher approves your design, create the device and use it to test the strength of the glue.

Elements Handbook



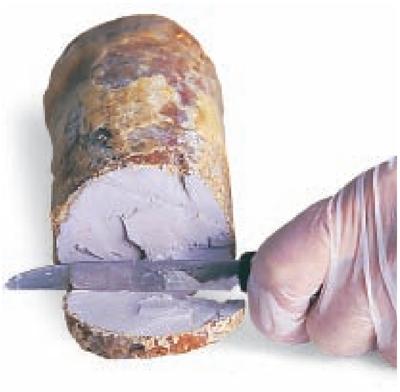
GROUP 1
ALKALI METALS



CHARACTERISTICS

- do not occur in nature as free elements
- are reactive metals and are obtained by reducing the 1+ ions in their natural compounds
- are stored under kerosene or other hydrocarbon solvent because they react with water vapor or oxygen in air
- consist of atoms with one electron in the outermost energy level
- form colorless ions in solution, each of which has a 1+ charge
- form ionic compounds
- form water-soluble bases
- are strong reducing agents
- consist of atoms that have low ionization energies
- are good heat and electrical conductors
- are ductile, malleable, and soft enough to be cut with a knife
- have a silvery luster, low density, and low melting point

Lithium was discovered in 1817. It is found in most igneous rocks and is used in batteries as an anode because it has a very low reduction potential. Lithium is soft and is stored in oil or kerosene to prevent it from reacting with the air.



Sodium derives its name from the word soda. It was first isolated in 1807 from the electrolysis of caustic soda, NaOH. Sodium is soft enough to be cut with a knife. It is shiny until it reacts with oxygen, which causes the surface to lose its luster.

Potassium was first isolated in 1807 from the electrolysis of caustic potash, KOH.

COMMON REACTIONS

With Water to Form Bases and Hydrogen Gas

Example: $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$

Li, K, Rb, and Cs also follow this pattern.

With Acids to Form Salts and Hydrogen Gas

Example: $2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$

Li, K, Rb, and Cs also follow this pattern.

With Halogens to Form Salts

Example: $2Na(s) + F_2(g) \longrightarrow 2NaF(s)$

Li, K, Rb, and Cs also follow this pattern in reacting with F_2, Cl_2, Br_2, and I_2.

With Oxygen to Form Oxides, Peroxides, or Superoxides

Lithium forms an oxide.

$4Li(s) + O_2(g) \longrightarrow 2Li_2O(s)$

Sodium also forms a peroxide.

$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$

Alkali metals with higher molecular masses can also form superoxides.

$K(s) + O_2(g) \longrightarrow KO_2(s)$

Rb and Cs also follow this pattern.

Alkali-Metal Oxides with Water to Form Bases

Oxides of Na, K, Rb, and Cs can be prepared indirectly. These basic anhydrides form hydroxides in water.

Example: $K_2O(s) + H_2O(l) \longrightarrow 2KOH(aq)$

Li, Na, Rb, and Cs also follow this pattern.

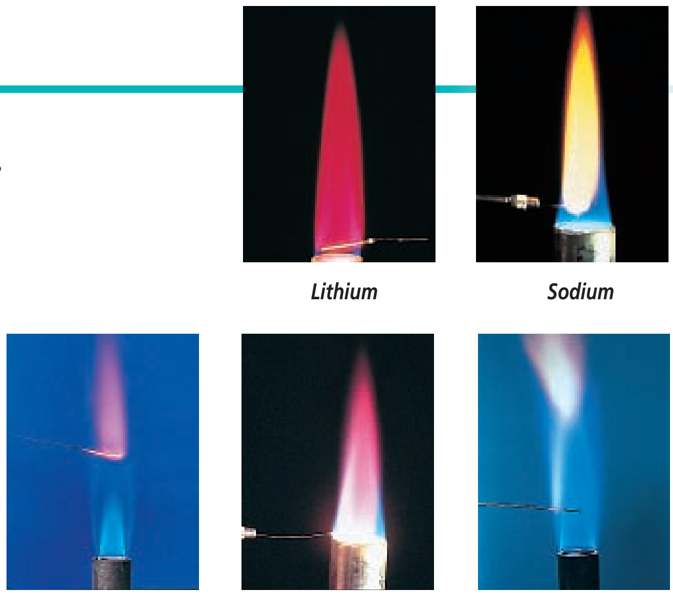
A small piece of potassium dropped into water will react explosively, releasing H_2 to form a strongly basic hydroxide solution. The energy of the reaction ignites the hydrogen gas that is produced.

Sodium reacts vigorously with chlorine to produce NaCl. Most salts of Group 1 metals are white crystalline compounds.

ANALYTICAL TEST

Alkali metals are easily detected by flame tests because each metal imparts a characteristic color to a flame.

When sodium and potassium are both present in a sample, the yellow color of the sodium masks the violet color of the potassium. The violet color can be seen only when the combined sodium-potassium flame is viewed through a cobalt-blue glass. The glass blocks the yellow flame of sodium and makes it possible to see the violet flame of potassium.



Lithium *Sodium*

Potassium *Rubidium* *Cesium*

PROPERTIES OF THE GROUP 1 ELEMENTS

	Li	Na	K	Rb	Cs	Fr
Melting point (°C)	180.5	97.8	63.25	38.89	28.5	27
Boiling point (°C)	1342	882.9	760	691	668	677
Density (g/cm^3)	0.534	0.971	0.862	1.53	1.87	—
Ionization energy (kJ/mol)	520	496	419	403	376	—
Atomic radius (pm)	152	186	227	248	265	270
Ionic radius (pm)	76	102	138	152	167	180
Common oxidation number in compounds	+1	+1	+1	+1	+1	—
Crystal structure	bcc*	bcc	bcc	bcc	bcc	—
Hardness (Mohs' scale)	0.6	0.4	0.5	0.3	0.2	—

*body-centered cubic

APPLICATION *Technology*

Sodium Vapor Lighting

The flame test for sodium shows two bright lines at 589.0 and 589.6 nm, which is the yellow range of the emission spectrum. Sodium can be vaporized at high temperatures in a sealed tube and made to give off light using two electrodes connected to a power source. Sodium vapor lighting is often used along highways and in parking lots because it provides good illumination while using less energy than other types of lighting.

Sodium vapor lighting comes in both low-pressure and high-pressure bulbs. Low-pressure lamps reach an internal temperature of 270°C to vaporize the sodium under a pressure of about 1 Pa. High-pressure lamps contain mercury and xenon in addition to sodium. These substances reach an internal temperature of 1100°C under a pressure of about 100 000 Pa. The high-pressure lamp provides a higher light intensity. The design of both types of lamps must take into account the high reactivity of sodium, which increases at high temperatures. Because ordinary glass will react with sodium at 250°C, a special sodium-resistant glass is used for low-pressure lamps. High-pressure lamps use an aluminum oxide material for the column containing the sodium, mercury, and xenon. Both types of lamps contain tungsten electrodes.

The light intensity per watt for sodium vapor lamps far exceeds that of fluorescent lamps, high-pressure mercury vapor lamps, tungsten halogen lamps, and incandescent bulbs.

APPLICATION *Health*

Electrolyte Balance in the Body

The elements of Group 1 are important to a person's diet and body maintenance because they form ionic compounds. These compounds are present in the body as solutions of the ions. All ions carry an electric charge, so they are electrolyte solutes. Two of the most important electrolyte solutes found in the body are K^+ and Na^+ ions. Both ions facilitate the transmission of nerve impulses and control the amount of water retained by cells.

During situations where the body is losing water rapidly through intense sweating or diarrhea for a prolonged period (more than 5 hours), a sports drink can hydrate the body and restore electrolyte balance.

TABLE 1A Sodium-Potassium Concentration in Body Fluids

Cation	Inside cells (mmol/L)	Outside cells or in plasma (mmol/L)
Na^+	12	145
K^+	140	4

The sodium and potassium ion concentrations in body fluids are shown in Table 1A. Sodium ions are found primarily in the fluid outside cells, while potassium ions are largely found in the fluid inside cells. Anions are present in the fluids to balance the electrical charge of the Na^+ and K^+ cations.

Abnormal electrolyte concentrations in blood serum can indicate the presence of disease. The ion concentrations that vary as a result of disease are Na^+, K^+, Cl^-, and HCO_3^-. Sodium ion concentration is a good indicator of the water balance between blood and tissue cells. Unusual potassium ion levels can indicate kidney or gastrointestinal problems. Chloride ion is the anion that balances the positive charge of the sodium ion in the fluid outside the cells. It also diffuses into a cell to maintain normal electrolyte balance when hydrogen carbonate ions diffuse out of the cell into the blood. Table 1B shows medical conditions associated with electrolyte imbalances.

TABLE 1B Electrolyte Imbalances

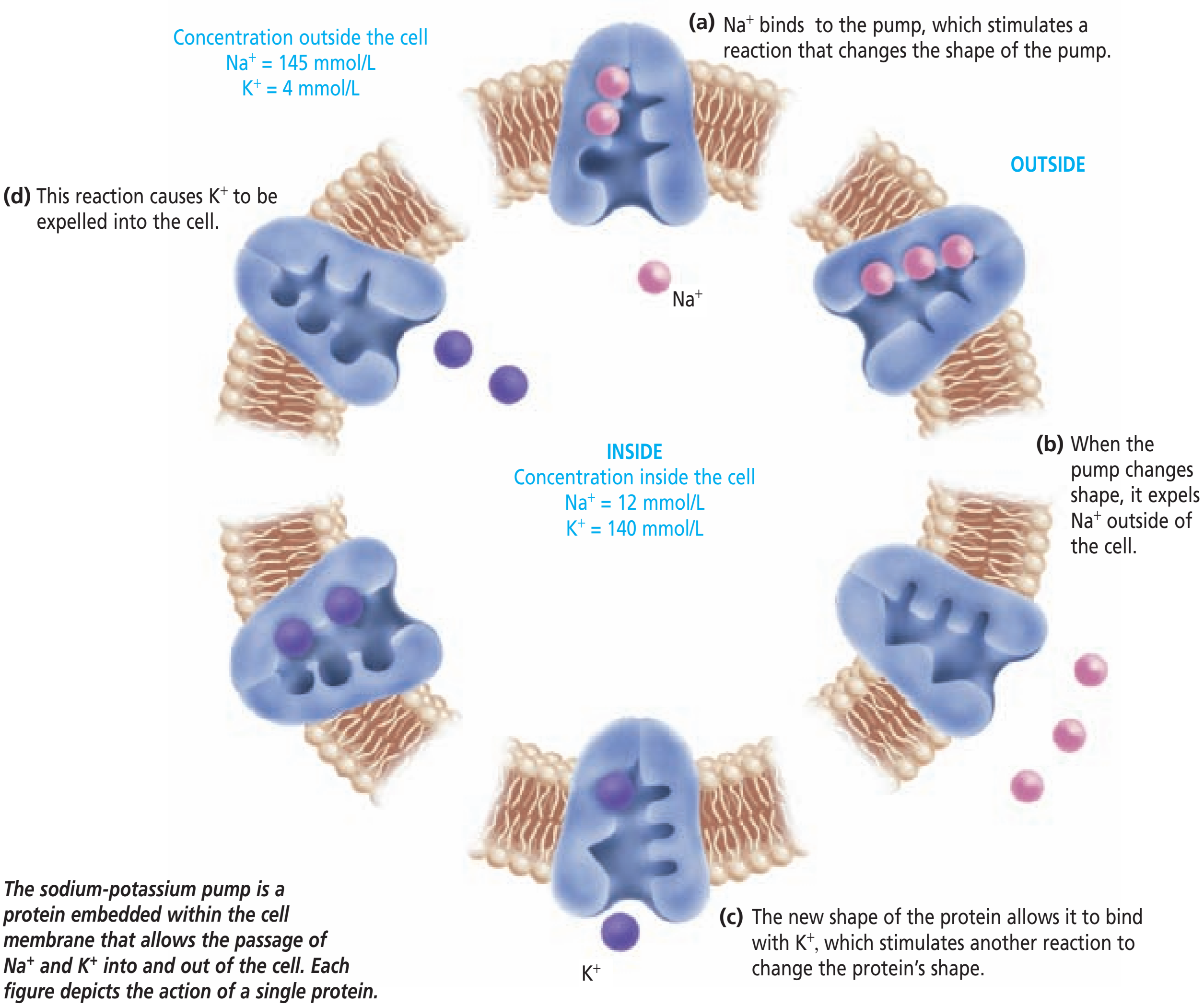
Electrolyte	Normal range (mmol/L)	Causes of imbalance: Excess	Causes of imbalance: Deficiency
Sodium, Na^+	135–145	hypernatremia (increased urine excretion; excess water loss)	hyponatremia (dehydration; diabetes-related low blood pH; vomiting; diarrhea)
Potassium, K^+	3.5–5.0	hyperkalemia (renal failure; low blood pH)	hypokalemia (gastrointestinal conditions)
Hydrogen carbonate, HCO_3^-	24–30	hypercapnia (high blood pH; hypoventilation)	hypocapnia (low blood pH; hyperventilation; dehydration)
Chloride, Cl^-	100–106	hyperchloremia (anemia; heart conditions; dehydration)	hypochloremia (acute infection; burns; hypoventilation)

Sodium-Potassium Pump in the Cell Membrane

The process of active transport allows a cell to maintain its proper electrolyte balance. To keep the ion concentrations at the proper levels shown in Table 1B, a sodium-potassium pump embedded in the cell membrane shuttles sodium ions out of the cell across the cell membrane. A model for the action of the sodium-potassium pump is shown below.

Nerve Impulses and Ion Concentration

The difference in Na^+ and K^+ concentrations inside and outside nerve cell membranes is essential for the normal operation of the nervous system. This unequal concentration of ions creates a voltage across nerve cell membranes. When a nerve cell is stimulated, sodium ions diffuse into the cell from the surrounding fluid, raising voltage across the nerve cell membrane from –70 mV to nearly +60 mV. Potassium ions then diffuse out of the cell into the surrounding fluid, restoring the voltage across the nerve cell membrane to –70 mV. This voltage fluctuation initiates the transmission of a nerve impulse. The amount of Na^+ inside the cell has increased slightly, and the amount of K^+ outside the cell has decreased. But the sodium-potassium pump will restore these ions to their proper concentrations.



The sodium-potassium pump is a protein embedded within the cell membrane that allows the passage of Na^+ and K^+ into and out of the cell. Each figure depicts the action of a single protein.

What's your sodium IQ?

Though sodium is an important mineral in your body, a diet that is high in sodium is one of several factors linked to high blood pressure, also known as hypertension. High Na^+ levels cause water retention, which results in increased blood pressure. Sodium is not the direct cause of all hypertension, but reducing sodium levels in the diet can affect individuals with a condition known as salt-sensitive hypertension. Therefore, the Dietary Guidelines for Americans recommend consuming salt and sodium in moderation. Test your knowledge about sodium in foods with the questions below.

1. Which of the following condiments do you think has the lowest salt content?
a. mustard c. catsup e. vinegar
b. steak sauce d. pickles

2. One-fourth of a teaspoon of salt contains about _____ of sodium.
a. 10 mg c. 500 mg e. 1 kg
b. 100 g d. 500 g

3. According to FDA regulations for food product labels, a food labeled *salt-free* must contain less than _____ mg of sodium ion per serving.
a. 100 c. 0.001 e. 0.00005
b. 5 d. 0.005

4. The Nutrition Facts label for a particular food reads "Sodium 15 mg." This is the amount of sodium ion per _____.
a. package c. serving e. RDA
b. teaspoon d. ounce

5. The recommended average daily intake of sodium ion for adults is 2400 mg. For a low-sodium diet the intake should be ______.
a. 200 mg c. 750 mg e. 150 mg
b. 2000 mg d. 500 mg

6. Each of the following ingredients can be found in the ingredients lists for some common food products. Which ones indicate that the product contains sodium?
a. trisodium phosphate d. sodium sulfate
b. sodium bicarbonate e. MSG
c. sodium benzoate f. baking soda

7. Which of the following spices is NOT a salt substitute?
a. caraway seeds c. ginger
b. dill d. onion salt

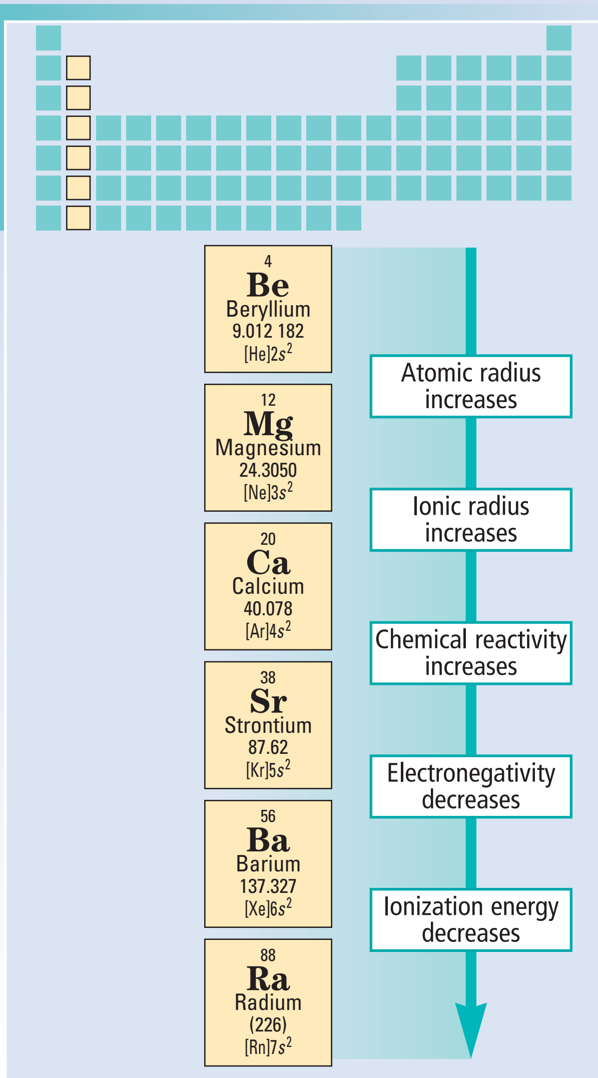
8. Most salt in the average American diet comes from salting foods too heavily at the dinner table.
a. true b. false

9. Which of the following foods are high in sodium?
a. potato chips c. doughnuts e. figs
b. pizza d. banana

10. Your body requires about 200 mg of sodium ion, or 500 mg of salt, per day. Why do these numbers differ?

Answers 1. e; 2. c; 3. b; 4. c; 5. c; 6. all of them; 7. d; 8. b, processed foods can contain very high levels of sodium; 9. a, b, c; 10. Salt is not pure sodium.

GROUP 2
ALKALINE EARTH METALS



CHARACTERISTICS

- do not occur naturally as free elements
- occur most commonly as the carbonates, phosphates, silicates, and sulfates
- occur naturally as compounds that are either insoluble or only slightly soluble in water
- consist of atoms that contain two electrons in their outermost energy level
- consist of atoms that tend to lose two electrons per atom, forming ions with a 2+ charge
- are less reactive than alkali metals
- form ionic compounds primarily
- react with water to form bases and hydrogen gas
- are good heat and electrical conductors
- are ductile and malleable
- have a silvery luster
- include the naturally radioactive element radium

Calcium carbonate is a major component of marble.

Beryllium is found in the mineral compound beryl. Beryl crystals include the dark green emerald and the blue-green aquamarine. The colors of these gems come from other metal impurities.

The mineral dolomite, $CaCO_3$•$MgCO_3$, is a natural source of both calcium and magnesium.

COMMON REACTIONS

With Water to Form Bases and Hydrogen Gas
Example: $Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$
Ca, Sr, and Ba also follow this pattern.

With Acids to Form Salts and Hydrogen Gas
Example: $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
Be, Ca, Sr, and Ba also follow this pattern.

With Halogens to Form Salts
Example: $Mg(s) + F_2(g) \longrightarrow MgF_2(s)$
Ca, Sr, and Ba also follow this pattern in reacting with F_2, Cl_2, Br_2, and I_2.

With Oxygen to Form Oxides or Peroxides
Magnesium forms an oxide.
$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$
Be and Ca also follow this pattern.

Strontium also forms a peroxide.
$Sr(s) + O_2(g) \longrightarrow SrO_2(s)$
Ba also reacts in this way.

With Hydrogen to Form Hydrides
Example: $Mg(s) + H_2(g) \longrightarrow MgH_2(s)$
Ca, Sr, and Ba also follow this pattern.

With Nitrogen to Form Nitrides
Example: $3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$
Be and Ca also follow this pattern.

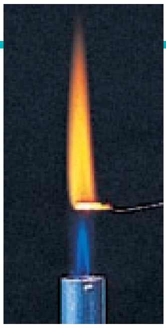
Calcium reacts with water to form hydrogen gas.

Magnesium burns in air to form MgO and Mg_3N_2.

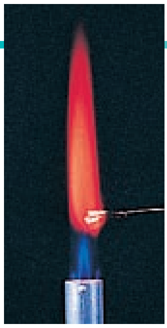
Magnesium reacts with HCl to produce $MgCl_2$(aq).

ANALYTICAL TEST

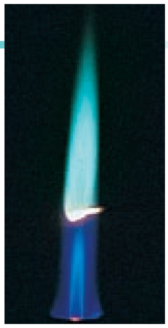
Flame tests can be used to identify three of the alkaline earth elements. The colors of both calcium and strontium can be masked by the presence of barium, which produces a green flame.



Calcium



Strontium



Barium

PROPERTIES OF THE GROUP 2 ELEMENTS

	Be	Mg	Ca	Sr	Ba	Ra
Melting point (°C)	1278 ± 5	649	839 ± 2	769	725	700
Boiling point (°C)	2467	1090	1484	1384	1640	1140
Density (g/cm^3)	1.85	1.74	1.54	2.6	3.51	5
Ionization energy (kJ/mol)	900	738	590	550	503	509
Atomic radius (pm)	112	160	197	215	222	220
Ionic radius (pm)	45	72	100	118	136	148
Common oxidation number in compounds	+2	+2	+2	+2	+2	+2
Crystal structure	hcp*	hcp	fcc**	fcc	bcc	bcc
Hardness (Mohs' scale)	4.0	2.0	1.5	1.8	1.5	—

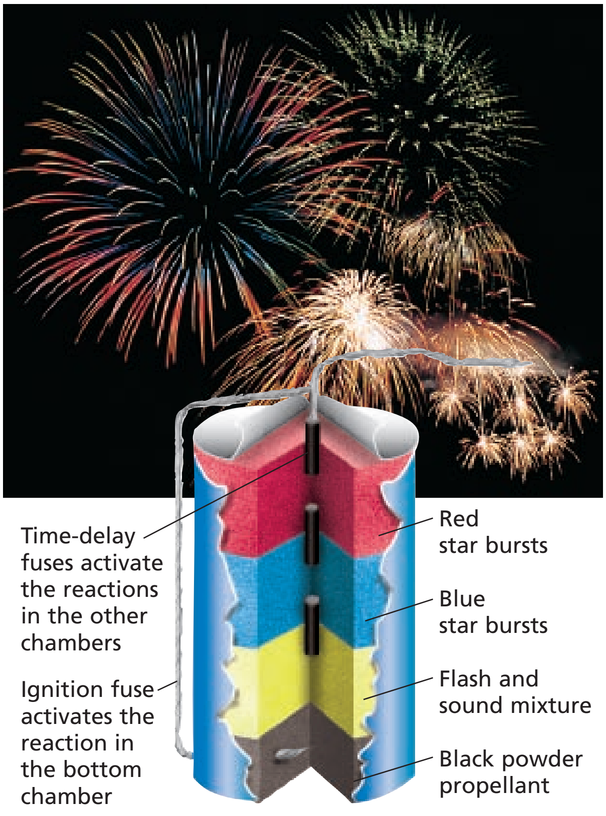
*hexagonal close-packed **face-centered cubic

APPLICATION *Technology*

Fireworks

Fireworks are made from pyrotechnics—chemical substances that produce light and smoke when they are ignited. Pyrotechnics are also used in flares, smoke bombs, explosives, and matches. An aerial fireworks device is a rocket made of a cylinder, chemicals inside the cylinder, and fuses attached to the cylinder. The illustration on the right shows how the device works. The lift charge at the bottom of the cylinder consists of a small amount of black gunpowder. When the side fuse ignites the gunpowder, it explodes like a small bomb. The gunpowder consists of potassium nitrate, charcoal, and sulfur. When these three chemicals react with one another, they produce gases. In this case, the gases produced are carbon monoxide, carbon dioxide, sulfur dioxide, and nitrogen monoxide. These hot gases expand very rapidly, providing the thrust that lifts the rocket into the sky.

About the time the shell reaches its maximum altitude and minimum speed, the time fuse ignites the chemicals contained in the cylinder. The chemicals inside the cylinder determine the color of the burst.



The cylinder of a multiple-burst rocket contains separate reaction chambers connected by fuses. A common fuse ignites the propellant and the time-delay fuse in the first reaction chamber.

Chemical Composition and Color

One of the characteristics of fireworks that we enjoy most is their variety of rich colors. These colors are created in much the same way as the colors produced during a flame test. In a fireworks device, the chloride salt is heated to a high temperature, causing the excited atoms to give off a burst of light. The color of light produced depends on the metal used. The decomposition of barium chloride, $BaCl_2$, for example, produces a burst of green light, whereas strontium chloride, $SrCl_2$, releases red light.

People who design fireworks combine artistry with a technical knowledge of chemical properties. They have found ways to combine different colors within a single cylinder and to make parts of the cylinder explode at different times. Fireworks designers have a technical knowledge of projectile motion that is used to determine the height, direction, and angle at which a fireworks device will explode to produce a fan, fountain, flower, stream, comet, spider, star, or other shape.

Strontium and the Visible Spectrum

When heated, some metallic elements and their compounds emit light at specific wavelengths that are characteristic of the element or compound. Visible light includes wavelengths between about 400 and 700 nanometers. The figure below shows the emission spectrum for strontium. When heated, strontium gives off the maximum amount of visible light at about 700 nanometers, which falls in the red-light region of the visible spectrum.

The emission spectrum for strontium shows strong bands in the red region of the visible light spectrum.

Flares

Flares operate on a chemical principle that is different from that of fireworks. A typical flare consists of finely divided magnesium metal and an oxidizing agent. When the flare is ignited, the oxidizing agent reacts with the magnesium metal to produce magnesium oxide. This reaction releases so much energy that it produces a glow like that of the filaments in a light bulb. The brilliant white light produced by the flare is caused by billions of tiny particles of magnesium that glow when they react. If slightly larger particles of magnesium metal are used in the flare, the system glows for a longer period of time because the particles' reaction with the oxidizing agent is slower.

A colored flare can be thought of as a combination of a white flare and a chemical that produces colored light when burned. For example, a red flare can be made from magnesium metal, an oxidizing agent, and a compound of strontium. When the flare is ignited, the oxidizing agent and magnesium metal react, heating the magnesium to white-hot temperatures. The energy from this reaction causes the strontium compound to give off its characteristic red color.

A flare is made up of billions of reacting magnesium particles.

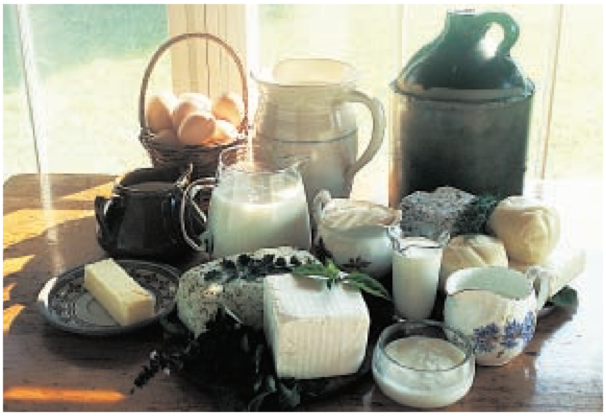
For safety reasons, some fireworks manufacturers store their products in metal sheds separated by sand banks. Also, people who work with fireworks are advised to wear cotton clothing because cotton is less likely than other fabrics to develop a static charge, which can cause a spark and accidentally ignite fireworks.

APPLICATION *Health*

Calcium: An Essential Mineral in the Diet

Calcium is the most abundant mineral in the body. It is the mineral that makes up a good portion of the teeth and the bone mass of the body. A small percentage of calcium in the body is used in the reactions by which cells communicate and in the regulation of certain body processes. Calcium is so important to normal body functioning that if the calcium level of the blood falls far below normal, hormones signal the release of calcium from bone and signal the gastrointestinal tract to absorb more calcium during the digestion process.

A prolonged diet that is low in calcium is linked to a disease characterized by a decrease in bone mass, a condition called osteoporosis. Reduced bone mass results in brittle bones that fracture easily. Osteoporosis generally occurs later in life and is more prevalent in females. However, because you achieve peak bone mass during the late teens or early twenties, it is critical that your diet meet the recommended requirements to increase your peak bone mass. The recommended dietary intake for calcium is 1000 mg per day. Maintaining that level in the diet along with regular exercise through adulthood are thought to reduce the rate of bone loss later in life. Excess calcium in the diet (consuming more than 2500 mg daily) can interfere with the absorption of other minerals.



Dairy products are generally good sources of calcium.

Magnesium: An Essential Mineral in the Diet

Though magnesium has several functions in the body, one of the more important functions is its role in the absorption of calcium by cells. Magnesium, like sodium and potassium, is involved in the transmission of nerve impulses. Like calcium, magnesium is a component of bone.

A major source of magnesium in the diet is plants. Magnesium is the central atom in the green plant pigment chlorophyll. The structure of chlorophyll in plants is somewhat similar to the structure of heme—the oxygen-carrying molecule in animals. (See page 816 for the heme structure.)

TABLE 2A Good Sources of Calcium in the Diet

Food	Serving size	Calcium present (mg)
Broccoli	6.3 oz	82
Cheddar cheese	1 oz	204
Cheese pizza, frozen	pizza for one	375
Milk, low-fat 1%	8 oz	300
Tofu, regular	4 oz	130
Vegetable pizza, frozen	pizza for one	500
Yogurt, low-fat	8 oz	415
Yogurt, plain whole milk	8 oz	274

Spinach is a good source of magnesium. Magnesium is the central atom in the green plant pigment chlorophyll. The chlorophyll structure is shown on the right.

The recommended dietary intake of magnesium is 400 mg per day. This is equivalent to just 4 oz of bran cereal. Because magnesium levels are easily maintained by a normal diet, it is unusual for anyone to have a magnesium deficiency. Most magnesium deficiencies are the result of factors that decrease magnesium absorption. People with gastrointestinal disorders, alcohol abusers, and the critically ill are most likely to have these types of absorption problems.

Excess magnesium in the diet is excreted by the kidneys, so there are no cumulative toxic effects.

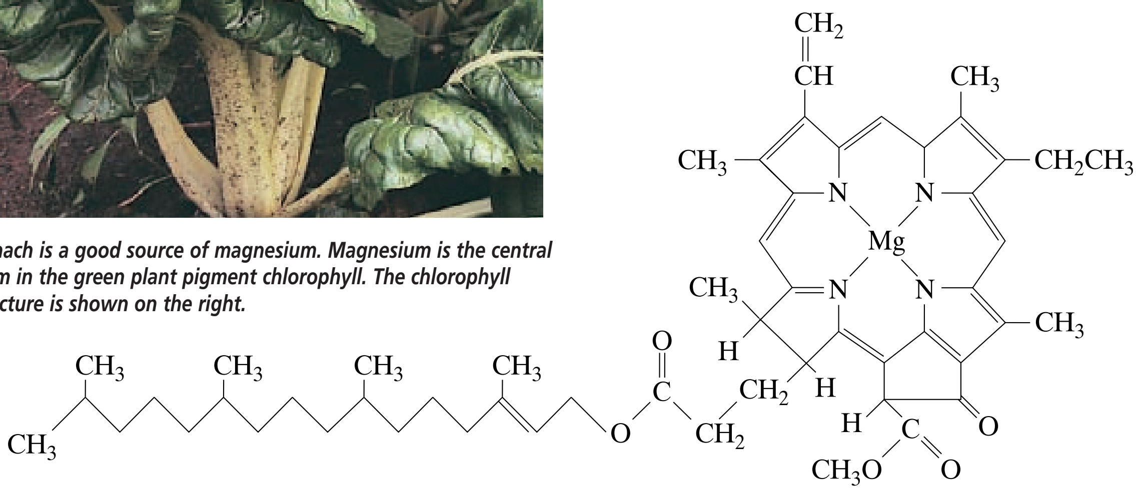
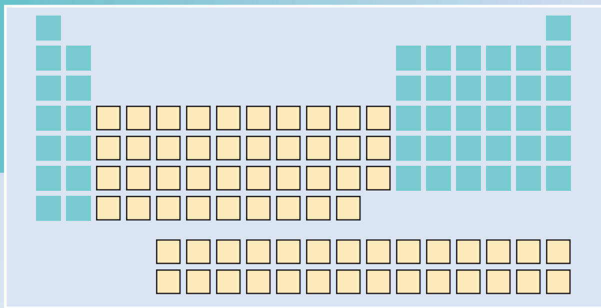


TABLE 2B Good Sources of Magnesium in the Diet

Food	Serving size	Magnesium present (mg)
Barley, raw	1 cup	244
Beef, broiled sirloin	4 oz	36
Cabbage, raw	1 med. head	134
Cashews, dry-roasted	1 oz	74
Chicken, roasted breast	4 oz	31
Lima beans, boiled	1/2 cup	63
Oatmeal	1 oz	39
Potato, baked	7.1 oz	115
Prunes, dried	4 oz	51
Rice bran	8 oz	648
Salmon, canned	4 oz	39
Spinach, raw	10 oz	161

GROUPS 3–12
TRANSITION METALS



CHARACTERISTICS

- consist of metals in Groups 3 through 12
- contain one or two electrons in their outermost energy level
- are usually harder and more brittle than metals in Groups 1 and 2
- have higher melting and boiling points than metals in Groups 1 and 2
- are good heat and electrical conductors
- are malleable and ductile
- have a silvery luster, except copper and gold
- include radioactive elements with numbers 89 through 112
- include mercury, the only liquid metal at room temperature
- have chemical properties that differ from each other
- tend to have two or more common oxidation states
- often form colored compounds
- may form complex ions

Iron ore is obtained from surface mines. Hematite, Fe_2O_3, is the most common iron ore.

Copper ores are also obtained from surface mines. Copper ore is shown here.

Gold, silver, platinum, palladium, iridium, rhodium, ruthenium, and osmium are sometimes referred to as the noble metals because they are not very reactive. These metals are found in coins, jewelry, and metal sculptures.

COMMON REACTIONS

Because this region of the periodic table is so large, you would expect great variety in the types of reaction characteristics of transition metals. For example, copper oxidizes in air to form the green patina you see on the Statue of Liberty. Copper reacts with concentrated HNO_3 but not with dilute HNO_3. Zinc, on the other hand, reacts readily with dilute HCl. Iron oxidizes in air to form rust, but chromium is generally unreactive in air. Some common reactions for transition elements are shown by the following.

May form two or more different ions

Example: $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$

Example: $Fe(s) \longrightarrow Fe^{3+}(aq) + 3e^-$

May react with oxygen to form oxides

Example: $4Cr(s) + 3O_2(g) \longrightarrow 2Cr_2O_3(s)$

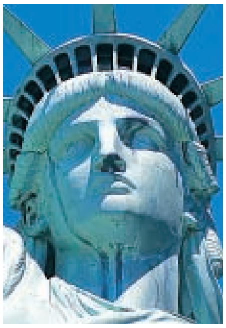
Example: $2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$

May react with halogens to form halides

Example: $Ni(s) + Cl_2(g) \longrightarrow NiCl_2(s)$

May form complex ions

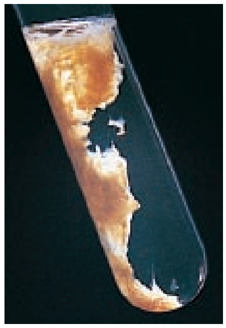
See examples in the lower right.



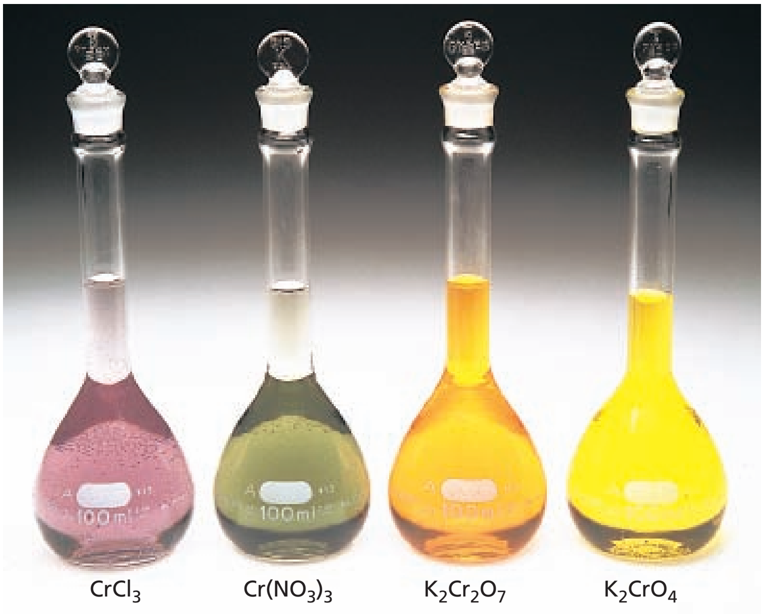
Copper reacts with oxygen in air.

Copper reacts with concentrated nitric acid.

Zinc reacts with dilute hydrochloric acid.



Soluble iron(III) salts form insoluble $Fe(OH)_3$ when they are reacted with a hydroxide base.

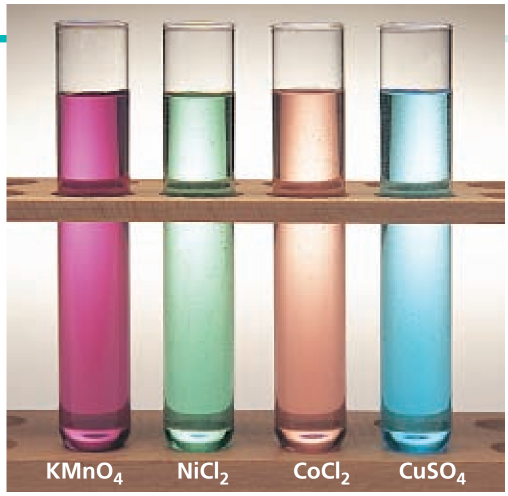


Chromium has several common oxidation states, represented here by aqueous solutions of its compounds. The violet and green solutions contain chromium in the +3 state, and the yellow and orange solutions contain chromium in the +6 oxidation state.

Complex ions belong to a class of compounds called coordination compounds. Coordination compounds show great variety in colors. Several transition-metal coordination compounds are shown.

ANALYTICAL TEST

Flame tests are not commonly used to identify transition metals. The presence of a certain transition-metal ion in a solution is sometimes obvious from the solution's color. Some transition-metal ions can be more accurately identified using a procedure called qualitative analysis. **Qualitative analysis** *is the identification of ions by their characteristic reactions.* The transition-metal ions most often identified through qualitative analysis include copper, nickel, zinc, chromium, iron cobalt, cadmium, manganese, and tin. Most tests to identify the presence of an ion in a mixture involve causing the ion to precipitate out of solution. Some of the more dramatic precipitation reactions for transition metals are shown.



Some transition metal ions can be identified by characteristic colors of their salt solutions.

Copper (formation of $[Cu(NH_3)_4](OH)_2$)

Cadmium (formation of CdS)

Zinc (formation of ZnS)

Chromium (formation of $PbCrO_4$)

Iron (formation of $[Fe(SCN)]^{2+}$)

Manganese (formation of MnO_4^-)

Nickel (formation of a nickel dimethylglyoxime complex)

PROPERTIES OF SOME TRANSITION METALS

	Cr	Fe	Co	Ni	Cu	Zn	Ag	Au	Hg
Melting point (°C)	1857 ± 20	1535	1495	1455	1083	420	962	1064	–38.8
Boiling point (°C)	2672	2750	2870	2732	2567	907	2212	2808 ± 2	356.6
Density (g/cm^3)	7.20	7.86	8.9	8.92	8.96	7.14	10.5	19.3	13.5
Ionization energy (kJ/mol)	653	762	760	737	746	906	731	890	1007
Atomic radius (pm)	128	126	125	124	128	134	144	144	151
Common oxidation numbers	+2, +3, +6	+2, +3	+2, +3	+2	+1, +2	+2	+1	+1, +3	+1, +2

APPLICATION *Geology*

Gemstones and Color

A gemstone is a mineral that can be cut and polished to make gems for an ornament or piece of jewelry. At one time, all gemstones were naturally occurring minerals mined from Earth's crust. Today, however, chemists can duplicate natural processes to produce artificial gemstones. Amethyst, emerald, jade, opal, ruby, sapphire, and topaz occur naturally and can also be produced synthetically.

The color of a gemstone is determined by the presence of small amounts of one or more transition metals. For example, aluminum oxide, Al_2O_3, often occurs naturally as corundum—a clear, colorless mineral. However, if as few as 1 to 2% of the aluminum ions, Al^{3+}, are replaced by chromium ions, Cr^{3+}, the corundum takes on a reddish color and is known as ruby. If a small fraction of aluminum ions in corundum are replaced by Fe^{3+} and Ti^{3+}, the corundum has a greenish color and is known as emerald. In another variation, if vanadium ions, V^{3+}, replace a few Al^{3+} ions in corundum, the result is a gemstone known as alexandrite. This gemstone appears green in reflected natural light and red in transmitted or artificial light.

Table 3A lists transition metals that are responsible for the colors of various gemstones. The table provides only a general overview, however, as most naturally occurring gemstones occur in a range of hues, depending on the exact composition of the stone.

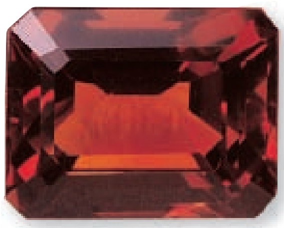
Artificial Gemstones

In 1902, the French chemist Auguste Verneuil found a way to melt a mixture of aluminum oxide and chromium salts and then cool the mixture very slowly to produce large crystals of reddish aluminum oxide—rubies.

Sapphire

Ruby

Peridot



Garnet

TABLE 3A Transition Metals and Gemstone Colors

Gemstone	Color	Element
Amethyst	purple	iron
Aquamarine	blue	iron
Emerald	green	iron/titanium
Garnet	red	iron
Peridot	yellow-green	iron
Ruby	red	chromium
Sapphire	blue	iron/titanium
Spinel	colorless to red to black	varies
Turquoise	blue	copper

Verneuil's method, although somewhat modified, is still the one most widely used today for the manufacture of colored gemstones. When magnesium oxide is substituted for aluminum oxide, a colorless spinel-like product is formed. The addition of various transition metals then adds a tint to the spinel that results in the formation of synthetic emerald, aquamarine, tourmaline, or other gemstones. Synthetic gems look very much like their natural counterparts.

Synthetic sapphire

Synthetic ruby

APPLICATION *Technology*

Alloys

An alloy is a mixture of a metal and one or more other elements. In most cases, the second component of the mixture is also a metal.

Alloys are desirable because mixtures of elements usually have properties different from and often superior to the properties of individual metals. For example, many alloys that contain iron are harder, stronger, and more resistant to oxidation than iron itself.

Amalgams are alloys that contain mercury. They are soft and pliable when first produced, but later become solid and hard. Dental fillings were once made of an amalgam of mercury and silver. Concerns about the possible toxicity of mercury led to the development of other filling materials.

Cast Iron and Steel

The term *steel* applies to any alloy consisting of iron and less than 1.5% carbon, and often other elements. When iron ore is treated with carbon in the form of coke to extract pure iron metal, some of the carbon also reacts with the iron to produce a form of iron carbide known as cementite. The reaction can be represented by the following equation.

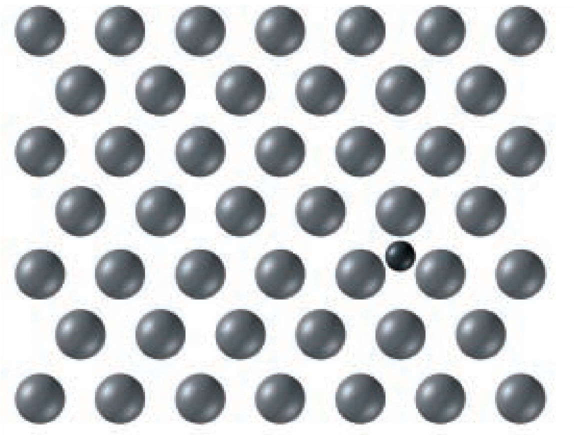
$$3Fe + C \longrightarrow Fe_3C$$

Cast iron is a mixture that consists of some pure iron, known as ferrite, some cementite, and some carbon atoms trapped within the crystalline structure of the iron and cementite. The rate at which cast iron is cooled changes the proportion of these three components. If the cast iron is cooled slowly, the ferrite and cementite tend to separate from each other, forming a banded product that is tough but not very hard. However, if the cast iron is cooled quickly, the components of the original mixture cannot separate from each other, forming a product that is both tough and hard.

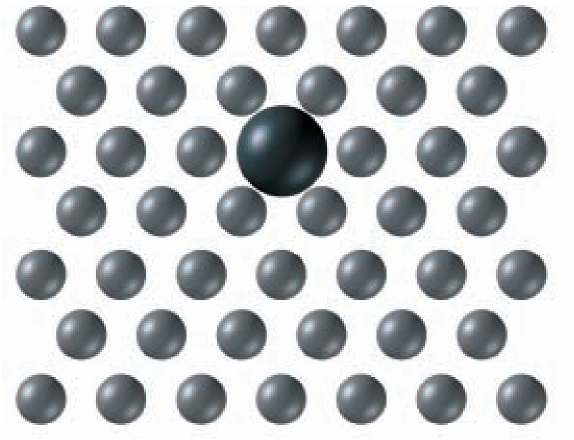
Stainless steel, which is hard and resists corrosion, is made of iron and chromium (12–30%). The properties of stainless steel make it a suitable alloy for making cutlery and utensils.

TABLE 3B Composition and Uses of Some Alloys

Name of alloy	Composition	Uses
Brass	copper with up to 50% zinc, some lead, and a small amount of tin	inexpensive jewelry; hose nozzles and couplings; piping; stamping dies
Bronze	copper with up to 12% tin	coins and medals; heavy gears; tools; electrical hardware
Coin metal	copper: 75% nickel: 25%	United States coins
Duralumin	aluminum: 95% copper: 4% magnesium: 0.5% manganese: <1%	aircraft, boats, railroad cars, and machinery because of its high strength and resistance to corrosion
Nichrome	nickel: 80–85% chromium: 15–20%	heating elements in toasters, electric heaters, etc.
Phosphor bronze	bronze with a small amount of phosphorus	springs, electrical springs, boat propellers
Solder	lead: 50%, tin: 50% or tin: 98%, silver: 2%	joining two metals to each other joining copper pipes
Sterling silver	silver: 92.5% copper: 7.5%	jewelry, art objects, flatware
Type metal	lead: 75–95% antimony: 2–18% tin: trace	used to make type for printing because it expands as it cools



Interstitial crystal
A smaller atom or ion fits into a small space between particles in the array.



Substitutional crystal
A larger atom or ion is substituted for a particle in the array.

Structures and Preparation of Alloys

Alloys generally crystallize in one of two ways, depending on relative sizes of atoms. If the atoms of one of the metals present are small enough to fit into the spaces between the atoms of the second metal, they form an alloy with an interstitial structure (*inter* means "between," and *stitial* means "to stand"). If atoms of the two metals are of similar size or if one is larger, the atoms of one metal can substitute for the atoms of the second metal in its crystalline structure. Such alloys are substitutional alloys. Models for both types of crystals are shown above.

Techniques for making alloys depend on the metals used in the mixture. In some cases, the two metals can simply be melted together to form a mixture. The composition of the mixture often varies within a range, evidence that the final product is indeed a mixture and not a compound. In other cases, one metal may be melted first and the second dissolved in it. Brass is prepared in this way. If copper and zinc were heated together to a high temperature, zinc (bp 907°C) would evaporate before copper (mp 1084°C) melted. Therefore, the copper is melted first, and the zinc is added to it.

Brass has a high luster and resembles gold when cleaned and polished. A brass object can be coated with a varnish to prevent reactions of the alloy with air and water.

Sterling silver is more widely used than pure silver because it is stronger and more durable.

APPLICATION *The Environment*

Mercury Poisoning

Mercury is the only metal that is liquid at room temperature. It has a very high density compared with most other common transition metals and has a very large surface tension and high vapor pressure. Mercury and many of its compounds must be handled with extreme care because they are highly toxic. Mercury spills are especially hazardous because the droplets scatter easily and are often undetected during cleanup. These droplets release toxic vapors into the air.

Overexposure to mercury vapor or its compounds can occur by absorption through the skin, respiratory tract, or digestive tract. Mercury is a cumulative poison, which means that its concentration in the body increases as exposure increases.

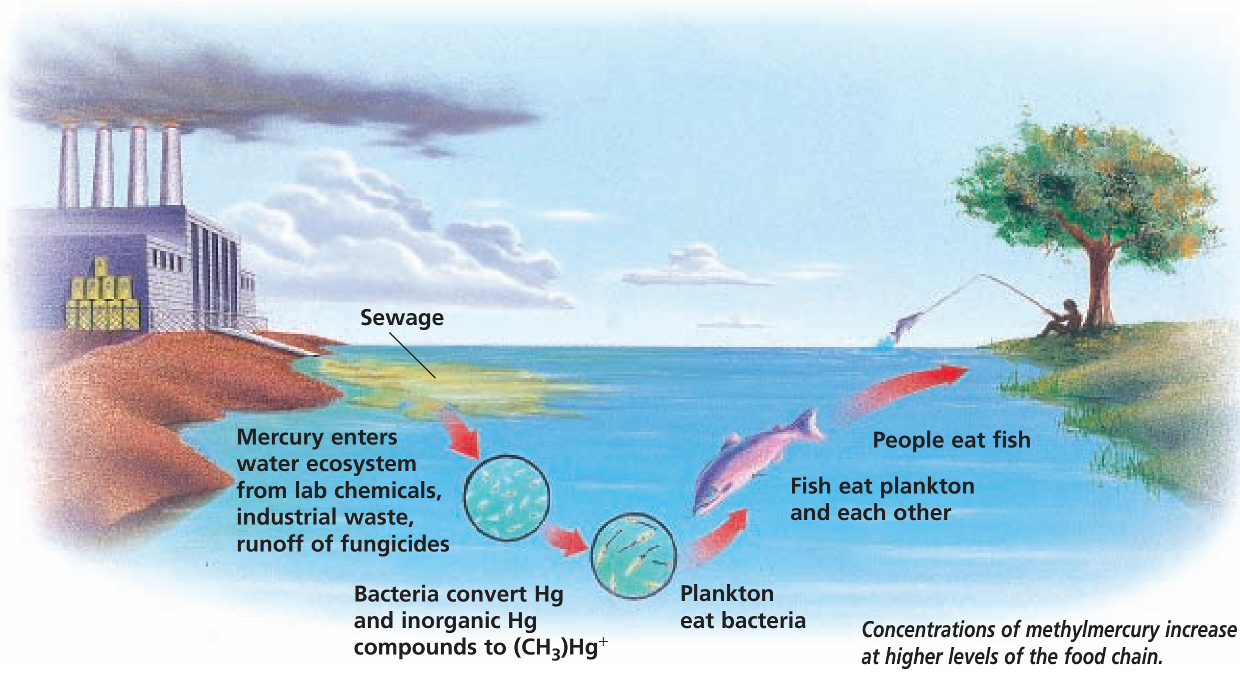
Mercury that enters the body damages the kidneys, heart, and brain. The action of mercury on the brain affects the nervous system. Symptoms of mercury poisoning include numbness, tunnel vision, garbled speech, bleeding and inflammation of the gums, muscle spasms, anemia, and emotional disorders, such as depression, irritability, and personality changes. The saying “mad as a hatter” probably came about because of mercury poisoning. Mercury salts were once routinely used to process the felt used in hats. Hatters often displayed the nerve and mental impairments associated with overexposure to mercury.

Methylmercury in Freshwater Ecosystems

Mercury, Hg, can be found in our environment and in our food supply. Fortunately, the body has some protective mechanisms to deal with trace amounts of mercury. However, levels of mercury and of methylmercury, $(CH_3)Hg^+$, are increasing in the environment due to mercury mining operations and runoff from the application of pesticides and fungicides.

Mercury is easily converted to methylmercury by bacteria. Methylmercury is more readily absorbed by cells than mercury itself. As a result, methylmercury accumulates in the food chain as shown in the diagram below. A serious incident of methylmercury poisoning occurred in Japan in the 1950s. People living in Minamata, Japan, were exposed to high levels of methylmercury from eating shellfish.

In the United States, there is concern about mercury levels in fish from some freshwater lakes. Though environmental regulations have reduced the level of lake pollutants, it takes time to see a reduction in the concentration of an accumulated poison.



Concentrations of methylmercury increase at higher levels of the food chain.

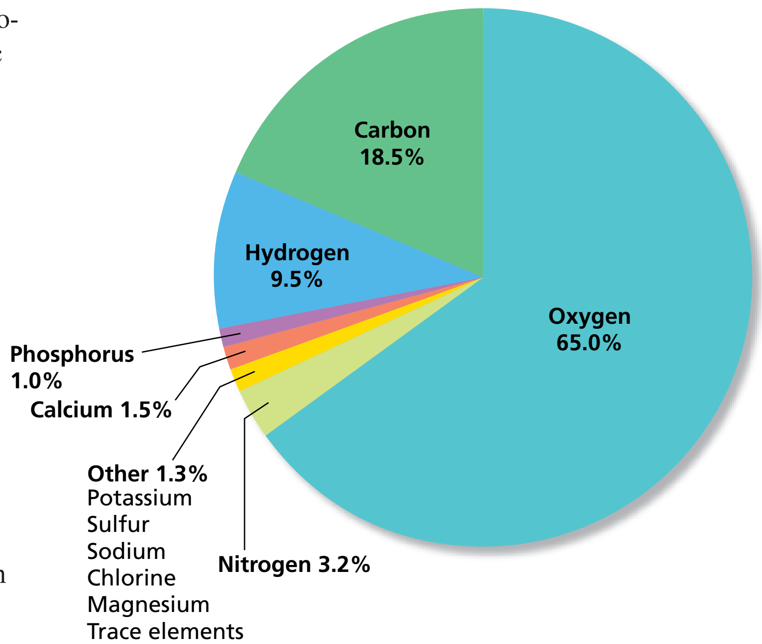
APPLICATION *Health*

Elements in the Body

The four most abundant elements in the body (oxygen, carbon, hydrogen, and nitrogen) are the major components of organic biomolecules, such as carbohydrates, proteins, fats, and nucleic acids. Other elements compose a dietary category of compounds called minerals. Minerals are considered the inorganic elements of the body. Minerals fall into two categories—the major minerals and the trace minerals, or trace elements, as they are sometimes called. Notice in the periodic table below that most elements in the trace elements category of minerals are transition metals.

Trace elements are minerals with dietary daily requirements of 100 mg or less. They are found in foods derived from both plants and animals. Though these elements are present in very small quantities, they perform a variety of essential functions in the body, as shown in Table 3C on the next page.

Abundance of Elements in the Body (by mass)



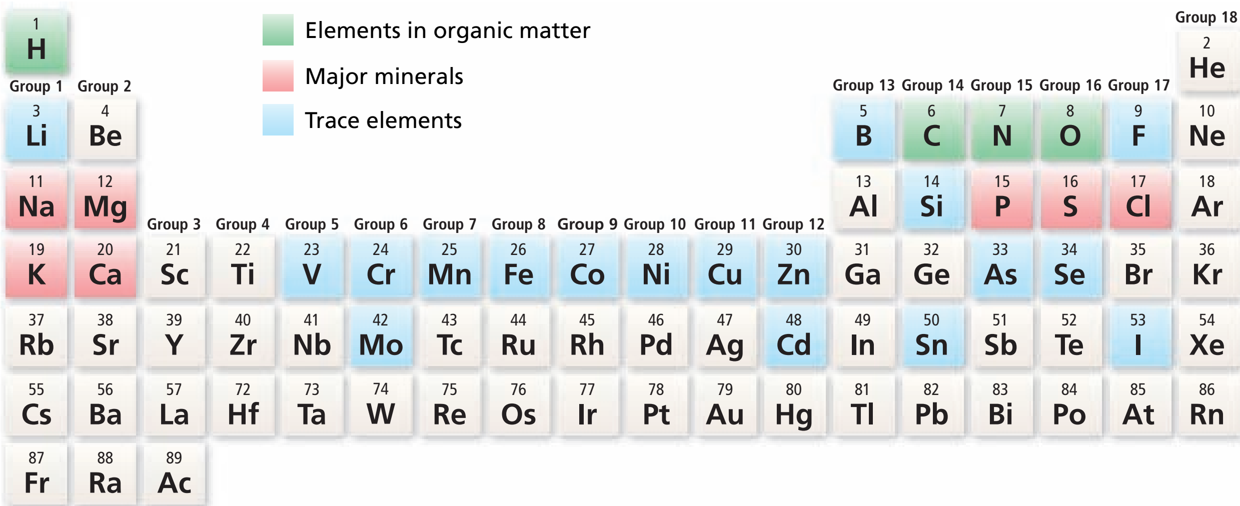


TABLE 3C Transition Metal Trace Elements

Transition metal	Function
Vanadium, Cadmium	function not fully determined, but linked to a reduced growth rate and impaired reproduction
Chromium	needed for glucose transport to cells
Manganese	used in the enzyme reactions that synthesize cholesterol and metabolize carbohydrates
Iron	central atom in the heme molecule—a component of hemoglobin, which binds oxygen in the blood for transport to cells
Cobalt	a component of vitamin B_{12}
Nickel	enzyme cofactor in the metabolism of fatty acids and amino acids
Copper	a major component of an enzyme that functions to protect cells from damage
Zinc	needed for tissue growth and repair and as an enzyme cofactor
Molybdenum	enzyme cofactor in the production of uric acid

Role of Iron

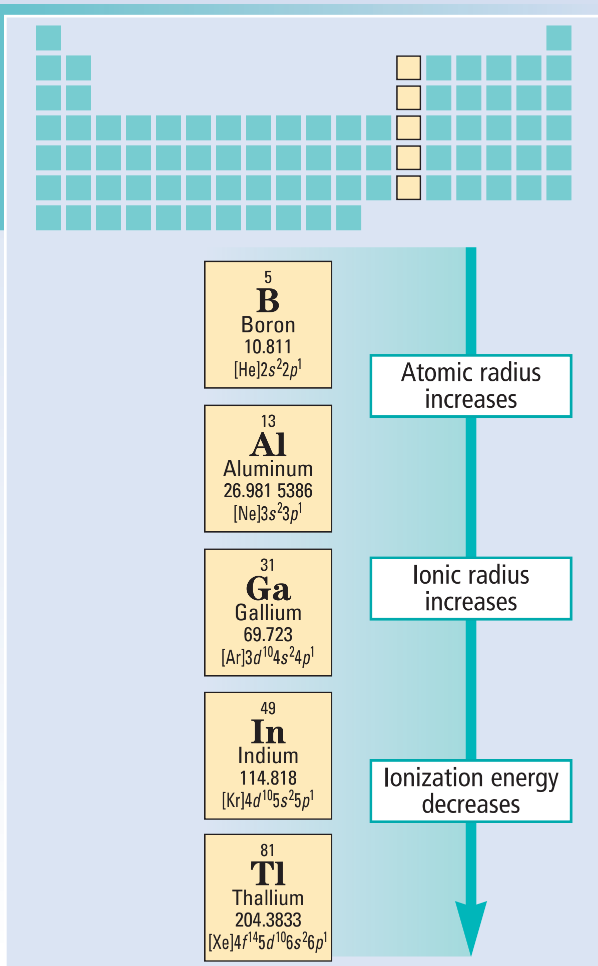
Most iron in the body is in hemoglobin. Fe^{3+} is the central ion in the heme molecule, which is a component of the proteins hemoglobin and myoglobin. Hemoglobin in red blood cells transports oxygen to cells and picks up carbon dioxide as waste. Myoglobin is a protein that stores oxygen to be used in muscle contraction. Iron is also in the proteins of the electron transport system and the immune system.

Mechanisms of the body control the rate of iron absorption from food in the diet. When iron reserves are low, chemical signals stimulate cells of the intestines to absorb more iron during digestion. If the diet is low in iron, causing a deficiency, hemoglobin production stops and a condition called iron-deficiency anemia results. The blood cells produced during this state are stunted and unable to deliver adequate oxygen to cells. As a result, a person with iron-deficiency anemia feels tired and weak and has difficulty maintaining normal body temperature. The recommended daily intake of iron is 15 mg. The recommended level doubles for pregnant women. Iron supplements are for people who do not get enough iron in their daily diets. Table 3D lists some foods that are good sources of iron in the diet. Too much iron can be toxic because the body stores iron once it is absorbed. Abusing iron supplements can cause severe liver and heart damage.

TABLE 3D Sources of Iron in Foods

Food	Serving size	Iron present (mg)
Beef roast (lean cut)	4 oz	3.55
Beef, T-bone steak (lean cut)	4 oz	3.40
Beef, ground (hamburger)	4 oz	2.78
Broccoli	6.3 oz	1.50
Chicken, breast	4 oz	1.35
Chicken, giblets	4 oz	7.30
Oatmeal, instant enriched	1 pkg	8.35
Pita bread, white enriched	6 1/2 in. diameter	1.40
Pork roast	4 oz	1.15
Prunes	4 oz	2.00
Raisins	4 oz	1.88

GROUP 13
BORON FAMILY



CHARACTERISTICS

- do not occur in nature as free elements
- are scarce in nature (except aluminum, which is the most abundant metallic element)
- consist of atoms that have three electrons in their outer energy level
- are metallic solids (except boron, which is a solid metalloid)
- are soft and have low melting points (except boron, which is hard and has a high melting point)
- are chemically reactive at moderate temperatures (except boron)

Boron is a covalent solid. Other members of the family are metallic solids.

The warmth of a person's hand will melt gallium. Gallium metal has the lowest melting point (29.77°C) of any metal except mercury.

Aluminum is the most abundant metal in Earth's crust. It exists in nature as an ore called bauxite.

COMMON REACTIONS

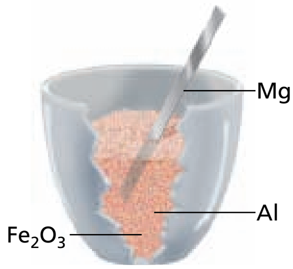
The reaction chemistry of boron differs greatly from that of the other members of this family. Pure boron is a covalent network solid, whereas the other members of the family are metallic crystals in pure form. Boron resembles silicon more closely than it resembles the other members of its family.

With Strong Bases to Form Hydrogen Gas and a Salt
Example: $2Al(s) + 2NaOH(aq) + 2H_2O(l) \longrightarrow 2NaAlO_2(aq) + 3H_2(g)$
Ga also follows this pattern.

With Dilute Acids to Form Hydrogen Gas and a Salt
Example: $2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$
Ga, In, and Tl follow this pattern in reacting with dilute HF, HCl, HBr, and HI.

With Halogens to Form Halides
Example: $2Al(s) + 3Cl_2(g) \longrightarrow 2AlCl_3(s)$
B, Al, Ga, In, and Tl also follow this pattern in reacting with F_2, Cl_2, Br_2, and I_2 (except BF_3).

With Oxygen to Form Oxides
Example: $4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$
Ga, In, and Tl also follow this pattern.

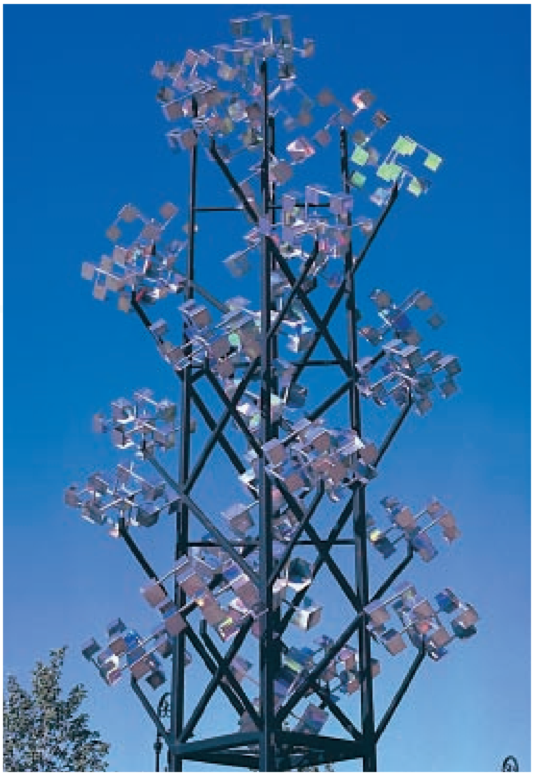


A mixture of powdered aluminum and iron(III) oxide is called thermite. Al reacts with Fe_2O_3 using Mg ribbon as a fuse to provide activation energy. The energy produced by the thermite reaction is sufficient to produce molten iron as a product.

ANALYTICAL TEST

Other than atomic absorption spectroscopy, there is no simple analytical test for all the members of the boron family.

The confirmatory test for the presence of aluminum in qualitative analysis is the red color formed by aluminum and the organic compound aluminon, $C_{22}H_{23}N_3O_9$.



Aluminum forms a thin layer of Al_2O_3, which protects the metal from oxidation and makes it suitable for outdoor use.

PROPERTIES OF THE GROUP 13 ELEMENTS

	B	Al	Ga	In	Tl
Melting point (°C)	2300	660.37	29.77	156.61	303.5
Boiling point (°C)	2550	2467	2403	2080	1457
Density (g/cm^3)	2.34	2.702	5.904	7.31	11.85
Ionization energy (kJ/mol)	801	578	579	558	589
Atomic radius (pm)	85	143	135	167	170
Ionic radius (pm)	—	54	62	80	89
Common oxidation number in compounds	+3	+3	+1, +3	+1, +3	+1, +3
Crystal structure	monoclinic	fcc	orthorhombic	fcc	hcp
Hardness (Mohs' scale)	9.3	2.75	1.5	1.2	1.2

APPLICATION *Technology*

Aluminum

Chemically, aluminum is much more active than copper, and it belongs to the category of *self-protecting metals*. These metals are oxidized when exposed to oxygen in the air and form a hard, protective metal oxide on the surface. The oxidation of aluminum is shown by the following reaction.

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

This oxide coating protects the underlying metal from further reaction with oxygen or other substances. Self-protecting metals are valuable in themselves or when used to coat iron and steel to keep them from corroding.

Aluminum is a very good conductor of electric current. Many years ago, most high-voltage electric power lines were made of copper. Although copper is a better conductor of electricity than aluminum, copper is heavier and more expensive. Today more than 90% of high-voltage transmission lines are made of relatively pure aluminum. The aluminum wire does not have to be self-supporting because steel cable is incorporated to bear the weight of the wire in the long spans between towers.

In the 1960s, aluminum electric wiring was used in many houses and other buildings. Over time, however, because the aluminum oxidized, Al_2O_3 built up and increased electric resistance at points where wires connected to outlets, switches, and other metals. As current flowed through the extra resistance, enough energy as heat was generated to cause a fire. Though some homes have been rewired, aluminum wiring is still prevalent in many homes.

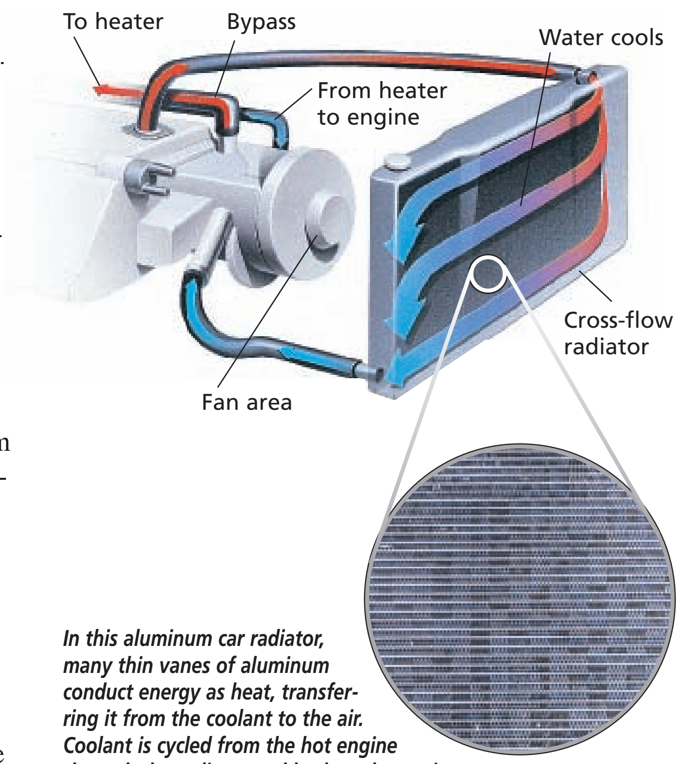
These high-voltage transmission lines are made of aluminum supported with steel cables.

Aluminum Alloys

Because aluminum has a low density and is inexpensive, it is used to construct aircraft, boats, sports equipment, and other lightweight, high-strength objects. The pure metal is not strong, so it is mixed with small quantities of other metals—usually manganese, copper, magnesium, zinc, or silicon—to produce strong low-density alloys. Typically, 80% of a modern aircraft frame consists of aluminum alloy.

Aluminum and its alloys are good heat conductors. An alloy of aluminum and manganese is used to make cookware. High-quality pots and pans made of stainless steel may have a plate of aluminum on the bottom to help conduct energy as heat quickly to the interior.

Automobile radiators made of aluminum conduct energy as heat as hot coolant from the engine enters the bottom of the radiator. The coolant is deflected into several channels. These channels are covered by thin vanes of aluminum, which conduct energy away from the coolant and transfer it to the cooler air rushing past. By the time the coolant reaches the top of the radiator, its temperature has dropped so that when it flows back into the engine it can absorb more energy as heat. To keep the process efficient, the outside of a radiator should be kept unobstructed and free of dirt buildup.



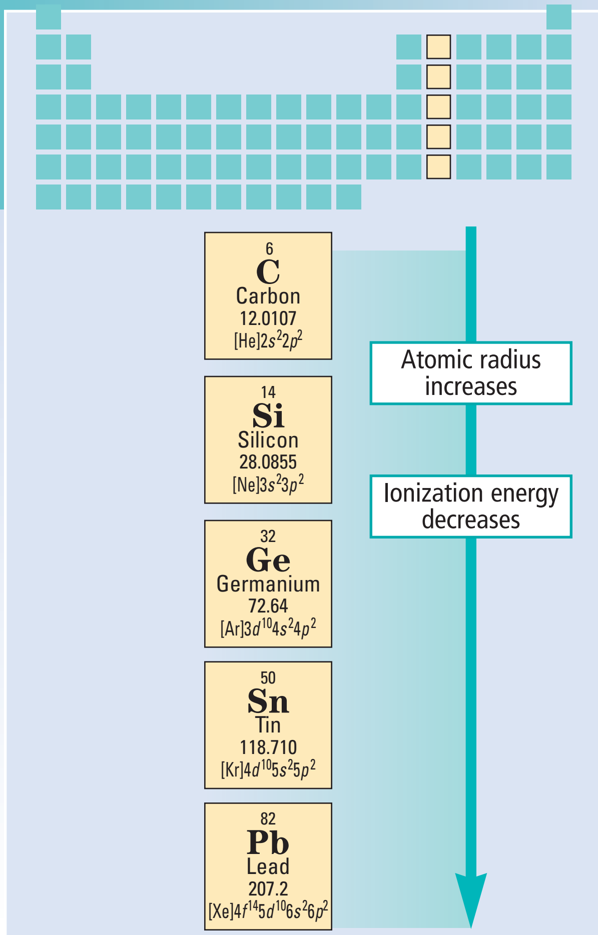
In this aluminum car radiator, many thin vanes of aluminum conduct energy as heat, transferring it from the coolant to the air. Coolant is cycled from the hot engine through the radiator and back to the engine.

TABLE 4A Alloys of Aluminum and Their Uses

Principal alloying element(s)*	Characteristics	Application examples
Manganese	moderately strong, easily worked	cookware, roofing, storage tanks, lawn furniture
Copper	strong, easily formed	aircraft structural parts; large, thin structural panels
Magnesium	strong, resists corrosion, easy to weld	parts for boats and ships, outdoor decorative objects, tall poles
Zinc and magnesium	very strong, resists corrosion	aircraft structural parts, vehicle parts, anything that needs high strength and low weight
Silicon	expands little on heating and cooling	aluminum castings
Magnesium and silicon	resists corrosion, easily formed	exposed parts of buildings, bridges

* All these alloys have small amounts of other elements.

GROUP 14
CARBON FAMILY



CHARACTERISTICS

- include a nonmetal (carbon), two metalloids (silicon and germanium), and two metals (tin and lead)
- vary greatly in both physical and chemical properties
- occur in nature in both combined and elemental forms
- consist of atoms that contain four electrons in the outermost energy level
- are relatively unreactive
- tend to form covalent compounds (tin and lead also form ionic compounds)

Lead has a low reactivity and is resistant to corrosion. It is very soft, highly ductile, and malleable. Lead is toxic and, like mercury, it is a cumulative poison.

Silicon has a luster but does not exhibit metallic properties. Most silicon in nature is a silicon oxide, which occurs in sand and quartz, which is shown here.

Tin, which is shown on the right, is a self-protecting metal like lead, but unlike lead it has a high luster. Tin occurs in nature in cassiterite ore, which is shown above.

COMMON REACTIONS

With Oxygen to Form Oxides

Example: $Sn(s) + O_2(g) \longrightarrow SnO_2(s)$

Pb follows this pattern, as do C, Si, and Ge at high temperatures.

With Acids to Form Salts and Hydrogen Gas

Only the metallic elements of this group react slowly with aqueous acids.

Example: $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

Both Sn and Pb can also react to form tin(IV) and lead(IV) salts, respectively.

With Halogens to Form Halides

Example: $Sn(s) + 2Cl_2(g) \longrightarrow SnCl_4(s)$

Si, Ge, and Pb follow this pattern, reacting with F_2, Cl_2, Br_2, and I_2.

ANALYTICAL TEST

Ionic compounds of tin and lead can be identified in aqueous solutions by adding a solution containing sulfide ions. The formation of a yellow precipitate indicates the presence of Sn^{4+}, and the formation of a black precipitate indicates the presence of Pb^{2+}.

$$Sn^{4+}(aq) + 2S^{2-}(aq) \longrightarrow SnS_2(s)$$
$$Pb^{2+}(aq) + S^{2-}(aq) \longrightarrow PbS(s)$$

PbS SnS$_2$

PROPERTIES OF THE GROUP 14 ELEMENTS

	C	Si	Ge	Sn	Pb
Melting point (°C)	3500/3652*	1410	937.4	231.88	327.502
Boiling point (°C)	4827	2355	2830	2260	1740
Density (g/cm^3)	3.51/2.25*	2.33 ± 0.01	5.323	7.28	11.343
Ionization energy (kJ/mol)	1086	787	762	709	716
Atomic radius (pm)	77	118	122	140	175
Ionic radius (pm)	260 (C^{4-} ion)	—	—	118 (Sn^{2+} ion)	119 (Pb^{2+} ion)
Common oxidation number in compounds	+4, −4	+4	+2, +4	+2, +4	+2, +4
Crystal structure	cubic/hexagonal*	cubic	cubic	tetragonal	fcc
Hardness (Mohs' scale)	10/0.5*	6.5	6.0	1.5	1.5

* The data are for two allotropic forms: diamond/graphite.

APPLICATION *Chemical Industry*

Carbon and the Reduction of Iron Ore

Some metals, especially iron, are separated from their ores through reduction reactions in a blast furnace. The blast furnace gets its name from the fact that air or pure oxygen is blown into the furnace, where it oxidizes carbon to form carbon monoxide, CO. Carbon and its compounds are important reactants in this process.

What happens inside the blast furnace to recover the iron from its ore? The actual chemical changes that occur are complex. A simplified explanation begins with the reaction of oxygen in hot air with coke, a form of carbon. Some of the coke burns to form carbon dioxide.

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

As the concentration of oxygen is decreased, the carbon dioxide comes in contact with pieces of hot coke and is reduced to carbon monoxide.

$$CO_2(g) + C(s) \longrightarrow 2CO(g)$$

The carbon monoxide now acts as a reducing agent to reduce the iron oxides in the ore to metallic iron.

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

The reduction is thought to occur in steps as the temperature in the furnace increases. The following are some of the possible steps.

$$Fe_2O_3 \longrightarrow Fe_3O_4 \longrightarrow FeO \longrightarrow Fe$$

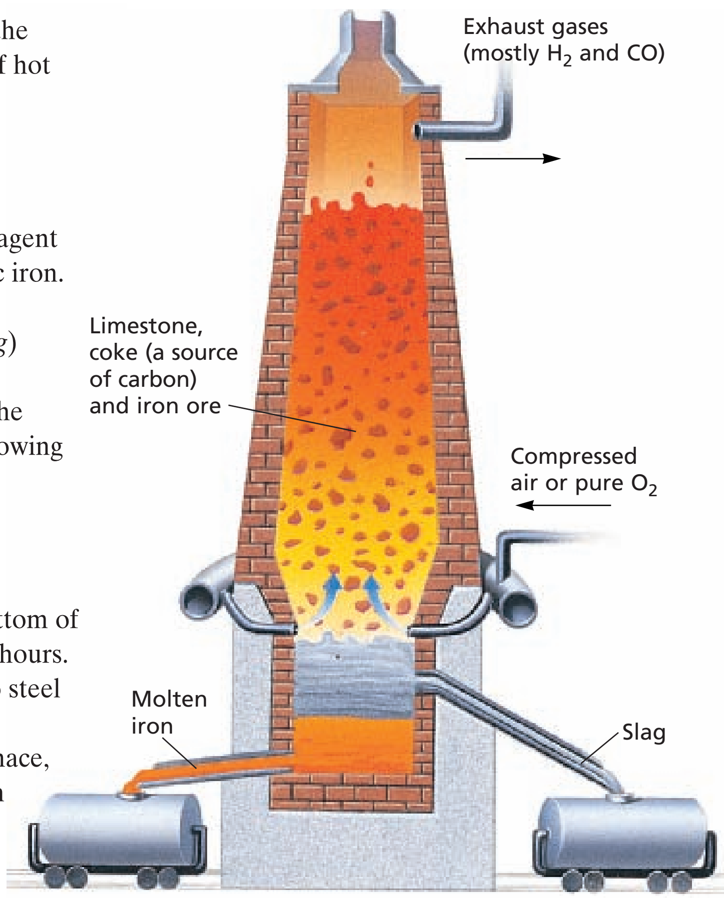
The white-hot liquid iron collects in the bottom of the furnace and is removed every four or five hours. The iron may be cast in molds or converted to steel in another process.

Limestone, present in the center of the furnace, decomposes to form calcium oxide and carbon dioxide.

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$$

The calcium oxide then combines with silica, a silicon compound, to form calcium silicate slag.

The relatively high carbon content of iron produced in a blast furnace makes the metal hard but brittle. It also has other impurities, like sulfur and phosphorus, that cause the recovered iron to be brittle. The conversion of iron to steel is essentially a purification process in which impurities are removed by oxidation. This purification process is carried out in another kind of furnace at very high temperatures. All steel contains 0.02 to 1.5% carbon. In fact, steels are graded by their carbon content. Low-carbon steels typically contain 0.02 to 0.3% carbon. Medium-carbon steels typically contain 0.03 to 0.7% carbon. High-carbon steels contain 0.7 to 1.5% carbon.



Molten iron flowing from the bottom of a blast furnace has been reduced from its ore through a series of reactions at high temperatures in different regions of the furnace.

Carbon Dioxide

Carbon dioxide is a colorless gas with a faintly irritating odor and a slightly sour taste. The sour taste is the result of the formation of carbonic acid when CO_2 dissolves in the water in saliva. It is a stable gas that does not burn or support combustion. At temperatures lower than 31°C and at pressures higher than 72.9 atm, CO_2 condenses to the liquid form. A phase diagram for CO_2 is found in the chapter review section of Chapter 10. At normal atmospheric pressure, solid CO_2 (dry ice) sublimes at –78.5°C. The linear arrangement of carbon dioxide molecules makes them nonpolar.

CO_2 is produced by the burning of organic fuels and from respiration processes in most living things. Most CO_2 released into the atmosphere is used by plants during photosynthesis. Recall that photosynthesis is the process by which green plants and some forms of algae and bacteria make food. During photosynthesis, CO_2 reacts with H_2O, using the energy from sunlight. The relationships among the various processes on Earth that convert carbon to carbon dioxide are summarized in the diagram of the carbon cycle, which is pictured below.

Carbon Monoxide

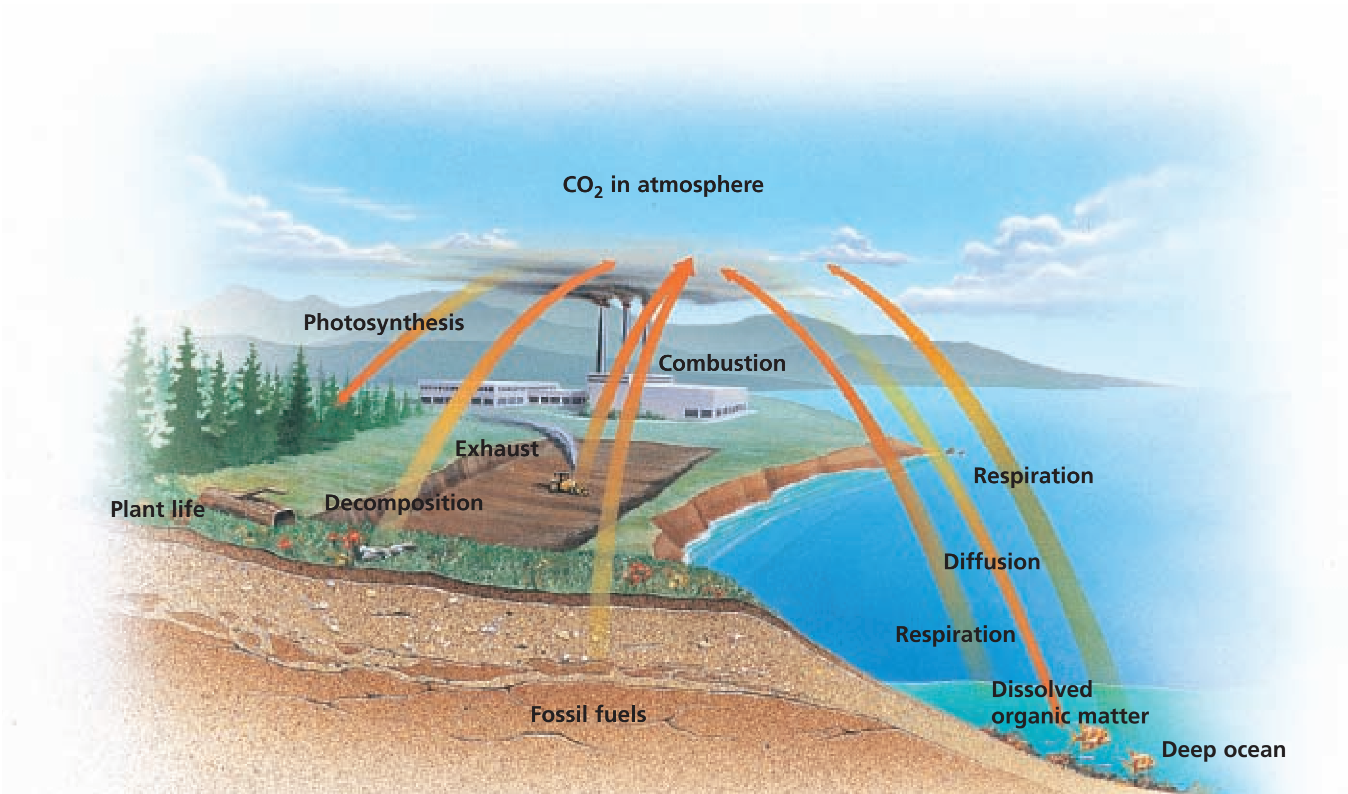
Carbon monoxide is a poisonous gas produced naturally by decaying plants, certain types of algae, volcanic eruptions, and the oxidation of methane in the atmosphere.

Because CO is colorless, odorless, and tasteless, it is difficult to detect. It is slightly less dense than air and slightly soluble in water. Its main chemical uses are in the reduction of iron, described on page 814, and the production of organic compounds, such as methanol.

$$CO(g) + 2H_2(g) \longrightarrow CH_3OH(l)$$

Carbon monoxide is also produced during the incomplete combustion of organic fuels. Incomplete combustion of methane occurs when the supply of oxygen is limited.

$$2CH_4(g) + 3O_2(g) \longrightarrow 2CO(g) + 4H_2O(g)$$



The processes that recycle CO_2 are collectively called the carbon cycle.

APPLICATION *Biochemistry*

Carbon Dioxide and Respiration

Many organisms, including humans, carry out cellular respiration. In this process, cells break down food molecules and release the energy used to build those molecules during photosynthesis. Glucose, $C_6H_{12}O_6$, is a common substance broken down in respiration. The following chemical equation expresses this process.

$$C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O + \text{energy}$$

In humans and most other vertebrate animals, the oxygen needed for this reaction is delivered to cells by hemoglobin found in red blood cells. Oxygen binds with hemoglobin as blood passes through capillaries in the lungs, as represented by the following reaction.

$$\text{Hb} + O_2 \longrightarrow \text{HbO}_2$$

Hb represents the hemoglobin molecule, and HbO_2 represents oxyhemoglobin, which is hemoglobin with bound oxygen. When the red blood cells pass through capillaries near cells that have depleted their oxygen supply through respiration, the reaction reverses and oxyhemoglobin gives up its oxygen.

$$\text{HbO}_2 \longrightarrow \text{Hb} + O_2$$

Carbon dioxide produced during respiration is a waste product that must be expelled from an organism. Various things happen when CO_2 enters the blood. Seven percent dissolves in the plasma, about 23% binds loosely to hemoglobin, and the remaining 70% reacts reversibly with water in plasma to form hydrogen carbonate, HCO_3^- ions. To form HCO_3^- ions, CO_2 first combines with H_2O to form carbonic acid, H_2CO_3, in a reversible reaction.

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

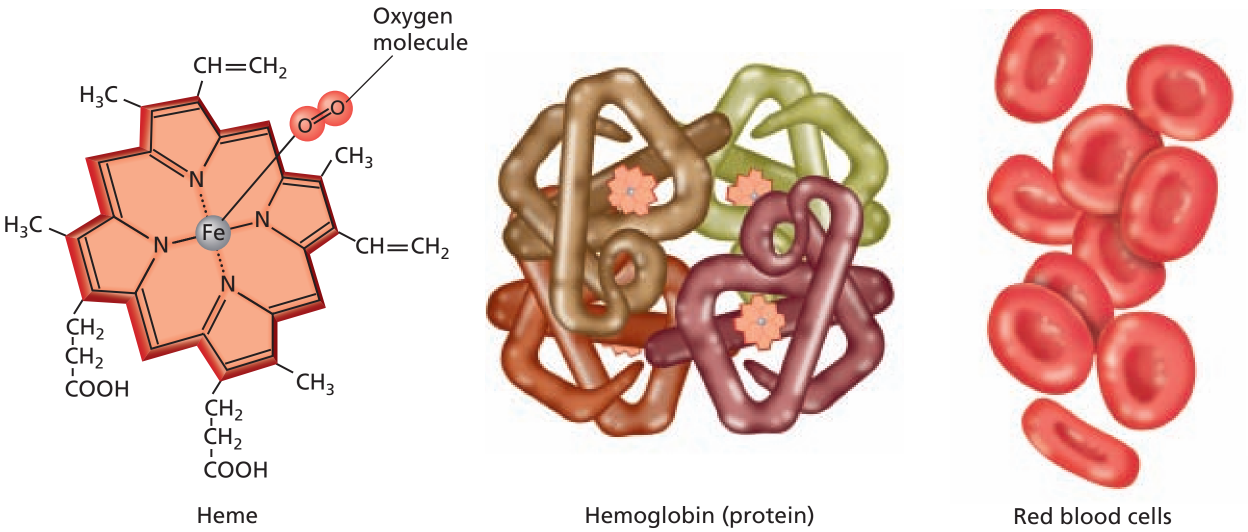
The dissolved carbonic acid ionizes to HCO_3^- ions and aqueous H^+ ions in the form of H_3O^+.

$$H_2CO_3(aq) + H_2O \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

The combined equilibrium reaction follows.

$$CO_2(aq) + 2H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

When the blood reaches the lungs, the reaction reverses and the blood releases CO_2, which is then exhaled to the surroundings.

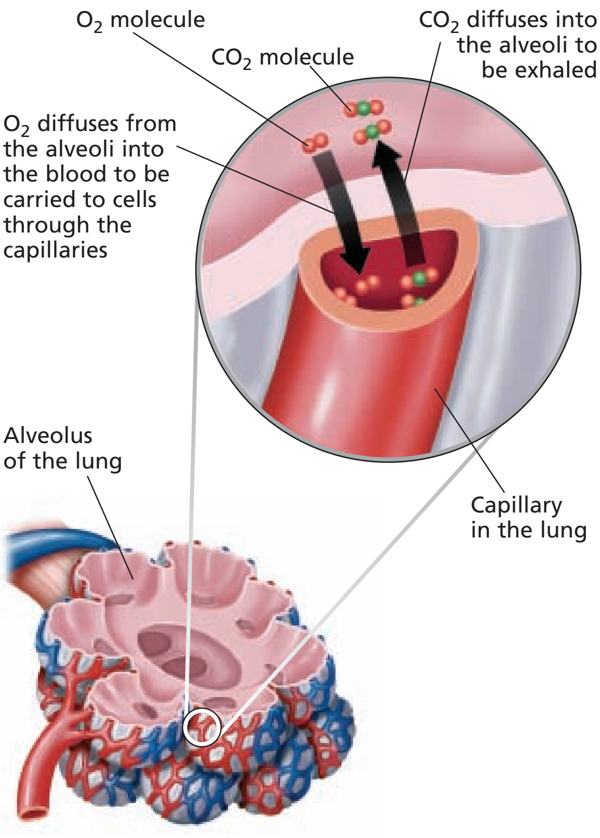


The oxygen carrier molecule, heme, is a component of the more-complex protein hemoglobin. Note that each hemoglobin molecule has four heme subunits. Hemoglobin is a component of red blood cells.

Exchange of CO_2 and O_2 in the Lungs

Why does CO_2 leave the blood as it passes through the lung's capillaries, and why does O_2 enter the blood? The exchange is caused by the difference in concentrations of CO_2 and O_2 in the blood and in the atmosphere. Oxygen is 21% of the atmosphere. Although the amount of CO_2 varies from place to place, it averages about 0.033% of the atmosphere. Thus, O_2 is about 640 times more concentrated in the atmosphere than is CO_2.

Substances tend to diffuse from regions of higher concentration toward regions of lower concentration. Thus, when blood reaches the capillaries of the lung, O_2 from the air diffuses into the blood, where its pressure is only 40 mm Hg, while CO_2 diffuses out of the blood, where its pressure is 45 mm Hg, and into the air. The diagram below summarizes the process.



The pressure of O_2 in the blood entering the lung is much lower than it is in the atmosphere. As a result, O_2 diffuses into the blood. The opposite situation exists for CO_2, so it diffuses from the blood into the air. Note that blood leaving the lung still contains a significant concentration of CO_2.

Acidosis and Alkalosis

In humans, blood is maintained between pH 7.3 and 7.5. The pH of blood is dependent on the concentration of CO_2 in the blood. Look again at this equilibrium system.

$$CO_2(aq) + 2H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

Notice that the right side of the equation contains the H_3O^+ ion, which determines the pH of the blood. If excess H_3O^+ enters the blood from tissues, the reverse reaction is favored. Excess H_3O^+ combines with HCO_3^- to produce more CO_2 and H_2O. If the H_3O^+ concentration begins to fall, the forward reaction is favored, producing additional H_3O^+ and HCO_3^-. To keep H_3O^+ in balance, adequate amounts of both CO_2 and HCO_3^- must be present. If something occurs that changes these conditions, a person can become very ill and can even die.

Hyperventilation occurs when a person breathes too rapidly for an extended time. Too much CO_2 is eliminated, causing the reverse reaction to be favored, and H_3O^+ and HCO_3^- are used up. As a result, the person develops a condition known as alkalosis because the pH of the blood rises to an abnormal alkaline level. The person begins to feel lightheaded and faint, and, unless treatment is provided, he or she may fall into a coma. Alkalosis is treated by having the victim breathe air that is rich in CO_2. One way to accomplish this is to have the person breathe with a bag held tightly over the nose and mouth. Alkalosis is also caused by fever, infection, intoxication, hysteria, and prolonged vomiting.

The reverse of alkalosis is a condition known as acidosis. This condition is often caused by a depletion of HCO_3^- ions from the blood, which can occur as a result of kidney dysfunction. The kidney controls the excretion of HCO_3^- ions. If there are too few HCO_3^- ions in solution, the forward reaction is favored and H_3O^+ ions accumulate, which lowers the blood's pH. Acidosis can also result from the body's inability to expel CO_2, which can occur during pneumonia, emphysema, and other respiratory disorders. Perhaps the single most common cause of acidosis is uncontrolled diabetes, in which acids normally excreted in the urinary system are instead retained by the body.

APPLICATION *The Environment*

Carbon Monoxide Poisoning

Standing on a street corner in any major city exposes a person to above-normal concentrations of carbon monoxide from automobile exhaust. Carbon monoxide also reacts with hemoglobin. The following reaction takes place in the capillaries of the lung.

$$Hb + CO \longrightarrow HbCO$$

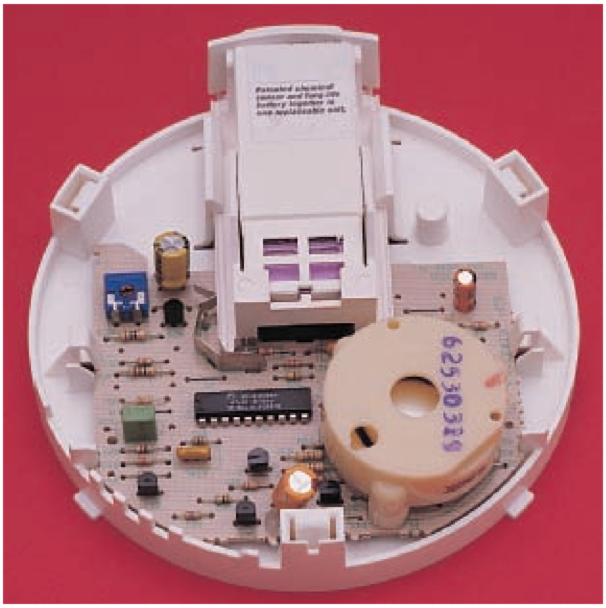
Unlike CO_2 or O_2, CO binds strongly to hemoglobin. Carboxyhemoglobin, HbCO, is 200 times more stable than oxyhemoglobin, HbO_2. So as blood circulates, more and more CO molecules bind to hemoglobin, reducing the amount of O_2 bond sites available for transport. Eventually, CO occupies so many hemoglobin binding sites that cells die from lack of oxygen. Symptoms of carbon monoxide poisoning include headache, mental confusion, dizziness, weakness, nausea, loss of muscular control, and decreased heart rate and respiratory rate. The victim loses consciousness and will die without treatment.

If the condition is caught in time, a victim of carbon monoxide poisoning can be revived by breathing pure oxygen. This treatment causes carboxyhemoglobin to be converted slowly to oxyhemoglobin according to the following chemical equation.

$$O_2 + HbCO \longrightarrow CO + HbO_2$$

Mild carbon monoxide poisoning usually does not have long-term effects. In severe cases, cells are destroyed. Damage to brain cells is irreversible.

The level of danger posed by carbon monoxide depends on two factors: the concentration of the gas in the air and the amount of time that a person is exposed to the gas. Table 5A shows the effects of increasing levels of carbon monoxide in the bloodstream. These effects vary considerably depending on a person's activity level and metabolic rate.



Carbon monoxide detectors are now available to reduce the risk of poisoning from defective home heating systems. The Consumer Products Safety Commission recommends that all homes have a CO detector with a UL label.

TABLE 5A Symptoms of CO Poisoning at Increasing Levels of CO Exposure and Concentration

Concentration of CO in air (ppm)*	Hemoglobin molecules as HbCO	Visible effects
100 for 1 hour or less	10% or less	no visible symptoms
500 for 1 hour or less	20%	mild to throbbing headache, some dizziness, impaired perception
500 for an extended period of time	30–50%	headache, confusion, nausea, dizziness, muscular weakness, fainting
1000 for 1 hour or less	50–80%	coma, convulsions, respiratory failure, death

* ppm is parts per million

APPLICATION *Biochemistry*

Macromolecules

Large organic polymers are called macromolecules (the prefix *macro* means “large”). Macromolecules play important roles in living systems. Most macromolecules essential to life belong to four main classes, three of which we know as nutrients in food:

1. **Proteins** Hair, tendons, ligaments, and silk are made of protein. Other proteins act as hormones, transport substances throughout the body, and fight infections. Enzymes are proteins that control the body’s chemical reactions. Proteins provide energy, yielding 17 kJ/g.
2. **Carbohydrates** Sugars, starches, and cellulose are carbohydrates. Carbohydrates are sources of energy, yielding 17 kJ/g.
3. **Lipids** Fats, oils, waxes, and steroids are lipids, nonpolar substances that do not dissolve in water. Fats are sources of energy, yielding 38 kJ/g.
4. **Nucleic acids** DNA and RNA are nucleic acids. In most organisms, DNA is used to store hereditary information and RNA helps to assemble proteins.

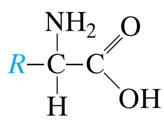
Proteins

Proteins are macromolecules formed by condensation reactions between amino acid monomers. Proteins contain carbon, oxygen, hydrogen, nitrogen, and usually some sulfur.

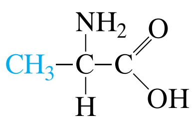
All amino acids have a carboxyl group, —COOH, and an amino group, —NH_2, attached to a central carbon atom, which is also attached to hydrogen, —H. Amino acids differ from one another at the fourth bond site of the central carbon, which is attached to a group of atoms (called an *R* group). *R* groups differ from one amino acid to another, as shown in the structures for several amino acids below. The proteins of all organisms contain a set of 20 common amino acids. The reaction that links amino acids is a condensation reaction, which is described in Chapter 22.

Each protein has its own unique sequence of amino acids. A complex organism has at least several thousand different proteins, each with a special structure and function. For instance, *insulin,* a hormone that helps the body regulate the level of sugar in the blood, is made up of two linked chains.

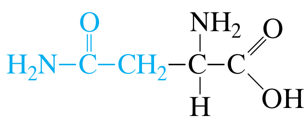
Amino acids have the same general structure. These examples show some of the variations within this class of compounds.



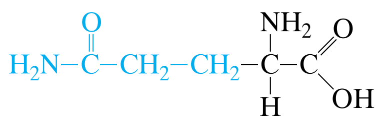
General structure



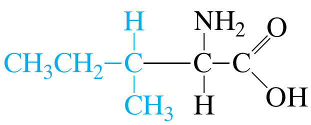
Alanine



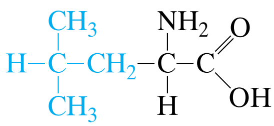
Asparagine



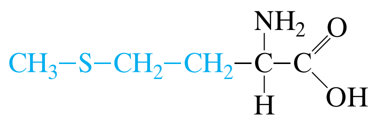
Glutamine



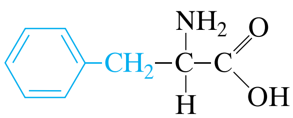
Isoleucine



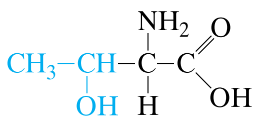
Leucine



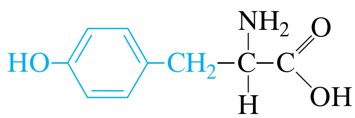
Methionine



Phenylalanine

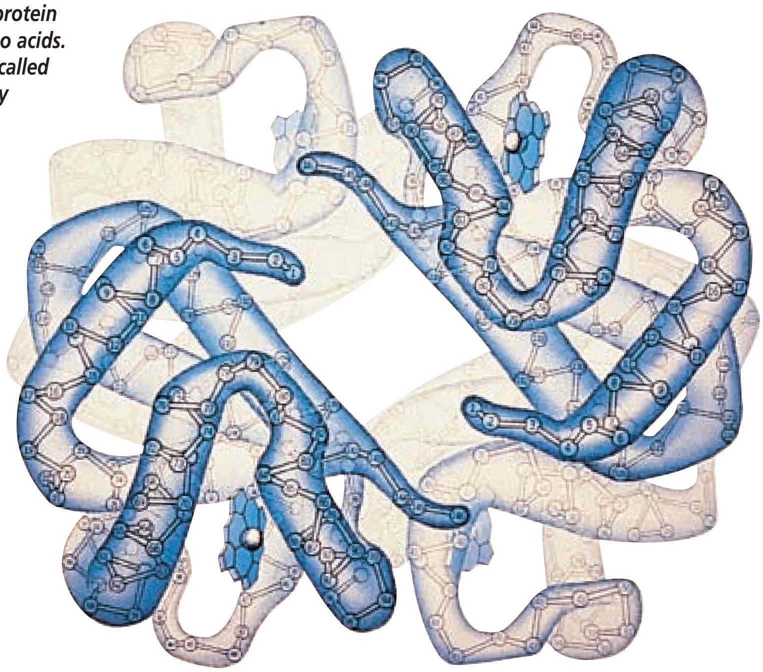


Threonine



Tyrosine

Hemoglobin is a complex protein made of hundreds of amino acids. Its 3-dimensional shape is called a tertiary structure. Tertiary structures break down when a protein is denatured.



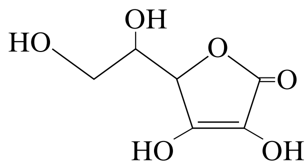
The chains are held together by S—S bonds between sulfur atoms in two cysteine amino acids. Insulin is one of the smaller proteins, containing only 51 amino acids. In contrast, hemoglobin, which carries oxygen in the blood, is a large protein consisting of four long chains with the complicated three-dimensional structures shown above. Proteins can lose their shape with increases in temperature or changes in the chemical composition of their environment. When they are returned to normal surroundings, they may fold or coil up again and re-form their original structure.

Changing even one amino acid can change a protein's structure and function. For example, the difference between normal hemoglobin and the hemoglobin that causes sickle cell anemia is just one amino acid substituted for another.

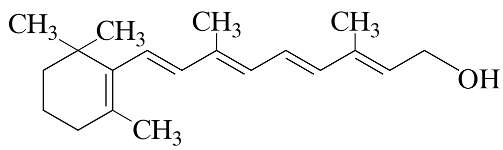
Enzymes

You learned how enzymes alter reaction rates in Chapter 17. Some enzymes cannot bind to their substrates without the help of additional molecules. These may be *minerals,* such as calcium or iron ions, or helper molecules called *coenzymes* that play accessory roles in enzyme-catalyzed reactions. Many vitamins are coenzymes or parts of coenzymes.

Vitamins are organic molecules that we cannot manufacture and hence need to eat in small amounts.



Vitamin C, $C_6H_8O_6$
Water-soluble



Vitamin A, $C_{20}H_{30}O$
Fat-soluble

You can see why we need vitamins and minerals in our diet—to enable our enzymes to work. You can also see why we need only small amounts of them. Minerals and coenzymes are not destroyed in biochemical reactions. Like enzymes, coenzymes and minerals can be used over and over again.

Temperature and pH have the most significant effects on the rates of reactions catalyzed by enzymes. Most enzymes work best in a solution of approximately neutral pH. Most body cells have a pH of 7.4. However, some enzymes function only in acidic or basic environments. For example, pepsin, the collective

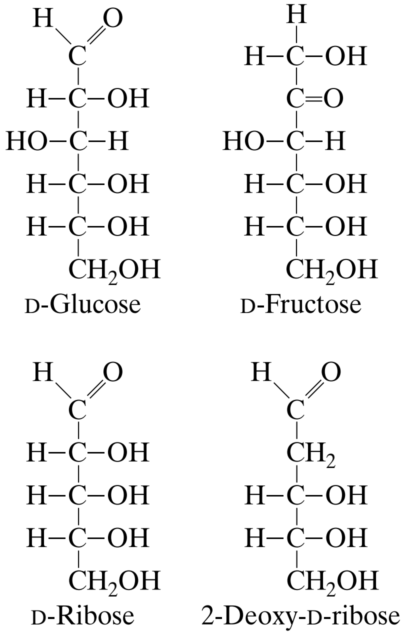
The protein in fish is denatured by the low pH of lime juice. Notice that the flesh shown with the limes has turned white compared with the flesh at normal pH.

term for the digestive enzymes found in the human stomach, works best at a very acidic pH of about 1.5. Cells that line the stomach secrete hydrochloric acid to produce this low pH environment. When food travels down the digestive tract, it carries these enzymes out of the stomach into the intestine. In the intestine, stomach enzymes stop working because sodium bicarbonate in the intestine raises the pH to about 8. Digestive enzymes in the intestine are formed by the pancreas and work best at pH 8.

Most chemical reactions, including enzyme reactions, speed up with increases in temperature. However, high temperatures (above about 60°C) destroy, or denature, protein by breaking up the three-dimensional structure. For example, the protein in an egg or a piece of meat denatures when the egg or meat is cooked. Proteins in the egg white become opaque when denatured. Heating can preserve food by denaturing the enzymes of organisms that cause decay. In milk pasteurization, the milk is heated to denature enzymes that would turn it sour. Refrigeration and freezing also help preserve food by slowing the enzyme reactions that cause decay.

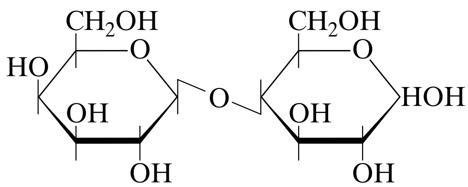
Carbohydrates

Carbohydrates are sugars, starches, and related compounds. The monomers of carbohydrates are monosaccharides, or simple sugars, such as fructose and glucose. A monosaccharide contains carbon, hydrogen, and oxygen in about a 1:2:1 ratio, which is an empirical formula of CH_2O.

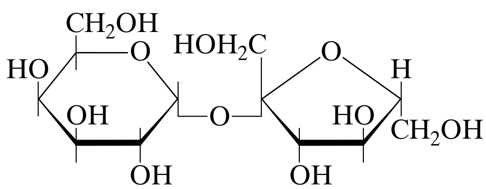


Monosaccharides chain representation

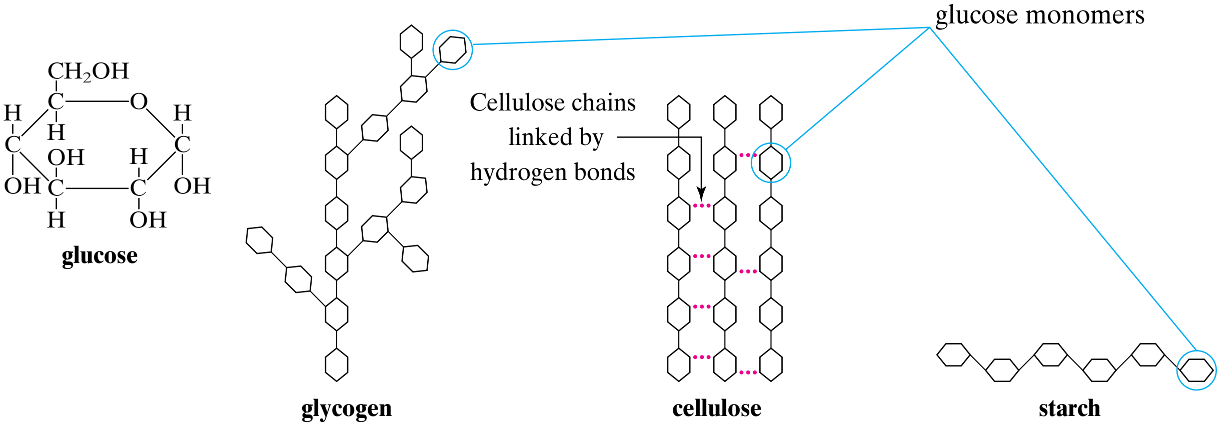
Two monosaccharides may be joined together to form a disaccharide. Sucrose, shown below, is a disaccharide. A disaccharide can be hydrolyzed to produce the monosaccharides that formed it. By a series of condensation reactions, many monosaccharides can be joined to form a polymer called a polysaccharide (commonly known as a complex carbohydrate).



Lactose—made from glucose and galactose



Sucrose—made from glucose and fructose



Glucose is the structural unit for glycogen, cellulose, and starch. Notice that these three polymers differ in the arrangement of glucose monomers.

Three important polysaccharides made of glucose monomers are glycogen, starch, and cellulose. Animals store energy in glycogen. The liver and muscles remove glucose from the blood and condense it into glycogen, which can later be hydrolyzed back into glucose and used to supply energy as needed.

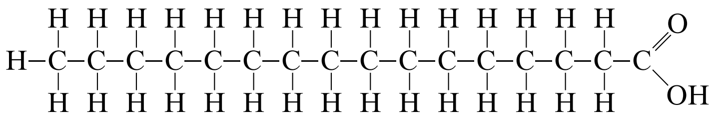
Starch consists of two kinds of glucose polymers. It is hydrolyzed in plants to form glucose for energy and for building material to produce more cells. The structural polysaccharide cellulose is probably the most common organic compound on Earth. Glucose monomers link cellulose chains together at the hydroxyl groups to form cellulose fibers. Cotton fibers consist almost entirely of cellulose.

Lipids

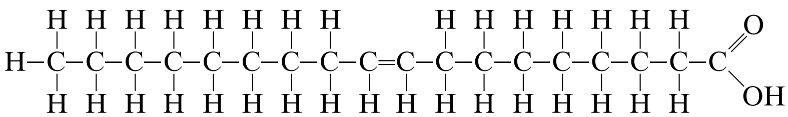
Lipids are a varied group of organic compounds that share one property: they are not very soluble in water. Lipids contain a high proportion of C—H bonds, and they dissolve in nonpolar organic solvents, such as ether, chloroform, and benzene.

Fatty acids are the simplest lipids. A fatty acid consists of an unbranched chain of carbon and hydrogen atoms with a carboxyl group at one end. Bonding within the carbon chain gives both saturated and unsaturated fatty acids, just as the simple hydrocarbons (see Chapter 22) can be saturated or unsaturated.

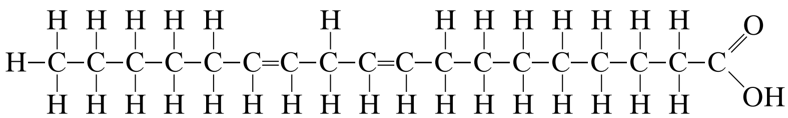
The bonds in a carboxyl group are polar, and so the carboxyl end of a fatty acid attracts water



Palmitic acid — saturated

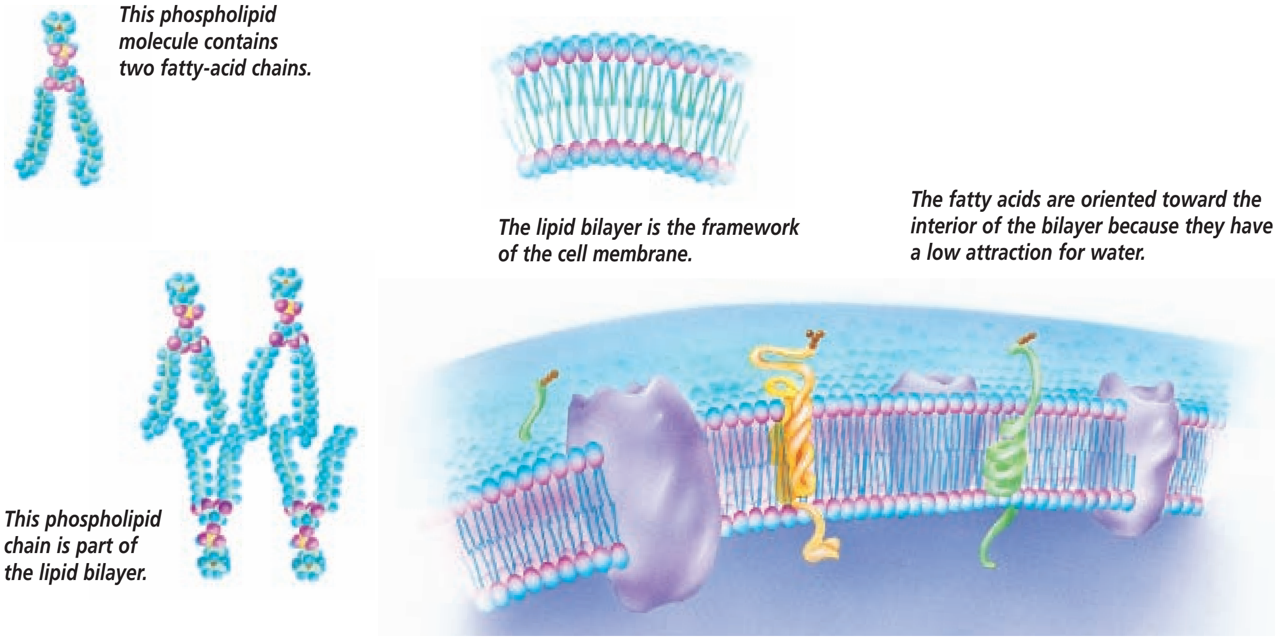


Oleic acid — monounsaturated



Linoleic acid — polyunsaturated

These examples of common fatty acids show the differences in saturation level.



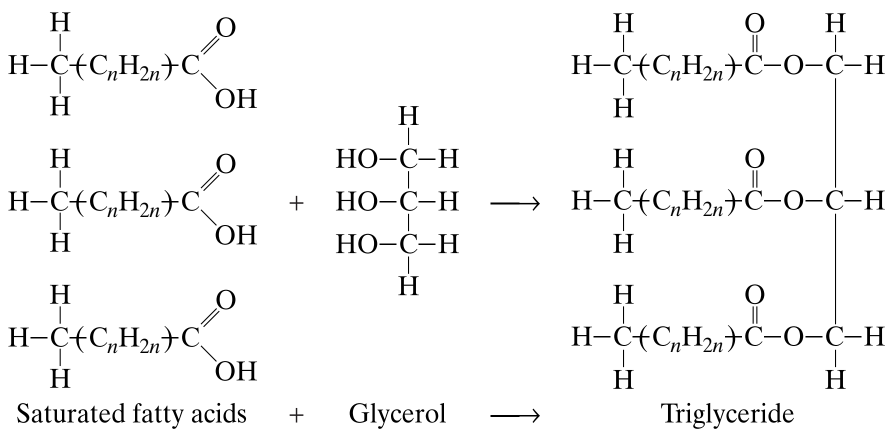
molecules. The carbon-hydrogen bonds of a lipid's hydrocarbon chain are nonpolar, however. The polar end will dissolve in water, and the other end will dissolve in nonpolar organic compounds. This behavior enables fatty acids to form membranes when they are dropped into water. It also gives soaps and detergents their cleaning power.

Lipids are the main compounds in biological membranes, such as the cell membrane. Because lipids are insoluble, the lipid bilayer of a cell membrane is adapted to keep the contents of the cell inside separated from the outer environment of the cell.

The structural component of a cell membrane is a phospholipid. The "head" of the phospholipid is polar, and the fatty acid tails are nonpolar, as shown in the model above.

Most fatty acids found in foods and soaps belong to a class of compounds called triglycerides. The fat content shown on a nutrition label for packaged food represents a mixture of the triglycerides in the food. Triglycerides have the general structure shown below.

Fatty acids are usually combined with other molecules to form classes of biomolecules called glycolipids (made from a carbohydrate and a lipid) or lipoproteins (made from a lipid and a protein). These compounds are also parts of more-complex lipids found in the body.



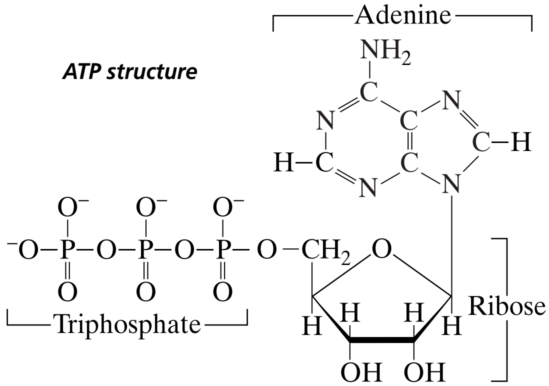
Triglycerides are made from three long-chain fatty acids bonded to a glycerol backbone.

Nucleic Acids

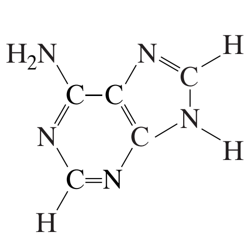
Nucleic acids are macromolecules that transmit genetic information. Deoxyribonucleic acid (DNA) is the material that contains the genetic information that all organisms pass on to their offspring during reproduction. This information includes instructions for making proteins as well as for making the other nucleic acid, ribonucleic acid (RNA). Ribonucleic acid (RNA) assists in protein synthesis by helping to coordinate the process of protein assembly.

Nucleotides are the monomers of nucleic acids. A nucleotide has three parts: one or more phosphate groups, a sugar containing five carbon atoms, and a ring-shaped nitrogen base, as shown below. RNA nucleotides contain the simple sugar ribose. DNA nucleotides contain deoxyribose (ribose stripped of one oxygen atom). Structures for both of these sugars are shown on page 821. Cells contain nucleotides with one, two, or three phosphate groups attached. Besides being the monomers of nucleic acids, several nucleotides play other roles. For example, adenosine triphosphate (ATP) is the nucleotide that supplies the energy for many metabolic reactions.

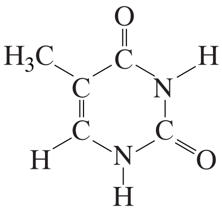
ATP structure



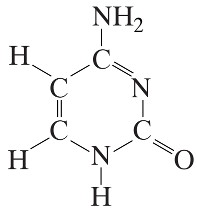
The bases in nucleic acids attract each other in pairs, a phenomenon known as base-pairing. DNA is made of four different nucleotides—those containing the bases adenine (A), thymine (T), guanine (G), and cytosine (C). The attraction between base pairs is hydrogen bonding. Adenine forms hydrogen bonds with thymine. Similarly, cytosine bonds to guanine. This base-pairing holds strands of DNA together.



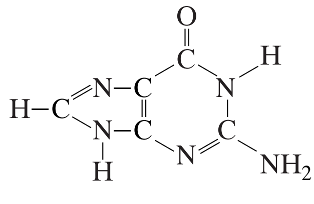
Adenine



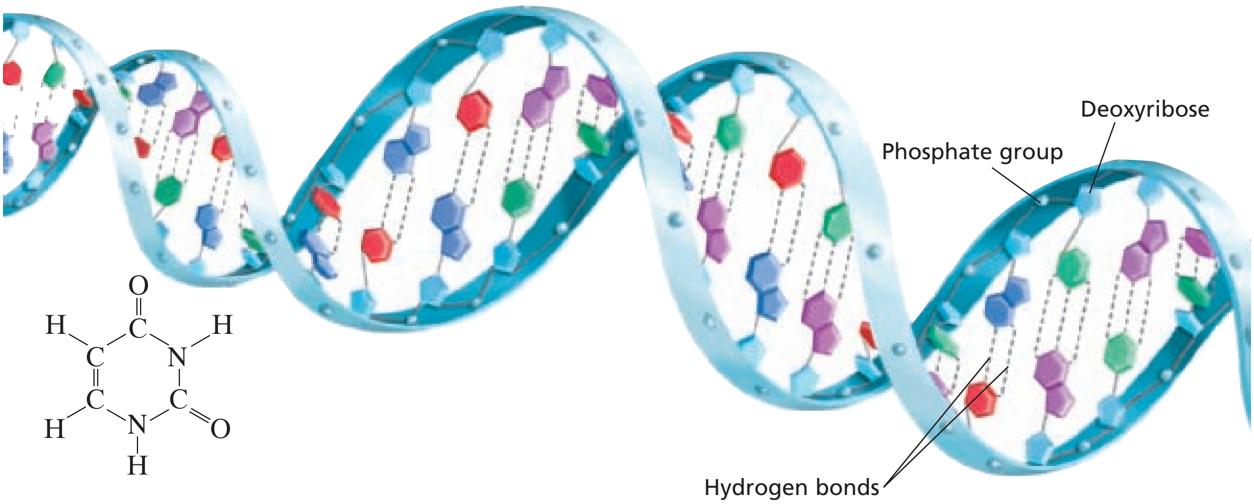
Thymine



Cytosine



Guanine



Uracil, RNA only

APPLICATION Chemical Industry

Silicon and Silicates

Silicon is as important in the mineral world as carbon is in living systems. Silicon dioxide and silicates make up about 87% of Earth's crust. Silicates are a class of compounds containing silicon, oxygen, one or more metals, and possibly hydrogen. Many mineral compounds are silicates. Sand is probably the most familiar silicate.

Glasses consist of 75% silicate. Borosilicate glass is the special heat-resistant glass used in making laboratory beakers and flasks. The addition of 5% boron oxide to the glass increases the softening temperature of the glass. Because boron and silicon atoms have roughly similar radii, these atoms can be substituted for one another to make borosilicate glass.

Asbestos is the name given to a class of fibrous magnesium silicate minerals. Asbestos is very strong and flexible, and it does not burn, so it was widely used as a heat-insulating material.

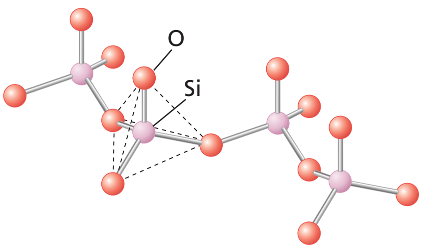
It is now known that asbestos is a carcinogen. When handled, asbestos releases dust particles that are easily inhaled and can cause lung cancer. Asbestos materials found in older homes and buildings should be removed by firms licensed by the Environmental Protection Agency (EPA).

Silicones

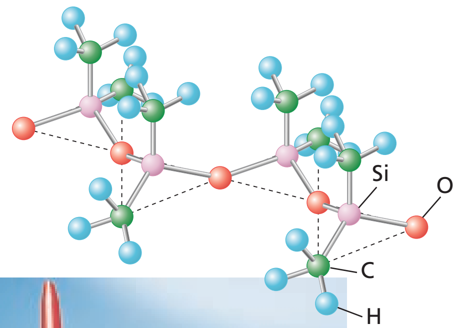
Silicones are a class of organic silicon polymers composed of silicon, carbon, oxygen, and hydrogen. The silicon chain is held together by bonding with the oxygen atoms. Each silicon atom is also bonded to different hydrocarbon groups to create a variety of silicone structures.

Silicones are widely used for their adhesive and protective properties. They have good electric insulating properties and are water-repellent. Some silicones have the character of oils or greases, so they are used as lubricants. Silicones are also used in automobile and furniture polishes as protective agents.

Silicon has the ability to form long chain compounds by bonding with oxygen. The SiO_4 subunit in this silicate is tetrahedral.



Silicones also have a tetrahedral structure. How does this structure differ from that of a silicate?



Silicates exist in a variety of mineral forms, including mica.

Because of their protective properties, silicones are used in a number of consumer products, from cosmetics to caulkings.

APPLICATION *Technology*

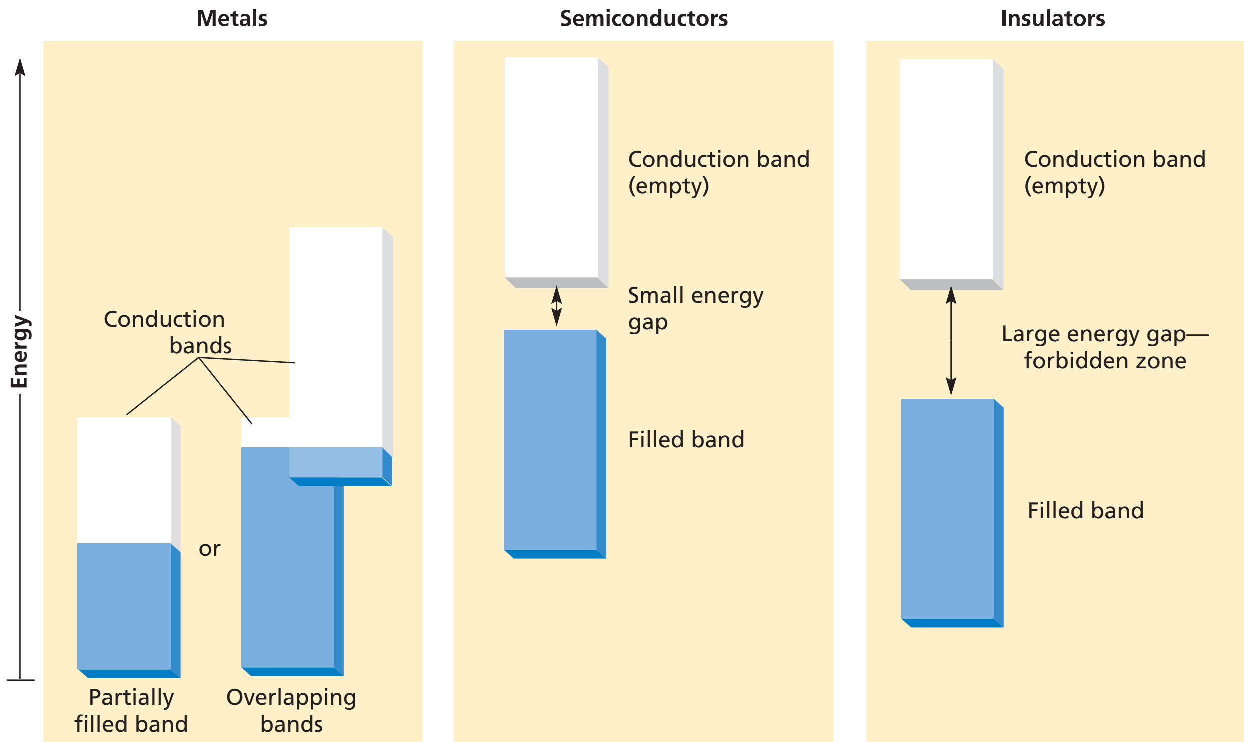
Semiconductors

When electrons can move freely through a material, the material is a conductor. The electrons in metals are loosely held and require little additional energy to move from one vacant orbital to the next. A set of overlapping orbitals is called a *conduction band*. Because electrons can easily jump to the conduction band, metals conduct electricity when only a very small voltage is applied.

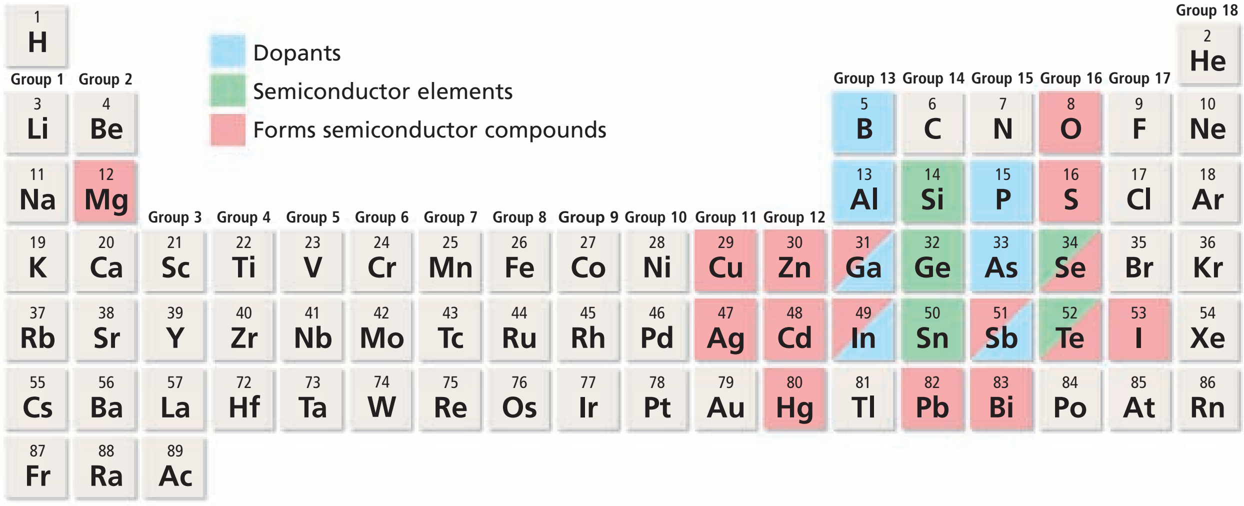
Semiconductors conduct a current if the voltage applied is large enough to excite the outer-level electrons of their atoms into the higher energy levels. With semiconductors, more energy, and thus a higher voltage, is required to cause conduction. By contrast, nonmetals are insulators because they do not conduct at ordinary voltages. Too much energy is needed to raise their outer electrons into conduction bands.

Semiconductor devices include transistors; diodes, including light-emitting diodes (LEDs); some lasers; and photovoltaic cells ("solar" cells). Though silicon is the basis of most semiconductor devices in the computer industry, pure silicon has little use as a semiconductor. Instead, small amounts of impurities are added to increase its conductive properties. Adding impurities to silicon is called *doping,* and the substances added are *dopants.* The dopant is usually incorporated into just the surface layer of a silicon chip. Typical dopants include the Group 15 elements phosphorus and arsenic and the Group 13 elements boron, aluminum, gallium, and indium.

A silicon atom has four electrons in its outer energy level whereas Group 13 atoms have three and Group 15 atoms have five. Adding boron to silicon creates a mix of atoms having four valence electrons and atoms having three valence electrons. Boron atoms form only three bonds with silicon, whereas silicon forms four bonds with other silicon atoms. The unbonded spot between a silicon atom



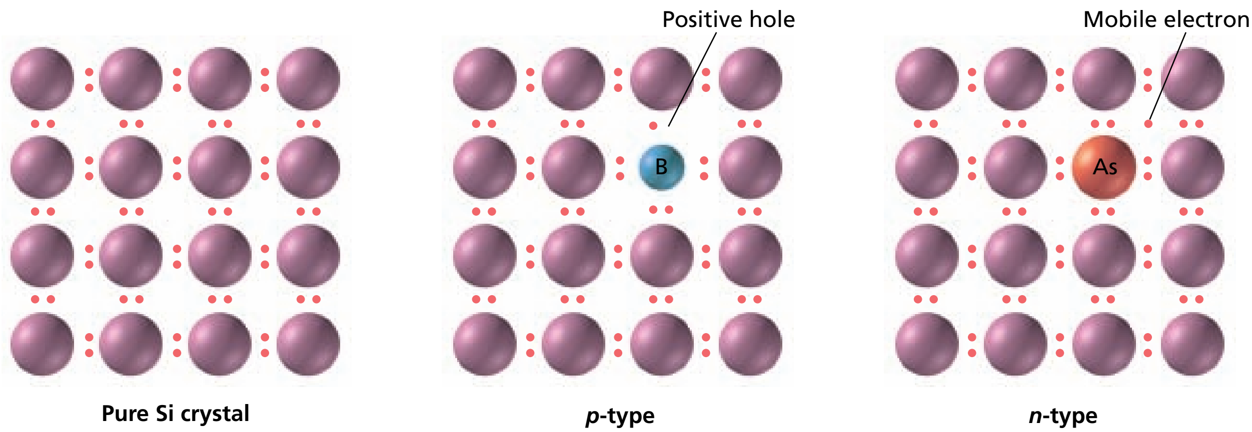
This model shows the difference in the levels of energy required to excite electrons into the conduction band in metals, semiconductors, and insulators. The forbidden zone is too great an energy gap in insulators for these elements to function as conductors. The energy gap for semiconductors is small enough that it can be crossed under certain conditions.



Semiconductor elements and dopants fall in the metalloid region of the periodic table. Semiconductor compounds often contain metals.

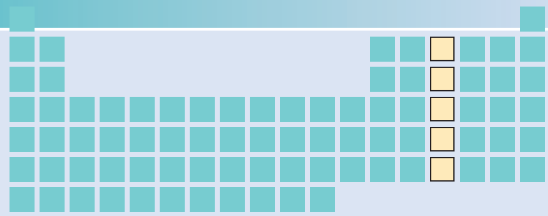
and a boron atom is a hole that a free electron can occupy. Because this hole "attracts" an electron, it is viewed as if it were positively charged. Semiconductors that are doped with boron, aluminum, or gallium are *p-type semiconductors,* the *p* standing for "positive." P-type semiconductors conduct electricity better than pure silicon because they provide spaces that moving electrons can occupy as they flow through the material.

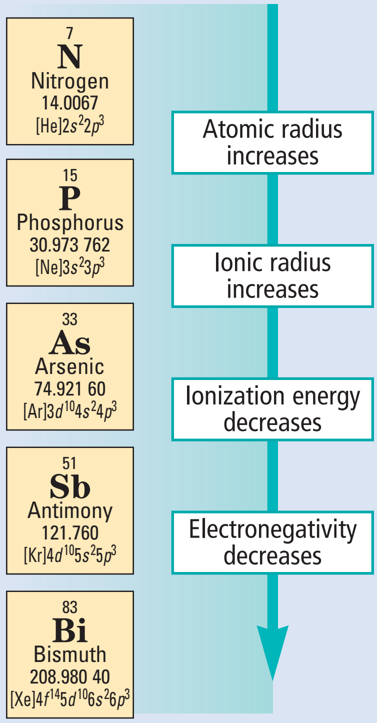
Doping silicon with phosphorus or arsenic produces the opposite effect. When phosphorus is added to silicon, it forms four bonds to silicon atoms and has a nonbonding electron left over. This extra electron is free to move through the material when a voltage is applied, thus increasing its conductivity compared with pure silicon. These extra electrons have a negative charge. Therefore, the material is an *n-type semiconductor*. Compare these two types of semiconductors in the models below.



Each silicon atom in the pure crystal is surrounded by four pairs of electrons. The **p*-type semiconductor model contains an atom of boron with a hole that an electron can occupy. The*** **n*-type semiconductor model contains an atom of arsenic, which provides the extra electron that can move through the crystal.***

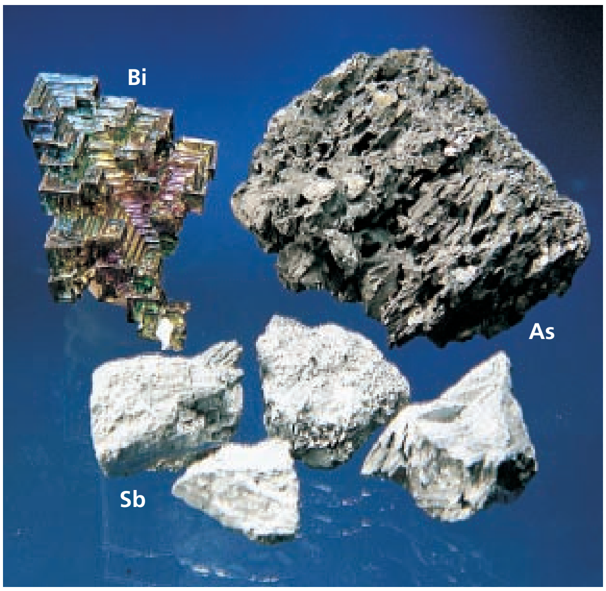
GROUP 15
NITROGEN FAMILY



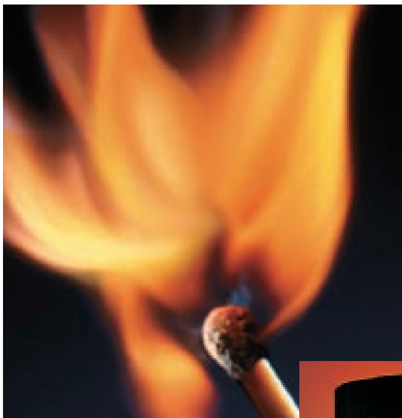


CHARACTERISTICS

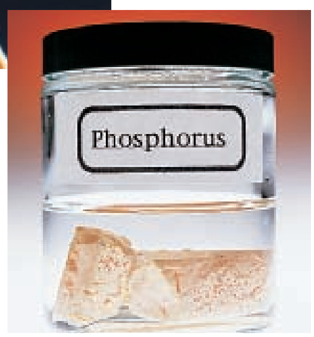
- consist of two nonmetals (nitrogen and phosphorus), two metalloids (arsenic and antimony), and one metal (bismuth)
- Nitrogen is most commonly found as atmospheric N_2; phosphorus as phosphate rock; and arsenic, antimony, and bismuth as sulfides or oxides. Antimony and bismuth are also found as free elements.
- range from very abundant elements (nitrogen and phosphorus) to relatively rare elements (arsenic, antimony, and bismuth)
- consist of atoms that contain five electrons in their outermost energy level
- tend to form covalent compounds, most commonly with oxidation numbers of +3 or +5
- exist in two or more allotropic forms, except nitrogen and bismuth
- are solids at room temperature, except nitrogen



You can see the contrast in physical properties among the elements of this family. Arsenic, antimony, and bismuth are shown.



Some matches contain phosphorus compounds in the match head. Safety matches contain phosphorus in the striking strip on the matchbox.



Phosphorus exists in three allotropic forms. White phosphorus must be kept underwater because it catches on fire when exposed to air. The red and black forms are stable in air.

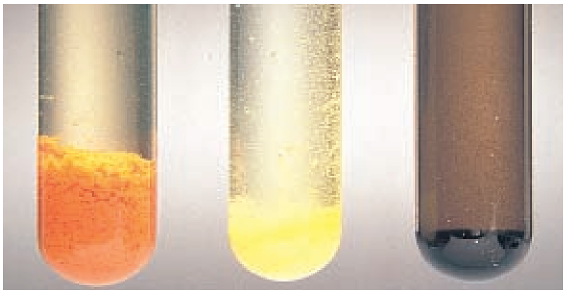
COMMON REACTIONS

With Oxygen to Form Oxides
Example: $P_4(s) + 5O_2(g) \longrightarrow P_4O_{10}(s)$
As, Sb, and Bi follow this reaction pattern, but as monatomic elements. N reacts to form NO and NO_2. It also reacts as N_2 to form N_2O_3 and N_2O_5.

With Metals to Form Binary Compounds
Example: $3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$

ANALYTICAL TEST

There are no simple analytical tests for the presence of nitrogen or phosphorus compounds in a sample. Antimony produces a pale green color in a flame test, and arsenic produces a light blue color. Arsenic, antimony, and bismuth are recognized in qualitative analyses by their characteristic sulfide colors.



Formation of sulfides is the confirmatory qualitative analysis test for the presence of bismuth, antimony, and arsenic.

Arsenic flame test

Antimony flame test

PROPERTIES OF THE GROUP 15 ELEMENTS

	N	P*	As	Sb	Bi
Melting point (°C)	–209.86	44.1	817 (28 atm)	630.5	271.3
Boiling point (°C)	–195.8	280	613 (sublimes)	1750	1560 ± 5
Density (g/cm^3)	1.25×10^{-3}	1.82	5.727	6.684	9.80
Ionization energy (kJ/mol)	1402	1012	947	834	703
Atomic radius (pm)	75	110	120	140	150
Ionic radius (pm)	146 (N^{3-})	212 (P^{3-})	—	76 (Sb^{3+})	103 (Bi^{3+})
Common oxidation number in compounds	–3, +3, +5	–3, +3, +5	+3, +5	+3, +5	+3
Crystal structure†	cubic (as a solid)	cubic	rhombohedral	hcp	rhombohedral
Hardness (Mohs' scale)	none (gas)	—	3.5	3.0	2.25

* Data given apply to white phosphorus.
† Crystal structures are for the most common allotropes.

APPLICATION *Biology*

Plants and Nitrogen

All organisms, including plants, require certain elements to survive and grow. These elements include carbon, hydrogen, oxygen, nitrogen, phosphorus, potassium, sulfur, and several other elements needed in small amounts. An organism needs nitrogen to synthesize structural proteins, enzymes, and the nucleic acids DNA and RNA.

Carbon, hydrogen, and oxygen are available to plants from carbon dioxide in the air and from water in both the air and the soil. Nitrogen is necessary for plants' survival. Although nitrogen gas, N_2, makes up 78% of air, most plants cannot take nitrogen out of the air and incorporate it into their cells, because the strong triple covalent bond in N_2 is not easily broken. Plants need nitrogen in the form of a compound that they can take in and use. The process of using atmospheric N_2 to make NH_3 is called *nitrogen fixation*. Several kinds of nitrogen-fixing bacteria live in the soil and in the root nodules of plants called legumes. Legumes obtain the nitrogen they need through a symbiotic relationship with nitrogen-fixing bacteria. Legumes include peas, beans, clover, alfalfa, and locust trees. The bacteria convert nitrogen into ammonia, NH_3, which is then absorbed by the host plants.

Because wheat, rice, corn, and potatoes cannot perform the same feat as legumes, these plants depend on nitrogen-fixing bacteria in the soil. Soil bacteria convert NH_3 into nitrate ions, NO_3^-, the form of nitrogen that can be absorbed and used by plants. These plants also often need nitrogen fertilizers to supplement the work of the bacteria. Besides supplying nitrogen, fertilizers are manufactured to contain phosphorus, potassium, and trace minerals.

Nitrogen-fixing bacteria, Rhizobium, *live in these small nodules that grow on the roots of soybeans.*

Soybeans are legumes that live in a symbiotic relationship with nitrogen-fixing bacteria.

APPLICATION *Chemical Industry*

Fertilizers

Fertilizers can supply nitrogen to plants in the form of ammonium sulfate, ammonium nitrate, and urea, all of which are made from NH_3. Now you know why there is such a demand for ammonia. Though some soils contain sufficient concentrations of phosphorus and potassium, most soils need additional nitrogen for adequate plant growth. Ammonia, ammonium nitrate, or urea can fill that need.

Most fertilizers contain all three major plant nutrients N, P, and K, and are called *complete fertilizers*. A typical complete fertilizer might contain ammonium nitrate or sodium nitrate to provide nitrogen. Calcium dihydrogen phosphate, $Ca(H_2PO_4)_2$, or the anhydrous form of phosphoric acid, P_4O_{10}, can provide phosphorus. Potassium chloride, KCl, or potassium oxide, K_2O, can provide potassium.

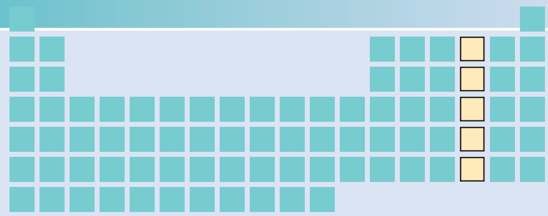
The proportion of each major nutrient in a fertilizer is indicated by a set of three numbers printed on the container. These numbers are the N-P-K formula of the fertilizer and indicate the percentage of N, P, and K, respectively. A fertilizer graded as 6-12-6, for example, contains 6% nitrogen, 12% phosphorus, and 6% potassium by weight.

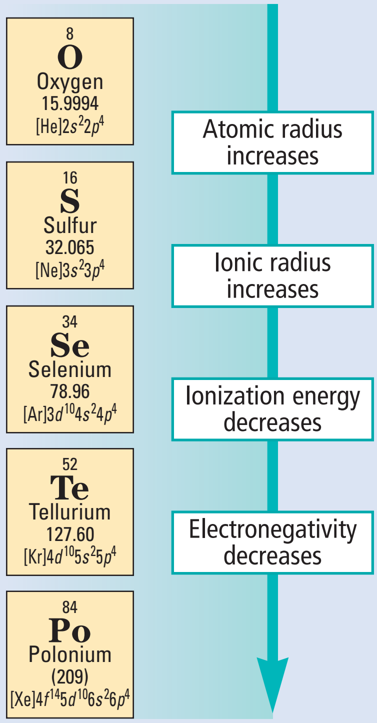
Nitrogen stimulates overall plant growth. Phosphorus promotes root growth and flowering. Potassium regulates the structures in leaves that allow CO_2 to enter the leaf and O_2 and H_2O to exit. Fertilizers are available in N-P-K formulas best suited for their intended use. For example, plants that produce large amounts of carbohydrates (sugars) need more potassium than most other types of plants. Grain crops need higher concentrations of phosphorus. Lawn fertilizers applied in the spring are generally high in nitrogen to stimulate shoot growth in grasses. Lawn fertilizers applied in the fall of the year should have a higher phosphorus content to stimulate root growth during the winter.

TABLE 6A Some Commercial Fertilizers and Uses

Fertilizer composition (N-P-K)	Uses
1-2-1 ratio 10-20-10 15-30-15	early-spring application for trees and shrubs with flowers and fruit; general-purpose feedings of the following: cucumbers, peppers, tomatoes
3-1-2 ratio 12-4-8 15-5-10 21-7-4 16-4-8 20-5-10	lawns and general-purpose feedings of the following: trees, shrubs, most berries, apple trees, grapes, vines, walnut trees, broccoli, cabbage, carrots, onions
High nitrogen 33-0-0 21-0-0 40-4-4 36-6-6	pecan trees, lawns, early feedings of corn
Balanced 13-13-13	general purpose feeding of the following: broccoli, cabbage, melons, potatoes
Special purpose: acid-loving flowering shrubs 12-10-4	azaleas, rhododendrons, camellias, gardenias
Special purpose 18-24-16	roses
Special purpose: flowering 12-55-6	flowering plants and shrubs (annuals and perennials)
Special purpose: root growth 5-20-10	starter fertilizer for transplants

GROUP 16
OXYGEN FAMILY

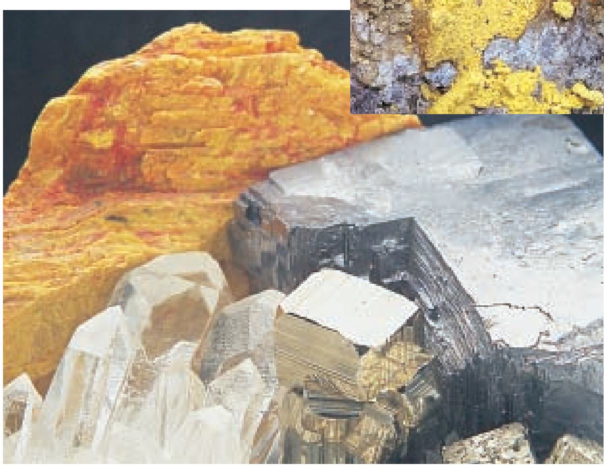




CHARACTERISTICS

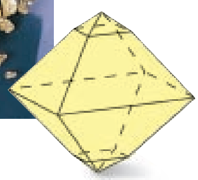
- occur naturally as free elements and in combined states
- consist of three nonmetals (oxygen, sulfur, and selenium), one metalloid (tellurium), and one metal (polonium)
- consist of atoms that have six electrons in their outermost energy level
- tend to form covalent compounds with other elements
- exist in several allotropic forms
- tend to exist as diatomic and polyatomic molecules, such as O_2, O_3, S_6, S_8, and Se_8
- commonly exist in compounds with the –2 oxidation state but often exhibit other oxidation states

Sulfur is found naturally in underground deposits and in the steam vents near volcanoes.

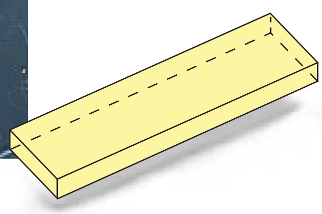


Sulfur exists in combined forms in many minerals. Iron pyrite, FeS_2, black galena, PbS, and yellow orpiment, As_2S_3, are shown.

Two allotropic forms of sulfur are orthorhombic and monoclinic. Each has a different crystal structure.



Orthorhombic



Monoclinic

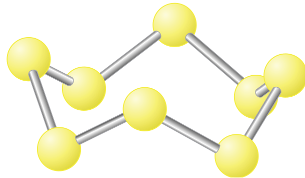
COMMON REACTIONS

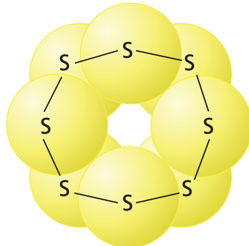
With Metals to Form Binary Compounds
Example: $8Mg(s) + S_8(l) \longrightarrow 8MgS(s)$
O_2, Se, and Te follow this pattern in reacting with Na, K, Ca, Mg, and Al.

With Oxygen to Form Oxides
Example: $Se(s) + O_2(g) \longrightarrow SeO_2(s)$
S, Te, and Po follow this pattern. S, Se, and Te can form SO_3, SeO_3, and TeO_3.

With Halogens to Form Binary Compounds
Example: $S_8(l) + 8Cl_2(g) \longrightarrow 8SCl_2(l)$
O, Se, Te, and Po follow this pattern in reacting with F_2, Cl_2, Br_2, and I_2.

With Hydrogen to Form Binary Compounds
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$



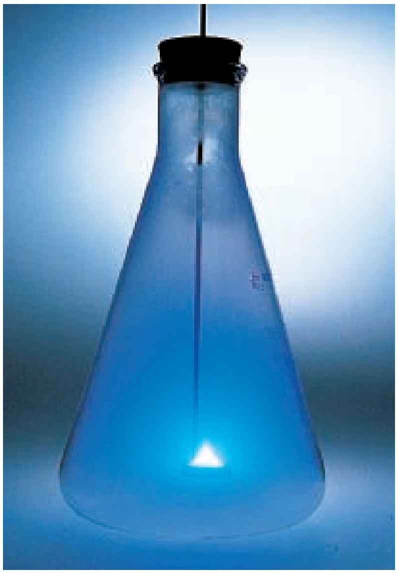


Sulfur exists as S_8 molecules in which the atoms are bonded in a ring, as shown by the ball-and-stick and space-filling models.

ANALYTICAL TEST

There is no simple analytical test to identify all elements of this family. Selenium and tellurium can be identified by flame tests. A light blue flame is characteristic of selenium, and a green flame is characteristic of tellurium. Oxygen can be identified by the splint test, in which a glowing splint bursts into flame when thrust into oxygen. Elemental sulfur is typically identified by its physical characteristics, especially its color and its properties when heated. It melts to form a viscous brown liquid and burns with a blue flame.

A glowing splint thrust into oxygen bursts into a bright flame.



Sulfur burns with a characteristically deep blue flame.

Molten sulfur returns to its orthorhombic form upon cooling.

PROPERTIES OF THE GROUP 16 ELEMENTS

	O	S	Se	Te	Po
Melting point (°C)	–218.4	119.0	217	449.8	254
Boiling point (°C)	–182.962	444.674	685	989.9	962
Density (g/cm^3)	1.429×10^{-3}	1.96	4.82	6.24	9.4
Ionization energy (kJ/mol)	1314	1000	941	869	812
Atomic radius (pm)	73	103	119	142	168
Ionic radius (pm)	140	184	198	221	—
Common oxidation number in compounds	–2	–2, +4, +6	–2, +2, +4, +6	–2, +2, +4, +6	–2, +2, +4, +6
Crystal structure*	orthorhombic, rhombohedral, cubic (when solid)	orthorhombic, monoclinic	hexagonal	hexagonal	cubic, rhombohedral
Hardness (Mohs' scale)	none (gas)	2.0	2.0	2.3	—

* Most elements of this family can have more than one crystal structure.

APPLICATION *Chemical Industry*

Oxides

Oxides of the reactive metals are ionic compounds. The oxide ion from any soluble oxide reacts immediately with water to form hydroxide ions as represented by the following equation.

$$O^{2-}(aq) + H_2O(l) \longrightarrow 2OH^-(aq)$$

The reactive metal oxides of Groups 1 and 2 react vigorously with water and release a large amount of energy as heat. The product of the reaction is a metal hydroxide. The following equation is an example of this reaction.

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

A basic oxide can be thought of as the dehydrated form of a hydroxide base. Oxides of the less reactive metals, such as magnesium, can be prepared by using thermal decomposition to drive off the water.

$$Mg(OH)_2(s) \xrightarrow{\text{heat}} MgO(s) + H_2O(g)$$

Hydroxides of the reactive metals of Group 1 are too stable to decompose in this manner.

If a hydroxide formed by a metal oxide is water-soluble, it dissolves to form a basic solution. An oxide that reacts with water to form a basic solution is called a basic oxide or a basic anhydride. Table 7A on the next page lists oxides that form basic solutions with water.

Molecular Oxides

Nonmetals, located on the right side of the periodic table, form molecular oxides. For example, sulfur forms two gaseous oxides: sulfur dioxide, SO_2, and sulfur trioxide, SO_3. In reactions typical of nonmetal oxides, each of the sulfur oxides reacts with water to form an oxyacid.

An oxide that reacts with water to form an acid is called an acidic oxide or an acid anhydride. As with the basic anhydrides, each acid anhydride can be thought of as the dehydrated form of the appropriate oxyacid. For example, when sulfuric acid decomposes, the loss of H_2O leaves the oxide SO_3, which is an anhydride.

$$H_2SO_4(aq) \xrightarrow{\text{heat}} H_2O(g) + SO_3(g)$$

Amphoteric Oxides

Table 7A lists some common oxides of main-group elements. You can see that the active metal oxides are basic and that the nonmetal oxides are acidic. Between these lies a group of oxides, the *amphoteric oxides*. The bonding in amphoteric oxides is intermediate between ionic and covalent bonding. As a result, oxides of this type show behavior intermediate between that of acidic oxides and basic oxides, and react as both acids and bases.

Aluminum oxide, Al_2O_3, is a typical amphoteric oxide. With hydrochloric acid, aluminum oxide acts as a base. The reaction produces a salt and water.

$$Al_2O_3(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2O(l)$$

With aqueous sodium hydroxide, aluminum oxide acts as an acid. The reaction forms a soluble ionic compound and water. That compound contains aluminate ions, AlO_2^-. (The AlO_2^- formula is used here rather than the more precise hydrated aluminate formula, $Al(OH)_4^-$.)

$$Al_2O_3(s) + 2NaOH(aq) \longrightarrow 2NaAlO_2(aq) + H_2O(l)$$

Reactions of Oxides

In the reaction between an acid and a metal oxide, the products are a salt and water—the same as the products in a neutralization reaction. For example, when magnesium oxide reacts with dilute sulfuric acid, magnesium sulfate and water are produced.

$$MgO(s) + H_2SO_4(dil.\ aq) \longrightarrow MgSO_4(aq) + H_2O(l)$$

The reaction between a basic metal oxide, such as MgO, and an acidic nonmetal oxide, such as CO_2, tends to produce an oxygen-containing salt. The dry oxides are mixed and heated without water. Salts such as metal carbonates, phosphates, and sulfates can be made by this synthesis reaction.

$$MgO(s) + CO_2(g) \longrightarrow MgCO_3(s)$$

$$6CaO(s) + P_4O_{10}(s) \longrightarrow 2Ca_3(PO_4)_2(s)$$

$$CaO(s) + SO_3(g) \longrightarrow CaSO_4(s)$$

Reactions of Hydroxides with Nonmetal Oxides

Nonmetal oxides tend to be acid anhydrides. The reaction of a hydroxide base with a nonmetal oxide is an acid-base reaction. The product is either a salt or a salt and water, depending on the identities and relative quantities of reactants. For example, 2 mol of the hydroxide base sodium hydroxide and 1 mol of the nonmetal oxide carbon dioxide form sodium carbonate, which is a salt, and water.

$$CO_2(g) + 2NaOH(aq) \longrightarrow Na_2CO_3(aq) + H_2O(l)$$

However, if sodium hydroxide is limited, only sodium hydrogen carbonate is produced.

$$CO_2(g) + NaOH(aq) \longrightarrow NaHCO_3(aq)$$

TABLE 7A Periodicity of Acidic and Basic Oxides of Main-Group Elements

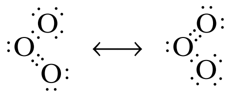
Group Number

1	2	13	14	15	16	17
Li_2O basic	BeO amphoteric	B_2O_3 acidic	CO_2 acidic	N_2O_5 acidic		
Na_2O basic	MgO basic	Al_2O_3 amphoteric	SiO_2 acidic	P_4O_{10} acidic	SO_3 acidic	Cl_2O acidic
K_2O basic	CaO basic	Ga_2O_3 amphoteric	GeO_2 amphoteric	As_4O_6 amphoteric	SeO_3 acidic	
Rb_2O basic	SrO basic	In_2O_3 basic	SnO_2 amphoteric	Sb_4O_6 amphoteric	TeO_3 acidic	I_2O_5 acidic
Cs_2O basic	BaO basic	Tl_2O_3 basic	PbO_2 amphoteric	Bi_2O_3 basic		

APPLICATION *The Environment*

Ozone

Ozone, O_3, is an allotrope of oxygen that is important for life on Earth. Like O_2, O_3 is a gas at room temperature. However, unlike O_2, O_3 is a poisonous bluish gas with an irritating odor at high concentrations. The triatomic ozone molecule is angular (bent) with a bond angle of about 116.5°. The O—O bonds in ozone are shorter and stronger than a single bond, but longer and weaker than a double bond. The ozone molecule is best represented by two resonance hybrid structures.



Ozone forms naturally in Earth's atmosphere more than 24 km above the Earth's surface in a layer called the stratosphere. There, O_2 molecules absorb energy from ultraviolet light and split into free oxygen atoms.

$$O_2(g) \xrightarrow{\text{ultraviolet light}} 2O$$

A free oxygen atom has an unpaired electron and is highly reactive. A chemical species that has one or more unpaired or unshared electrons is referred to as a *free radical*. A free radical is a short-lived fragment of a molecule. The oxygen free radical can react with a molecule of O_2 to produce an ozone molecule.

$$O + O_2(g) \longrightarrow O_3(g)$$

A molecule of O_3 can then absorb ultraviolet light and split to produce O_2 and a free oxygen atom.

$$O_3(g) \xrightarrow{\text{ultraviolet light}} O_2(g) + O$$

The production and breakdown of ozone in the stratosphere are examples of *photochemical* processes, in which light causes a chemical reaction.

In this way, O_3 is constantly formed and destroyed in the stratosphere, and its concentration is determined by the balance among these reactions. The breakdown of ozone absorbs the sun's intense ultraviolet light in the range of wavelengths between 290 nm and 320 nm. Light of these wavelengths damages and kills living cells, so if these wavelengths were to reach Earth's surface in large amounts, life would be impossible. Even now, the normal amount of ultraviolet light reaching Earth's surface is a major cause of skin cancer and the damage to DNA molecules that causes mutations. One life-form that is very sensitive to ultraviolet radiation is the phytoplankton in the oceans. These organisms carry out photosynthesis and are the first level of oceanic food webs.

Ozone and Air Pollution

Ozone in the lower atmosphere is a harmful pollutant. Ozone is highly reactive and can oxidize organic compounds. The products of these reactions are harmful substances that, when mixed with air, water vapor, and dust, make up *photochemical smog*. This mixture is the smog typically found in cities.

Typically, ozone is produced in a complex series of reactions involving unburned hydrocarbons and nitrogen oxides given off from engines in the form of exhaust and from fuel-burning power plants. When fuel burns explosively in the cylinder of an internal-combustion engine, some of the nitrogen in the cylinder also combines with oxygen to form NO, a very reactive nitrogen oxide free radical.

$$N_2(g) + O_2(g) \longrightarrow 2NO$$

When the free radical reaches the air, it reacts with oxygen to produce NO_2 radicals, which react with water in the air to produce HNO_3.

$$2NO + O_2(g) \longrightarrow 2NO_2$$
$$3NO_2 + H_2O(l) \longrightarrow NO + 2HNO_3(aq)$$

In sunlight, nitrogen dioxide decomposes to give nitric oxide and an atom of oxygen. Note that the NO produced is free to undergo the previous reaction once more.

$$NO_2 \xrightarrow{\text{sunlight}} NO + O$$

Just as it is in the stratosphere, a free oxygen atom in the lower atmosphere is highly reactive and reacts with a molecule of diatomic oxygen to form ozone.

$$O + O_2(g) \longrightarrow O_3(g)$$

APPLICATION *Chemical Industry*

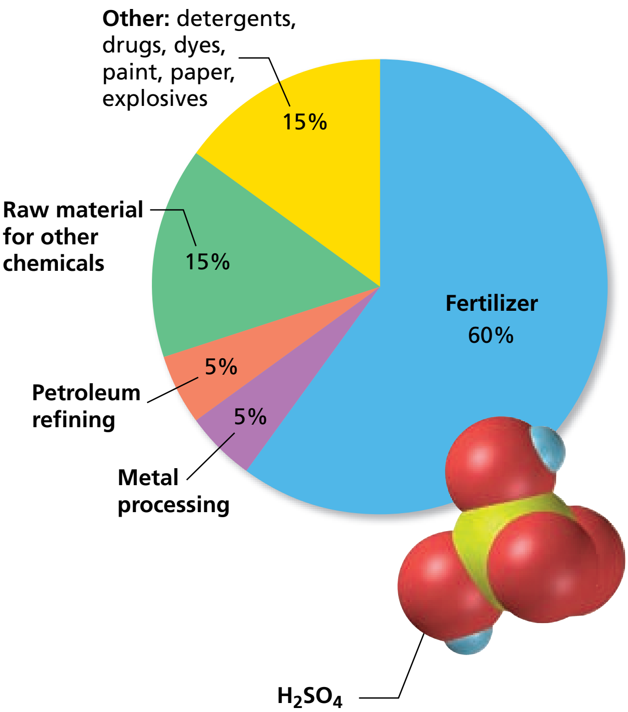
Sulfuric Acid

Sulfuric acid is the so-called "king of chemicals" because it is produced in the largest volume in the United States. It is produced by the contact process. This process starts with the production of SO_2 by burning sulfur or roasting iron pyrite, FeS_2. The purified sulfur dioxide is mixed with air and passed through hot iron pipes containing a catalyst. The contact between the catalyst, SO_2, and O_2 produces sulfur trioxide, SO_3, and gives the contact process its name. SO_3 is dissolved in concentrated H_2SO_4 to produce pyrosulfuric acid, $H_2S_2O_7$.

$$SO_3(g) + H_2SO_4(aq) \longrightarrow H_2S_2O_7(aq)$$

The pyrosulfuric acid is then diluted with water to produce sulfuric acid.

$$H_2S_2O_7(aq) + H_2O(l) \longrightarrow 2H_2SO_4(aq)$$



Important uses of the U.S. supply of sulfuric acid

Properties and Uses of Sulfuric Acid

Concentrated sulfuric acid is a good oxidizing agent. During the oxidation process, sulfur is reduced from +6 to +4 or −2. The change in oxidation state for a reaction depends on the concentration of the acid and on the nature of the reducing agent used in the reaction.

Sulfuric acid is also an important dehydrating agent. Gases that do not react with H_2SO_4 can be dried by being bubbled through concentrated sulfuric acid. Organic compounds, like sucrose, are dehydrated to leave carbon, as shown by the following reaction.

$$C_{12}H_{22}O_{11}(s) + 11H_2SO_4(aq) \longrightarrow 12C(s) + 11H_2SO_4{\cdot}H_2O(l)$$

The decomposition of sucrose proceeds rapidly, as shown in Figure 17 on page 737.

About 60% of the sulfuric acid produced in this country is used to make superphosphate, which is a mixture of phosphate compounds used in fertilizers.

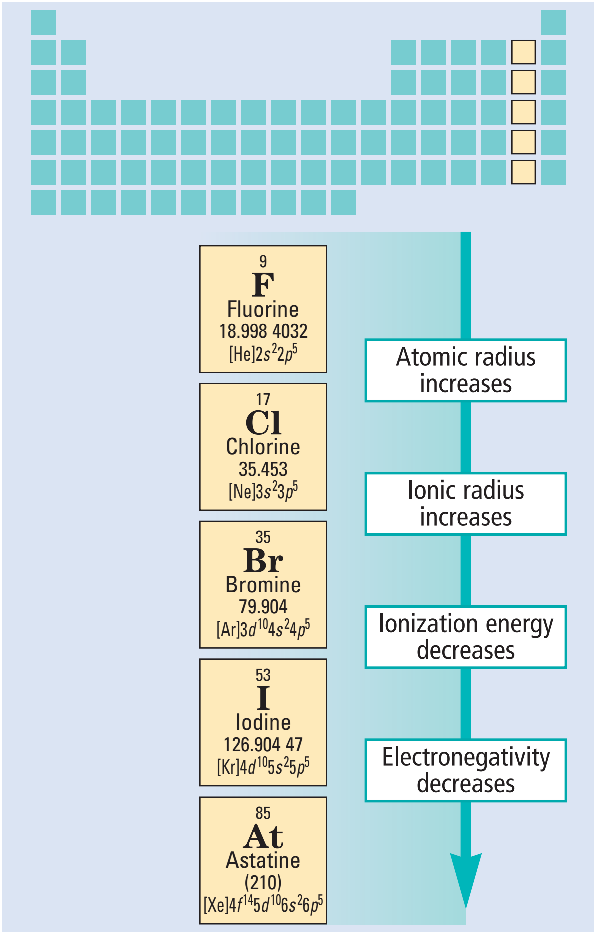
TABLE 7B Top Ten Chemicals Produced in the U.S.

Rank	Chemical	Physical state	Formula
1	sulfuric acid	*l*	H_2SO_4
2	nitrogen	*g*	N_2
3	oxygen	*g*	O_2
4	ethylene	*g*	C_2H_4
5	calcium oxide (lime)	*s*	CaO
6	ammonia	*g*	NH_3
7	phosphoric acid	*l*	H_3PO_4
8	sodium hydroxide	*s*	NaOH
9	propylene	*g*	C_3H_6
10	chlorine	*g*	Cl_2

GROUP 17
HALOGEN FAMILY

CHARACTERISTICS

- are all nonmetals and occur in combined form in nature, mainly as metal halides
- are found in the rocks of Earth's crust and dissolved in sea water
- range from fluorine, the 13th most abundant element, to astatine, which is one of the rarest elements
- exist at room temperature as a gas (F_2 and Cl_2), a liquid (Br_2), and a solid (I_2 and At)
- consist of atoms that have seven electrons in their outermost energy level
- tend to gain one electron to form a halide, X^- ion, but also share electrons and have positive oxidation states
- are reactive, with fluorine being the most reactive of all nonmetals



Halogens are the only family that contains elements representing all three states of matter at room temperature. Chlorine is a yellowish green gas; bromine is a reddish brown liquid; and iodine is a purple-black solid.

Iodine sublimes to produce a violet vapor that recrystallizes on the bottom of the evaporating dish filled with ice.

COMMON REACTIONS*

With Metals to Form Halides

Example: $Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$
Example: $Sn(s) + 2F_2(g) \longrightarrow SnF_4(s)$
The halide formula depends on the oxidation state of the metal.

With Hydrogen to Form Hydrogen Halides

Example: $H_2(g) + F_2(g) \rightarrow 2HF(g)$
Cl_2, Br_2, and I_2 also follow this pattern.

With Nonmetals and Metalloids to Form Halides

Example: $Si(s) + 2Cl_2(g) \longrightarrow SiCl_4(s)$
Example: $N_2(g) + 3F_2(g) \longrightarrow 2NF_3(g)$
Example: $P_4(s) + 6Br_2(l) \longrightarrow 4PBr_3(s)$
The formula of the halide depends on the oxidation state of the metalloid or nonmetal.

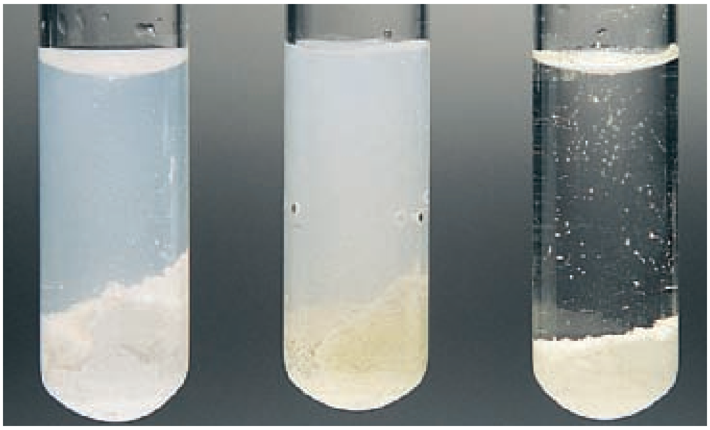
With Other Halogens to Form Interhalogen Compounds

Example: $Br_2(l) + 3F_2(g) \longrightarrow 2BrF_3(l)$

* Chemists assume that astatine undergoes similar reactions, but few chemical tests have been made.

Chlorine combines readily with iron wool, which ignites in chlorine gas to form $FeCl_3$.

Hydrofluoric acid is used to etch patterns into glass.



Shown here from left to right are precipitates of AgCl, AgBr, and AgI.

ANALYTICAL TEST

As with most elements, the presence of each of the halogens can be determined by atomic absorption spectroscopy. Fluorides react with concentrated sulfuric acid, H_2SO_4, to release hydrogen fluoride gas. Three of the halide ions can be identified in solution by their reactions with silver nitrate.

$$Cl^-(aq) + Ag^+(aq) \longrightarrow AgCl(s)$$
$$Br^-(aq) + Ag^+(aq) \longrightarrow AgBr(s)$$
$$I^-(aq) + Ag^+(aq) \longrightarrow AgI(s)$$

PROPERTIES OF THE GROUP 17 ELEMENTS

	F	Cl	Br	I	At
Melting point (°C)	–219.62	–100.98	–7.2	113.5	302
Boiling point (°C)	–188.14	–34.6	58.78	184.35	337
Density (g/cm^3)	1.69×10^{-3}	3.214×10^{-3}	3.119	4.93	not known
Ionization energy (kJ/mol)	1681	1251	1140	1008	—
Atomic radius (pm)	72	100	114	133	140
Ionic radius (pm)	133	181	196	220	—
Common oxidation number in compounds	–1	–1, +1, +3, +5, +7	–1, +1, +3, +5, +7	–1, +1, +3, +5, +7	–1, +5
Crystal structure	cubic	orthorhombic	orthorhombic	orthorhombic	not known

APPLICATION *The Environment*

Chlorine in Water Treatment

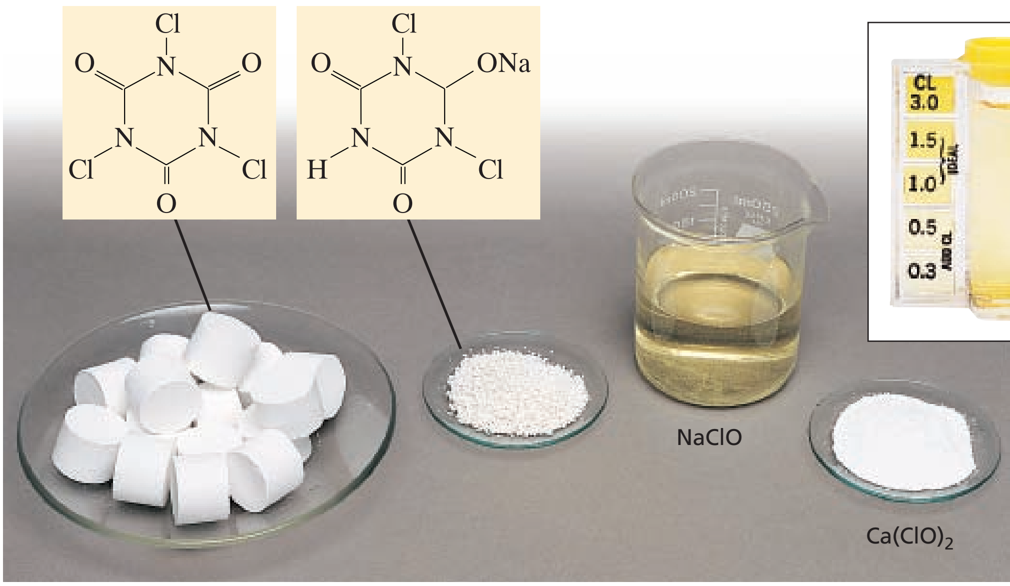
For more than a century, communities have treated their water to prevent disease. A treatment process widely used in the United States is chlorination. All halogens kill bacteria and other microorganisms. Chlorine, however, is the only halogen acceptable for large-scale treatment of public water supplies.

When chlorine is added to water, the following reaction produces HCl and hypochlorous acid, HOCl.

$$Cl_2(g) + H_2O(l) \longrightarrow HCl(aq) + HOCl(aq)$$

Hypochlorous acid is a weak acid that ionizes to give hydrogen ions and hypochlorite ions, OCl^-.

$$HOCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + OCl^-(aq)$$



Swimming pools are routinely tested to be sure the chlorine level is safe.

The "chlorine" used in swimming pools is really the compounds shown above and not chlorine at all.

The OCl^- ions are strong oxidizing agents that can destroy microorganisms.

In some water-treatment plants, calcium hypochlorite, $Ca(ClO)_2$, a salt of hypochlorous acid, is added to water to provide OCl^- ions. Similar treatments are used in swimming pools.

Nearly a hundred cities in the United States and thousands of communities in Europe use chlorine in the form of chlorine dioxide, ClO_2, as their primary means of disinfecting water. The main drawback to the use of ClO_2 is that it is unstable and cannot be stored. Instead, ClO_2 must be prepared on location by one of the following reactions involving sodium chlorite, $NaClO_2$.

$$10NaClO_2(aq) + 5H_2SO_4(aq) \longrightarrow 8ClO_2(g) + 5Na_2SO_4(aq) + 2HCl(aq) + 4H_2O(l)$$

$$2NaClO_2(aq) + Cl_2(g) \longrightarrow 2ClO_2(g) + 2NaCl(aq)$$

The expense of using ClO_2 makes it less desirable than Cl_2 in water-treatment systems unless there are other considerations. For example, the use of ClO_2 is likely to result in purified water with less of the aftertaste and odor associated with water purified by Cl_2.

Fluoride and Tooth Decay

In the 1940s, scientists noticed that people living in communities that have natural water supplies with high concentrations of fluoride ions, F^-, have significantly lower rates of dental caries (tooth decay) than most of the population.

In June 1944, a study on the effects of water fluoridation began in two Michigan cities, Muskegon and Grand Rapids, where the natural level of fluoride in drinking water was low (about 0.05 ppm). In Grand Rapids, sodium fluoride, NaF, was added to the drinking water to raise levels to 1.0 ppm. In Muskegon, no fluoride was added. Also included in the study was Aurora, Illinois, a city that was similar to Grand Rapids and Muskegon, except that it had a natural F^- concentration of 1.2 ppm in the water supply. After 10 years, the rate of tooth decay in Grand Rapids had dropped far below that in Muskegon and was about the same as it was in Aurora.

Tooth enamel is made of a strong, rocklike material consisting mostly of calcium hydroxyphosphate, $Ca_5(PO_4)_3(OH)$, also known as apatite. Apatite is an insoluble and very hard compound—ideal for tooth enamel. Sometimes, however, saliva becomes more acidic, particularly after a person eats a high-sugar meal. Acids ionize to produce hydronium ions, which react with the hydroxide ion, OH^-, in the apatite to form water. The loss of OH^- causes the apatite to dissolve.

$$Ca_5(PO_4)_3(OH)(s) + H_3O^+(aq) \longrightarrow 5Ca^{2+}(aq) + 3PO_4^{3-}(aq) + 2H_2O(l)$$

Saliva supplies more OH^- ions, and new apatite is formed, but slowly.

If fluoride ions are present in saliva, some fluorapatite, $Ca_5(PO_4)_3F$, also forms.

$$5Ca^{2+}(aq) + 3PO_4^{3-}(aq) + F^-(aq) \longrightarrow Ca_5(PO_4)_3F(s)$$

Fluorapatite resists attack by acids, so the tooth enamel resists decay better than enamel containing no fluoride.

When the beneficial effect of fluoride had been established, public health authorities proposed that fluoride compounds be added to water supplies in low-fluoride communities. Fluoridation started in the 1950s, and by 1965, nearly every medical and dental association in the United States had endorsed fluoridation of water supplies. In the past decade, however, that trend slowed as opposition to fluoridation grew.

Preparing for Chemistry Lab

Performing experiments in the chemistry laboratory provides you with the opportunity to learn important lab techniques and observe interesting chemical reactions. Taking the time to prepare for your lab activity will help ensure that you understand the procedures you are to follow and that your experiment runs smoothly and safely. You can prepare for each lab activity by reviewing the lab preparation tips below.

- Read the experiment thoroughly, and familiarize yourself with what you will do at each step in the procedure. If there is any part of the experiment that you do not understand, ask your teacher before you start the experiment.
- Do any assigned pre-lab exercises. These will generally cover any calculations or important observations that you will make.
- Prepare data tables ahead of time. If you have the data table ready before you begin, you will be able to record your observations as they happen, in the appropriate spaces.
- Review the materials list for the experiment. Make sure that you have all the items you need to perform the experiment. If you are missing any items or if any items are broken or unusable, let your teacher know before you begin the lab.
- Review all safety guidelines and safety icons at the beginning of each experiment. Read “Safety in the Chemistry Laboratory” in the front of your book for a complete explanation of the safety rules that you should follow in the lab.
- Know where the emergency eyewash station, safety shower, and fire extinguisher are located, and be sure you know how to use them. If you cannot locate the lab safety equipment or you do not know how to use any equipment, ask your teacher for help.
- Know the proper disposal procedures for the experiment. Your teacher will tell you what to do with any substances that need to be disposed of.

Pre-Laboratory Procedures

The Pre-Laboratory Procedures help you develop solid laboratory skills before you do an actual experiment.

Extraction and Filtration

Extraction, the separation of substances in a mixture by using a solvent, depends on solubility. For example, sand can be separated from salt by adding water to the mixture. The salt dissolves in the water, and the sand settles to the bottom of the container. The sand can be recovered by decanting the water. The salt can then be recovered by evaporating the water.

Filtration separates substances based on differences in their physical states or in the size of their particles. For example, a liquid can be separated from a solid by pouring the mixture through a paper-lined funnel, or if the solid is denser than the liquid, the solid will settle to the bottom of the container, which will leave the liquid on top. The liquid can then be decanted, which will leave the solid.

SETTLING AND DECANTING

1. Fill an appropriate-sized beaker with the solid-liquid mixture provided by your teacher. Allow the beaker to sit until the bottom is covered with solid particles and the liquid is clear.
2. Grasp the beaker with one hand. With the other hand, pick up a stirring rod and hold it along the lip of the beaker. Tilt the beaker slightly so that liquid begins to pour out in a slow, steady stream, as shown in Figure A.

FIGURE A
Settling and decanting

GRAVITY FILTRATION

1. Prepare a piece of filter paper as shown in Figure B. Fold it in half and then in half again. Tear the corner of the filter paper, and open the filter paper into a cone. Place it in the funnel.

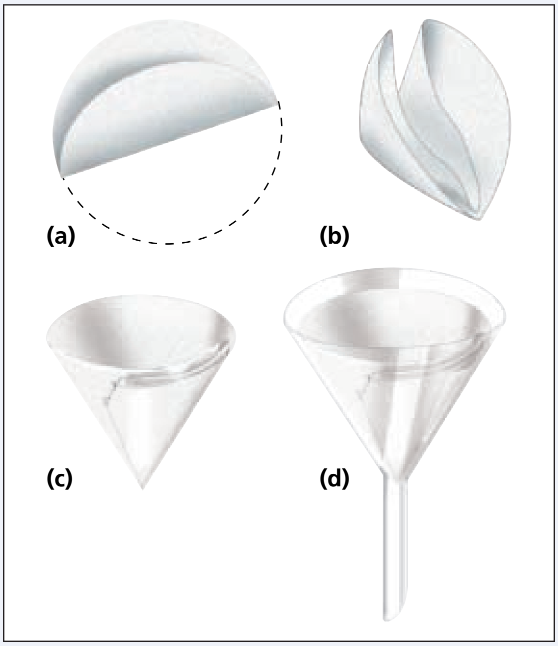


FIGURE B Folding the filter paper

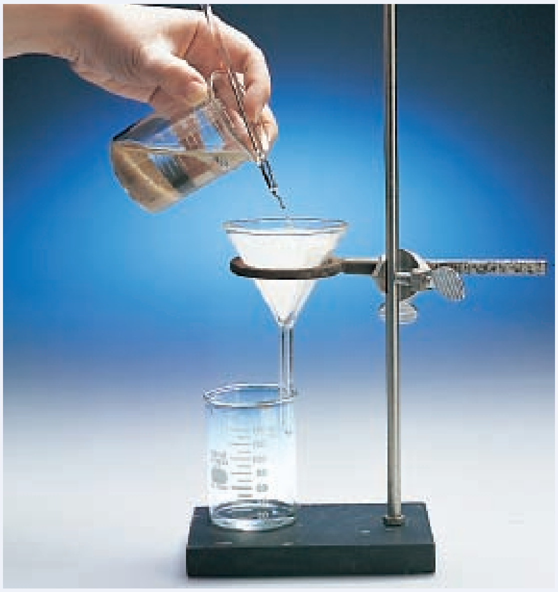


FIGURE C Gravity filtration

2. Put the funnel, stem first, into a filtration flask, or suspend it over a beaker by using an iron ring, as shown in Figure C.
3. Wet the filter paper with distilled water from a wash bottle. The paper should adhere to the sides of the funnel, and the torn corner should prevent air pockets from forming between the paper and the funnel.
4. Pour the mixture to be filtered down a stirring rod into the filter. The stirring rod directs the mixture into the funnel and reduces splashing.
5. Do not let the level of the mixture in the funnel rise above the edge of the filter paper.
6. Use a wash bottle to rinse all of the mixture from the beaker into the funnel.

VACUUM FILTRATION

1. Check the T attachment to the faucet. Turn on the water. Water should run without overflowing the sink or spitting while creating a vacuum. To test for a vacuum, cover the opening of the horizontal arm of the T with your thumb or index finger. If you feel your thumb being pulled inward, there is a vacuum. Note the number of turns of the knob that are needed to produce the flow of water that creates a vacuum.
2. Turn the water off. Attach the pressurized rubber tubing to the *horizontal* arm of the T. (You do not want water to run through the tubing.)
3. Attach the free end of the rubber tubing to the side arm of a filter flask. Check for a vacuum. Turn on the water so that it rushes out of the faucet (refer to step 1). Place the palm of your hand over the opening of the Erlenmeyer flask. You should feel the vacuum pull your hand inward. If you do not feel any pull or if the pull is weak, increase the flow of water. If increasing the flow of water fails to work, shut off the water and make sure your tubing connections are tight.
4. Insert the neck of a Büchner funnel into a one-hole rubber stopper until the stopper is about two-thirds to three-fourths up the neck of the funnel. Place the funnel stem into the Erlenmeyer flask so that the stopper rests in the mouth of the flask, as shown in Figure D.
5. Obtain a piece of round filter paper. Place it inside the Büchner funnel over the holes. Turn on the water as in step 1. Hold the filter flask with one hand, place the palm of your hand over the mouth of the funnel, and check for a vacuum.
6. Pour the mixture to be filtered into the funnel. Use a wash bottle to rinse all of the mixture from the beaker into the funnel.

FIGURE D Vacuum filtration

Gravimetric Analysis

Gravimetric analytical methods are based on accurate and precise mass measurements. They are used to determine the amount or percentage of a compound or element in a sample material. For example, if we want to determine the percentage of iron in an ore or the percentage of chloride ion in drinking water, gravimetric analysis would be used.

A gravimetric procedure generally involves reacting the sample to produce a reaction product that can be used to calculate the mass of the element or compound in the original sample. For example, to calculate the percentage of iron in a sample of iron ore, determine the mass of the ore. The ore is then dissolved in hydrochloric acid to produce $FeCl_3$. The $FeCl_3$ precipitate is converted to a hydrated form of Fe_2O_3 by adding water and ammonia to the system. The mixture is then filtered to separate the hydrated Fe_2O_3 from the mixture. The hydrated Fe_2O_3 is heated in a crucible to drive the water from the hydrate, producing anhydrous Fe_2O_3. The mass of the crucible and its contents is determined after successive heating steps to ensure that the product has reached constant mass and that all of the water has been driven off. The mass of Fe_2O_3 produced can be used to calculate the mass and percentage of iron in the original ore sample.

Gravimetric procedures require accurate and precise techniques and measurements to obtain suitable results. Possible sources of error are the following:

1. **The product (precipitate) that is formed is contaminated.**
2. **Some product is lost when transferring the product from a filter to a crucible.**
3. **The empty crucible is not clean or is not at constant mass for the initial mass measurement.**
4. **The system is not heated sufficiently to obtain an anhydrous product.**

GENERAL SAFETY


Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the location of the emergency lab shower and eyewash station and the procedure for using them.

When using a Bunsen burner, confine long hair and loose clothing. Do not heat glassware that is broken, chipped, or cracked. Use tongs or a hot mitt to handle heated glassware and other equipment; heated glassware does not always look hot. If your clothing catches fire, WALK to the emergency lab shower and use it to put out the fire.

Never put broken glass or ceramics in a regular waste container. Broken glass and ceramics should be disposed of in a separate container designated by your teacher.

SETTING UP THE EQUIPMENT

1. The general setup for heating a sample in a crucible is shown in Figure A. Attach a metal ring clamp to a ring stand, and lay a clay triangle on the ring.

CLEANING THE CRUCIBLE

2. Wash and dry a metal or ceramic crucible and lid. Cover the crucible with its lid, and use a balance to obtain its mass. If the balance is located far from your working station, use crucible tongs to place the crucible and lid on a piece of wire gauze. Carry the crucible to the balance, using the wire gauze as a tray.

HEATING THE CRUCIBLE TO OBTAIN A CONSTANT MASS

3. After recording the mass of the crucible and lid, suspend the crucible over a Bunsen burner by placing it on the clay triangle, as shown in Figure B. Then, place the lid on the crucible so that the entire contents are covered.

4. Light the Bunsen burner. Heat the crucible for 5 min with a gentle flame, and then adjust the burner to produce a strong flame. Heat for 5 min more. Shut off the gas to the burner. Allow the crucible and lid to cool. Using crucible tongs, carry the crucible and lid to the balance, as shown in Figure C. Measure and record the mass. If the mass differs from the mass before heating, repeat the process until mass data from heating trials are within 1% of each other. This measurement assumes that the crucible has a constant mass. The crucible is now ready to be used in a gravimetric analysis procedure. Details will be found in the following experiments.

FIGURE A

FIGURE B

FIGURE C

Gravimetric methods are used in Experiment 7 to synthesize magnesium oxide and in Experiment 9 to separate $SrCO_3$ from a solution.

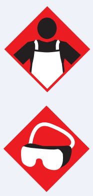
PRE-LABORATORY PROCEDURE

Paper Chromatography

Chromatography is a technique used to separate substances dissolved in a mixture. The Latin roots of the word are *chromato,* which means "color," and *graphy,* which means "to write." Paper is one medium used to separate the components of a solution.

Paper is made of cellulose fibers that are pressed together. As a solution passes over the fibers and through the pores, the paper acts as a filter and separates the mixture's components. Particles of the same component group together, producing a colored band. Properties such as particle size, molecular mass, and charge of the different solute particles in the mixture affect the distance the components will travel on the paper. The components of the mixture that are the most soluble in the solvent and the least attracted to the paper will travel the farthest. Their band of color will be closest to the edge of the paper.

GENERAL SAFETY



Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the location of the emergency lab shower and eyewash station and the procedure for using them.

PROCEDURE

1. Use a lead pencil to sketch a circle about the size of a quarter in the center of a piece of circular filter paper that is 12 cm in diameter.
2. Write one numeral for each substance, including any unknowns, around the inside of this circle. In this experiment, six mixtures are to be separated, so the circle is labeled "1" through "6," as shown in Figure A.
3. Use a micropipet to place a spot of each substance to be separated next to a number. Make one spot per number. If the spot is too large, you will get a broad, tailing trace with little or no detectable separation.
4. Use the pencil to poke a small hole in the center of the spotted filter paper. Insert a wick through the hole. A wick can be made by rolling a triangular piece of filter paper into a cylinder: start

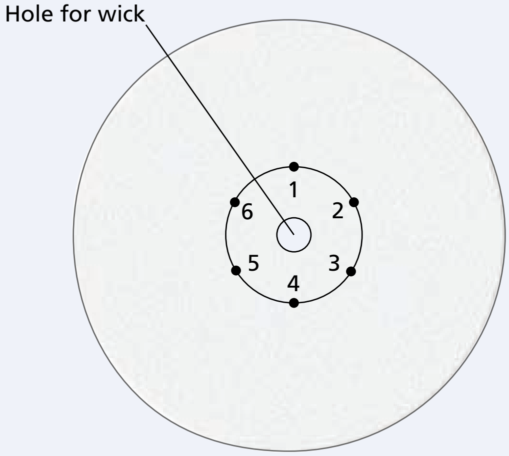


FIGURE A Filter paper used in paper chromatography is spotted with the mixtures to be separated. Each spot is labeled with a numeral or a name that identifies the mixture to be separated. A hole punched in the center of the paper will attach to a wick that delivers the solvent to the paper.

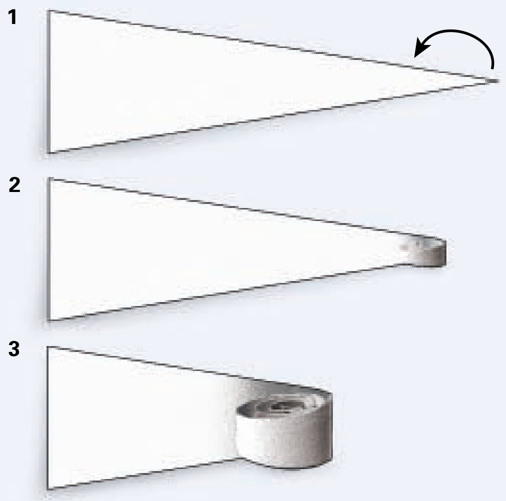


FIGURE B Cut the triangle from filter paper. Roll the paper into a cylinder, starting at the narrow end.

at the point of the triangle, and roll toward its base. See Figure B.

5. Fill a Petri dish or lid two-thirds full of solvent (usually water or alcohol).

6. Set the bottom of the wick in the solvent so that the filter paper rests on the top of the Petri dish. See Figure C.

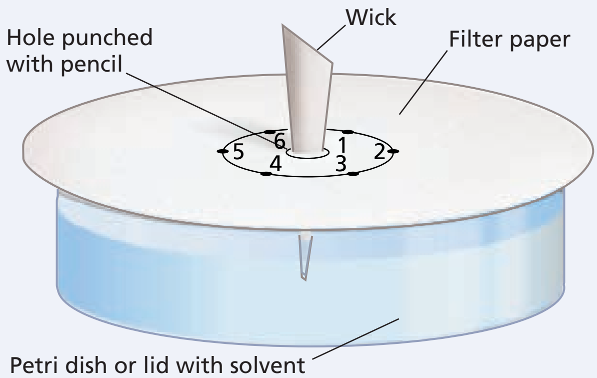


FIGURE C The wick is inserted through the hole of the spotted filter paper. The filter paper with the wick is then placed on top of a Petri dish or lid filled two-thirds full of water or another solvent.

7. When the solvent is 1 cm from the outside edge of the paper, remove the paper from the Petri dish and allow the chromatogram to dry. Sample chromatograms are shown in Figures D and E.

Most writing or drawing inks are mixtures of various components that give them specific properties. Therefore, paper chromatography can be used to study the composition of an ink. Experiment 12 investigates the composition of ball-point-pen ink.

FIGURE D Each of the original spots has migrated along with the solvent toward the outer edge of the filter paper. For each substance that was a mixture, you should see more than one distinct spot of color in its trace.

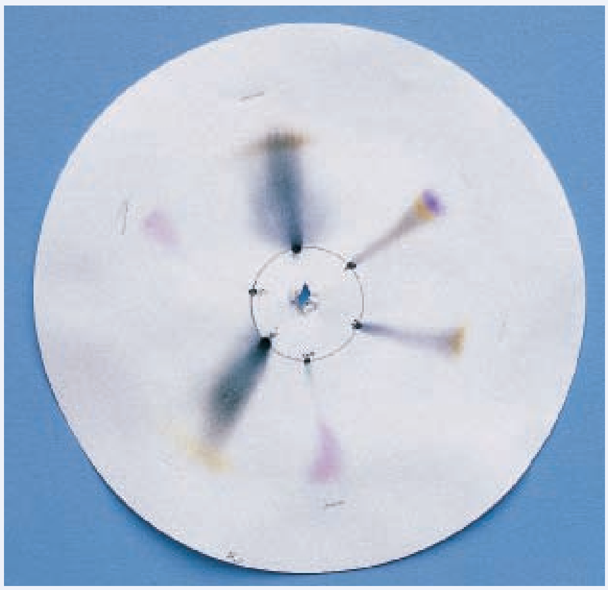


FIGURE E This chromatogram is unacceptable due to bleeding, or spread of the pigment front. Too much ink was used in the spot, or the ink was too soluble in the solvent.

PRE-LABORATORY PROCEDURE

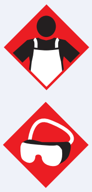
Volumetric Analysis

Volumetric analysis, the quantitative determination of the concentration of a solution, is achieved by adding a substance of known concentration to a substance of unknown concentration until the reaction between them is complete. The most common application of volumetric analysis is titration.

A buret is used in titrations. The solution with the known concentration is usually in the buret. The solution with the unknown concentration is usually in the Erlenmeyer flask. A few drops of a visual indicator are also added to the flask. The solution in the buret is then added to the flask until the indicator changes color, which signals that the reaction between the two solutions is complete. Then, the volumetric data obtained and the balanced chemical equation for the reaction are used to calculate the unknown concentration.

FIGURE A

GENERAL SAFETY



Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the location of the emergency lab shower and eyewash station and the procedure for using them.

The general setup for a titration is shown in Figure A. The steps for setting up this technique follow.

ASSEMBLING THE APPARATUS

1. Attach a buret clamp to a ring stand.
2. Thoroughly wash and rinse a buret. If water droplets cling to the walls of the buret, wash it again and gently scrub the inside walls with a buret brush.
3. Attach the buret to one side of the buret clamp.
4. Place an Erlenmeyer flask for waste solutions under the buret tip, as shown in Figure A.

FIGURE B

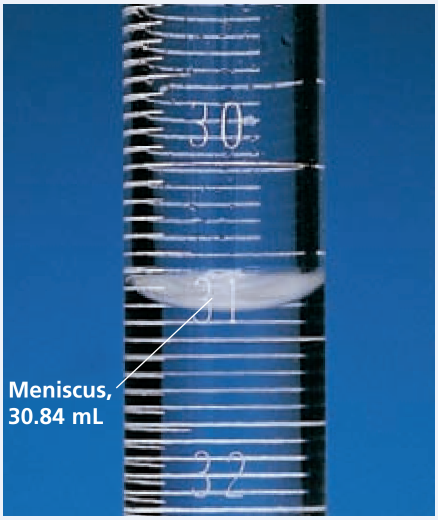


FIGURE C

OPERATING THE STOPCOCK

1. The stopcock should be operated with the left hand. This method gives better control but may prove awkward at first for right-handed students. The handle should be moved with the thumb and first two fingers of the left hand, as shown in Figure B.

2. Rotate the stopcock back and forth. It should move freely and easily. If it sticks or will not move, ask your teacher for assistance. Turn the stopcock to the closed position. Use a wash bottle to add 10 mL of distilled water to the buret. Rotate the stopcock to the open position. The water should come out in a steady stream. If no water comes out or if the stream of water is blocked, ask your teacher to check the stopcock for clogs.

FILLING THE BURET

1. To fill the buret, place a funnel in the top of the buret. Slowly and carefully pour the solution of known concentration from a beaker into the funnel. Open the stopcock, and allow some of the solution to drain into the waste beaker. Then, add enough solution to the buret to raise the level above the zero mark, but do not allow the solution to overflow.

READING THE BURET

1. Drain the buret until the bottom of the meniscus is on the zero mark or within the calibrated portion of the buret. If the solution level is not at zero, record the exact reading. If you start from the zero mark, your final buret reading will equal the amount of solution added. Remember, burets can be read to the second decimal place. Burets are designed to read the volume of liquid delivered to the flask, so numbers increase as you read downward from the top. For example, the meniscus in Figure C is at 30.84 mL, not 31.16 mL.

2. Replace the waste beaker with an Erlenmeyer flask containing a measured amount of the solution of unknown concentration.

Experiment 15 is an example of a back-titration applied to an acid-base reaction; it can be performed on a larger scale if micropipets are replaced with burets.

PRE-LABORATORY PROCEDURE

Calorimetry

Calorimetry, the measurement of the transfer of energy as heat, allows chemists to determine thermal constants, such as the specific heat of metals and the enthalpy of solution.

When two substances at different temperatures touch one another, energy as heat flows from the warmer substance to the cooler substance until the two substances are at the same temperature. The amount of energy transferred is measured in joules. (One joule equals 4.184 calories.)

A device used to measure the transfer of energy as heat is a calorimeter. Calorimeters vary in construction depending on the purpose and the accuracy of the energy measurement required. No calorimeter is a perfect insulator; some energy is always lost to the surroundings as heat. Therefore, every calorimeter must be calibrated to obtain its calorimeter constant.

GENERAL SAFETY

Always wear safety goggles and a lab apron to protect your eyes and clothing. If you get a chemical in your eyes, immediately flush the chemical out at the eyewash station while calling to your teacher. Know the location of the emergency lab shower and eyewash station and the procedure for using them.

Turn off hot plates and other heat sources when not in use. Do not touch a hot plate after it has just been turned off; it is probably hotter than you think. Use tongs when handling heated containers. Never hold or touch containers with your hands while heating them.

The general setup for a calorimeter made from plastic-foam cups is shown in Figure A. The steps for constructing this setup follow.

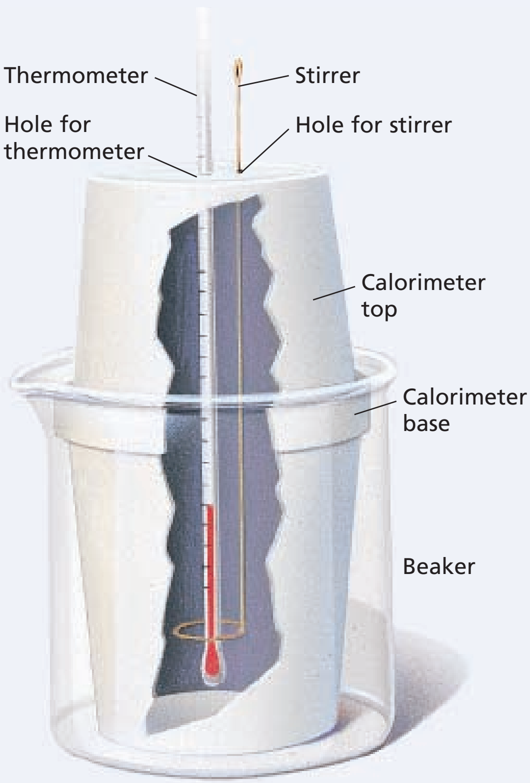


FIGURE A Position the hole for the stirrer so that the thermometer is in the center of the wire ring.

CONSTRUCTING THE CALORIMETER

1. Trim the lip of one plastic-foam cup, and use that cup as the top of your calorimeter. The other cup will be used as the base.
2. Use the pointed end of a pencil to gently make a hole in the center of the calorimeter top. The hole should be large enough to insert a thermometer. Make a hole for the wire stirrer. As you can see in Figure A, this hole should be positioned so that the wire stirrer can be raised and lowered without interfering with the thermometer.
3. Place the calorimeter in a beaker to prevent it from tipping over.

CALIBRATING A PLASTIC-FOAM CUP CALORIMETER

1. Measure 50 mL of distilled water in a graduated cylinder. Pour it into the calorimeter. Measure and record the temperature of the water in the plastic-foam cup.
2. Pour another 50 mL of distilled water into a beaker. Set the beaker on a hot plate, and warm the water to about 60°C, as shown in Figure B. Measure and record the temperature of the water.
3. Immediately pour the warm water into the cup, as shown in Figure C. Cover the cup, and move the stirrer gently up and down to mix the contents thoroughly. **Take care not to break the thermometer.**

FIGURE B Heat the distilled water to approximately 60°C.

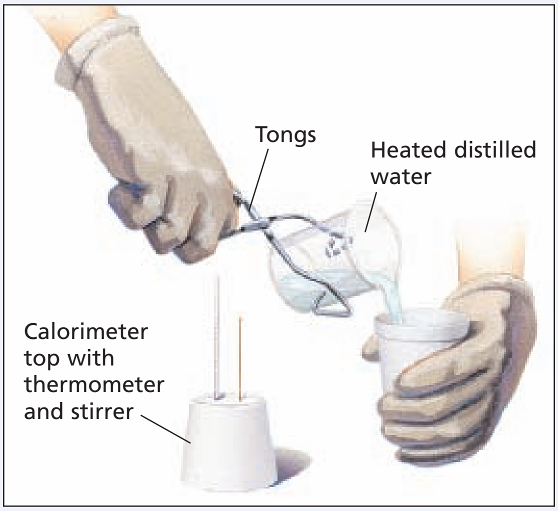


FIGURE C When transferring warm water from the beaker to the calorimeter, hold the bottom of the calorimeter steady so it does not tip over.

4. Watch the thermometer, and record the highest temperature attained (usually after about 30 s).
5. Empty the calorimeter.
6. The derivation of the equation to find the calorimeter constant starts with the following relationship.

   Energy lost by the warm water = Energy gained by the cool water + Energy gained by the calorimeter

   $$q_{warm\ H_2O} = q_{cool\ H_2O} + q_{calorimeter}$$

   The energy lost as heat by the warm water is calculated by the following equation.

   $$q_{warm\ H_2O} = mass_{warm\ H_2O} \times 4.184\ \text{J/g•°C} \times \Delta t$$

   The energy gained as heat by the calorimeter system equals the energy lost as heat by the warm water. You can use the following equation to calculate the calorimeter constant C' for your calorimeter.

   $$q_{calorimeter} = q_{\ warm\ H_2O} = (mass_{cool\ H_2O})\ (4.184\ \text{J/g•°C})\ (\Delta t_{cool\ H_2O}) + C'(\Delta t_{cool\ H_2O})$$

   Substitute the data from your calibration, and solve for C'.

APPENDIX A

TABLE A-1 SI MEASUREMENT

Metric Prefixes

Prefix	Symbol	Factor of Base Unit
giga	G	1 000 000 000
mega	M	1 000 000
kilo	k	1 000
hecto	h	100
deka	da	10
deci	d	0.1
centi	c	0.01
milli	m	0.001
micro	μ	0.000 001
nano	n	0.000 000 001
pico	p	0.000 000 000 001

Mass

1 kilogram (kg)	= SI base unit of mass
1 gram (g)	= 0.001 kg
1 milligram (mg)	= 0.000 001 kg
1 microgram (μg)	= 0.000 000 001 kg

Length

1 kilometer (km)	=	1 000 m
1 meter (m)	=	SI base unit of length
1 centimeter (cm)	=	0.01 m
1 millimeter (mm)	=	0.001 m
1 micrometer (μm)	=	0.000 001 m
1 nanometer (nm)	=	0.000 000 001 m
1 picometer (pm)	=	0.000 000 000 001 m

Area

1 square kilometer (km^2)	=	100 hectares (ha)
1 hectare (ha)	=	10 000 square meters (m^2)
1 square meter (m^2)	=	10 000 square centimeters (cm^2)
1 square centimeter (cm^2)	=	100 square millimeters (mm^2)

Volume

1 liter (L)	=	common unit for liquid volume (not SI)
1 cubic meter (m^3)	=	1000 L
1 kiloliter (kL)	=	1000 L
1 milliliter (mL)	=	0.001 L
1 milliliter (mL)	=	1 cubic centimeter (cm^3)

TABLE A-2 UNIT SYMBOLS

amu	=	atomic mass unit (mass)
atm	=	atmosphere (pressure, non-SI)
Bq	=	becquerel (nuclear activity)
°C	=	degree Celsius (temperature)
J	=	joule (energy)
K	=	kelvin (temperature, thermodynamic)
mol	=	mole (quantity)
M	=	molarity (concentration)
N	=	newton (force)
Pa	=	pascal (pressure)
s	=	second (time)
V	=	volt (electric potential difference)

TABLE A-3 SYMBOLS

Symbol		Meaning
α	=	helium nucleus (also ${}^{4}_{2}He$) emission from radioactive materials
β	=	electron (also ${}^{0}_{-1}e$) emission from radioactive materials
γ	=	high-energy photon emission from radioactive materials
Δ	=	change in a given quantity (e.g., ΔH for change in enthalpy)
c	=	speed of light in vacuum
c_p	=	specific heat capacity (at constant pressure)
D	=	density
E_a	=	activation energy
E^0	=	standard electrode potential
E^0 cell	=	standard potential of an electrochemical cell
G	=	Gibbs free energy
ΔG^0	=	standard free energy of reaction
$\Delta G^0{}_f$	=	standard molar free energy of formation
H	=	enthalpy
ΔH^0	=	standard enthalpy of reaction
$\Delta H^0{}_f$	=	standard molar enthalpy of formation
K_a	=	ionization constant (acid)
K_b	=	dissociation constant (base)
K_{eq}	=	equilibrium constant
K_{sp}	=	solubility-product constant
KE	=	kinetic energy
m	=	mass
N_A	=	Avogadro's number
n	=	number of moles
P	=	pressure
pH	=	measure of acidity ($-\log[H_3O^+]$)
R	=	ideal gas law constant
S	=	entropy
S^0	=	standard molar entropy
T	=	temperature (thermodynamic, in kelvins)
t	=	temperature (± degrees Celsius)
V	=	volume
v	=	velocity

TABLE A-4 PHYSICAL CONSTANTS

Quantity	Symbol	Value
Atomic mass unit	amu	$1.660\,5389 \times 10^{-27}$ kg
Avogadro's number	N_A	$6.022\,142 \times 10^{23}$/mol
Electron rest mass	m_e	$9.109\,3826 \times 10^{-31}$ kg 5.4858×10^{-4} amu
Ideal gas law constant	R	8.314 L • kPa/(mol • K) 0.0821 L • atm/(mol • K)
Molar volume of ideal gas at STP	V_M	22.414 10 L/mol
Neutron rest mass	m_n	$1.674\,9273 \times 10^{-27}$ kg 1.008 665 amu
Normal boiling point of water	T_b	373.15 K = 100.0°C
Normal freezing point of water	T_f	273.15 K = 0.00°C
Planck's constant	h	$6.626\,069 \times 10^{-34}$ J • s
Proton rest mass	m_p	$1.672\,6217 \times 10^{-27}$ kg 1.007 276 amu
Speed of light in a vacuum	c	$2.997\,924\,58 \times 10^{8}$ m/s
Temperature of triple point of water		273.16 K = 0.01°C

TABLE A-5 ENTHALPY OF COMBUSTION

Substance	Formula	State	ΔH_c
hydrogen	H_2	*g*	–285.8
graphite	C	*s*	–393.5
carbon monoxide	CO	*g*	–283.0
methane	CH_4	*g*	–890.8
ethane	C_2H_6	*g*	–1560.7
propane	C_3H_8	*g*	–2219.2
butane	C_4H_{10}	*g*	–2877.6
pentane	C_5H_{12}	*g*	–3535.6
hexane	C_6H_{14}	*l*	–4163.2
heptane	C_7H_{16}	*l*	–4817.0
octane	C_8H_{18}	*l*	–5470.5
ethene (ethylene)	C_2H_4	*g*	–1411.2
propene (propylene)	C_3H_6	*g*	–2058.0
ethyne (acetylene)	C_2H_2	*g*	–1301.1
benzene	C_6H_6	*l*	–3267.6
toluene	C_7H_8	*l*	–3910.3
naphthalene	$C_{10}H_8$	*s*	–5156.3
anthracene	$C_{14}H_{10}$	*s*	–7163.0
methanol	CH_3OH	*l*	–726.1
ethanol	C_2H_5OH	*l*	–1366.8
ether	$(C_2H_5)_2O$	*l*	–2751.1
formaldehyde	CH_2O	*g*	–570.7
glucose	$C_6H_{12}O_6$	*s*	–2803.0
sucrose	$C_{12}H_{22}O_{11}$	*s*	–5640.9

ΔH_c = enthalpy of combustion of the given substance. All values of ΔH_c are expressed as kJ/mol of substance oxidized to $H_2O(l)$ and/or $CO_2(g)$ at constant pressure and 25°C.
s = solid, *l* = liquid, *g* = gas

TABLE A-6 THE ELEMENTS—SYMBOLS, ATOMIC NUMBERS, AND ATOMIC MASSES

Name of element	Symbol	Atomic number	Atomic mass
actinium	Ac	89	[227]
aluminum	Al	13	26.9815386
americium	Am	95	[243]
antimony	Sb	51	121.760
argon	Ar	18	39.948
arsenic	As	33	74.92160
astatine	At	85	[210]
barium	Ba	56	137.327
berkelium	Bk	97	[247]
beryllium	Be	4	9.012182
bismuth	Bi	83	208.98040
bohrium	Bh	107	[264]
boron	B	5	10.811
bromine	Br	35	79.904
cadmium	Cd	48	112.411
calcium	Ca	20	40.078
californium	Cf	98	[251]
carbon	C	6	12.0107
cerium	Ce	58	140.116
cesium	Cs	55	132.9054519
chlorine	Cl	17	35.453
chromium	Cr	24	51.9961
cobalt	Co	27	58.933195
copper	Cu	29	63.546
curium	Cm	96	[247]
darmstadtium	Ds	110	[271]
dubnium	Db	105	[262]
dysprosium	Dy	66	162.500
einsteinium	Es	99	[252]
erbium	Er	68	167.259
europium	Eu	63	151.964
fermium	Fm	100	[257]
fluorine	F	9	18.9984032
francium	Fr	87	[223]
gadolinium	Gd	64	157.25
gallium	Ga	31	69.723
germanium	Ge	32	72.64
gold	Au	79	196.966569
hafnium	Hf	72	178.49
hassium	Hs	108	[277]
helium	He	2	4.00260
holmium	Ho	67	164.93032
hydrogen	H	1	1.00794
indium	In	49	114.818
iodine	I	53	126.90447
iridium	Ir	77	192.217

TABLE A-6 CONTINUED

Name of element	Symbol	Atomic number	Atomic mass
iron	Fe	26	55.845
krypton	Kr	36	83.798
lanthanum	La	57	138.90547
lawrencium	Lr	103	[262]
lead	Pb	82	207.2
lithium	Li	3	6.941
lutetium	Lu	71	174.967
magnesium	Mg	12	24.3050
manganese	Mn	25	54.938045
meitnerium	Mt	109	[268]
mendelevium	Md	101	[258]
mercury	Hg	80	200.59
molybdenum	Mo	42	95.94
neodymium	Nd	60	144.242
neon	Ne	10	20.1797
neptunium	Np	93	[237]
nickel	Ni	28	58.6934
niobium	Nb	41	92.90638
nitrogen	N	7	14.0067
nobelium	No	102	[259]
osmium	Os	76	190.23
oxygen	O	8	15.9994
palladium	Pd	46	106.42
phosphorus	P	15	30.973762
platinum	Pt	78	195.084
plutonium	Pu	94	[244]
polonium	Po	84	[209]
potassium	K	19	39.0983
praseodymium	Pr	59	140.90765
promethium	Pm	61	[145]
protactinium	Pa	91	231.03588
radium	Ra	88	[226]
radon	Rn	86	[222]
rhenium	Re	75	186.207
rhodium	Rh	45	102.90550
roentgenium	Rg	111	[272]
rubidium	Rb	37	85.4678
ruthenium	Ru	44	101.07
rutherfordium	Rf	104	[261]
samarium	Sm	62	150.36
scandium	Sc	21	44.955912
seaborgium	Sg	106	[266]
selenium	Se	34	78.96
silicon	Si	14	28.0855
silver	Ag	47	107.8682
sodium	Na	11	22.98976928
strontium	Sr	38	87.62
sulfur	S	16	32.065
tantalum	Ta	73	180.94788
technetium	Tc	43	[98]
tellurium	Te	52	127.60
terbium	Tb	65	158.92535
thallium	Tl	81	204.3833
thorium	Th	90	232.03806
thulium	Tm	69	168.93421
tin	Sn	50	118.710
titanium	Ti	22	47.867
tungsten	W	74	183.84
uranium	U	92	238.02891
vanadium	V	23	50.9415
xenon	Xe	54	131.293
ytterbium	Yb	70	173.04
yttrium	Y	39	88.90585
zinc	Zn	30	65.409
zirconium	Zr	40	91.224

A value given in brackets denotes the mass number of the most stable or most common isotope. The atomic masses of most of these elements are believed to have an error no greater than ±1 in the last digit given.

TABLE A-7 COMMON IONS

Cation	Symbol	Anion	Symbol
aluminum	Al^{3+}	**acetate**	CH_3COO^-
ammonium	$NH_4{}^+$	**bromide**	Br^-
arsenic(III)	As^{3+}	**carbonate**	CO_3^{2-}
barium	Ba^{2+}	**chlorate**	ClO_3^-
calcium	Ca^{2+}	**chloride**	Cl^-
chromium(II)	Cr^{2+}	**chlorite**	ClO_2^-
chromium(III)	Cr^{3+}	**chromate**	CrO_4^{2-}
cobalt(II)	Co^{2+}	**cyanide**	CN^-
cobalt(III)	Co^{3+}	**dichromate**	$Cr_2O_7^{2-}$
copper(I)	Cu^+	**fluoride**	F^-
copper(II)	Cu^{2+}	**hexacyanoferrate(II)**	$Fe(CN)_6^{4-}$
hydronium	H_3O^+	**hexacyanoferrate(III)**	$Fe(CN)_6^{3-}$
iron(II)	Fe^{2+}	**hydride**	H^-
iron(III)	Fe^{3+}	**hydrogen carbonate**	HCO_3^-
lead(II)	Pb^{2+}	**hydrogen sulfate**	HSO_4^-
magnesium	Mg^{2+}	**hydroxide**	OH^-
mercury(I)	$Hg_2{}^{2+}$	**hypochlorite**	ClO^-
mercury(II)	Hg^{2+}	**iodide**	I^-
nickel(II)	Ni^{2+}	**nitrate**	NO_3^-
potassium	K^+	**nitrite**	NO_2^-
silver	Ag^+	**oxide**	O^{2-}
sodium	Na^+	**perchlorate**	ClO_4^-
strontium	Sr^{2+}	**permanganate**	MnO_4^-
tin(II)	Sn^{2+}	**peroxide**	O_2^{2-}
tin(IV)	Sn^{4+}	**phosphate**	PO_4^{3-}
titanium(III)	Ti^{3+}	**sulfate**	SO_4^{2-}
titanium(IV)	Ti^{4+}	**sulfide**	S^{2-}
zinc	Zn^{2+}	**sulfite**	SO_3^{2-}

TABLE A-8 WATER-VAPOR PRESSURE

Temperature (°C)	Pressure (mm Hg)	Pressure (kPa)	Temperature (°C)	Pressure (mm Hg)	Pressure (kPa)
0.0	4.6	0.61	**23.0**	21.1	2.81
5.0	6.5	0.87	**23.5**	21.7	2.90
10.0	9.2	1.23	**24.0**	22.4	2.98
15.0	12.8	1.71	**24.5**	23.1	3.10
15.5	13.2	1.76	**25.0**	23.8	3.17
16.0	13.6	1.82	**26.0**	25.2	3.36
16.5	14.1	1.88	**27.0**	26.7	3.57
17.0	14.5	1.94	**28.0**	28.3	3.78
17.5	15.0	2.00	**29.0**	30.0	4.01
18.0	15.5	2.06	**30.0**	31.8	4.25
18.5	16.0	2.13	**35.0**	42.2	5.63
19.0	16.5	2.19	**40.0**	55.3	7.38
19.5	17.0	2.27	**50.0**	92.5	12.34
20.0	17.5	2.34	**60.0**	149.4	19.93
20.5	18.1	2.41	**70.0**	233.7	31.18
21.0	18.6	2.49	**80.0**	355.1	47.37
21.5	19.2	2.57	**90.0**	525.8	70.12
22.0	19.8	2.64	**95.0**	633.9	84.53
22.5	20.4	2.72	**100.0**	760.0	101.32

TABLE A-9 DENSITIES OF GASES AT STP

Gas	Density (g/L)
air, dry	1.293
ammonia	0.771
carbon dioxide	1.997
carbon monoxide	1.250
chlorine	3.214
dinitrogen monoxide	1.977
ethyne (acetylene)	1.165
helium	0.1785
hydrogen	0.0899
hydrogen chloride	1.639
hydrogen sulfide	1.539
methane	0.7168
nitrogen	1.2506
nitrogen monoxide (at 10°C)	1.340
oxygen	1.429
sulfur dioxide	2.927

TABLE A-10 DENSITY OF WATER

Temperature (°C)	Density (g/cm^3)
0	0.999 84
2	0.999 94
3.98 (maximum)	0.999 973
4	0.999 97
6	0.999 94
8	0.999 85
10	0.999 70
14	0.999 24
16	0.998 94
20	0.998 20
25	0.997 05
30	0.995 65
40	0.992 22
50	0.988 04
60	0.983 20
70	0.977 77
80	0.971 79
90	0.965 31
100	0.958 36

TABLE A-11 SOLUBILITIES OF GASES IN WATER

Volume of gas (in liters) at STP that can be dissolved in 1 L of water at the temperature (°C) indicated.

Gas	0°C	10°C	20°C	60°C
air	0.029 18	0.022 84	0.018 68	0.012 16
ammonia	1130	870	680	200
carbon dioxide	1.713	1.194	0.878	0.359
carbon monoxide	0.035 37	0.028 16	0.023 19	0.014 88
chlorine	—	3.148	2.299	1.023
hydrogen	0.021 48	0.019 55	0.018 19	0.016 00
hydrogen chloride	512	475	442	339
hydrogen sulfide	4.670	3.399	2.582	1.190
methane	0.055 63	0.041 77	0.033 08	0.019 54
nitrogen*	0.023 54	0.018 61	0.015 45	0.010 23
nitrogen monoxide	0.073 81	0.057 09	0.047 06	0.029 54
oxygen	0.048 89	0.038 02	0.031 02	0.019 46
sulfur dioxide	79.789	56.647	39.374	—

*Atmospheric nitrogen–98.815% N_2, 1.185% inert gases

TABLE A-12 SOLUBILITY CHART

	acetate	bromide	carbonate	chlorate	chloride	chromate	hydroxide	iodine	nitrate	oxide	phosphate	silicate	sulfate	sulfide
aluminum	S	S	—	S	S	—	A	S	S	a	A	I	S	d
ammonium	S	S	S	S	S	S	S	S	S	—	S	—	S	S
barium	S	S	P	S	S	A	S	S	S	S	A	S	a	d
calcium	S	S	P	S	S	S	S	S	S	P	P	P	S	S
copper(II)	S	S	—	S	S	—	A	—	S	A	A	A	S	A
hydrogen	S	S	—	S	S	—	—	S	S	S	S	I	S	S
iron(II)	—	S	P	S	S	—	A	S	S	A	A	—	S	A
iron(III)	—	S	—	S	S	A	A	S	S	A	P	—	P	d
lead(II)	S	S	A	S	S	A	P	P	S	P	A	A	P	A
magnesium	S	S	P	S	S	S	A	S	S	A	P	A	S	d
manganese(II)	S	S	P	S	S	—	A	S	S	A	P	I	S	A
mercury(I)	P	A	A	S	a	P	—	A	S	A	A	—	P	I
mercury(II)	S	S	—	S	S	P	A	P	S	P	A	—	d	I
potassium	S	S	S	S	S	S	S	S	S	S	S	S	S	S
silver	P	a	A	S	a	P	—	I	S	P	A	—	P	A
sodium	S	S	S	S	S	S	S	S	S	d	S	S	S	S
strontium	S	S	P	S	S	P	S	S	S	S	A	A	P	S
tin(II)	d	S	—	S	S	A	A	S	d	A	A	—	S	A
tin(IV)	S	S	—	—	S	S	P	d	—	A	—	—	S	A
zinc	S	S	P	S	S	P	A	S	S	P	A	A	S	A

S = soluble in water. A = soluble in acids, insoluble in water. P = partially soluble in water, soluble in dilute acids.
I = insoluble in dilute acids and in water. a = slightly soluble in acids, insoluble in water. d = decomposes in water.

TABLE A-13 SOLUBILITY OF COMPOUNDS

Solubilities are given in grams of solute that can be dissolved in 100 g of water at the temperature (°C) indicated.

Compound	Formula	0°C	20°C	60°C	100°C
aluminum sulfate	$Al_2(SO_4)_3$	31.2	36.4	59.2	89.0
ammonium chloride	NH_4Cl	29.4	37.2	55.3	77.3
ammonium nitrate	NH_4NO_3	118	192	421	871
ammonium sulfate	$(NH_4)_2SO_4$	70.6	75.4	88	103
barium carbonate	$BaCO_3$	—*	$0.0022^{18°}$	—*	0.0065
barium chloride dihydrate	$BaCl_2 \bullet 2H_2O$	31.2	35.8	46.2	59.4
barium hydroxide	$Ba(OH)_2$	1.67	3.89	20.94	$101.40^{80°}$
barium nitrate	$Ba(NO_3)_2$	4.95	9.02	20.4	34.4
barium sulfate	$BaSO_4$	—*	$0.000\ 246^{25°}$	—*	0.000 413
calcium carbonate	$CaCO_3$	—*	$0.0014^{25°}$	—*	$0.0018^{75°}$
calcium fluoride	CaF_2	$0.0016^{18°}$	$0.0017^{26°}$	—*	—*
calcium hydrogen carbonate	$Ca(HCO_3)_2$	16.15	16.60	17.50	18.40
calcium hydroxide	$Ca(OH)_2$	0.189	0.173	0.121	0.076
calcium sulfate	$CaSO_4$	—*	$0.209^{30°}$	—*	0.1619
copper(II) chloride	$CuCl_2$	68.6	73.0	96.5	120
copper(II) sulfate pentahydrate	$CuSO_4 \bullet 5H_2O$	23.1	32.0	61.8	114
lead(II) chloride	$PbCl_2$	0.67	1.00	1.94	3.20
lead(II) nitrate	$Pb(NO_3)_2$	37.5	54.3	91.6	133
lithium chloride	$LiCl$	69.2	83.5	98.4	128
lithium sulfate	Li_2SO_4	36.1	34.8	32.6	$30.9^{90°}$
magnesium hydroxide	$Mg(OH)_2$	—*	$0.0009^{18°}$	—*	0.004
magnesium sulfate	$MgSO_4$	22.0	33.7	54.6	68.3
mercury(I) chloride	Hg_2Cl_2	—*	$0.000\ 20^{25°}$	$0.001^{43°}$	—*
mercury(II) chloride	$HgCl_2$	3.63	6.57	16.3	61.3
potassium bromide	KBr	53.6	65.3	85.5	104
potassium chlorate	$KClO_3$	3.3	7.3	23.8	56.3
potassium chloride	KCl	28.0	34.2	45.8	56.3
potassium chromate	K_2CrO_4	56.3	63.7	70.1	$74.5^{90°}$
potassium iodide	KI	128	144	176	206
potassium nitrate	KNO_3	13.9	31.6	106	245
potassium permanganate	$KMnO_4$	2.83	6.34	22.1	—*
potassium sulfate	K_2SO_4	7.4	11.1	18.2	24.1
silver acetate	$AgC_2H_3O_2$	0.73	1.05	1.93	$2.59^{80°}$
silver chloride	$AgCl$	$0.000\ 089^{10°}$	—*	—*	0.0021
silver nitrate	$AgNO_3$	122	216	440	733
sodium acetate	$NaC_2H_3O_2$	36.2	46.4	139	170
sodium chlorate	$NaClO_3$	79.6	95.9	137	204
sodium chloride	$NaCl$	35.7	35.9	37.1	39.2
sodium nitrate	$NaNO_3$	73.0	87.6	122	180
sucrose	$C_{12}H_{22}O_{11}$	179.2	203.9	287.3	487.2

*Dashes indicate that values are not available.

TABLE A-14 ENTHALPY OF FORMATION

Substance	State	ΔH_f	Substance	State	ΔH_f
ammonia	*g*	−45.9	lead(IV) oxide	*s*	−274.5
ammonium chloride	*s*	−314.4	lead(II) nitrate	*s*	−451.9
ammonium sulfate	*s*	−1180.9	lead(II) sulfate	*s*	−919.94
barium chloride	*s*	−858.6	lithium chloride	*s*	−408.6
barium nitrate	*s*	−768.2	lithium nitrate	*s*	−483.1
barium sulfate	*s*	−1473.2	magnesium chloride	*s*	−641.5
benzene	*g*	+82.88	magnesium oxide	*s*	−601.6
benzene	*l*	+49.080	magnesium sulfate	*s*	−1261.79
calcium carbonate	*s*	−1207.6	manganese(IV) oxide	*s*	−520.0
calcium chloride	*s*	−795.4	manganese(II) sulfate	*s*	−1065.3
calcium hydroxide	*s*	−983.2	mercury(I) chloride	*s*	−264.2
calcium nitrate	*s*	−938.2	mercury(II) chloride	*s*	−230.0
calcium oxide	*s*	−634.9	mercury(II) oxide (red)	*s*	−90.8
calcium sulfate	*s*	−1434.5	methane	*g*	−74.9
carbon (diamond)	*s*	+1.9	nitrogen dioxide	*g*	+33.2
carbon (graphite)	*s*	0.00	nitrogen monoxide	*g*	+90.29
carbon dioxide	*g*	−393.5	dinitrogen monoxide	*g*	+82.1
carbon monoxide	*g*	−110.5	dinitrogen tetroxide	*g*	+9.2
copper(II) nitrate	*s*	−302.9	oxygen (O_2)	*g*	0.00
copper(II) oxide	*s*	−157.3	ozone (O_3)	*g*	+142.7
copper(II) sulfate	*s*	−771.4	tetraphosphorus decoxide	*s*	−3009.9
ethane	*g*	−83.8	potassium bromide	*s*	−393.8
ethyne (acetylene)	*g*	+228.2	potassium chloride	*s*	−436.49
hydrogen (H_2)	*g*	0.00	potassium hydroxide	*s*	−424.58
hydrogen bromide	*g*	−36.29	potassium nitrate	*s*	−494.6
hydrogen chloride	*g*	−92.3	potassium sulfate	*s*	−1437.8
hydrogen fluoride	*g*	−273.3	silicon dioxide (quartz)	*s*	−910.7
hydrogen iodide	*g*	+26.5	silver chloride	*s*	−127.01 ± 0.5
hydrogen oxide (water)	*g*	−241.8	silver nitrate	*s*	−120.5
hydrogen oxide (water)	*l*	−285.8	silver sulfide	*s*	−32.59
hydrogen peroxide	*g*	−136.3	sodium bromide	*s*	−361.8
hydrogen peroxide	*l*	−187.8	sodium chloride	*s*	−385.9
hydrogen sulfide	*g*	−20.6	sodium hydroxide	*s*	−425.9
iodine (I_2)	*s*	0.00	sodium nitrate	*s*	−467.9
iodine (I_2)	*g*	+62.4	sodium sulfate	*l*	−1387.1
iron(II) chloride	*s*	−399.4	sulfur dioxide	*g*	−296.8
iron(II) oxide	*s*	−272.0	sulfur trioxide	*g*	−395.7
iron(III) oxide	*s*	−824.2	tin(IV) chloride	*l*	−511.3
iron(II) sulfate	*s*	−928.4	zinc nitrate	*s*	−483.7
iron(II) sulfide	*s*	−100.0	zinc oxide	*s*	−350.5
lead(II) oxide	*s*	−217.3	zinc sulfate	*s*	−980.14

ΔH_f is enthalpy of formation of the given substance from its elements. All values of ΔH_f are expressed as kJ/mol at 25ºC. Negative values of ΔH_f indicate exothermic reactions. *s* = solid, *l* = liquid, *g* = gas

TABLE A-15 PROPERTIES OF COMMON ELEMENTS

Name	Form/color at room temperature	Density (g/cm^3)†	Melting point (°C)	Boiling point (°C)	Common oxidation states
aluminum	silver metal	2.702	660.37	2467	3+
arsenic	gray metalloid	5.727^{14}	817 (28 atm)	613 *(sublimes)*	3−, 3+, 5+
barium	bluish white metal	3.51	725	1640	2+
bromine	red-brown liquid	3.119	−7.2	58.78	1−, 1+, 3+, 5+, 7+
calcium	silver metal	1.54	839 ± 2	1484	2+
carbon	diamond graphite	3.51 2.25	3500 (63.5 atm) 3652 *(sublimes)*	3930 —	2+, 4+
chlorine	green-yellow gas	3.214*	−100.98	−34.6	1−, 1+, 3+, 5+, 7+
chromium	gray metal	7.2028	1857 ± 20	2672	2+, 3+, 6+
cobalt	gray metal	8.9	1495	2870	2+, 3+
copper	red metal	8.92	1083.4 ± 0.2	2567	1+, 2+
fluorine	yellow gas	1.69‡	−219.62	−188.14	1−
germanium	gray metalloid	5.323^{25}	937.4	2830	4+
gold	yellow metal	19.31	1064.43	2808 ± 2	1+, 3+
helium	colorless gas	0.1785*	−272.2 (26 atm)	−268.9	0
hydrogen	colorless gas	0.0899*	−259.14	−252.8	1−, 1+
iodine	blue-black solid	4.93	113.5	184.35	1−, 1+, 3+, 5+, 7+
iron	silver metal	7.86	1535	2750	2+, 3+
lead	bluish white metal	11.343716	327.502	1740	2+, 4+
lithium	silver metal	0.534	180.54	1342	1+
magnesium	silver metal	1.745	648.8	1107	2+
manganese	gray-white metal	7.20	1244 ± 3	1962	2+, 3+, 4+, 6+, 7+
mercury	silver liquid metal	13.5462	−38.87	356.58	1+, 2+
neon	colorless gas	0.9002*	−248.67	−245.9	0
nickel	silver metal	8.90	1455	2732	2+, 3+
nitrogen	colorless gas	1.2506*	−209.86	−195.8	3−, 3+, 5+
oxygen	colorless gas	1.429*	−218.4	−182.962	2−
phosphorus	yellow solid	1.82	44.1	280	3−, 3+, 5+
platinum	silver metal	21.45	1772	3827 ± 100	2+, 4+
potassium	silver metal	0.86	63.25	760	1+
silicon	gray metalloid	2.33 ± 0.01	1410	2355	2+, 4+
silver	white metal	10.5	961.93	2212	1+
sodium	silver metal	0.97	97.8	882.9	1+
strontium	silver metal	2.6	769	1384	2+
sulfur	yellow solid	1.96	119.0	444.674	2−, 4+, 6+
tin	white metal	7.28	231.88	2260	2+, 4+
titanium	white metal	4.5	1660 ± 10	3287	2+, 3+, 4+
uranium	silver metal	19.05 ± 0.02^{25}	1132.3 ± 0.8	3818	3+, 4+, 6+
zinc	blue-white metal	7.14	419.58	907	2+

† Densities obtained at 20°C unless otherwise noted (superscript)
‡ Density of fluorine given in g/L at 1 atm and 15°C
* Densities of gases given in g/L at STP

Study Skills for Chemistry Table of Contents

Study Skills for Chemistry

Succeeding in Your Chemistry Class

Your success in this course will depend on your ability to apply some basic study skills to learning the material. Studying chemistry can be difficult, but you can make it easier using simple strategies for dealing with the concepts and problems. Becoming skilled in using these strategies will be your keys to success in this and many other courses.

▶ Reading the Text

- **Read the assigned material before class** so that the class lecture makes sense. Use a dictionary to help you build and interpret vocabulary. Remember that, while reading, one of your tasks is to figure out what information is important.

  Working together with others using Paired Reading and Discussion strategies can help you decide what is important and clarify the material. (For more discussion, see Other Reading Strategies on page 876.)
- **Select a quiet setting** away from distractions so that you can concentrate on what you are reading.
- **Have a pencil and paper nearby to jot down notes and questions** you may have. Be sure to get these questions answered in class. Power Notes (see page 871) can help you organize the notes you take and prepare you for class.
- **Use the Objectives in the beginning of each section as a list of what you need to know from the section.** Teachers generally make their tests based on the text objectives or their own objectives. Using the objectives to focus your reading can make your learning more efficient.

Using the K/W/L strategy (see page 874) can help you relate new material to what you already know and what you need to learn.

▶ Taking Notes in Class

- **Be prepared to take notes during class.** Have your materials organized in a notebook. Separate sheets of paper can be easily lost.
- **Don't write down everything your teacher says.** Try to tell which parts of the lecture are important and which are not. Reading the text before class will help in this. You will not be able to write down everything, so you must try to write down only the important things.
- **Recopying notes later is a waste of time** and does not help you learn material for a test. Do

it right the first time. Organize your notes as you are writing them down so that you can make sense of your notes when you review them without needing to recopy them.

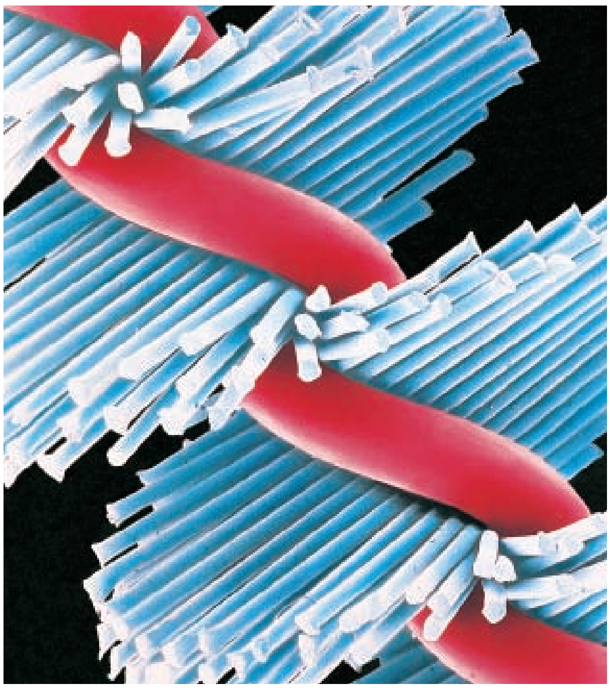
Reviewing Class Notes

- **Review your notes as soon as possible after class.** Write down any questions you may have about the material covered that day. Be sure to get these questions answered during the next class. You can work with friends to use strategies such as Paired Summarizing and L.I.N.K. (See page 878.)
- **Do not wait until the test to review.** By then you will have forgotten a good portion of the material.
- **Be selective about what you memorize.** You cannot memorize everything in a chapter. First of all, it is too time consuming. Second, memorizing and understanding are not the same thing. Memorizing topics as they appear in your notes or text does not guarantee that you will be able to correctly answer questions that require understanding of those topics. You should only memorize material that you understand. Concept Maps and other Reading Organizers, Sequencing/Pattern Puzzles, and Prediction Guides can help you understand key ideas and major concepts. (See pages 868, 875, and 877.)

Working Problems

In addition to understanding the concepts, the ability to solve problems will be a key to your success in chemistry. You will probably spend a lot of time working problems in class and at home. The ability to solve chemistry problems is a skill, and like any skill, it requires practice.

- **Always review the Sample Problems in the chapter.** The Sample Problems in the text provide road maps for solving certain types of problems. Cover the solution while trying to work the problem yourself.
- **The problems in the Chapter Review are similar to the Sample Problems.** If you can relate an assigned problem to one of the Sample Problems in the chapter, it shows that you understand the material.
- **The four steps: Analyze, Plan, Compute, and Evaluate should be the steps you go through when working assigned problems.** These steps will allow you to organize your thoughts and help you develop your problem-solving skills.
- **Never spend more than 15 minutes trying to solve a problem.** If you have not been able to come up with a plan for the solution after 15 minutes, additional time spent will only cause you to become frustrated. What do you do? Get help! See your teacher or a classmate. Find out what it is that you do not understand.
- **Do not try to memorize the Sample Problems; spend your time trying to understand how the solution develops.** Memorizing a particular sample problem will not ensure that you understand it well enough to solve a similar problem.
- **Always look at your answer and ask yourself if it is reasonable and makes sense.** Check to be sure you have the correct units and numbers of significant figures.



► Completing Homework

Your teacher will probably assign questions and problems from the Section Reviews and Chapter Reviews or assign *Modern Chemistry* Daily Homework. The purpose of these assignments is to review what you have covered in class and to see if you can use the information to answer questions or solve problems. As in reviewing class notes, do your homework as soon after class as possible while the topics are still fresh in your mind. Do not wait until late at night, when you are more likely to be tired and to become frustrated.

► Preparing for and Taking Exams

Reviewing for an exam

- **Don't panic and don't cram! It takes longer to learn if you are under pressure.** If you have followed the strategies listed here and reviewed along the way, studying for the exam should be less stressful.
- **When looking over your notes and concept maps, recite ideas out loud.** There are two reasons for reciting:
  1. You are hearing the information, which is effective in helping you learn.
  2. If you cannot recite the ideas, it should be a clue that you do not understand the material, and you should begin rereading or reviewing the material again.
- **Studying with a friend provides a good opportunity for recitation.** If you can explain ideas to your study partner, you know the material.

Taking an exam

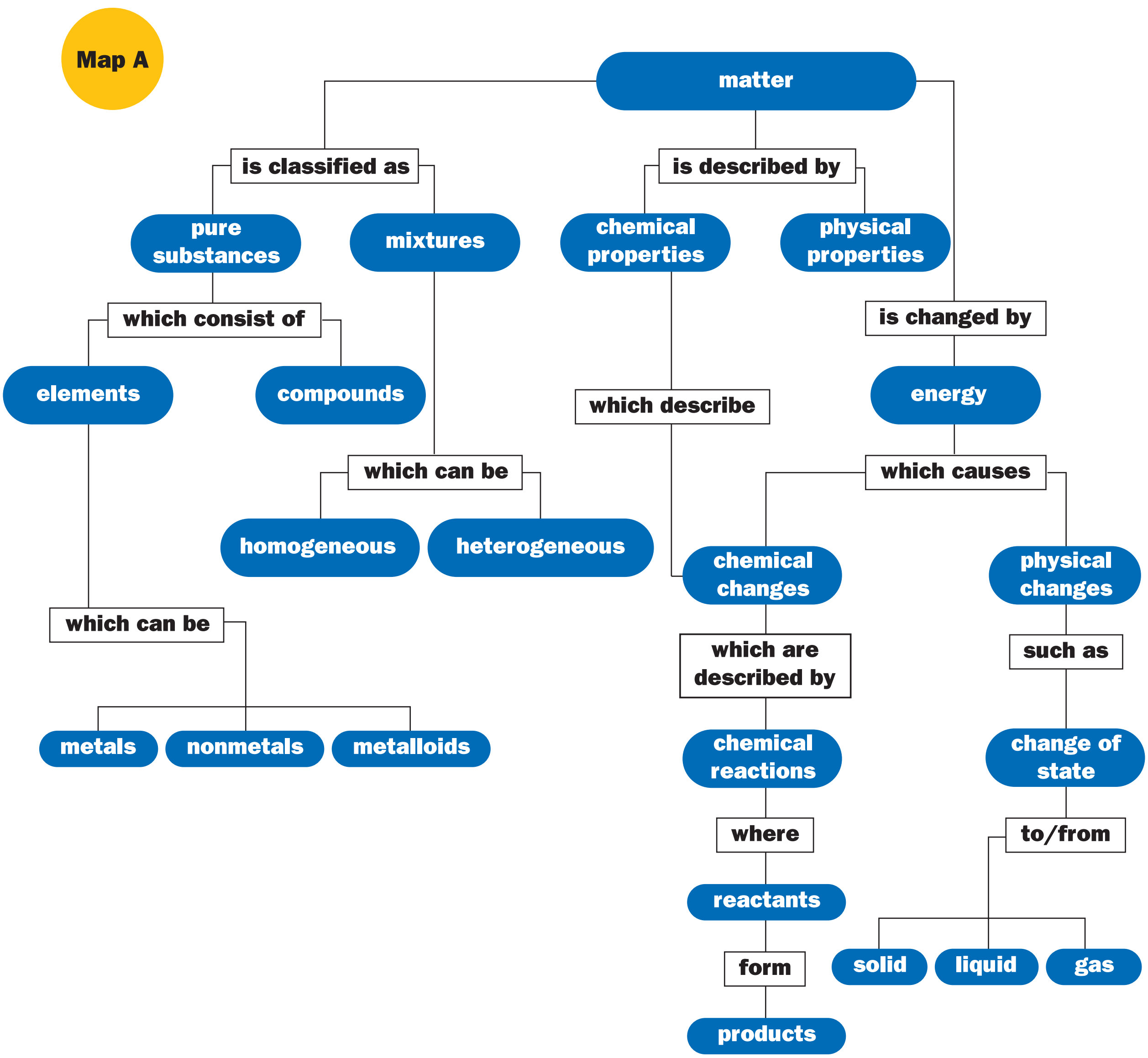
- **Get plenty of rest before the exam so that you can think clearly.** If you have been awake all night studying, you are less likely to succeed than if you had gotten a full night of rest.
- **Start with the questions you know.** If you get stuck on a question, save it for later. As time passes and you work through the exam, you may recall the information you need to answer a difficult question or solve a difficult problem.

Good luck!

Making Concept Maps

Making concept maps can help you decide what material in a chapter is important and how to efficiently learn that material. A concept map presents key ideas, meanings, and relationships for the concepts being studied. It can be thought of as a visual road map of the chapter. Learning happens efficiently when you use concept maps because you work with only the key ideas and how they fit together.

The concept map shown as **Map A** was made from vocabulary terms in Chapter 1. Vocabulary terms are generally labels for concepts, and concepts are generally nouns. In a concept map, linking words are used to form propositions that connect concepts and give them meaning in context. For example, on the map below, "matter is described by physical properties" is a proposition.



Studies show that people are better able to remember materials presented visually. A concept map is better than an outline because you can see the relationships among many ideas. Because outlines are linear, there is no way of linking the ideas from various sections of the outline. Read through the map to become familiar with the information presented. Then look at the map in relation to all of the text pages in Chapter 1; which gives a better picture of the important concepts—the map or the full chapter?

▶ To Make a Concept Map

1. List all the important concepts.
We'll use some of the boldfaced and italicized terms from Chapter 1, Section 2.

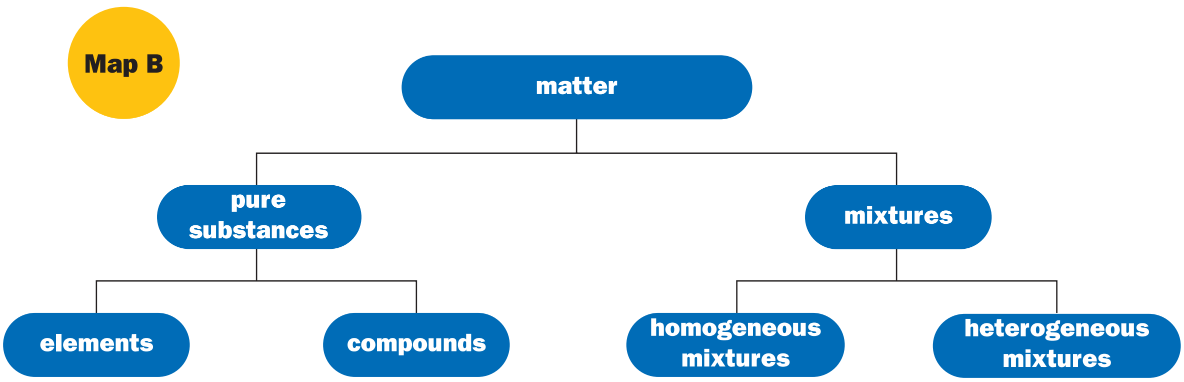
matter	mixture
compound	pure substance
element	
homogenous mixture	
heterogeneous mixture	

- From this list, group similar concepts together. For example, one way to group these concepts would be into two groups—one that is related to mixtures and one that is related to pure substances.

mixture	*pure substance*
heterogeneous mixture	compound
homogeneous mixture	element

2. Select a main concept for the map.
We will use *matter* as the main concept for this map.

3. Build the map by placing the concepts according to their importance under the main concept, *matter*.
One way of arranging the concepts is shown in **Map B.**



4. Add linking words to give meaning to the arrangement of concepts.
When adding the links, be sure that each proposition makes sense. To distinguish concepts from links, place your concepts in circles, ovals, or rectangles, as shown in the maps. Then make cross-links. Cross-links are made of propositions and lines connecting concepts across the map. Links that apply in only one direction are indicated with an arrowhead. **Map C** is a finished map covering the main ideas listed in Step 1.

Making maps might seem difficult at first, but the process forces you to think about the meanings and relationships among the concepts. If you do not understand those relationships, you can get help early on.

Practice mapping by making concept maps about topics you know. For example, if you know a lot about a particular sport, such as basketball, or if you have a particular hobby, such as playing a musical instrument, you can use that topic to make a practice map. By perfecting your skills with information that you know very well, you will begin to feel more confident about making maps from the information in a chapter.

Remember, the time you devote to mapping will pay off when it is time to review for an exam.

Practice

1. Classify each of the following as either a concept or linking word(s).

a. classification ______________________

b. is classified as ______________________

c. forms ______________________

d. is described by ______________________

e. reaction ______________________

f. reacts with ______________________

g. metal ______________________

h. defines ______________________

2. Write three propositions from the information in **Map A.** ______________________

____

____

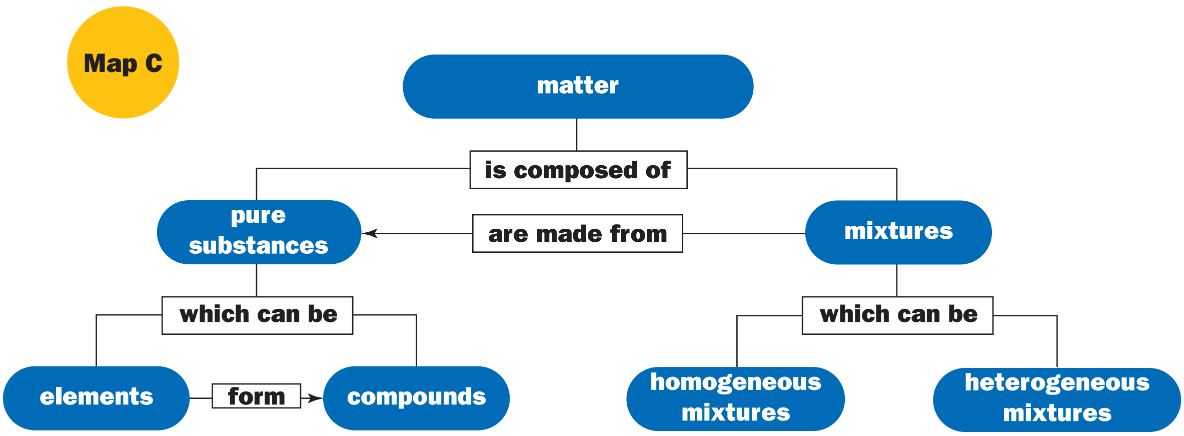
____

3. List two cross-links shown on **Map C.**

____

____

____



Making Power Notes

Power notes help you organize the chemical concepts you are studying by distinguishing main ideas from details. Similar to outlines, power notes are linear in form and provide you with a framework of important concepts. Power notes are easier to use than outlines because their structure is simpler. Using the power notes numbering system you assign a *1* to each main idea and a *2, 3,* or *4* to each detail.

Power notes are an invaluable asset to the learning process, and they can be used frequently throughout your chemistry course. You can use power notes to organize ideas while reading your text or to restructure your class notes for studying purposes.

To learn to make power notes, practice first by using single-word concepts and a subject you are especially interested in, such as animals, sports, or movies. As you become comfortable with structuring power notes, integrate their use into your study of chemistry. For an easier transition, start with a few boldfaced or italicized terms. Later you can strengthen your notes by expanding these single-word concepts into more-detailed phrases and sentences. Use the following general format to help you structure your power notes.

Power 1: Main idea
- **Power 2: Detail or support for power 1**
  - **Power 3: Detail or support for power 2**
    - **Power 4: Detail or support for power 3**

1. Pick a Power 1 word from the text.
The text you choose does not have to come straight from your chemistry textbook. You may be making power notes from your lecture notes or from an outside source. We'll use the term *atom* found in Chapter 3, Section 2 of your textbook.

Power 1: Atom

2. Using the text, select some Power 2 words to support your Power 1 word.
We'll use the terms *nucleus* and *electrons,* which are two parts of an atom.

Power 1: Atom
- **Power 2: Nucleus**
- **Power 2: Electrons**

3. Select some Power 3 words to support your Power 2 words.
We'll use the terms *positively charged* and *negatively charged,* two terms that describe the Power 2 words.

Power 1: Atom
- **Power 2: Nucleus**
  - **Power 3: Positively charged**
- **Power 2: Electrons**
  - **Power 3: Negatively charged**

4. Continue to add powers to support and detail the main idea as necessary.
There are no restrictions on how many power numbers you can use in your notes. If you have a main idea that requires a lot of support, add more powers to help you extend and organize your ideas. Be sure that words having the same power number have a similar relationship to the power above. Power 1 terms do not have to be related to each other. You can use power notes to organize the material in an entire section or chapter of your text. Doing so will provide you with an invaluable study guide for your classroom quizzes and tests.

Power 1: Atom
- **Power 2: Nucleus**
  - **Power 3: Positively charged**
  - **Power 3: Protons**
    - **Power 4: Positively charged**
  - **Power 3: Neutrons**
    - **Power 4: No charge**
- **Power 2: Electrons**
  - **Power 3: Negatively charged**

Practice

1. Use a periodic table and the power notes structure below to organize the following terms: *alkaline-earth metals, nonmetals, calcium, sodium, halogens, metals, alkali metals, chlorine, barium,* and *iodine.*

1 ____________________
- **2** ____________________
  - **3** ____________________
- **2** ____________________
  - **3** ____________________
  - **3** ____________________

1 ____________________
- **2** ____________________
  - **3** ____________________
  - **3** ____________________

Making Two-Column Notes

Two-column notes can be used to learn and review definitions of vocabulary terms, examples of multiple-step processes, or details of specific concepts. The two-column-note strategy is simple: write the term, main idea, step-by-step process, or concept in the left-hand column, and the definition, example, or detail on the right.

One strategy for using two-column notes is to organize main ideas and their details. The main ideas from your reading are written in the left-hand column of your paper and can be written as questions, key words, or a combination of both. Details describing these main ideas are then written in the right-hand column of your paper.

1. **Identify the main ideas.** The main ideas for a chapter are listed in the section objectives. However, you decide which ideas to include in your notes. For example, the table below shows some main ideas from the objectives in Chapter 5, Section 2.
   - Describe the locations in the periodic table and the general properties of the alkali metals, alkaline-earth metals, the halogens, and the noble gases.
2. **Divide a blank sheet of paper into two columns and write the main ideas in the left-hand column.** Summarize your ideas using quick phrases that are easy for you to understand and remember. Decide how many details you need for each main idea, and write that number in parentheses under the main idea.
3. **Write the detail notes in the right-hand column.** Be sure you list as many details as you designated in the main-idea column. The table below shows some details that correspond to the main ideas in Chapter 5, Section 2.

The two-column method of review is perfect whether you use it to study for a short quiz or for a test on the material in an entire chapter. Just cover the information in the right-hand column with a sheet of paper, and after reciting what you know, uncover the notes to check your answers. Then ask yourself what else you know about that topic. Linking ideas in this way will help you to gain a more complete picture of chemistry.

Main Idea	Detail Notes
• Alkali metals (4 details)	• Group 1 • highly reactive • ns^1 electron configuration • soft, silvery
• Alkaline-earth metals (4 details)	• Group 2 • reactive • ns^2 electron configuration • harder than alkali metals
• Halogens (3 details)	• Group 17 • reactive • nonmetallic
• Noble gases (3 details)	• Group 18 • low reactivity • stable ns^2np^6 configuration

Using the K/W/L Strategy

The K/W/L strategy stands for "what I **K**now—what I **W**ant to know—what I **L**earned." You start by brainstorming about the subject matter before reading the assigned material. Relating new ideas and concepts to those you have learned previously will help you better understand and apply the new knowledge you obtain. The section objectives throughout your textbook are ideal for using the K/W/L strategy.

1. **Read the section objectives.** You may also want to scan headings, boldfaced terms, and illustrations before reading. Here are two of the objectives from Chapter 1, Section 2 to use as an example.
   - Explain the gas, liquid, and solid states in terms of particles.
   - Distinguish between a mixture and a pure substance.
2. **Divide a sheet of paper into three columns, and label the columns "What I Know," "What I Want to Know," and "What I Learned."**
3. **Brainstorm about what you know about the information in the objectives, and write these ideas in the first column.** Because this chart is designed primarily to help you integrate your own knowledge with new information, it is not necessary to write complete sentences.
4. **Think about what you want to know about the information in the objectives, and write these ideas in the second column.** Include information from both the section objectives and any other objectives your teacher has given you.
5. **While reading the section or afterwards, use the third column to write down the information you learned.** While reading, pay close attention to any information about the topics you wrote in the "What I Want to Know" column. If you do not find all of the answers you are looking for, you may need to reread the section or reference a second source. Be sure to ask your teacher if you still cannot find the information after reading the section a second time.

It is also important to review your brainstormed ideas when you have completed reading the section. Compare your ideas in the first column with the information you wrote down in the third column. If you find that some of your brainstormed ideas are incorrect, cross them out. It is extremely important to identify and correct any misconceptions you had prior to reading before you begin studying for your test.

What I Know	What I Want to Know	What I Learned
• gas has no definite shape or volume • liquid has no definite shape, but has definite volume • solid has definite shape and volume • mixture is combination of substances • pure substance has only one component	• how gas, liquid, and solid states are related to particles • how mixtures and pure substances are different	• molecules in solid and liquid states are close together, but are far apart in gas state • molecules in solid state have fixed positions, but molecules in liquid and gas states can flow • mixtures are combinations of pure substances • pure substances have fixed compositions and definite properties

Using Sequencing/Pattern Puzzles

You can use pattern puzzles to help you remember sequential information. Pattern puzzles are not just a tool for memorization. They also promote a greater understanding of a variety of chemical processes, from the steps in solving a mass-mass stoichiometry problem to the procedure for making a solution of specified molarity.

1. **Write down the steps of a process in your own words.** For an example, we will use the process for converting the amount of a substance in moles to mass in grams. (See Sample Problem B on page 84.) On a sheet of notebook paper, write down one step per line, and do not number the steps. Also, do not copy the process straight from your textbook.

   Writing the steps in your own words promotes a more thorough understanding of the process. You may want to divide longer steps into two or three shorter steps.

- List the given and unknown information.
- Look at the periodic table to determine the molar mass of the substance.
- Write the correct conversion factor to convert moles to grams.
- Multiply the amount of substance by the conversion factor.
- Solve the equation and check your answer.

2. **Cut the sheet of paper into strips with only one step per strip of paper.** Shuffle the strips of paper so that they are out of sequence.

- Look at the periodic table to determine the molar mass of the substance.
- Solve the equation and check your answer.
- List the given and unknown information.
- Multiply the amount of substance by the conversion factor.
- Write the correct conversion factor to convert moles to grams.

3. **Place the strips in their proper sequence.** Confirm the order of the process by checking your text or your class notes.

- List the given and unknown information.
- Look at the periodic table to determine the molar mass of the substance.
- Write the correct conversion factor to convert moles to grams.
- Multiply the amount of substance by the conversion factor.
- Solve the equation and check your answer.

Pattern puzzles are especially helpful when you are studying for your chemistry tests. Before tests, use your puzzles to practice sequencing and to review the steps of chemistry processes. You and a classmate can also take turns creating your own pattern puzzles of different chemical processes and putting each other's puzzles in the correct sequence. Studying with a classmate in this manner will help make studying fun and will enable you to help each other.

Other Reading Strategies

Brainstorming

Brainstorming is a strategy that helps you recognize and evaluate the knowledge you already have before you start reading. It works well individually or in groups. When you brainstorm, you start with a central term or idea, then quickly list all the words, phrases, and other ideas that you think are related to it.

Because there are no "right" or "wrong" answers, you can use the list as a basis for classifying terms, developing a general explanation, or speculating about new relationships. For example, you might brainstorm a list of terms related to the word *element* before you read Chapter 1, Section 2. The list might include gold, metals, chemicals, silver, carbon, oxygen, and water. As you read the textbook, you might decide that some of the terms you listed are *not* elements. Later, you might use that information to help you distinguish between elements and compounds.

Building/Interpreting Vocabulary

Using a dictionary to look up the meanings of prefixes and suffixes as well as word origins and meanings helps you build your vocabulary and interpret what you read. If you know the meaning of prefixes like *kilo-* (one thousand) and *milli-* (one thousandth), you have a good idea what kilograms, kilometers, milligrams, and millimeters are and how they are different. (See page 35 for a list of SI Prefixes.)

Knowledge of prefixes, suffixes, and word origins can help you understand the meaning of new words. For example, if you know the suffix *-protic* comes from the same word as *proton,* it will help you understand what monoprotic and polyprotic acids are (see page 479).

Reading Hints

Reading hints help you identify and bookmark important charts, tables, and illustrations for easy reference. For example, you may want to use a self-adhesive note to bookmark the periodic table on pages 140–141 or on the inside back cover of your book so you can easily locate it and use it for reference as you study different aspects of chemistry and solve problems involving elements and compounds.

Interpreting Graphic Sources of Information

Charts, tables, photographs, diagrams, and other illustrations are graphic, or visual, sources of information. The labels and captions, together with the illustrations help you make connections between the words and the ideas presented in the text.

Reading Response Logs

Keeping a reading response log helps you interpret what you read and gives you a chance to express your reactions and opinions about what you have read. Draw a vertical line down the center of a piece of paper. In the left-hand column, write down or make notes about passages you read to which you have reactions, thoughts, feelings, questions, or associations. In the right-hand column, write what those reactions, thoughts, feelings, questions, or associations are. For example, you might keep a reading response log when studying about Nuclear Energy in Chapter 21.

Other Studying Strategies

▶ Comparing and Contrasting

Comparing and contrasting is a strategy that helps you note similarities and differences between two or more objects or events. When you determine similarities, you are comparing. When you determine differences, you are contrasting.

You can use comparing and contrasting to help you classify objects or properties, differentiate between similar concepts, and speculate about new relationships. For example, as you read Chapter 1 you might begin to make a table in which you compare and contrast metals, nonmetals, and metalloids. As you continue to learn about these substances in Chapters 4 and 5, you can add to your table, giving you a better understanding of the similarities and differences among elements.

▶ Identifying Cause and Effect

Identifying causes and effects as you read helps you understand the material and builds logical reasoning skills. An effect is an event or the result of some action. A cause is the reason the event or action occurred. Signal words, such as *because, so, since, therefore, as a result,* and *depends on,* indicate a cause-and-effect relationship.

You can use arrows to show cause and effect. For example, you might write this cause-and-effect relationship as you read Chapter 11, Section 2: At constant pressure, increase in temperature (cause) → increase in gas volume (effect).

▶ Making a Prediction Guide

A prediction guide is a list of statements about which you express and try to justify your opinions based on your current knowledge. After reading the material, you re-evaluate your opinion in light of what you learned. Using prediction guides helps you evaluate your knowledge, identify assumptions you may have that could lead to mistaken conclusions, and form an idea of expected results.

1. **Read the statements your teacher writes on the board.** For example, look at the five statements from Dalton's theory listed on page 68 of your textbook.
2. **Decide whether you think each statement is true or false and discuss reasons why you think so.**
3. **After reading the section, re-evaluate your opinion of each statement. Discuss why your opinion changed or remained the same. Find passages in the text that account for the change of reinforcement of your opinions.** For example, you might have agreed with all five statements from Dalton's theory before reading the text. Then, after reading about atoms and subatomic particles, you might have changed your opinion about the first statement.

Cooperative Learning Techniques

▶ Reading with a Partner

Reading with a partner is a strategy that can help you understand what you read and point out where more explanation is needed.

1. **First read the text silently by yourself. Use self-adhesive notes to mark those parts of the text that you do not understand.** For example, you might have difficulty with some of the material about quantum numbers in Section 2 of Chapter 4, while another student understands quantum numbers but has trouble with electron configurations in Section 3.
2. **Work with a partner to discuss the passages each of you marked.** Take turns listening and trying to clarify the difficult passages for each other. Together, study the related tables and illustrations and explain to each other how they relate to the text.
3. **For concepts that need further explanation, work together to formulate questions for class discussion or for your teacher to answer.**

▶ Using L.I.N.K.

The L.I.N.K. strategy stands for **L**ist, **I**nquire, **N**otes, **K**now. It is similar to the K/W/L strategy, but you work as a class or in groups.

1. **Brainstorm all the words, phrases, and ideas associated with the term your teacher provides.** Volunteers can keep track of contributions on the board or on a separate sheet of paper.
2. **Your teacher will direct you in a class or group discussion about the words and ideas listed.** Now is the time to inquire, or ask your teacher and other students for clarification of the listed ideas.
3. **At the end of the discussion, make notes about everything you can remember.** Look over your notes to see if you have left anything out.
4. **See what you now know about the given concept based on your own experience and the discussion.**

▶ Summarizing/Paired Summarizing

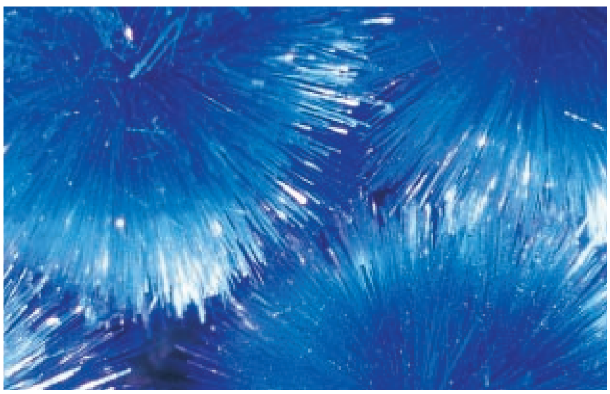
A summary is a brief statement of main ideas or important concepts. Making a summary of what you have read provides you with a way to review what you have learned, see what information needs further clarification, and helps you make connections to previously studied material.

Paired summarizing helps strengthen your ability to read, listen, and understand. It is especially useful when a section of text has several subdivisions, each dealing with different concepts, such as Chapter 2, Section 3 in your textbook.

1. **First read the material silently by yourself.**
2. **Then you and your partner take turns being the "listener" and the "reteller."** The reteller summarizes the material for the listener, who does not interrupt until the reteller has finished. If necessary, the reteller may consult the text, and the listener may ask for clarification. The listener then states any inaccuracies or omissions made by the reteller.
3. **Work together to refine the summary.** Make sure the summary states the important ideas in a clear and concise manner.

▶ Discussing Ideas

Discussing ideas with a partner or in a group before you read is a strategy that can help you broaden your knowledge base and decide what concepts to focus on as you are reading. Discussing ideas after you have read a section or chapter can help you check your understanding, clarify difficult concepts, and lead you to speculate about new ideas.



Graphing Calculator Technology

Charts, graphs, and data analysis are essential elements of chemistry. To be successful in your study of chemistry, you must know how to make and interpret graphs and must understand the relationships between different variables.

Your graphing calculator can be a powerful tool for analyzing chemical data. In addition to using your calculator to organize and graph data, you can program your calculator to perform specialized functions.

In the text, you can use your graphing calculator to help solve the Chapter Review exercises. In addition, specific Graphing Calculator exercises referred to in the Chapter Reviews will help you explore the capabilities of your calculator while enhancing your understanding of mathematical relationships that are important in chemistry. The CBL 2™ and LabPro® probeware experiments will enable you to become adept at analyzing experimental data.

Graphing Calculator Exercises

The **go.hrw.com** site provides downloadable programs for the TI-83 Plus and the TI-84 Plus graphing calculator families. These programs include data sets to analyze. Using these programs will improve your ability to handle scientific data.

You will learn to use your calculator to graph data. Then, you will interpret the graphs and will extract the information required to answer the questions in the exercises. You will gain experience with simple linear relationships, such as the relationship between energy and temperature, and complex relationships, such as the relationship between pH and titrant volume, which is represented by a titration curve.

To solve the Graphing Calculator exercises in the Chapter Reviews and to download the programs, you will need

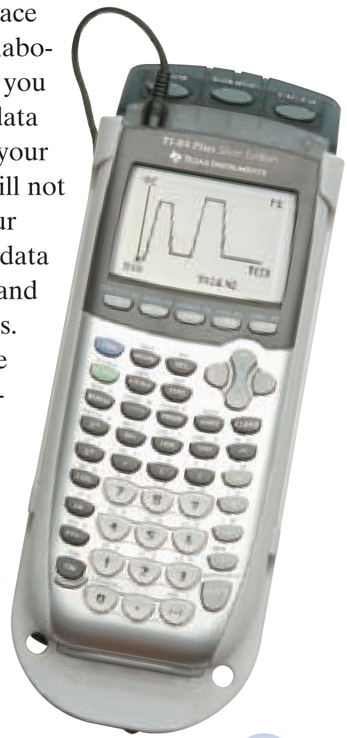
- a graphing calculator (TI-83 Plus or TI-84 Plus)
- appropriate TI Connectivity computer-to-calculator cable and TI Connect™ software
- a computer that has Internet access

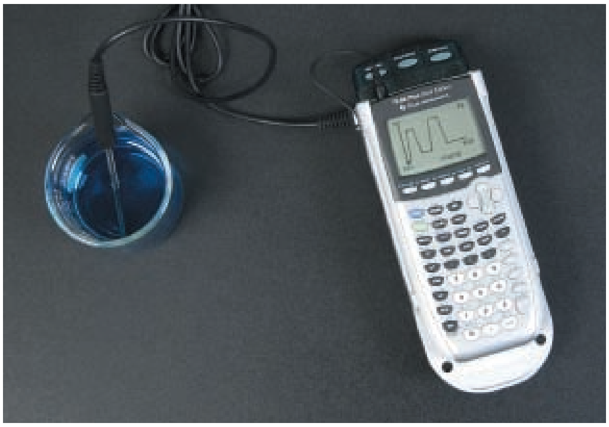
The detailed instructions on how to download the calculator programs can be found at **go.hrw.com** (keyword **HC6 CALC**).

Calculator-Based Laboratories

Analyzing your data properly is as important as using good experimental technique in the chemistry lab. Your results will be meaningless if you do not know how to interpret them. The Calculator-Based Laboratory 2™ (CBL 2™) data-collection interface by Texas Instruments and the Vernier LabPro® data-collection interface by Vernier Software & Technology can simplify the process of obtaining and analyzing experimental data.

The data-collection interface mimics expensive electronic laboratory equipment and allows you to collect real experimental data that are stored directly onto your calculator. As a result, you will not have to record and graph your data manually. Instead, your data are automatically tabulated, and you can view real-time graphs. Consequently, you have more time to interpret your experimental results. The interface and probes allow you to collect experimental data, and with the DataMate™ App the information coming from the probes is automatically recognized.





The CBL 2™ set-up is shown above. In some labs, you will use a temperature sensor, a pressure sensor, a voltage sensor, and a colorimeter to collect data. You will then analyze the data on your calculator to obtain the results of your experiment. To perform the calculator-based probeware experiments, you will need

- a graphing calculator (TI-83 Plus or TI-84 Plus)
- the CBL 2™ or LabPro® data collection interface
- the appropriate Texas Instruments or Vernier probe for the experiment
- the probeware experiment

The probeware experiments are available at **go.hrw.com** (keyword **HC6 CBL**). For additional information about the CBL 2 and LabPro hardware or software, visit **education.ti.com** or **www.vernier.com**

Making the Most of your Calculator

In addition to organizing and graphing scientific data, Texas Instruments calculators have other applications that are available for use in chemistry.

The Periodic Table App provides an electronic version of the periodic table as well as information about the elements.

The Science Tool App provides various physical constants and allows you to convert between different units and to determine the number of significant figures for calculations.

Additional information about these and other applications for your calculator is available at **education.ti.com.** These applications are standard for some models of TI calculators and can be obtained from Texas Instruments and downloaded onto other models.

Troubleshooting

- Calculator instructions in the *Modern Chemistry* program are written for the TI-84 Plus. You may use other graphing calculators, but some of the programs and instructions may require minor adjustments.
- Calculator-based probeware experiments are written for the CBL 2™ and LabPro® interfaces. Older versions of the experiments for the original CBL are available at **go.hrw.com.**
- The DataMate™ App is not compatible with the original CBL™ system. For more information and user manuals for the CBL 2™ and the CBL™ systems, visit **education.ti.com/guides**
- If you have problems loading programs or applications onto your calculator, you may need to clear programs or other data from your calculator's memory.
- Always make sure that you are downloading the version of the software that is correct for your calculator and that you have the latest operating system for your calculator and your CBL 2™.
- If you need additional help, Texas Instruments and Vernier Software & Technology can provide technical support. Contact TI at **education.ti.com** or 1-800-TI-CARES and Vernier at **info@vernnier.com** or 1-888-837-6437.

go.hrw.com
Keyword: **HC6 CALC**
This Web site contains links for downloading programs and applications that you will need for the Graphing Calculator exercises.

Problem Bank

Conversions: Chap. 2, Sec. 2

Converting Simple SI Units

1. State the following measured quantities in the units indicated.
   a. 5.2 cm of magnesium ribbon in millimeters
   b. 0.049 kg of sulfur in grams
   c. 1.60 mL of ethanol in microliters
   d. 0.0025 g of vitamin A in micrograms
   e. 0.020 kg of tin in milligrams
   f. 3 kL of saline solution in liters
2. State the following measured quantities in the units indicated.
   a. 150 mg of aspirin in grams
   b. 2500 mL of hydrochloric acid in liters
   c. 0.5 g of sodium in kilograms
   d. 55 L of carbon dioxide gas in kiloliters
   e. 35 mm in centimeters
   f. 8740 m in kilometers
   g. 209 nm in millimeters
   h. 500 000 μg in kilograms
3. The greatest distance between Earth and the sun during Earth's revolution is 152 million kilometers. What is this distance in megameters?
4. How many milliliters of water will it take to fill a 2.00 L bottle that already contains 1.87 L of water?
5. A piece of copper wire is 150 cm long. How long is the wire in millimeters? How many 50 mm segments of wire can be cut from the length?
6. The ladle at an iron foundry can hold 8500 kg of molten iron; 646 metric tons of iron are needed to make rails. How many ladlefuls of iron will it take to make 646 metric tons of iron? (1 metric ton = 1000 kg)

Converting Derived SI Units

7. State the following measured quantities in the units indicated.
   a. 310 000 cm^3 of concrete in cubic meters
   b. 6.5 m^2 of steel sheet in square centimeters
   c. 0.035 m^3 of chlorine gas in cubic centimeters
   d. 0.49 cm^2 of copper in square millimeters
   e. 1200 dm^3 of acetic acid solution in cubic meters
   f. 87.5 mm^3 of actinium in cubic centimeters
   g. 250 000 cm^2 of polyethylene sheet in square meters
8. How many palisade cells from plant leaves would fit in a volume of 1.0 cm^3 of cells if the average volume of a palisade cell is 0.0147 mm^3?

Mixed Review

9. Convert each of the following quantities to the required unit.
   a. 12.75 Mm to kilometers
   b. 277 cm to meters
   c. 30 560 m^2 to hectares (1 ha = 10 000 m^2)
   d. 81.9 cm^2 to square meters
   e. 300 000 km to megameters
10. Convert each of the following quantities to the required unit.
   a. 0.62 km to meters
   b. 3857 g to milligrams
   c. 0.0036 mL to microliters
   d. 0.342 metric tons to kg (1 metric ton = 1000 kg)
   e. 68.71 kL to liters
11. Convert each of the following quantities to the required unit.
   a. 856 mg to kilograms
   b. 1 210 000 μg to kilograms
   c. 6598 μL to cubic centimeters (1 mL = 1 cm^3)
   d. 80 600 nm to millimeters
   e. 10.74 cm^3 to liters
12. Convert each of the following quantities to the required unit.
   a. 7.93 L to cubic centimeters
   b. 0.0059 km to centimeters
   c. 4.19 L to cubic decimeters
   d. 7.48 m^2 to square centimeters
   e. 0.197 m^3 to liters
13. An automobile uses 0.05 mL of oil for each kilometer it is driven. How much oil in liters is consumed if the automobile is driven 20 000 km?
14. How many microliters are there in a volume of 370 mm^3 of cobra venom?
15. A baker uses 1.5 tsp of vanilla extract in each cake. How much vanilla extract in liters should the baker order to make 800 cakes? (1 tsp = 5 mL)
16. A person drinks eight glasses of water each day, and each glass contains 300 mL. How many liters of water will that person consume in a year? What is the mass of this volume of water in kilograms? (Assume one year has 365 days and the density of water is 1.00 kg/L.)
17. At the equator Earth rotates with a velocity of about 465 m/s.

a. What is this velocity in kilometers per hour?
b. What is this velocity in kilometers per day?

18. A chemistry teacher needs to determine what quantity of sodium hydroxide to order. If each student will use 130 g and there are 60 students, how many kilograms of sodium hydroxide should the teacher order?

19. The teacher in item 18 also needs to order plastic tubing. If each of the 60 students needs 750 mm of tubing, what length of tubing in meters should the teacher order?

20. Convert the following to the required units.
a. 550 μL/h to milliliters per day
b. 9.00 metric tons/h to kilograms per minute
c. 3.72 L/h to cubic centimeters per minute
d. 6.12 km/h to meters per second

21. Express the following in the units indicated.
a. 2.97 kg/L as grams per cubic centimeter
b. 4128 g/dm^2 as kilograms per square centimeter
c. 5.27 g/cm^3 as kilograms per cubic decimeter
d. 6.91 kg/m^3 as milligrams per cubic millimeter

22. A gas has a density of 5.56 g/L.
a. What volume in milliliters would 4.17 g of this gas occupy?
b. What would be the mass in kilograms of 1 m^3 of this gas?

23. The average density of living matter on Earth's land areas is 0.10 g/cm^2. What mass of living matter in kilograms would occupy an area of 0.125 ha?

24. A textbook measures 250. mm long, 224 mm wide, and 50.0 mm thick. It has a mass of 2.94 kg.
a. What is the volume of the book in cubic meters?
b. What is the density of the book in grams per cubic centimeter?
c. What is the area of one cover in square meters?

25. A glass dropper delivers liquid so that 25 drops equal 1.00 mL.
a. What is the volume of one drop in milliliters?
b. How many milliliters are in 37 drops?
c. How many drops would be required to get 0.68 L?

26. Express each of the following in kilograms and grams.
a. 504 700 mg
b. 9 200 000 μg
c. 122 mg
d. 7195 cg

27. Express each of the following in liters and milliliters.
a. 582 cm^3
b. 0.0025 m^3
c. 1.18 dm^3
d. 32 900 μL

28. Express each of the following in grams per liter and kilograms per cubic meter.
a. 1.37 g/cm^3
b. 0.692 kg/dm^3
c. 5.2 kg/L
d. 38 000 g/m^3
e. 5.79 mg/mm^3
f. 1.1 μg/mL

29. An industrial chemical reaction is run for 30.0 h and produces 648.0 kg of product. What is the average rate of product production in the stated units?
a. grams per minute
b. kilograms per day
c. milligrams per millisecond

30. What is the speed of a car in meters per second when it is moving at 100. km/h?

31. A heater gives off energy as heat at a rate of 330 kJ/min. What is the rate of energy output in kilocalories per hour? (1 cal = 4.184 J)

32. The instructions on a package of fertilizer tell you to apply it at the rate of 62 g/m^2. How much fertilizer in kilograms would you need to apply to 1.0 ha? (1 ha = 10 000 m^2)

33. A water tank leaks water at the rate of 3.9 mL/h. If the tank is not repaired, what volume of water in liters will it leak in a year? Show your setup for solving this. Hint: Use one conversion factor to convert hours to days and another to convert days to years, and assume that one year has 365 days.

34. A nurse plans to give flu injections of 50 μL each from a bottle containing 2.0 mL of vaccine. How many doses are in the bottle?

Significant Figures: Chap. 2, Sec. 2

35. Determine the number of significant figures in the following measurements.
a. 640 cm^3
b. 200.0 mL
c. 0.5200 g
d. 1.005 kg
e. 10 000 L
f. 20.900 cm
g. 0.000 000 56 g/L
h. 0.040 02 kg/m^3
i. 790 001 cm^2
j. 665.000 kg•m/s^2

36. Perform the following calculations, and express the result in the correct units and number of significant figures.
a. 47.0 m ÷ 2.2 s
b. 140 cm × 35 cm
c. 5.88 kg ÷ 200 m^3
d. 0.00 50 m^2 × 0.042 m
e. 300.3 L ÷ 180. s
f. 33.00 cm^2 × 2.70 cm
g. 35 000 kJ ÷ 0.250 min

37. Perform the following calculations and express the results in the correct units and number of significant figures.
a. 22.0 m + 5.28 m + 15.5 m
b. 0.042 kg + 1.229 kg + 0.502 kg
c. 170 cm^2 + 3.5 cm^2 − 28 cm^2
d. 0.003 L + 0.0048 L + 0.100 L
e. 24.50 dL + 4.30 dL + 10.2 dL
f. 3200 mg + 325 mg − 688 mg
g. 14 000 kg + 8000 kg + 590 kg

Mixed Review

38. Determine the number of significant figures in the following measurements.
a. 0.0120 m
b. 100.5 mL
c. 101 g
d. 350 cm^2
e. 0.97 km
f. 1000 kg
g. 180. mm
h. 0.4936 L
i. 0.020 700 s

39. Round the following quantities to the specified number of significant figures.
a. 5 487 129 m to three significant figures
b. 0.013 479 265 mL to six significant figures

c. 31 947.972 cm^2 to four significant figures
d. 192.6739 m^2 to five significant figures
e. 786.9164 cm to two significant figures
f. 389 277 600 J to six significant figures
g. 225 834.762 cm^3 to seven significant figures

40. Perform the following calculations, and express the answer in the correct units and number of significant figures.
a. 651 cm × 75 cm
b. 7.835 kg ÷ 2.5 L
c. 14.75 L ÷ 1.20 s
d. 360 cm × 51 cm × 9.07 cm
e. 5.18 m × 0.77 m × 10.22 m
f. 34.95 g ÷ 11.169 cm^3

41. Perform the following calculations, and express the answer in the correct units and number of significant figures.
a. 7.945 J + 82.3 J − 0.02 J
b. 0.0012 m − 0.000 45 m − 0.000 11 m
c. 500 g + 432 g + 2 g
d. 31.2 kPa + 0.0035 kPa − 0.147 kPa
e. 312 dL − 31.2 dL − 3.12 dL
f. 1701 kg + 50 kg + 43 kg

42. A rectangle measures 87.59 cm by 35.1 mm. Express its area with the proper number of significant figures in the specified unit.
a. in cm^2
b. in mm^2
c. in m^2

43. A box measures 900. mm by 31.5 mm by 6.3 cm. State its volume with the proper number of significant figures in the specified unit.
a. in cm^3
b. in m^3
c. in mm^3

44. A 125 mL sample of liquid has a mass of 0.16 kg. What is the density of the liquid in the following measurements?
a. kg/m^3
b. g/mL
c. kg/dm^3

45. Perform the following calculations, and express the results in the correct units and with the proper number of significant figures.
a. 13.75 mm × 10.1 mm × 0.91 mm
b. 89.4 cm^2 × 4.8 cm
c. 14.9 m^3 ÷ 3.0 m^2
d. 6.975 m × 30 m × 21.5 m

46. What is the volume of a region of space that measures 752 m × 319 m × 110 m? Give your answer in the correct unit and with the proper number of significant figures.

47. Perform the following calculations, and express the results in the correct units and with the proper number of significant figures.
a. 7.382 g + 1.21 g + 4.7923 g
b. 51.3 mg + 83 mg − 34.2 mg
c. 0.007 L − 0.0037 L + 0.012 L
d. 253.05 cm^2 + 33.9 cm^2 + 28 cm^2
e. 14.77 kg + 0.086 kg − 0.391 kg
f. 319 mL + 13.75 mL + 20. mL

48. A container measures 30.5 mm × 202 mm × 153 mm. When it is full of a liquid, it has a mass of 1.33 kg. When it is empty, it has a mass of 0.30 kg. What is the density of the liquid in kilograms per liter?

49. If 7.76 km of wire has a mass of 3.3 kg, what is the mass of the wire in g/m? What length in meters would have a mass of 1.0 g?

50. A container of plant food recommends an application rate of 52 kg/ha. If the container holds 10 kg of plant food, how many square meters will it cover? (1 ha = 10 000 m^2)

51. A chemical process produces 974 550 kJ of energy as heat in 37.0 min. What is the rate in kilojoules per minute? What is the rate in kilojoules per second?

52. A water pipe fills a container that measures 189 cm × 307 cm × 272 cm in 97 s.
a. What is the volume of the container in cubic meters?
b. What is the rate of flow in the pipe in liters per minute?
c. What is the rate of flow in cubic meters per hour?

53. Perform the following calculations, and express the results in the correct units and with the proper number of significant figures. Note, in problems with multiple steps, it is better to perform the entire calculation and then round to significant figures.
a. (0.054 kg + 1.33 kg) × 5.4 m^2
b. 67.35 cm^2 ÷ (1.401 cm − 0.399 cm)
c. 4.198 kg × (1019 m^2 − 40 m^2) ÷ (54.2 s × 31.3 s)
d. 3.14159 m × (4.17 m + 2.150 m)
e. 690 000 m ÷ (5.022 h − 4.31 h)
f. (6.23 cm + 3.111 cm − 0.05 cm) × 14.99 cm

Scientific Notation: Chap. 2, Sec. 3

Converting Quantities to Scientific Notation

54. Express the following quantities in scientific notation.
a. 8 800 000 000 m
b. 0.0015 kg
c. 0.000 000 000 06 kg/m^3
d. 8 002 000 Hz
e. 0.009 003 A
f. 70 000 000 000 000 000 km
g. 6028 L
h. 0.2105 g
i. 600 005 000 kJ/h
j. 33.8 m^2

Calculating with Quantities in Scientific Notation

55. Carry out the following calculations. Express the results in scientific notation and with the correct number of significant figures.
a. 4.74×10^4 km + 7.71×10^3 km + 1.05×10^3 km
b. 2.75×10^{-4} m + 8.03×10^{-5} m + 2.122×10^{-3} m
c. 4.0×10^{-5} m^3 + 6.85×10^{-6} m^3 − 1.05×10^{-5} m^3
d. 3.15×10^2 mg + 3.15×10^3 mg + 3.15×10^4 mg
e. 3.01×10^{22} atoms + 1.19×10^{23} atoms + 9.80×10^{21} atoms
f. 6.85×10^7 nm + 4.0229×10^8 nm − 8.38×10^6 nm

56. Carry out the following computations, and express the result in scientific notation.
   a. 7.20×10^3 cm $\times$ 8.08×10^3 cm
   b. 3.7×10^4 mm $\times$ 6.6×10^4 mm $\times$ 9.89×10^3 mm
   c. 8.27×10^2 m $\times$ 2.5×10^{-3} m $\times$ 3.00×10^{-4} m
   d. 4.44×10^{-35} m $\times$ 5.55×10^{19} m $\times$ 7.69×10^{-12} kg
   e. 6.55×10^4 dm $\times$ 7.89×10^9 dm $\times$ 4.01893×10^5 dm

57. Carry out the following computations, and express the result in scientific notation.
   a. 2.290×10^7 cm $\div$ 4.33×10^3 s
   b. 1.788×10^{-5} L $\div$ 7.111×10^{-3} m^2
   c. 5.515×10^4 L $\div$ 6.04×10^3 km
   d. 3.29×10^{-4} km $\div$ 1.48×10^{-2} min
   e. 4.73×10^{-4} g $\div$ (2.08×10^{-3} km $\times$ 5.60×10^{-4} km)

Mixed Review

58. Express the following quantities in scientific notation.
   a. 158 000 km
   b. 0.000 009 782 L
   c. 837 100 000 cm^3
   d. 6 500 000 000 mm^2
   e. 0.005 93 g
   f. 0.000 000 006 13 m
   g. 12 552 000 J
   h. 0.000 008 004 g/L
   i. 0.010 995 kg
   j. 1 050 000 000 Hz

59. Perform the following calculations, and express the result in scientific notation with the correct number of significant figures.
   a. 2.48×10^2 kg + 9.17×10^3 kg + 7.2×10^1 kg
   b. 4.07×10^{-5} mg + 3.966×10^{-4} mg + 7.1×10^{-2} mg
   c. 1.39×10^4 m^3 + 6.52×10^2 m^3 − 4.8×10^3 m^3
   d. 7.70×10^{-9} m − 3.95×10^{-8} m + 1.88×10^{-7} m
   e. 1.111×10^5 J + 5.82×10^4 J + 3.01×10^6 J
   f. 9.81×10^{27} molecules + 3.18×10^{25} molecules − 2.09×10^{26} molecules
   g. 1.36×10^7 cm + 3.456×10^6 cm − 1.01×10^7 cm + 5.122×10^5 cm

60. Perform the following computations, and express the result in scientific notation with the correct number of significant figures.
   a. 1.54×10^{-1} L $\div$ 2.36×10^{-4} s
   b. 3.890×10^4 mm $\times$ 4.71×10^2 mm^2
   c. 9.571×10^3 kg $\div$ 3.82×10^{-1} m^2
   d. 8.33×10^3 km $\div$ 1.97×10^2 s
   e. 9.36×10^2 m $\times$ 3.82×10^3 m $\times$ 9.01×10^{-1} m
   f. 6.377×10^4 J $\div$ 7.35×10^{-3} s

61. Your electric company charges you for the electric energy you use, measured in kilowatt-hours (kWh). One kWh is equivalent to 3 600 000 J. Express this quantity in scientific notation.

62. The pressure in the deepest part of the ocean is 11 200 000 Pa. Express this pressure in scientific notation.

63. Convert 1.5 km to millimeters, and express the result in scientific notation.

64. Light travels at a speed of about 300 000 km/s.
   a. Express this value in scientific notation.
   b. Convert this value to meters per hour.
   c. What distance in centimeters does light travel in 1 μs?

65. There are 7.11×10^{24} molecules in 100.0 cm^3 of a certain substance.
   a. What is the number of molecules in 1.09 cm^3 of the substance?
   b. What would be the number of molecules in 2.24×10^4 cm^3 of the substance?
   c. What number of molecules are in 9.01×10^{-6} cm^3 of the substance?

66. The number of transistors on a particular integrated circuit is 3 578 000, and the integrated circuit measures 9.5 mm $\times$ 8.2 mm.
   a. What is the area occupied by each transistor?
   b. Using your answer from (a), how many transistors could be formed on a silicon sheet that measures 353 mm $\times$ 265 mm?

67. A solution has 0.0501 g of a substance in 1.00 L. Express this concentration in grams per microliter.

68. Cesium atoms are the largest of the naturally occurring elements. They have a diameter of 5.30×10^{-10} m. Calculate the number of cesium atoms that would have to be lined up to give a row of cesium atoms 2.54 cm (1 in.) long.

69. The neutron has a volume of approximately 1.4×10^{-44} m^3 and a mass of 1.675×10^{-24} g. Calculate the density of the neutron in g/m^3. What is the mass of 1.0 cm^3 of neutrons in kilograms?

70. The pits in a compact disc are some of the smallest things ever mass-produced mechanically by humans. These pits represent the 1s and 0s of digital information on a compact disc. These pits are only 1.6×10^{-8} m deep (1/4 the wavelength of red laser light). How many of these pits would have to be stacked on top of each other to make a hole 0.305 m deep?

71. 22 400 mL of oxygen gas contains 6.022×10^{23} oxygen molecules at 0°C and standard atmospheric pressure.
   a. How many oxygen molecules are in 0.100 mL of gas?
   b. How many oxygen molecules are in 1.00 L of gas?
   c. What is the average space in milliliters occupied by one oxygen molecule?

72. The mass of the atmosphere is calculated to be 5.136×10^{18} kg, and there are 6 500 000 000 people living on Earth. Calculate the following values.
   a. The mass of atmosphere in kilograms per person.
   b. The mass of atmosphere in metric tons per person.
   c. If the number of people increases to 9 500 000 000, what is the mass in kilograms per person?

73. The mass of the sun is 1.989×10^{30} kg, and the mass of Earth is 5.974×10^{24} kilograms. How many Earths would be needed to equal the mass of the sun?

74. A new landfill has dimensions of 2.3 km $\times$ 1.4 km $\times$ 0.15 km.

a. What is the volume in cubic kilometer?
b. What is the volume in cubic meters?
c. If 250 000 000 objects averaging 0.060 m^3 each are placed into the landfill each year, how many years will it take to fill the landfill?

75. A dietary calorie (C) is exactly equal to 1000 cal. If your daily intake of food gives you 2400 C, what is your intake in joules per day? (1 cal = 4.184 J)

Four Steps for Solving Quantitative Problems: Chap. 2, Sec. 3

76. Gasoline has a density of 0.73 g/cm^3. How many liters of gasoline would be required to increase the mass of an automobile from 1271 kg to 1305 kg?

77. A swimming pool measures 9.0 m long by 3.5 m wide by 1.75 m deep. What mass of water in metric tons (1 metric ton = 1000 kg) does the pool contain when filled? The density of the water in the pool is 0.997 g/cm^3.

78. A tightly packed box of crackers contains 250 g of crackers and measures 7.0 cm × 17.0 cm × 19.0 cm. What is the average density in kilograms per liter of the crackers in the package? Assume that the unused volume is negligible.

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Solve these problems by using the Four Steps for Solving Quantitative Problems.

79. The aluminum foil on a certain roll has a total area of 18.5 m^2 and a mass of 1275 g. Using a density of 2.7 g per cubic centimeter for aluminum, determine the thickness in millimeters of the aluminum foil.

80. If a liquid has a density of 1.17 g/cm^3, how many liters of the liquid have a mass of 3.75 kg?

81. A stack of 500 sheets of paper measuring 28 cm × 21 cm is 44.5 mm high and has a mass of 2090 g. What is the density of the paper in grams per cubic centimeter?

82. A triangular-shaped piece of a metal has a mass of 6.58 g. The triangle is 0.560 mm thick and measures 36.4 mm on the base and 30.1 mm in height. What is the density of the metal in grams per cubic centimeter?

83. A packing crate measures 0.40 m × 0.40 m × 0.25 m. You must fill the crate with boxes of cookies that each measure 22.0 cm × 12.0 cm × 5.0 cm. How many boxes of cookies can fit into the crate?

84. Calculate the unknown quantities in the following table. Use the following relationships for volumes of the various shapes.

Volume of a cube = $l \times l \times l$
Volume of a rectangle = $l \times w \times h$
Volume of a sphere = $4/3\pi r^3$
Volume of a cylinder = $\pi r^2 \times h$

D	*m*	*V*	*Shape*	*Dimensions*
a. 2.27 g/cm^3	3.93 kg	? L	cube	? m × ? m × ? m
b. 1.85 g/cm^3	? g	? cm^3	rectangle	33 mm × 21 mm × 7.2 mm
c. 3.21 g/L	? kg	? dm^3	sphere	3.30 m diameter
d. ? g/cm^3	497 g	? m^3	cylinder	7.5 cm diameter × 12 cm
e. 0.92 g/cm^3	? kg	? cm^3	rectangle	3.5 m × 1.2 m × 0.65 m

85. When a sample of a metal alloy that has a mass of 9.65 g is placed into a graduated cylinder containing water, the volume reading in the cylinder increases from 16.0 mL to 19.5 mL. What is the density of the alloy sample in grams per cubic centimeter?

86. Pure gold can be made into extremely thin sheets called gold leaf. Suppose that 50. kg of gold is made into gold leaf having an area of 3620 m^2. The density of gold is 19.3 g/cm^3.
a. How thick in micrometers is the gold leaf?
b. A gold atom has a radius of 1.44×10^{-10} m. How many atoms thick is the gold leaf?

87. A chemical plant process requires that a cylindrical reaction tank be filled with a certain liquid in 238 s. The tank is 1.2 m in diameter and 4.6 m high. What flow rate in liters per minute is required to fill the reaction tank in the specified time?

88. The radioactive decay of 2.8 g of plutonium-238 generates 1.0 joule of energy as heat every second. Plutonium has a density of 19.86 g/cm^3. How many calories (1 cal = 4.184 J) of energy as heat will a rectangular piece of plutonium that is 4.5 cm × 3.05 cm × 15 cm generate per hour?

89. The mass of Earth is 5.974×10^{24} kg. Assume that Earth is a sphere of diameter 1.28×10^4 km and calculate the average density of Earth in g/cm^3.

90. What volume of magnesium in cubic centimeters would have the same mass as 1.82 dm^3 of platinum? The density of magnesium is 1.74 g/cm^3, and the density of platinum is 21.45 g/cm^3.

91. A roll of transparent tape has 66 m of tape on it. If an average of 5.0 cm of tape is needed each time the tape is used, how many uses can you get from a case of tape containing 24 rolls?

92. An automobile can travel 38 km on 4.0 L of gasoline. If the automobile is driven 75% of the days in a year and the average distance traveled each day is 86 km, how many liters of gasoline will be consumed in one year (assume the year has 365 days)?

93. A hose delivers water to a swimming pool that measures 9.0 m long by 3.5 m wide by 1.75 m deep. It requires 97 h to fill the pool. At what rate in liters per minute will the hose fill the pool?

94. Automobile batteries are filled with a solution of sulfuric acid, which has a density of 1.285 g/cm^3. The solution used to fill the battery is 38% (by mass) sulfuric acid. How many grams of sulfuric acid are present in 500 mL of battery acid?

Mole Concept: Chap. 3, Sec. 3; Chap. 7, Sec. 3

Problems Involving Atoms and Elements

95. Calculate the number of moles in each of the following masses.
a. 64.1 g of aluminum
b. 28.1 g of silicon
c. 0.255 g of sulfur
d. 850.5 g of zinc

96. Calculate the mass of each of the following amounts.
a. 1.22 mol sodium
b. 14.5 mol copper
c. 0.275 mol mercury
d. 9.37×10^{-3} mol magnesium

97. Calculate the amount in moles in each of the following quantities.
a. 3.01×10^{23} atoms of rubidium
b. 8.08×10^{22} atoms of krypton
c. 5 700 000 000 atoms of lead
d. 2.997×10^{25} atoms of vanadium

98. Calculate the number of atoms in each of the following amounts.
a. 1.004 mol bismuth
b. 2.5 mol manganese
c. 0.000 0002 mol helium
d. 32.6 mol strontium

99. Calculate the number of atoms in each of the following masses.
a. 54.0 g of aluminum
b. 69.45 g of lanthanum
c. 0.697 g of gallium
d. 0.000 000 020 g beryllium

100. Calculate the mass of the following numbers of atoms.
a. 6.022×10^{24} atoms of tantalum
b. 3.01×10^{21} atoms of cobalt
c. 1.506×10^{24} atoms of argon
d. 1.20×10^{25} atoms of helium

Problems Involving Molecules, Formula Units, and Ions

101. Calculate the number of moles in each of the following masses.
a. 3.00 g of boron tribromide, BBr_3
b. 0.472 g of sodium fluoride, NaF
c. 7.50×10^{2} g of methanol, CH_3OH
d. 50.0 g of calcium chlorate, $Ca(ClO_3)_2$

102. Determine the mass of each of the following amounts.
a. 1.366 mol of NH_3
b. 0.120 mol of glucose, $C_6H_{12}O_6$
c. 6.94 mol barium chloride, $BaCl_2$
d. 0.005 mol of propane, C_3H_8

103. Calculate the number of molecules in each of the following amounts.
a. 4.99 mol of methane, CH_4
b. 0.005 20 mol of nitrogen gas, N_2
c. 1.05 mol of phosphorus trichloride, PCl_3
d. 3.5×10^{-5} mol of vitamin C, ascorbic acid, $C_6H_8O_6$

104. Calculate the number of formula units in the following amounts.
a. 1.25 mol of potassium bromide, KBr
b. 5.00 mol of magnesium chloride, $MgCl_2$
c. 0.025 mol of sodium carbonate, Na_2CO_3
d. 6.82×10^{-6} mol of lead(II) nitrate, $Pb(NO_3)_2$

105. Calculate the amount in moles of the following numbers of molecules or formula units.
a. 3.34×10^{34} formula units of $Cu(OH)_2$
b. 1.17×10^{16} molecules of H_2S
c. 5.47×10^{21} formula units of nickel(II) sulfate, $NiSO_4$
d. 7.66×10^{19} molecules of hydrogen peroxide, H_2O_2

106. Calculate the mass of each of the following quantities.
a. 2.41×10^{24} molecules of hydrogen, H_2
b. 5.00×10^{21} formula units of aluminum hydroxide, $Al(OH)_3$
c. 8.25×10^{22} molecules of bromine pentafluoride, BrF_5
d. 1.20×10^{23} formula units of sodium oxalate, $Na_2C_2O_4$

107. Calculate the number of molecules or formula units in each of the following masses.
a. 22.9 g of sodium sulfide, Na_2S
b. 0.272 g of nickel(II) nitrate, $Ni(NO_3)_2$
c. 260 mg of acrylonitrile, CH_2CHCN

Mixed Review

108. Calculate the number of moles in each of the following masses.
a. 0.039 g of palladium
b. 8200 g of iron
c. 0.0073 kg of tantalum
d. 0.006 55 g of antimony
e. 5.64 kg of barium

109. Calculate the mass in grams of each of the following amounts.
a. 1.002 mol of chromium
b. 550 mol of aluminum
c. 4.08×10^{-8} mol of neon
d. 7 mol of titanium
e. 0.0086 mol of xenon
f. 3.29×10^{4} mol of lithium

110. Calculate the number of atoms in each of the following amounts.
a. 17.0 mol of germanium
b. 0.6144 mol of copper
c. 3.02 mol of tin
d. 2.0×10^{6} mol of carbon
e. 0.0019 mol of zirconium
f. 3.227×10^{-10} mol of potassium

111. Calculate the number of moles in each of the following quantities.
a. 6.022×10^{24} atoms of cobalt
b. 1.06×10^{23} atoms of tungsten
c. 3.008×10^{19} atoms of silver

d. 950 000 000 atoms of plutonium
e. 4.61×10^{17} atoms of radon
f. 8 trillion atoms of cerium

112. Calculate the number of atoms in each of the following masses.
a. 0.0082 g of gold
b. 812 g of molybdenum
c. 2.00×10^{2} mg of americium
d. 10.09 kg of neon
e. 0.705 mg of bismuth
f. 37 μg of uranium

113. Calculate the mass of each of the following.
a. 8.22×10^{23} atoms of rubidium
b. 4.05 Avogadro's constants of manganese atoms
c. 9.96×10^{26} atoms of tellurium
d. 0.000 025 Avogadro's constants of rhodium atoms
e. 88 300 000 000 000 atoms of radium
f. 2.94×10^{17} atoms of hafnium

114. Calculate the number of moles in each of the following masses.
a. 45.0 g of acetic acid, CH_3COOH
b. 7.04 g of lead(II) nitrate, $Pb(NO_3)_2$
c. 5000 kg of iron(III) oxide, Fe_2O_3
d. 12.0 mg of ethylamine, $C_2H_5NH_2$
e. 0.003 22 g of stearic acid, $C_{17}H_{35}COOH$
f. 50.0 kg of ammonium sulfate, $(NH_4)_2SO_4$

115. Calculate the mass of each of the following amounts.
a. 3.00 mol of selenium oxybromide, $SeOBr_2$
b. 488 mol of calcium carbonate, $CaCO_3$
c. 0.0091 mol of retinoic acid, $C_{20}H_{28}O_2$
d. 6.00×10^{-8} mol of nicotine, $C_{10}H_{14}N_2$
e. 2.50 mol of strontium nitrate, $Sr(NO_3)_2$
f. 3.50×10^{-6} mol of uranium hexafluoride, UF_6

116. Calculate the number of molecules or formula units in each of the following amounts.
a. 4.27 mol of tungsten(VI) oxide, WO_3
b. 0.003 00 mol of strontium nitrate, $Sr(NO_3)_2$
c. 72.5 mol of toluene, $C_6H_5CH_3$
d. 5.11×10^{-7} mol of α-tocopherol (vitamin E), $C_{29}H_{50}O_2$
e. 1500 mol of hydrazine, N_2H_4
f. 0.989 mol of nitrobenzene $C_6H_5NO_2$

117. Calculate the number of molecules or formula units in each of the following masses.
a. 285 g of iron(III) phosphate, $FePO_4$
b. 0.0084 g of C_5H_5N
c. 85 mg of 2-methyl-1-propanol, $(CH_3)_2CHCH_2OH$
d. 4.6×10^{-4} g of mercury(II) acetate, $Hg(C_2H_3O_2)_2$
e. 0.0067 g of lithium carbonate, Li_2CO_3

118. Calculate the mass of each of the following quantities.
a. 8.39×10^{23} molecules of fluorine, F_2
b. 6.82×10^{24} formula units of beryllium sulfate, $BeSO_4$
c. 7.004×10^{26} molecules of chloroform, $CHCl_3$
d. 31 billion formula units of chromium(III) formate, $Cr(CHO_2)_3$
e. 6.3×10^{18} molecules of nitric acid, HNO_3
f. 8.37×10^{25} molecules of freon 114, $C_2Cl_2F_4$

119. Precious metals are commonly measured in troy ounces. A troy ounce is equivalent to 31.1 g. How many moles are in a troy ounce of gold? How many moles are in a troy ounce of platinum? of silver?

120. A chemist needs 22.0 g of phenol, C_6H_5OH, for an experiment. How many moles of phenol is this?

121. A student needs 0.015 mol of iodine crystals, I_2, for an experiment. What mass of iodine crystals should the student obtain?

122. The weight of a diamond is given in carats. One carat is equivalent to 200. mg. A pure diamond is made up entirely of carbon atoms. How many carbon atoms make up a 1.00 carat diamond?

123. 8.00 g of calcium chloride, $CaCl_2$, is dissolved in 1.000 kg of water.
a. How many moles of $CaCl_2$ are in solution? How many moles of water are present?
b. Assume that the ionic compound, $CaCl_2$, separates completely into Ca^{2+} and Cl^- ions when it dissolves in water. How many moles of each ion are present in the solution?

124. How many moles are in each of the following masses?
a. 453.6 g (1.000 pound) of sucrose (table sugar), $C_{12}H_{22}O_{11}$
b. 1.000 pound of table salt, NaCl

125. When the ionic compound NH_4Cl dissolves in water, it breaks into one ammonium ion, NH_4^+, and one chloride ion, Cl^-. If you dissolved 10.7 g of NH_4Cl in water, how many moles of ions would be in solution?

126. What is the total amount in moles of atoms in a jar that contains 2.41×10^{24} atoms of chromium, 1.51×10^{23} atoms of nickel, and 3.01×10^{23} atoms of copper?

127. The density of liquid water is 0.997 g/mL at 25°C.
a. Calculate the mass of 250.0 mL (about a cupful) of water.
b. How many moles of water are in 250.0 mL of water? Hint: Use the result of (a).
c. Calculate the volume that would be occupied by 2.000 mol of water at 25°C.
d. What mass of water is 2.000 mol of water?

128. An Avogadro's constant (1 mol) of sugar molecules has a mass of 342 g, but an Avogadro's constant (1 mol) of water molecules has a mass of only 18 g. Explain why there is such a difference between the mass of 1 mol of sugar and the mass of 1 mol of water.

129. Calculate the mass of aluminum that would have the same number of atoms as 6.35 g of cadmium.

130. A chemist weighs a steel cylinder of compressed oxygen, O_2, and finds that it has a mass of 1027.8 g. After some of the oxygen is used in an experiment, the cylinder has a mass of 1023.2 g. How many moles of oxygen gas are used in the experiment?

131. Suppose that you could decompose 0.250 mol of Ag_2S into its elements.
a. How many moles of silver would you have? How many moles of sulfur would you have?
b. How many moles of Ag_2S are there in 38.8 g of Ag_2S? How many moles of silver and sulfur would be produced from this amount of Ag_2S?
c. Calculate the masses of silver and sulfur produced in (b).

Percentage Composition: Chap. 7, Sec. 3

132. Determine the percentage composition of each of the following compounds.
a. sodium oxalate, $Na_2C_2O_4$
b. ethanol, C_2H_5OH
c. aluminum oxide, Al_2O_3
d. potassium sulfate, K_2SO_4

133. Suppose that a laboratory analysis of white powder showed 42.59% Na, 12.02% C, and 44.99% oxygen. Would you report that the compound is sodium oxalate or sodium carbonate? (Use 43.38% Na, 11.33% C, and 45.29% O for sodium carbonate, and 34.31% Na, 17.93% C, and 47.76% O for sodium oxalate.)

134. Calculate the mass of the given element in each of the following compounds.
a. bromine in 50.0 g potassium bromide, KBr
b. chromium in 1.00 kg sodium dichromate, $Na_2Cr_2O_7$
c. nitrogen in 85.0 mg of the amino acid lysine, $C_6H_{14}N_2O_2$
d. cobalt in 2.84 g cobalt(II) acetate, $Co(C_2H_3O_2)_2$

Hydrates

135. Calculate the percentage of water in each of the following hydrates.
a. sodium carbonate decahydrate, $Na_2CO_3{\bullet}10H_2O$
b. nickel(II) iodide hexahydrate, $NiI_2{\bullet}6H_2O$
c. ammonium hexacyanoferrate(III) trihydrate (commonly called ammonium ferricyanide), $(NH_4)_2Fe(CN)_6{\bullet}3H_2O$
d. aluminum bromide hexahydrate

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136. Write formulas for the following compounds and determine the percentage composition of each.
a. nitric acid
b. ammonia
c. mercury(II) sulfate
d. antimony(V) fluoride

137. Calculate the percentage composition of the following compounds.
a. lithium bromide, LiBr
b. anthracene, $C_{14}H_{10}$
c. ammonium nitrate, NH_4NO_3
d. nitrous acid, HNO_2
e. silver sulfide, Ag_2S
f. iron(II) thiocyanate, $Fe(SCN)_2$
g. lithium acetate
h. nickel(II) formate

138. Calculate the percentage of the given element in each of the following compounds.
a. nitrogen in urea, NH_2CONH_2
b. sulfur in sulfuryl chloride, SO_2Cl_2
c. thallium in thallium(III) oxide, Tl_2O_3
d. oxygen in potassium chlorate, $KClO_3$
e. bromine in calcium bromide, $CaBr_2$
f. tin in tin(IV) oxide, SnO_2

139. Calculate the mass of the given element in each of the following quantities.
a. oxygen in 4.00 g of manganese dioxide, MnO_2
b. aluminum in 50.0 metric tons of aluminum oxide, Al_2O_3
c. silver in 325 g silver cyanide, AgCN
d. gold in 0.780 g of gold(III) selenide, Au_2Se_3
e. selenium in 683 g sodium selenite, Na_2SeO_3
f. chlorine in 5.0×10^4 g of 1,1-dichloropropane, $CHCl_2CH_2CH_3$

140. Calculate the percentage of water in each of the following hydrates.
a. strontium chloride hexahydrate, $SrCl_2{\bullet}6H_2O$
b. zinc sulfate heptahydrate, $ZnSO_4{\bullet}7H_2O$
c. calcium fluorophosphate dihydrate, $CaFPO_3{\bullet}2H_2O$
d. beryllium nitrate trihydrate, $Be(NO_3)_2{\bullet}3H_2O$

141. Calculate the percentage of the given element in each of the following hydrates. You must first determine the formulas of the hydrates.
a. nickel in nickel(II) acetate tetrahydrate
b. chromium in sodium chromate tetrahydrate
c. cerium in cerium(IV) sulfate tetrahydrate

142. Cinnabar is a mineral that is mined in order to produce mercury. Cinnabar is mercury(II) sulfide, HgS. What mass of mercury can be obtained from 50.0 kg of cinnabar?

143. The minerals malachite, $Cu_2(OH)_2CO_3$, and chalcopyrite, $CuFeS_2$, can be mined to obtain copper metal. How much copper could be obtained from 1.00×10^3 kg of each? Which of the two has the greater copper content?

144. Calculate the percentage of the given element in each of the following hydrates.
a. vanadium in vanadium oxysulfate dihydrate, $VOSO_4{\bullet}2H_2O$
b. tin in potassium stannate trihydrate, $K_2SnO_3{\bullet}3H_2O$
c. chlorine in calcium chlorate dihydrate, $CaClO_3{\bullet}2H_2O$

145. Heating copper sulfate pentahydrate will evaporate the water from the crystals, leaving anhydrous copper sulfate, a white powder. *Anhydrous* means "without water." What mass of anhydrous $CuSO_4$ would be produced by heating 500.0 g of $CuSO_4{\bullet}5H_2O$?

146. Silver metal may be precipitated from a solution of silver nitrate by placing a copper strip into the solution. What mass of $AgNO_3$ would you dissolve in water in order to get 1.00 g of silver?

147. A sample of Ag_2S has a mass of 62.4 g. What mass of each element could be obtained by decomposing this sample?

148. A quantity of epsom salts, magnesium sulfate heptahydrate, $MgSO_4{\bullet}7H_2O$, is heated until all the water is driven off. The sample loses 11.8 g in the process. What was the mass of the original sample?

149. The process of manufacturing sulfuric acid begins with the burning of sulfur. What mass of sulfur would have to be burned in order to produce 1.00 kg of H_2SO_4? Assume that all of the sulfur ends up in the sulfuric acid.

Empirical Formulas: Chap. 7, Sec. 4

150. Determine the empirical formula for compounds that have the following analyses.
- **a.** 28.4% copper, 71.6% bromine
- **b.** 39.0% potassium, 12.0% carbon, 1.01% hydrogen, and 47.9% oxygen
- **c.** 77.3% silver, 7.4% phosphorus, 15.3% oxygen
- **d.** 0.57% hydrogen, 72.1% iodine, 27.3% oxygen

151. Determine the simplest formula for compounds that have the following analyses. The data may not be exact.
- **a.** 36.2% aluminum and 63.8% sulfur
- **b.** 93.5% niobium and 6.50% oxygen
- **c.** 57.6% strontium, 13.8% phosphorus, and 28.6% oxygen
- **d.** 28.5% iron, 48.6% oxygen, and 22.9% sulfur

152. Determine the molecular formula of each of the following unknown substances.
- **a.** empirical formula CH_2
  experimental molar mass 28 g/mol
- **b.** empirical formula B_2H_5
  experimental molar mass 54 g/mol
- **c.** empirical formula C_2HCl
  experimental molar mass 179 g/mol
- **d.** empirical formula C_6H_8O
  experimental molar mass 290 g/mol
- **e.** empirical formula C_3H_2O
  experimental molar mass 216 g/mol

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153. Determine the empirical formula for compounds that have the following analyses.
- **a.** 66.0% barium and 34.0% chlorine
- **b.** 80.38% bismuth, 18.46% oxygen, and 1.16% hydrogen
- **c.** 12.67% aluminum, 19.73% nitrogen, and 67.60% oxygen
- **d.** 35.64% zinc, 26.18% carbon, 34.88% oxygen, and 3.30% hydrogen
- **e.** 2.8% hydrogen, 9.8% nitrogen, 20.5% nickel, 44.5% oxygen, and 22.4% sulfur
- **f.** 8.09% carbon, 0.34% hydrogen, 10.78% oxygen, and 80.78% bromine

154. Sometimes, instead of percentage composition, you will have the composition of a sample by mass. Using the actual mass of the sample, determine the empirical formula for compounds that have the following analyses.
- **a.** a 0.858 g sample of an unknown substance is composed of 0.537 g of copper and 0.321 g of fluorine
- **b.** a 13.07 g sample of an unknown substance is composed of 9.48 g of barium, 1.66 g of carbon, and 1.93 g of nitrogen
- **c.** a 0.025 g sample of an unknown substance is composed of 0.0091 g manganese, 0.0106 g oxygen, and 0.0053 g sulfur

155. Determine the empirical formula for compounds that have the following analyses.
- **a.** a 0.0082 g sample contains 0.0015 g of nickel and 0.0067 g of iodine
- **b.** a 0.470 g sample contains 0.144 g of manganese, 0.074 g of nitrogen, and 0.252 g of oxygen
- **c.** a 3.880 g sample contains 0.691 g of magnesium, 1.824 g of sulfur, and 1.365 g of oxygen
- **d.** a 46.25 g sample contains 14.77 g of potassium, 9.06 g of oxygen, and 22.42 g of tin

156. Determine the empirical formula for compounds that have the following analyses:
- **a.** 60.9% As and 39.1% S
- **b.** 76.89% Re and 23.12% O
- **c.** 5.04% H, 35.00% N, and 59.96% O
- **d.** 24.3% Fe, 33.9% Cr, and 41.8% O
- **e.** 54.03% C, 37.81% N, and 8.16% H
- **f.** 55.81% C, 3.90% H, 29.43% F, and 10.85% N

157. Determine the molecular formulas for compounds having the following empirical formulas and molar masses.
- **a.** C_2H_4S; experimental molar mass 179
- **b.** C_2H_4O; experimental molar mass 176
- **c.** $C_2H_3O_2$; experimental molar mass 119
- **d.** C_2H_2O, experimental molar mass 254

158. Use the experimental molar mass to determine the molecular formula for compounds having the following analyses.
- **a.** 41.39% carbon, 3.47% hydrogen, and 55.14% oxygen; experimental molar mass 116.07
- **b.** 54.53% carbon, 9.15% hydrogen, and 36.32% oxygen; experimental molar mass 88
- **c.** 64.27% carbon, 7.19% hydrogen, and 28.54% oxygen; experimental molar mass 168.19

159. A 0.400 g sample of a white powder contains 0.141 g of potassium, 0.115 g of sulfur, and 0.144 g of oxygen. What is the empirical formula for the compound?

160. A 10.64 g sample of a lead compound is analyzed and found to be made up of 9.65 g of lead and 0.99 g of oxygen. Determine the empirical formula for this compound.

161. A 2.65 g sample of a salmon-colored powder contains 0.70 g of chromium, 0.65 g of sulfur, and 1.30 g of oxygen. The molar mass is 392.2. What is the formula of the compound?

162. Ninhydrin is a compound that reacts with amino acids and proteins to produce a dark-colored complex. It is used by forensic chemists and detectives to see fingerprints that might otherwise be invisible. Ninhydrin's composition is 60.68% carbon, 3.40% hydrogen, and 35.92% oxygen. What is the empirical formula for ninhydrin?

163. Histamine is a substance that is released by cells in response to injury, infection, stings, and materials that cause allergic responses, such as pollen. Histamine causes dilation of blood vessels and swelling due to accumulation of fluid in the tissues. People sometimes take *anti*histamine drugs to counteract the effects of histamine. A sample of histamine having a mass of 385 mg is composed of 208 mg of carbon, 31 mg of hydrogen, and 146 mg of nitrogen. The molar mass of histamine is 111 g/mol. What is the molecular formula for histamine?

164. You analyze two substances in the laboratory and discover that each has the empirical formula CH_2O. You can easily see that they are different substances because one is a liquid with a sharp, biting odor and the other is an odorless, crystalline solid. How can you account for the fact that both have the same empirical formula?

Stoichiometry: Chap. 9, Sec. 1–2

165. How many moles of sodium will react with water to produce 4.0 mol of hydrogen in the following reaction?

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

166. How many moles of lithium chloride will be formed by the reaction of chlorine with 0.046 mol of lithium bromide in the following reaction?

$$2LiBr(aq) + Cl_2(g) \rightarrow 2LiCl(aq) + Br_2(l)$$

167. Aluminum will react with sulfuric acid in the following reaction.

$$2Al(s) + 3H_2SO_4(l) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$$

a. How many moles of H_2SO_4 will react with 18 mol Al?
b. How many moles of each product will be produced?

168. Propane burns in excess oxygen according to the following reaction.

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

a. How many moles each of CO_2 and H_2O are formed from 3.85 mol of propane?
b. If 0.647 mol of oxygen are used in the burning of propane, how many moles each of CO_2 and H_2O are produced? How many moles of C_3H_8 are consumed?

169. Phosphorus burns in air to produce a phosphorus oxide in the following reaction:

$$4P(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$$

a. What mass of phosphorus will be needed to produce 3.25 mol of P_4O_{10}?
b. If 0.489 mol of phosphorus burns, what mass of oxygen is used? What mass of P_4O_{10} is produced?

170. Hydrogen peroxide breaks down, releasing oxygen, in the following reaction.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

a. What mass of oxygen is produced when 1.840 mol of H_2O_2 decompose?
b. What mass of water is produced when 5.0 mol O_2 is produced by this reaction?

171. Sodium carbonate reacts with nitric acid according to the following equation:

$$Na_2CO_3(s) + 2HNO_3 \rightarrow 2NaNO_3 + CO_2 + H_2O$$

a. How many moles of Na_2CO_3 are required to produce 100.0 g of $NaNO_3$?
b. If 7.50 g of Na_2CO_3 reacts, how many moles of CO_2 are produced?

172. Hydrogen is generated by passing hot steam over iron, which oxidizes to form Fe_3O_4, in the following equation:

$$3Fe(s) + 4H_2O(g) \rightarrow 4H_2(g) + Fe_3O_4(s)$$

a. If 625 g of Fe_3O_4 is produced in the reaction, how many moles of hydrogen are produced at the same time?
b. How many moles of iron would be needed to generate 27 g of hydrogen?

173. Calculate the mass of silver bromide produced from 22.5 g of silver nitrate in the following reaction:

$$2AgNO_3(aq) + MgBr_2(aq) \rightarrow 2AgBr(s) + Mg(NO_3)_2(aq)$$

174. What mass of acetylene, C_2H_2, will be produced from the reaction of 90. g of calcium carbide, CaC_2, with water in the following reaction?

$$CaC_2(s) + 2H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(s)$$

175. Chlorine gas can be produced in the laboratory by adding concentrated hydrochloric acid to manganese(IV) oxide in the following reaction:

$$MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + 2H_2O(l) + Cl_2(g)$$

a. Calculate the mass of MnO_2 needed to produce 25.0 g of Cl_2.
b. What mass of $MnCl_2$ is produced when 0.091 g of Cl_2 is generated?

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176. How many moles of ammonium sulfate can be made from the reaction of 30.0 mol of NH_3 with H_2SO_4 according to the following equation:

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

177. In a very violent reaction called a thermite reaction, aluminum metal reacts with iron(III) oxide to form iron metal and aluminum oxide according to the following equation:

$$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$$

a. What mass of Al will react with 150 g of Fe_2O_3?
b. If 0.905 mol Al_2O_3 is produced in the reaction, what mass of Fe is produced?
c. How many moles of Fe_2O_3 will react with 99.0 g of Al?

178. The reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ is used to produce ammonia commercially. If 1.40 g of N_2 are used in the reaction, how many grams of H_2 will be needed?

179. What mass of sulfuric acid, H_2SO_4, is required to react with 1.27 g of potassium hydroxide, KOH? The products of this reaction are potassium sulfate and water.

180. Ammonium hydrogen phosphate, $(NH_4)_2HPO_4$, a common fertilizer; is made from reacting phosphoric acid, H_3PO_4, with ammonia.
a. Write the equation for this reaction.
b. If 10.00 g of ammonia react, how many moles of fertilizer will be produced?

c. What mass of ammonia will react with 2800 kg of H_3PO_4?

181. The following reaction shows the synthesis of zinc citrate, a ingredient in toothpaste, from zinc carbonate and citric acid:

$$3ZnCO_3(s) + 2C_6H_8O_7(aq) \rightarrow Zn_3(C_6H_5O_7)_2(aq) + 3H_2O(l) + 3CO_2(g)$$

a. How many moles of $ZnCO_3$ and $C_6H_8O_7$ are required to produce 30.0 mol of $Zn_3(C_6H_5O_7)_2$?

b. What quantities, in kilograms, of H_2O and CO_2 are produced by the reaction of 500. mol of citric acid?

182. Methyl butanoate, an oily substance with a strong fruity fragrance can be made by reacting butanoic acid with methanol according to the following equation:

$$C_3H_7COOH + CH_3OH \rightarrow C_3H_7COOCH_3 + H_2O$$

a. What mass of methyl butanoate is produced from the reaction of 52.5 g of butanoic acid?

b. In order to purify methyl butanoate, water must be removed. What mass of water is produced from the reaction of 5800. g of methanol?

183. Ammonium nitrate decomposes to yield nitrogen gas, water, and oxygen gas in the following reaction:

$$2NH_4NO_3 \rightarrow 2N_2 + O_2 + 4H_2O$$

a. How many moles of nitrogen gas are produced when 36.0 g of NH_4NO_3 reacts?

b. If 7.35 mol of H_2O are produced in this reaction, what mass of NH_4NO_3 reacted?

184. Lead(II) nitrate reacts with potassium iodide to produce lead(II) iodide and potassium nitrate. If 1.23 mg of lead nitrate are consumed, what is the mass of the potassium nitrate produced?

185. A car battery produces electrical energy with the following chemical reaction:

$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

If the battery loses 0.34 kg of lead in this reaction, how many moles of lead(II) sulfate are produced?

186. In a space shuttle, the CO_2 that the crew exhales is removed from the air by a reaction within canisters of lithium hydroxide. On average, each astronaut exhales about 20.0 mol of CO_2 daily. What mass of water will be produced when this amount reacts with LiOH? The other product of the reaction is Li_2CO_3.

187. Water is sometimes removed from the products of a reaction by placing them in a closed container with excess P_4O_{10}. Water is absorbed by the following reaction:

$$P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$$

a. What mass of water can be absorbed by 1.00×10^2 g of P_4O_{10}?

b. If the P_4O_{10} in the container absorbs 0.614 mol of water, what mass of H_3PO_4 is produced?

c. If the mass of the container of P_4O_{10} increases from 56.64 g to 63.70 g, how many moles of water are absorbed?

188. Ethanol, C_2H_5OH, is considered a clean fuel because it burns in oxygen to produce carbon dioxide and water with few trace pollutants. If 95.0 g of H_2O are produced during the combustion of ethanol, how many grams of ethanol were present at the beginning of the reaction?

189. Sulfur dioxide is one of the major contributors to acid rain. Sulfur dioxide can react with oxygen and water in the atmosphere to form sulfuric acid, as shown in the following equation:

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

If 50.0 g of sulfur dioxide from pollutants reacts with water and oxygen found in the air, how many grams of sulfuric acid can be produced? How many grams of oxygen are used in the process?

190. When heated, sodium bicarbonate, $NaHCO_3$, decomposes into sodium carbonate, Na_2CO_3, water, and carbon dioxide. If 5.00 g of $NaHCO_3$ decomposes, what is the mass of the carbon dioxide produced?

191. A reaction between hydrazine, N_2H_4, and dinitrogen tetroxide, N_2O_4, has been used to launch rockets into space. The reaction produces nitrogen gas and water vapor.

a. Write a balanced chemical equation for this reaction.

b. What is the mole ratio of N_2O_4 to N_2?

c. How many moles of N_2 will be produced if 20 000 mol of N_2H_4 are used by a rocket?

d. How many grams of H_2O are made when 450. kg of N_2O_4 are consumed?

192. Joseph Priestley is credited with the discovery of oxygen. He produced O_2 by heating mercury(II) oxide, HgO, to decompose it into its elements. How many moles of oxygen could Priestley have produced if he had decomposed 517.84 g of mercury oxide?

193. Iron(III) chloride, $FeCl_3$, can be made by the reaction of iron with chlorine gas. How much iron, in grams, will be needed to completely react with 58.0 g of Cl_2?

194. Sodium sulfide and cadmium nitrate undergo a double-displacement reaction as shown by the following equation:

$$Na_2S + Cd(NO_3)_2 \rightarrow 2NaNO_3 + CdS$$

What is the mass, in milligrams, of cadmium sulfide that can be made from 5.00 mg of sodium sulfide?

195. Potassium permanganate and glycerin react explosively according to the following equation:

$$14KMnO_4 + 4C_3H_5(OH)_3 \rightarrow 7K_2CO_3 + 7Mn_2O_3 + 5CO_2 + 16H_2O$$

a. How many moles of carbon dioxide can be produced from 4.44 mol of $KMnO_4$?

b. If 5.21 g of H_2O are produced, how many moles of glycerin, $C_3H_5(OH)_3$, were used?

c. If 3.39 mol of potassium carbonate are made, how many grams of manganese(III) oxide are also made?

d. How many grams of glycerin will be needed to react with 50.0 g of $KMnO_4$? How many grams of CO_2 will be produced in the same reaction?

196. Calcium carbonate found in limestone and marble reacts with hydrochloric acid to form calcium chloride, carbon dioxide, and water according to the following equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

a. What mass of HCl will be needed to produce 5.00×10^3 kg of $CaCl_2$?

b. What mass of CO_2 could be produced from the reaction of 750 g of $CaCO_3$?

197. The fuel used to power the booster rockets on the space shuttle is a mixture of aluminum metal and ammonium perchlorate. The following balanced equation represents the reaction of these two ingredients:

$$3Al(s) + 3NH_4ClO_4(s) \rightarrow Al_2O_3(s) + AlCl_3(g) + 3NO(g) + 6H_2O(g)$$

a. If 1.50×10^5 g of Al react, what mass of NH_4ClO_4, in grams, is required?

b. If aluminum reacts with 620 kg of NH_4ClO_4, what mass of nitrogen monoxide is produced?

198. Phosphoric acid is typically produced by the action of sulfuric acid on rock that has a high content of calcium phosphate according to the following equation:

$$3H_2SO_4 + Ca_3(PO_4)_2 + 6H_2O \rightarrow 3[CaSO_4 \cdot 2H_2O] + 2H_3PO_4$$

a. If 2.50×10^5 kg of H_2SO_4 react, how many moles of H_3PO_4 can be made?

b. What mass of calcium sulfate dihydrate is produced by the reaction of 400. kg of calcium phosphate?

c. If the rock being used contains 78.8% $Ca_3(PO_4)_2$, how many metric tons of H_3PO_4 can be produced from 68 metric tons of rock?

199. Rusting of iron occurs in the presence of moisture according to the following equation:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

Suppose that 3.19% of a heap of steel scrap with a mass of 1650 kg rusts in a year. What mass will the heap have after one year of rusting?

Limiting Reactants: Chap. 9, Sec. 3

200. Aluminum oxidizes according to the following equation:

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

Powdered Al (0.048 mol) is placed into a container containing 0.030 mol O_2. What is the limiting reactant?

201. A process by which zirconium metal can be produced from the mineral zirconium(IV) orthosilicate, $ZrSiO_4$, starts by reacting it with chlorine gas to form zirconium(IV) chloride:

$$ZrSiO_4 + 2Cl_2 \rightarrow ZrCl_4 + SiO_2 + O_2$$

What mass of $ZrCl_4$ can be produced if 862 g of $ZrSiO_4$ and 950. g of Cl_2 are available? You must first determine the limiting reactant.

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202. Heating zinc sulfide in the presence of oxygen yields the following:

$$ZnS + O_2 \rightarrow ZnO + SO_2$$

If 1.72 mol of ZnS is heated in the presence of 3.04 mol of O_2, which reactant will be used up? Balance the equation first.

203. Use the following equation for the oxidation of aluminum in the following problems:

$$4Al + 3O_2 \rightarrow 2Al_2O_3$$

a. Which reactant is limiting if 0.32 mol Al and 0.26 mol O_2 are available?

b. How many moles of Al_2O_3 are formed from the reaction of 6.38×10^{-3} mol of O_2 and 9.15×10^{-3} mol of Al?

c. If 3.17 g of Al and 2.55 g of O_2 are available, which reactant is limiting?

204. In the production of copper from ore containing copper(II) sulfide, the ore is first roasted to change it to the oxide according to the following equation:

$$2CuS + 3O_2 \rightarrow 2CuO + 2SO_2$$

a. If 100 g of CuS and 56 g of O_2 are available, which reactant is limiting?

b. What mass of CuO can be formed from the reaction of 18.7 g of CuS and 12.0 g of O_2?

205. A reaction such as the one shown here is often used to demonstrate a single-displacement reaction:

$$3CuSO_4(aq) + 2Fe(s) \rightarrow 3Cu(s) + Fe_2(SO_4)_3(aq)$$

If you place 0.092 mol of iron filings in a solution containing 0.158 mol of $CuSO_4$, what is the limiting reactant? How many moles of Cu will be formed?

206. In the reaction $BaCO_3 + 2HNO_3 \rightarrow Ba(NO_3)_2 + CO_2 + H_2O$, what mass of $Ba(NO_3)_2$ can be formed by combining 55 g $BaCO_3$ and 26 g HNO_3?

207. Bromine replaces iodine in magnesium iodide by the following process:

$$MgI_2 + Br_2 \rightarrow MgBr_2 + I_2$$

a. Which is the excess reactant when 560 g of MgI_2 and 360 g of Br_2 react, and what mass remains?

b. What mass of I_2 is formed in the same process?

208. Nickel replaces silver from silver nitrate in solution according to the following equation:

$$2AgNO_3 + Ni \rightarrow 2Ag + Ni(NO_3)_2$$

a. If you have 22.9 g of Ni and 112 g of $AgNO_3$, which reactant is in excess?

b. What mass of nickel(II) nitrate would be produced given the quantities above?

209. Carbon disulfide, CS_2, is an important industrial substance. Its fumes can burn explosively in air to form sulfur dioxide and carbon dioxide:

$$CS_2(g) + O_2(g) \rightarrow SO_2(g) + CO_2(g)$$

If 1.60 mol of CS_2 burns with 5.60 mol of O_2, how many moles of the excess reactant will still be present when the reaction is over?

210. Although poisonous, mercury compounds were once used to kill bacteria in wounds and on the skin. One was called "ammoniated mercury" and is made from mercury(II) chloride according to the following equation:

$$HgCl_2(aq) + 2NH_3(aq) \rightarrow Hg(NH_2)Cl(s) + NH_4Cl(aq)$$

a. What mass of $Hg(NH_2)Cl$ could be produced from 0.91 g of $HgCl_2$ assuming plenty of ammonia is available?
b. What mass of $Hg(NH_2)Cl$ could be produced from 0.91 g of $HgCl_2$ and 0.15 g of NH_3 in solution?

211. Aluminum chips are sometimes added to sodium hydroxide-based drain cleaners because they react to generate hydrogen gas which bubbles and helps loosen material in the drain. The equation follows:

$$Al(s) + NaOH(aq) + H_2O(l) \rightarrow NaAlO_2(aq) + H_2(g)$$

a. Balance the equation.
b. How many moles of H_2 can be generated from 0.57 mol Al and 0.37 mol NaOH in excess water?
c. Which reactant should be limiting in order for the mixture to be most effective as a drain cleaner? Explain your choice.

212. Copper is changed to copper(II) ions by nitric acid according to the following equation:

$$4HNO_3 + Cu \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$

a. How many moles each of HNO_3 and Cu must react in order to produce 0.0845 mol of NO_2?
b. If 5.94 g of Cu and 23.23 g of HNO_3 are combined, which reactant is in excess?

213. One industrial process for producing nitric acid begins with the following reaction:

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

a. If 2.90 mol NH_3 and 3.75 mol O_2 are available, how many moles of each product are formed?
b. Which reactant is limiting if 4.20×10^4 g of NH_3 and 1.31×10^5 g of O_2 are available?
c. What mass of NO is formed in the reaction of 869 kg of NH_3 and 2480 kg O_2?

214. Acetaldehyde, CH_3CHO, is manufactured by the reaction of ethanol with copper(II) oxide according to the following equation:

$$CH_3CH_2OH + CuO \rightarrow CH_3CHO + H_2O + Cu$$

What mass of acetaldehyde can be produced by the reaction between 620 g of ethanol and 1020 g of CuO? What mass of which reactant will be left over?

215. Hydrogen bromide can be produced by a reaction among bromine, sulfur dioxide, and water as follows:

$$SO_2 + Br_2 + H_2O \rightarrow 2HBr + H_2SO_4$$

If 250 g of SO_2 and 650 g of Br_2 react in the presence of excess water, what mass of HBr will be formed?

216. Sulfur dioxide can be produced in the laboratory by the reaction of hydrochloric acid and a sulfite salt such as sodium sulfite:

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + SO_2 + H_2O$$

What mass of SO_2 can be made from 25.0 g of Na_2SO_3 and 22.0 g of HCl?

217. The rare-earth metal terbium is produced from terbium(III) fluoride and calcium metal by the following single-displacement reaction:

$$2TbF_3 + 3Ca \rightarrow 3CaF_2 + 2Tb$$

a. Given 27.5 g of TbF_3 and 6.96 g of Ca, how many grams of terbium could be produced?
b. How many grams of the excess reactant is left over?

Percentage Yield: Chap. 9, Sec. 3

218. Calculate the percentage yield in each of the following cases.
a. theoretical yield is 50.0 g of product; actual yield is 41.9 g
b. theoretical yield is 290 kg of product; actual yield is 270 kg
c. theoretical yield is 6.05×10^4 kg of product; actual yield is 4.18×10^4 kg
d. theoretical yield is 0.00192 g of product; actual yield is 0.00089 g

219. In the commercial production of the element arsenic, arsenic(III) oxide is heated with carbon, which reduces the oxide to the metal according to the following equation:

$$2As_2O_3 + 3C \rightarrow 3CO_2 + 4As$$

a. If 8.87 g of As_2O_3 is used in the reaction and 5.33 g of As is produced, what is the percentage yield?
b. If 67 g of carbon is used up in a different reaction and 425 g of As is produced, calculate the percentage yield of this reaction.

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220. Ethyl acetate is a sweet-smelling solvent used in varnishes and fingernail-polish remover. It is produced industrially by heating acetic acid and ethanol together in the presence of sulfuric acid, which is added to speed up the reaction. The ethyl acetate is distilled off as it is formed. The equation for the process is as follows:

$$\underset{\text{acetic acid}}{CH_3COOH} + \underset{\text{ethanol}}{CH_3CH_2OH} \xrightarrow{H_2SO_4} \underset{\text{ethyl acetate}}{CH_3COOCH_2CH_3} + H_2O$$

Determine the percentage yield in the following cases.
a. 68.3 g of ethyl acetate should be produced but only 43.9 g is recovered.
b. 0.0419 mol of ethyl acetate is produced but 0.0722 mol is expected. (Hint: Percentage yield can also be calculated by dividing the actual yield in moles by the theoretical yield in moles.)
c. 4.29 mol of ethanol is reacted with excess acetic acid, but only 2.98 mol of ethyl acetate is produced.
d. A mixture of 0.58 mol ethanol and 0.82 mol acetic acid is reacted and 0.46 mol ethyl acetate is produced. (Hint: What is the limiting reactant?)

221. Assume the following hypothetical reaction takes place:

$$2A + 7B \rightarrow 4C + 3D$$

Calculate the percentage yield in each of the following cases.
a. The reaction of 0.0251 mol of A produces 0.0349 mol of C.

b. The reaction of 1.19 mol of A produces 1.41 mol of D.
c. The reaction of 189 mol of B produces 39 mol of D.
d. The reaction of 3500 mol of B produces 1700 mol of C.

222. Elemental phosphorus can be produced by heating calcium phosphate from rocks with silica sand (SiO_2) and carbon in the form of coke. The following reaction takes place:

$$Ca_3(PO_4)_2 + 3SiO_2 + 5C \rightarrow 3CaSiO_3 + 2P + 5CO$$

a. If 57 mol of $Ca_3(PO_4)_2$ is used and 101 mol of $CaSiO_3$ is obtained, what is the percentage yield?
b. Determine the percentage yield obtained if 1280 mol of carbon is consumed and 622 mol of $CaSiO_3$ is produced.
c. The engineer in charge of this process expects a yield of 81.5%. If 1.4×10^5 mol of $Ca_3(PO_4)_2$ is used, how many moles of phosphorus will be produced?

223. Tungsten (W) can be produced from its oxide by reacting the oxide with hydrogen at a high temperature according to the following equation:

$$WO_3 + 3H_2 \rightarrow W + 3H_2O$$

a. What is the percentage yield if 56.9 g of WO_3 yields 41.4 g of tungsten?
b. How many moles of tungsten will be produced from 3.72 g of WO_3 if the yield is 92.0%?
c. A chemist carries out this reaction and obtains 11.4 g of tungsten. If the percentage yield is 89.4%, what mass of WO_3 was used?

224. Carbon tetrachloride, CCl_4, is a solvent that was once used in large quantities in dry cleaning. Because it is a dense liquid that does not burn, it was also used in fire extinguishers. Unfortunately, its use was discontinued because it was found to be a carcinogen. It was manufactured by the following reaction:

$$CS_2 + 3Cl_2 \rightarrow CCl_4 + S_2Cl_2$$

The reaction was economical because the byproduct disulfur dichloride, S_2Cl_2, could be used by industry in the manufacture of rubber products and other materials.

a. What is the percentage yield of CCl_4 if 719 kg is produced from the reaction of 410. kg of CS_2?
b. If 67.5 g of Cl_2 are used in the reaction and 39.5 g of S_2Cl_2 is produced, what is the percentage yield?
c. If the percentage yield of the industrial process is 83.3%, how many kilograms of CS_2 should be reacted to obtain 5.00×10^4 kg of CCl_4? How many kilograms of S_2Cl_2 will be produced, assuming the same yield for that product?

225. Nitrogen dioxide, NO_2, can be converted to dinitrogen pentoxide, N_2O_5, by reacting it with ozone, O_3. The reaction of NO_2 takes place according to the following equation:

$$2NO_2(g) + O_3(g) \rightarrow N_2O_5(s \text{ or } g) + O_2(g)$$

a. Calculate the percentage yield for a reaction in which 0.38 g of NO_2 reacts and 0.36 g of N_2O_5 is recovered.
b. What mass of N_2O_5 will result from the reaction of 6.0 mol of NO_2 if there is a 61.1% yield in the reaction?

226. In the past, hydrogen chloride, HCl, was made using the *salt-cake* method as shown in the following equation:

$$2NaCl(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(s) + 2HCl(g)$$

If 30.0 g of NaCl and 0.250 mol of H_2SO_4 are available, and 14.6 g of HCl is made, what is the percentage yield?

227. Cyanide compounds such as sodium cyanide, NaCN, are especially useful in gold refining because they will react with gold to form a stable compound that can then be separated and broken down to retrieve the gold. Ore containing only small quantities of gold can be used in this form of "chemical mining." The equation for the reaction follows:

$$4Au + 8NaCN + 2H_2O + O_2 \rightarrow 4NaAu(CN)_2 + 4NaOH$$

a. What percentage yield is obtained if 410 g of gold produces 540 g of $NaAu(CN)_2$?
b. Assuming a 79.6% yield in the conversion of gold to $NaAu(CN)_2$, what mass of gold would produce 1.00 kg of $NaAu(CN)_2$?
c. Given the conditions in (b), what mass of gold ore that is 0.001% gold would be needed to produce 1.00 kg of $NaAu(CN)_2$?

228. Diiodine pentoxide is useful in devices such as respirators because it reacts with the dangerous gas carbon monoxide, CO, to produce relatively harmless CO_2 according to the following equation:

$$I_2O_5 + 5CO \rightarrow I_2 + 5CO_2$$

a. In testing a respirator, 2.00 g of carbon monoxide gas is passed through diiodine pentoxide. Upon analyzing the results, it is found that 3.17 g of I_2 was produced. Calculate the percentage yield of the reaction.
b. Assuming that the yield in (a) resulted because some of the CO did not react, calculate the mass of CO that passed through.

229. Sodium hypochlorite, NaClO, the main ingredient in household bleach, is produced by bubbling chlorine gas through a strong lye (sodium hydroxide, NaOH) solution. The following equation shows the reaction that occurs:

$$2NaOH(aq) + Cl_2(g) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$$

a. What is the percentage yield of the reaction if 1.2 kg of Cl_2 reacts to form 0.90 kg of NaClO?
b. If a plant operator wants to make 25 metric tons of NaClO per day at a yield of 91.8%, how many metric tons of chlorine gas must be on hand each day?
c. What mass of NaCl is formed per mole of chlorine gas at a yield of 81.8%?
d. At what rate in kg per hour must NaOH be replenished if the reaction produces 370 kg/h of NaClO at a yield of 79.5%? Assume that all of the NaOH reacts to produce this yield.

230. Magnesium burns in oxygen to form magnesium oxide. However, when magnesium burns in air, which is only about one-fifth oxygen, side reactions form other products, such as magnesium nitride, Mg_3N_2.

a. Write a balanced equation for the burning of magnesium in oxygen.
b. If enough magnesium burns in air to produce 2.04 g of magnesium oxide but only 1.79 g is obtained, what is the percentage yield?
c. Magnesium will react with pure nitrogen to form the nitride, Mg_3N_2. Write a balanced equation for this reaction.
d. If 0.097 mol of Mg react with nitrogen and 0.027 mol of Mg_3N_2 is produced, what is the percentage yield of the reaction?

231. Some alcohols can be converted to organic acids by using sodium dichromate and sulfuric acid. The following equation shows the reaction of 1-propanol to propanoic acid:

$$3CH_3CH_2CH_2OH + 2Na_2Cr_2O_7 + 8H_2SO_4 \rightarrow 3CH_3CH_2COOH + 2Cr_2(SO_4)_3 + 2Na_2SO_4 + 11H_2O$$

a. If 0.89 g of 1-propanol reacts and 0.88 g of propanoic acid is produced, what is the percentage yield?
b. A chemist uses this reaction to obtain 1.50 mol of propanoic acid. The reaction consumes 136 g of propanol. Calculate the percentage yield.
c. Some 1-propanol of uncertain purity is used in the reaction. If 116 g of $Na_2Cr_2O_7$ are consumed in the reaction and 28.1 g of propanoic acid are produced, what is the percentage yield?

232. Acrylonitrile, $C_3H_3N(g)$, is an important ingredient in the production of various fibers and plastics. Acrylonitrile is produced from the following reaction:

$$C_3H_6(g) + NH_3(g) + O_2(g) \rightarrow C_3H_3N(g) + H_2O(g)$$

If 850. g of C_3H_6 is mixed with 300. g of NH_3 and unlimited O_2, to produce 850. g of acrylonitrile, what is the percentage yield? You must first balance the equation.

233. Methanol, CH_3OH, is frequently used in race cars as fuel. It is produced as the sole product of the combination of carbon monoxide gas and hydrogen gas.
a. If 430. kg of hydrogen react, what mass of methanol could be produced?
b. If 3.12×10^3 kg of methanol are actually produced, what is the percentage yield?

234. The compound, $C_6H_{16}N_2$, is one of the starting materials in the production of nylon. It can be prepared from the following reaction involving adipic acid, $C_6H_{10}O_4$:

$$C_6H_{10}O_4(l) + 2NH_3(g) + 4H_2(g) \rightarrow C_6H_{16}N_2(l) + 4H_2O$$

What is the percentage yield if 750. g of adipic acid results in the production of 578 g of $C_6H_{16}N_2$?

235. Plants convert carbon dioxide to oxygen during photosynthesis according to the following equation:

$$CO_2 + H_2O \rightarrow C_6H_{12}O_6 + O_2$$

Balance this equation, and calculate how much oxygen would be produced if 1.37×10^4 g of carbon dioxide reacts with a percentage yield of 63.4%.

236. Lime, CaO, is frequently added to streams and lakes which have been polluted by acid rain. The calcium oxide reacts with the water to form a base that can neutralize the acid as shown in the following reaction:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$$

If 2.67×10^2 mol of base are needed to neutralize the acid in a lake, and the above reaction has a percentage yield of 54.3%, what is the mass, in kilograms, of lime that must be added to the lake?

Gas Laws: Chap. 11, Sec. 2

Boyle's Law

In each of the following problems, assume that the temperature and molar quantity of gas do not change.

237. Calculate the unknown quantity in each of the following measurements of gases.

P_1	V_1	P_2	V_2
a. 3.0 atm	25 mL	6.0 atm	? mL
b. 99.97 kPa	550. mL	? kPa	275 mL
c. 0.89 atm	? L	3.56 atm	20.0 L
d. ? kPa	800. mL	500. kPa	160. mL
e. 0.040 atm	? L	250 atm	1.0×10^{-2} L

238. A sample of neon gas occupies a volume of 2.8 L at 1.8 atm. What will its volume be at 1.2 atm?

239. To what pressure would you have to compress 48.0 L of oxygen gas at 99.3 kPa in order to reduce its volume to 16.0 L?

240. A chemist collects 59.0 mL of sulfur dioxide gas on a day when the atmospheric pressure is 0.989 atm. On the next day, the pressure has changed to 0.967 atm. What will the volume of the SO_2 gas be on the second day?

241. 2.2 L of hydrogen at 6.5 atm pressure is used to fill a balloon at a final pressure of 1.15 atm. What is its final volume?

Charles's Law

In each of the following problems, assume that the pressure and molar quantity of gas do not change.

242. Calculate the unknown quantity in each of the following measurements of gases:

V_1	T_1	V_2	T_2
a. 40.0 mL	280. K	? mL	350. K
b. 0.606 L	300. K	0.404 L	? K
c. ? mL	292 K	250. mL	365 K
d. 100. mL	? K	125 mL	305 K
e. 0.0024 L	22°C	? L	−14°C

243. A balloon full of air has a volume of 2.75 L at a temperature of 18°C. What is the balloon's volume at 45°C?

244. A sample of argon has a volume of 0.43 mL at 24°C. At what temperature in degrees Celsius will it have a volume of 0.57 mL?

Gay-Lussac's Law

In each of the following problems, assume that the volume and molar quantity of gas do not change.

245. Calculate the unknown quantity in each of the following measurements of gases.

P_1	T_1	P_2	T_2
a. 1.50 atm	273 K	? atm	410 K
b. 0.208 atm	300. K	0.156 atm	? K
c. ? kPa	52°C	99.7 kPa	77°C
d. 5.20 atm	?°C	4.16 atm	−13°C
e. 8.33×10^{-4} atm	−84°C	3.92×10^{-3} atm	? °C

246. A cylinder of compressed gas has a pressure of 4.882 atm on one day. The next day, the same cylinder of gas has a pressure of 4.690 atm, and its temperature is 8°C. What was the temperature on the previous day in °C?

247. A mylar balloon is filled with helium gas to a pressure of 107 kPa when the temperature is 22°C. If the temperature changes to 45°C, what will be the pressure of the helium in the balloon?

The Combined Gas Law

In each of the following problems, it is assumed that the molar quantity of gas does not change.

248. Calculate the unknown quantity in each of the following measurements of gases.

P_1	V_1	T_1	P_2	V_2	T_2
a. 99.3 kPa	225 mL	15°C	102.8 kPa	? mL	24°C
b. 0.959 atm	3.50 L	45°C	? atm	3.70 L	37°C
c. 0.0036 atm	62 mL	373 K	0.0029 atm	64 mL	? K
d. 100. kPa	43.2 mL	19°C	101.3 kPa	? mL	0°C

249. A student collects 450. mL of HCl(*g*) hydrogen chloride gas at a pressure of 100. kPa and a temperature of 17°C. What is the volume of the HCl at 0°C and 101.3 kPa?

Dalton's Law of Partial Pressures

250. A chemist collects a sample of $H_2S(g)$ over water at a temperature of 27°C. The total pressure of the gas that has displaced a volume of 15 mL of water is 207.33 kPa. What is the pressure of the H_2S gas collected?

In each of the following problems, assume that the molar quantity of gas does not change.

251. Some hydrogen is collected over water at 10°C and 105.5 kPa pressure. The total volume of the sample was 1.93 L. Calculate the volume of the hydrogen corrected to STP.

252. One student carries out a reaction that gives off methane gas and obtains a total volume by water displacement of 338 mL at a temperature of 19°C and a pressure of 0.9566 atm. Another student does the identical experiment on another day at a temperature of 26°C and a pressure of 0.989 atm. Which student collected more CH_4?

Mixed Review

In each of the following problems, assume that the molar quantity of gas does not change.

253. Calculate the unknown quantity in each of the following measurements of gases.

P_1	V_1	P_2	V_2
a. 127.3 kPa	796 cm^3	? kPa	965 cm^3
b. 7.1×10^2 atm	? mL	9.6×10^{-1} atm	3.7×10^3 mL
c. ? kPa	1.77 L	30.79 kPa	2.44 L
d. 114 kPa	2.93 dm^3	4.93×10^4 kPa	? dm^3
e. 1.00 atm	120. mL	? atm	97.0 mL
f. 0.77 atm	3.6 m^3	1.90 atm	? m^3

254. A gas cylinder contains 0.722 m^3 of hydrogen gas at a pressure of 10.6 atm. If the gas is used to fill a balloon at a pressure of 0.96 atm, what is the volume in m^3 of the filled balloon?

255. A weather balloon has a maximum volume of 7.50×10^3 L. The balloon contains 195 L of helium gas at a pressure of 0.993 atm. What will be the pressure when the balloon is at maximum volume?

256. A rubber ball contains 5.70×10^{-1} dm^3 of gas at a pressure of 1.05 atm. What volume will the gas occupy at 7.47 atm?

257. Calculate the unknown quantity in each of the following measurements of gases.

V_1	T_1	V_2	T_2
a. 26.5 mL	? K	32.9 mL	290. K
b. ? dm^3	100.°C	0.83 dm^3	29°C
c. 7.44×10^4 mm^3	870.°C	2.59×10^2 mm^3	? °C
d. 5.63×10^{-2} L	132 K	? L	190. K
e. ? cm^3	243 K	819 cm^3	409 K
f. 679 m^3	−3°C	? m^3	−246°C

258. A bubble of carbon dioxide gas in some unbaked bread dough has a volume of 1.15 cm^3 at a temperature of 22°C. What volume will the bubble have when the bread is baked and the bubble reaches a temperature of 99°C?

259. A perfectly elastic balloon contains 6.75 dm^3 of air at a temperature of 40.°C. What is the temperature if the balloon has a volume of 5.03 dm^3?

260. Calculate the unknown quantity in each of the following measurements of gases.

P_1	T_1	P_2	T_2
a. 0.777 atm	?°C	5.6 atm	192°C
b. 152 kPa	302 K	? kPa	11 K
c. ? atm	−76°C	3.97 atm	27°C
d. 395 atm	46°C	706 atm	?°C
e. ? atm	−37°C	350. atm	2050°C
f. 0.39 atm	263 K	0.058 atm	? K

261. A 2 L bottle containing only air is sealed at a temperature of 22°C and a pressure of 0.982 atm. The bottle is placed in a freezer and allowed to cool to −3°C. What is the pressure in the bottle?

262. The pressure in a car tire is 2.50 atm at a temperature of 33°C . What would the pressure be if the tire were allowed to cool to 0°C? Assume that the tire does not change volume.

263. A container filled with helium gas has a pressure of 127.5 kPa at a temperature of 290. K. What is the temperature when the pressure is 3.51 kPa?

264. Calculate the unknown quantity in each of the following measurements of gases.

P_1	V_1	T_1	P_2	V_2	T_2
a. 1.03 atm	1.65 L	19°C	0.920 atm	? L	46°C
b. 107.0 kPa	3.79 dm^3	73°C	? kPa	7.58 dm^3	217°C
c. 0.029 atm	249 mL	? K	0.098 atm	197 mL	293 K
d. 113 kPa	? mm^3	12°C	149 kPa	3.18×10^3 mm^3	−18°C
e. 1.15 atm	0.93 m^3	−22°C	1.01 atm	0.85 m^3	?°C
f. ? atm	156 cm^3	195 K	2.25 atm	468 cm^3	585 K

265. A scientist has a sample of gas that was collected several days earlier. The sample has a volume of 392 cm^3 at a pressure of 0.987 atm and a temperature of 21°C. On the day the gas was collected, the temperature was 13°C and the pressure was 0.992 atm. What volume did the gas have on the day it was collected?

266. Hydrogen gas is collected by water displacement. Total volume collected is 0.461 L at a temperature of 17°C and a pressure of 0.989 atm. What is the pressure of dry hydrogen gas collected?

267. One container with a volume of 1.00 L contains argon at a pressure of 1.77 atm, and a second container of 1.50 L volume contains argon at a pressure of 0.487 atm. They are then connected to each other so that the pressure can become equal in both containers. What is the equalized pressure? Hint: Each sample of gas now occupies the total space. Dalton's law of partial pressures applies here.

268. Oxygen gas is collected over water at a temperature of 10.°C and a pressure of 1.02 atm. The volume of gas plus water vapor collected is 293 mL. What volume of oxygen at STP was collected?

269. A 500 mL bottle is partially filled with water so that the total volume of gases (water vapor and air) remaining in the bottle is 325 cm^3, measured at 20.°C and 101.3 kPa. The bottle is sealed and taken to a mountaintop where the pressure is 76.24 kPa and the temperature is 10°C. If the bottle is upside down and the seal leaks, how much water will leak out? The key to this problem is to determine the pressure in the 325 cm^3 space when the bottle is at the top of the mountain.

270. An air thermometer can be constructed by using a glass bubble attached to a piece of small-diameter glass tubing. The tubing contains a small amount of colored water that rises when the temperature increases and the trapped air expands. You want a 0.20 cm^3 change in volume to equal a 1°C change in temperature. What total volume of air at 20.°C should be trapped in the apparatus below the liquid?

271. A sample of nitrogen gas is collected over water, yielding a total volume of 62.25 mL at a temperature of 22°C and a total pressure of 97.7 kPa. At what pressure will the nitrogen alone occupy a volume of 50.00 mL at the same temperature?

272. The theoretical yield of a reaction that gives off nitrogen trifluoride gas is 844 mL at STP. What total volume of NF_3 plus water vapor will be collected over water at 25°C and a total pressure of 1.017 atm?

273. A weather balloon is inflated with 2.94 kL of helium at a location where the pressure is 1.06 atm and the temperature is 32°C. What will be the volume of the balloon at an altitude where the pressure is 0.092 atm and the temperature is −35°C?

274. The safety limit for a certain can of aerosol spray is 95°C. If the pressure of the gas in the can is 2.96 atm when it is 17°C, what will the pressure be at the safety limit?

275. A chemistry student collects a sample of ammonia gas at a temperature of 39°C. Later, the student measures the volume of the ammonia as 108 mL, but its temperature is now 21°C. What was the volume of the ammonia when it was collected?

276. A quantity of CO_2 gas occupies a volume of 624 L at a pressure of 1.40 atm. If this CO_2 is pumped into a gas cylinder that has a volume of 80.0 L, what pressure will the CO_2 exert on the cylinder?

The Ideal Gas Law: Chap. 11, Sec. 3

277. Use the ideal-gas-law equation to calculate the unknown quantity in each of the following sets of measurements. You will need to convert Celsius temperatures to Kelvin temperatures and volume units to liters.

P	*V*	*n*	*T*
a. 1.09 atm	? L	0.0881 mol	302 K
b. 94.9 kPa	0.0350 L	? mol	55°C
c. ? kPa	15.7 L	0.815 mol	−20.°C
d. 0.500 atm	629 mL	0.0337 mol	? K
e. 0.950 atm	? L	0.0818 mol	19°C
f. 107 kPa	39.0 mL	? mol	27°C

278. A student collects 425 mL of oxygen at a temperature of 24°C and a pressure of 0.899 atm. How many moles of oxygen did the student collect?

Applications of the Ideal Gas Law

279. A sample of an unknown gas has a mass of 0.116 g. It occupies a volume of 25.0 mL at a temperature of 127°C and has a pressure of 155.3 kPa. Calculate the molar mass of the gas.

280. Determine the mass of CO_2 gas that has a volume of 7.10 L at a pressure of 1.11 atm and a temperature of 31°C. Hint: Solve the equation for *m*, and calculate the molar mass using the chemical formula and the periodic table.

281. What is the density of silicon tetrafluoride gas at 72°C and a pressure of 144.5 kPa?

282. At what temperature will nitrogen gas have a density of 1.13 g/L at a pressure of 1.09 atm?

Mixed Review

283. Use the ideal-gas-law equation to calculate the unknown quantity in each of the following sets of measurements.

P	*V*	*n*	*t*
a. 0.0477 atm	15 200 L	? mol	−15°C
b. ? kPa	0.119 mL	0.000 350 mol	0°C
c. 500.0 kPa	250. mL	0.120 mol	?°C
d. 19.5 atm	?	4.7×10^4 mol	300.°C

284. Use the ideal-gas-law equation to calculate the unknown quantity in each of the following sets of measurements.

P	*V*	*m*	*M*	*t*
a. 0.955 atm	3.77 L	8.23 g	? g/mol	25°C
b. 105.0 kPa	50.0 mL	? g	48.02 g/mol	0°C
c. 0.782 atm	? L	3.20×10^{-3} g	2.02 g/mol	−5°C
d. ? atm	2.00 L	7.19 g	159.8 g/mol	185°C
e. 107.2 kPa	26.1 mL	0.414 g	? g/mol	45°C

285. Determine the volume of one mole of an ideal gas at 25°C and 0.915 kPa.

286. Calculate the unknown quantity in each of the following sets of measurements.

P	*Molar Mass*	*Density*	*t*
a. 1.12 atm	? g/mol	2.40 g/L	2°C
b. 7.50 atm	30.07 g/mol	? g/L	20.°C
c. 97.4 kPa	104.09 g/mol	4.37 g/L	? °C
d. ? atm	77.95 g/mol	6.27 g/L	66°C

287. What pressure in atmospheres will 1.36 kg of N_2O gas exert when it is compressed in a 25.0 L cylinder and is stored in an outdoor shed where the temperature can reach 59°C during the summer?

288. Aluminum chloride sublimes at high temperatures. What density will the vapor have at 225°C and 0.939 atm pressure?

289. An unknown gas has a density of 0.0262 g/mL at a pressure of 0.918 atm and a temperature of 10.°C. What is the molar mass of the gas?

290. A large balloon contains 11.7 g of helium. What volume will the helium occupy at an altitude of 10 000 m, where the atmospheric pressure is 0.262 atm and the temperature is −50.°C?

291. A student collects ethane by water displacement at a temperature of 15°C (vapor pressure of water is 1.5988 kPa) and a total pressure of 100.0 kPa. The volume of the collection bottle is 245 mL. How many moles of ethane are in the bottle?

292. A reaction yields 3.75 L of nitrogen monoxide. The volume is measured at 19°C and at a pressure of 1.10 atm. What mass of NO was produced by the reaction?

293. A reaction has a theoretical yield of 8.83 g of ammonia. The reaction gives off 10.24 L of ammonia measured at 52°C and 105.3 kPa. What was the percent yield of the reaction?

294. An unknown gas has a density of 0.405 g/L at a pressure of 0.889 atm and a temperature of 7°C. Calculate its molar mass.

295. A paper label has been lost from an old tank of compressed gas. To help identify the unknown gas, you must calculate its molar mass. It is known that the tank has a capacity of 90.0 L and weighs 39.2 kg when empty. You find its current mass to be 50.5 kg. The gauge shows a pressure of 1780 kPa when the temperature is 18°C. What is the molar mass of the gas in the cylinder?

296. What is the pressure inside a tank that has a volume of 1.20×10^3 L and contains 12.0 kg of HCl gas at a temperature of 18°C?

297. What pressure in kPa is exerted at a temperature of 20.°C by compressed neon gas that has a density of 2.70 g/L?

298. A tank with a volume of 658 mL contains 1.50 g of neon gas. The maximum safe pressure that the tank can withstand is 4.50×10^2 kPa. At what temperature will the tank have that pressure?

299. The atmospheric pressure on Mars is about 6.75 millibars (1 bar = 100 kPa = 0.9869 atm), and the nighttime temperature can be about −75°C on the same day that the daytime temperature goes up to −8°C. What volume would a bag containing 1.00 g of H_2 gas have at both the daytime and nighttime temperatures?

300. What is the pressure in kPa of 3.95 mol of Cl_2 gas if it is compressed in a cylinder with a volume of 850. mL at a temperature of 15°C?

301. What volume in mL will 0.00660 mol of hydrogen gas occupy at a pressure of 0.907 atm and a temperature of 9°C?

302. What volume will 8.47 kg of sulfur dioxide gas occupy at a pressure of 89.4 kPa and a temperature of 40.°C?

303. A cylinder contains 908 g of compressed helium. It is to be used to inflate a balloon to a final pressure of 128.3 kPa at a temperature of 2°C. What will the volume of the balloon be under these conditions?

304. The density of dry air at 27°C and 100.0 kPa is 1.162 g/L. Use this information to calculate the molar mass of air (calculate as if air were a pure substance).

Stoichiometry of Gases: Chap. 11, Sec. 3

305. In one method of manufacturing nitric acid, ammonia is oxidized to nitrogen monoxide and water:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

What volume of oxygen will be used in a reaction of 2800 L of NH_3? What volume of NO will be produced? All volumes are measured under the same conditions.

306. Fluorine gas reacts violently with water to produce hydrogen fluoride and ozone according to the following equation:

$$3F_2(g) + 3H_2O(l) \rightarrow 6HF(g) + O_3(g)$$

What volumes of O_3 and HF gas would be produced by the complete reaction of 3.60×10^4 mL of fluorine gas? All gases are measured under the same conditions.

307. A sample of ethanol burns in O_2 to form CO_2 and H_2O according to the following equation:

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

If the combustion uses 55.8 mL of oxygen measured at 2.26 atm and 40.°C, what volume of CO_2 is produced when measured at STP?

308. Dinitrogen pentoxide decomposes into nitrogen dioxide and oxygen. If 5.00 L of N_2O_5 reacts at STP, what volume of NO_2 is produced when measured at 64.5°C and 1.76 atm?

309. Complete the table below using the following equation, which represents a reaction that produces aluminum chloride:

$$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$$

Mass Al	*Volume Cl_2*	*Conditions*	*Mass $AlCl_3$*
a. excess	? L	STP	7.15 g
b. 19.4 g	? L	STP	NA
c. 1.559 kg	? L	20.°C and 0.945 atm	NA
d. excess	920. L	STP	? g
e. ? g	1.049 mL	37°C and 5.00 atm	NA
f. 500.00 kg	? m^3	15°C and 83.0 kPa	NA

Mixed Review

310. The industrial production of ammonia proceeds according to the following equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

a. What volume of nitrogen at STP is needed to react with 57.0 mL of hydrogen measured at STP?
b. What volume of NH_3 at STP can be produced from the complete reaction of 6.39×10^4 L of hydrogen?
c. If 20.0 mol of nitrogen is available, what volume of NH_3 at STP can be produced?
d. What volume of H_2 at STP will be needed to produce 800. L of ammonia, measured at 55°C and 0.900 atm?

311. Propane burns according to the following equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

a. What volume of water vapor measured at 250.°C and 1.00 atm is produced when 3.0 L of propane at STP is burned?
b. What volume of oxygen at 20.°C and 102.6 kPa is used if 640. L of CO_2 is produced? The CO_2 is also measured at 20.°C and 102.6 kPa.
c. If 465 mL of oxygen at STP is used in the reaction, what volume of CO_2, measured at 37°C and 0.973 atm, is produced?
d. When 2.50 L of C_3H_8 at STP burns, what total volume of gaseous products is formed? The volume of the products is measured at 175°C and 1.14 atm.

312. Carbon monoxide will burn in air to produce CO_2 according to the following equation:

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

What volume of oxygen at STP will be needed to react with 3500. L of CO measured at 20.°C and a pressure of 0.953 atm?

313. Silicon tetrafluoride gas can be produced by the action of HF on silica according to the following equation:

$$SiO_2(s) + 4HF(g) \rightarrow SiF_4(g) + 2H_2O(l)$$

1.00 L of HF gas under pressure at 3.48 atm and a temperature of 25°C reacts completely with SiO_2 to form SiF_4. What volume of SiF_4, measured at 15°C and 0.940 atm, is produced by this reaction?

314. One method used in the eighteenth century to generate hydrogen was to pass steam through red-hot steel tubes. The following reaction takes place:

$$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$$

a. What volume of hydrogen at STP can be produced by the reaction of 6.28 g of iron?
b. What mass of iron will react with 500. L of steam at 250.°C and 1.00 atm pressure?
c. If 285 g of Fe_3O_4 are formed, what volume of hydrogen, measured at 20.°C and 1.06 atm, is produced?

315. Sodium reacts vigorously with water to produce hydrogen and sodium hydroxide according to the following equation:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

If 0.027 g of sodium reacts with excess water, what volume of hydrogen at STP is formed?

316. Diethyl ether burns in air according to the following equation:

$$C_4H_{10}O(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$$

If 7.15 L of CO_2 is produced at a temperature of 125°C and a pressure of 1.02 atm, what volume of oxygen, measured at STP, was consumed and what mass of diethyl ether was burned?

317. When nitroglycerin detonates, it produces large volumes of hot gases almost instantly according to the following equation:

$$4C_3H_5N_3O_9(l) \rightarrow 6N_2(g) + 12CO_2(g) + 10H_2O(g) + O_2(g)$$

a. When 0.100 mol of nitroglycerin explodes, what volume of each gas measured at STP is produced?
b. What total volume of gases is produced at 300.°C and 1.00 atm when 10.0 g of nitroglycerin explodes?

318. Dinitrogen monoxide can be prepared by heating ammonium nitrate, which decomposes according to the following equation:

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$$

What mass of ammonium nitrate should be decomposed in order to produce 250. mL of N_2O, measured at STP?

319. Phosphine, PH_3, is the phosphorus analogue to ammonia, NH_3. It can be produced by the reaction between calcium phosphide and water according to the following equation:

$$Ca_3P_2(s) + 6H_2O(l) \rightarrow 3Ca(OH)_2(s \text{ and } aq) + 2PH_3(g)$$

What volume of phosphine, measured at 18°C and 102.4 kPa, is produced by the reaction of 8.46 g of Ca_3P_2?

320. In one method of producing aluminum chloride, HCl gas is passed over aluminum and the following reaction takes place:

$$2Al(s) + 6HCl(g) \rightarrow 2AlCl_3(g) + 3H_2(g)$$

What mass of Al should be on hand in order to produce 6.0×10^3 kg of $AlCl_3$? What volume of compressed HCl at 4.71 atm and a temperature of 43°C should be on hand at the same time?

321. Urea, $(NH_2)_2CO$, is an important fertilizer that is manufactured by the following reaction:

$$2NH_3(g) + CO_2(g) \rightarrow (NH_2)_2CO(s) + H_2O(g)$$

What volume of NH_3 at STP will be needed to produce 8.50×10^4 kg of urea if there is an 89.5% yield in the process?

322. An obsolete method of generating oxygen in the laboratory involves the decomposition of barium peroxide by the following equation:

$$2BaO_2(s) \rightarrow 2BaO(s) + O_2(g)$$

What mass of BaO_2 reacted if 265 mL of O_2 is collected by water displacement at 0.975 atm and 10.°C?

323. It is possible to generate chlorine gas by dripping concentrated HCl solution onto solid potassium permanganate according to the following equation:

$$2KMnO_4(aq) + 16HCl(aq) \rightarrow 2KCl(aq) + 2MnCl_2(aq) + 8H_2O(l) + 5Cl_2(g)$$

If excess HCl is dripped onto 15.0 g of $KMnO_4$, what volume of Cl_2 will be produced? The Cl_2 is measured at 15°C and 0.959 atm.

324. Ammonia can be oxidized in the presence of a platinum catalyst according to the following equation:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

The NO that is produced reacts almost immediately with additional oxygen according to the following equation:

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

If 35.0 kL of oxygen at STP react in the first reaction, what volume of NH_3 at STP reacts with it? What volume of NO_2 at STP will be formed in the second reaction, assuming there is excess oxygen that was not used up in the first reaction?

325. Oxygen can be generated in the laboratory by heating potassium chlorate. The reaction is represented by the following equation:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

What mass of $KClO_3$ must be used in order to generate 5.00 L of O_2, measured at STP?

326. One of the reactions in the Solvay process is used to make sodium hydrogen carbonate. It occurs when car-

bon dioxide and ammonia are passed through concentrated salt brine. The following equation represents the reaction:

$$NaCl(aq) + H_2O(l) + CO_2(g) + NH_3(g) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$$

a. What volume of NH_3 at 25°C and 1.00 atm pressure will be required if 38 000 L of CO_2, measured under the same conditions, react to form $NaHCO_3$?
b. What mass of $NaHCO_3$ can be formed when the gases in (a) react with NaCl?
c. If this reaction forms 46.0 kg of $NaHCO_3$, what volume of NH_3, measured at STP, reacted?
d. What volume of CO_2, compressed in a tank at 5.50 atm and a temperature of 42°C, will be needed to produce 100.00 kg of $NaHCO_3$?

327. The combustion of butane is represented in the following equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

a. If 4.74 g of butane react with excess oxygen, what volume of CO_2, measured at 150.°C and 1.14 atm, will be formed?
b. What volume of oxygen, measured at 0.980 atm and 75°C, will be consumed by the complete combustion of 0.500 g of butane?
c. A butane-fueled torch has a mass of 876.2 g. After burning for some time, the torch has a mass of 859.3 g. What volume of CO_2, at STP, was formed while the torch burned?
d. What mass of H_2O is produced when butane burns and produces 3720 L of CO_2, measured at 35°C and 0.993 atm pressure?

Concentration of Solutions: Chap. 12, Sec. 3

Percentage Concentration

328. What is the percentage concentration of 75.0 g of ethanol dissolved in 500.0 g of water?

329. A chemist dissolves 3.50 g of potassium iodate and 6.23 g of potassium hydroxide in 805.05 g of water. What is the percentage concentration of each solute in the solution?

330. A student wants to make a 5.00% solution of rubidium chloride using 0.377 g of the substance. What mass of water will be needed to make the solution?

331. What mass of lithium nitrate would have to be dissolved in 30.0 g of water in order to make an 18.0% solution?

Molarity

332. Determine the molarity of a solution prepared by dissolving 141.6 g of citric acid, $C_3H_5O(COOH)_3$, in water and then diluting the resulting solution to 3500.0 mL.

333. What is the molarity of a salt solution made by dissolving 280.0 mg of NaCl in 2.00 mL of water? Assume the final volume is the same as the volume of the water.

334. What is the molarity of a solution that contains 390.0 g of acetic acid, CH_3COOH, dissolved in enough acetone to make 1000.0 mL of solution?

335. What mass of glucose, $C_6H_{12}O_6$, would be required to prepare 5.000×10^3 L of a 0.215 M solution?

336. What mass of magnesium bromide would be required to prepare 720. mL of a 0.0939 M aqueous solution?

337. What mass of ammonium chloride is dissolved in 300. mL of a 0.875 M solution?

Molality

338. Determine the molality of a solution of 560 g of acetone, CH_3COCH_3, in 620 g of water.

339. What is the molality of a solution of 12.9 g of fructose, $C_6H_{12}O_6$, in 31.0 g of water?

340. How many moles of 2-butanol, $CH_3CHOHCH_2CH_3$, must be dissolved in 125 g of ethanol in order to produce a 12.0 *m* 2-butanol solution? What mass of 2-butanol is this?

Mixed Review

341. Complete the table below by determining the missing quantity in each example. All solutions are aqueous. Any quantity that is not applicable to a given solution is marked NA.

Solution Made	*Mass of Solute Used*	*Quantity of Solution Made*	*Quantity of Solvent Used*
a. 12.0% $KMnO_4$	? g $KMnO_4$	500.0 g	? g H_2O
b. 0.60 M $BaCl_2$	? g $BaCl_2$	1.750 L	NA
c. 6.20 *m* glycerol, $HOCH_2CHOHCH_2OH$	? g glycerol	NA	800.0 g H_2O
d. ? M $K_2Cr_2O_7$	12.27 g $K_2Cr_2O_7$	650. mL	NA
e. ? *m* $CaCl_2$	288 g $CaCl_2$	NA	2.04 kg H_2O
f. 0.160 M NaCl	? g NaCl	25.0 mL	NA
g. 2.00 *m* glucose, $C_6H_{12}O_6$	? g glucose	? g solution	1.50 kg H_2O

342. How many moles of H_2SO_4 are in 2.50 L of a 4.25 M aqueous solution?

343. Determine the molal concentration of 71.5 g of linoleic acid, $C_{18}H_{32}O_2$, in 525 g of hexane, C_6H_{14}.

344. You have a solution that is 16.2% sodium thiosulfate, $Na_2S_2O_3$, by mass.
a. What mass of sodium thiosulfate is in 80.0 g of solution?
b. How many moles of sodium thiosulfate are in 80.0 g of solution?
c. If 80.0 g of the sodium thiosulfate solution is diluted to 250.0 mL with water, what is the molarity of the resulting solution?

345. What mass of anhydrous cobalt(II) chloride would be needed in order to make 650.00 mL of a 4.00 M cobalt(II) chloride solution?

346. A student wants to make a 0.150 M aqueous solution of silver nitrate, $AgNO_3$, and has a bottle containing 11.27 g of silver nitrate. What should be the final volume of the solution?

347. What mass of urea, NH_2CONH_2, must be dissolved in 2250 g of water in order to prepare a 1.50 *m* solution?

348. What mass of barium nitrate is dissolved in 21.29 mL of a 3.38 M solution?

349. Describe what you would do to prepare 100.0 g of a 3.5% solution of ammonium sulfate in water.

350. What mass of anhydrous calcium chloride should be dissolved in 590.0 g of water in order to produce a 0.82 *m* solution?

351. How many moles of ammonia are in 0.250 L of a 5.00 M aqueous ammonia solution? If this solution were diluted to 1.000 L, what would be the molarity of the resulting solution?

352. What is the molar mass of a solute if 62.0 g of the solute in 125 g of water produce a 5.3 *m* solution?

353. A saline solution is 0.9% NaCl. What masses of NaCl and water would be required to prepare 50. L of this saline solution? Assume that the density of water is 1.000 g/mL and that the NaCl does not add to the volume of the solution.

354. A student weighs an empty beaker on a balance and finds its mass to be 68.60 g. The student weighs the beaker again after adding water and finds the new mass to be 115.12 g. A mass of 4.08 g of glucose is then dissolved in the water. What is the percentage concentration of glucose in the solution?

355. The density of ethyl acetate at 20°C is 0.902 g/mL. What volume of ethyl acetate at 20°C would be required to prepare a 2.0% solution of cellulose nitrate using 25 g of cellulose nitrate?

356. Aqueous cadmium chloride reacts with sodium sulfide to produce bright-yellow cadmium sulfide. Write the balanced equation for this reaction and answer the following questions.
   a. How many moles of $CdCl_2$ are in 50.00 mL of a 3.91 M solution?
   b. If the solution in (a) reacted with excess sodium sulfide, how many moles of CdS would be formed?
   c. What mass of CdS would be formed?

357. What mass of H_2SO_4 is contained in 60.00 mL of a 5.85 M solution of sulfuric acid?

358. A truck carrying 22.5 kL of 6.83 M aqueous hydrochloric acid used to clean brick and masonry has overturned. The authorities plan to neutralize the acid with sodium carbonate. How many moles of HCl will have to be neutralized?

359. A chemist wants to produce 12.00 g of barium sulfate by reacting a 0.600 M $BaCl_2$ solution with excess H_2SO_4, as shown in the reaction below. What volume of the $BaCl_2$ solution should be used?

$$BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$$

360. Many substances are hydrates. Whenever you make a solution, it is important to know whether or not the solute you are using is a hydrate and, if it is a hydrate, how many molecules of water are present per formula unit of the substance. This water must be taken into account when weighing out the solute. Something else to remember when making aqueous solutions from hydrates is that once the hydrate is dissolved, the water of hydration is considered to be part of the solvent. A common hydrate used in the chemistry laboratory is copper sulfate pentahydrate, $CuSO_4•5H_2O$. Describe how you would make each of the following solutions using $CuSO_4•5H_2O$. Specify masses and volumes as needed.
   a. 100. g of a 6.00% solution of $CuSO_4$
   b. 1.00 L of a 0.800 M solution of $CuSO_4$
   c. a 3.5 *m* solution of $CuSO_4$ in 1.0 kg of water

361. What mass of calcium chloride hexahydrate is required in order to make 700.0 mL of a 2.50 M solution?

362. What mass of the amino acid arginine, $C_6H_{14}N_4O_2$, would be required to make 1.250 L of a 0.00205 M solution?

363. How much water would you have to add to 2.402 kg of nickel(II) sulfate hexahydrate in order to prepare a 25.00% solution?

364. What mass of potassium aluminum sulfate dodecahydrate, $KAl(SO_4)_2•12H_2O$, would be needed to prepare 35.00 g of a 15.00% $KAl(SO_4)_2$ solution? What mass of water would be added to make this solution?

Dilutions: Chap. 12, Sec. 3

365. Complete the table below by calculating the missing value in each row.

Molarity of Stock Solution	*Volume of Stock Solution*	*Molarity of Dilute Solution*	*Volume of Dilute Solution*
a. 0.500 M KBr	20.00 mL	? M KBr	100.00 mL
b. 1.00 M LiOH	? mL	0.075 M LiOH	500.00 mL
c. ? M HI	5.00 mL	0.0493 M HI	100.00 mL
d. 12.0 M HCl	0.250 L	1.8 M HCl	? L
e. 7.44 M NH_3	? mL	0.093 M NH_3	4.00 L

366. What volume of water would be added to 16.5 mL of a 0.0813 M solution of sodium borate in order to get a 0.0200 M solution?

Mixed Review

367. What is the molarity of a solution of ammonium chloride prepared by diluting 50.00 mL of a 3.79 M NH_4Cl solution to 2.00 L?

368. A student takes a sample of KOH solution and dilutes it with 100.00 mL of water. The student determines that the diluted solution is 0.046 M KOH, but has forgotten to record the volume of the original sample. The concentration of the original solution is 2.09 M. What was the volume of the original sample?

369. A chemist wants to prepare a stock solution of H_2SO_4 so that samples of 20.00 mL will produce a solution with a concentration of 0.50 M when added to 100.0 mL of water.

a. What should the molarity of the stock solution be?

b. If the chemist wants to prepare 5.00 L of the stock solution from concentrated H_2SO_4, which is 18.0 M, what volume of concentrated acid should be used?

c. The density of 18.0 M H_2SO_4 is 1.84 g/mL. What mass of concentrated H_2SO_4 should be used to make the stock solution in (b)?

370. To what volume should 1.19 mL of an 8.00 M acetic acid solution be diluted in order to obtain a final solution that is 1.50 M?

371. What volume of a 5.75 M formic acid solution should be used to prepare 2.00 L of a 1.00 M formic acid solution?

372. A 25.00 mL sample of ammonium nitrate solution produces a 0.186 M solution when diluted with 50.00 mL of water. What is the molarity of the stock solution?

373. Given a solution of known percentage concentration by mass, a laboratory worker can often measure out a calculated mass of the solution in order to obtain a certain mass of solute. Sometimes, though, it is impractical to use the mass of a solution, especially with fuming solutions, such as concentrated HCl and concentrated HNO_3. Measuring these solutions by volume is much more practical. In order to determine the volume that should be measured, a worker would need to know the density of the solution. This information usually appears on the label of the solution bottle.

a. Concentrated hydrochloric acid is 36% HCl by mass and has a density of 1.18 g/mL. What is the volume of 1.0 kg of this HCl solution? What volume contains 1.0 g of HCl? What volume contains 1.0 mol of HCl?

b. The density of concentrated nitric acid is 1.42 g/mL, and its concentration is 71% HNO_3 by mass. What volume of concentrated HNO_3 would be needed to prepare 10.0 L of a 2.00 M solution of HNO_3?

c. What volume of concentrated HCl solution would be needed to prepare 4.50 L of 3.0 M HCl? See (a) for data.

374. A 3.8 M solution of $FeSO_4$ solution is diluted to eight times its original volume. What is the molarity of the diluted solution?

375. A chemist prepares 480. mL of a 2.50 M solution of $K_2Cr_2O_7$ in water. A week later, the chemist wants to use the solution, but the stopper has been left off the flask and 39 mL of water has evaporated. What is the new molarity of the solution?

376. You must write out procedures for a group of lab technicians. One test they will perform requires 25.00 mL of a 1.22 M solution of acetic acid. You decide to use a 6.45 M acetic acid solution that you have on hand. What procedure should the technicians use in order to get the solution they need?

377. A chemical test has determined the concentration of a solution of an unknown substance to be 2.41 M. A 100.0 mL volume of the solution is evaporated to dryness, leaving 9.56 g of crystals of the unknown solute. Calculate the molar mass of the unknown substance.

378. Tincture of iodine can be prepared by dissolving 34 g of I_2 and 25 g of KI in 25 mL of distilled water and diluting the solution to 500. mL with ethanol. What is the molarity of I_2 in the solution?

379. Phosphoric acid is commonly supplied as an 85% solution. What mass of this solution would be required to prepare 600.0 mL of a 2.80 M phosphoric acid solution?

380. Commercially available concentrated sulfuric acid is 18.0 M H_2SO_4. What volume of concentrated H_2SO_4 would you use in order to make 3.00 L of a 4.0 M stock solution?

381. Describe how to prepare 1.00 L of a 0.495 M solution of urea, NH_2CONH_2, starting with a 3.07 M stock solution.

382. Honey is a solution consisting almost entirely of a mixture of the hexose sugars fructose and glucose; both sugars have the formula $C_6H_{12}O_6$, but they differ in molecular structure.

a. A sample of honey is found to be 76.2% $C_6H_{12}O_6$ by mass. What is the molality of the hexose sugars in honey? Consider the sugars to be equivalent.

b. The density of the honey sample is 1.42 g/mL. What mass of hexose sugars are in 1.00 L of honey? What is the molarity of the mixed hexose sugars in honey?

383. Industrial chemicals used in manufacturing are almost never pure, and the content of the material may vary from one batch to the next. For these reasons, a sample is taken from each shipment and sent to a laboratory, where its makeup is determined. This procedure is called assaying. Once the content of a material is known, engineers adjust the manufacturing process to account for the degree of purity of the starting chemicals.

Suppose you have just received a shipment of sodium carbonate, Na_2CO_3. You weigh out 50.00 g of the material, dissolve it in water, and dilute the solution to 1.000 L. You remove 10.00 mL from the solution and dilute it to 50.00 mL. By measuring the amount of a second substance that reacts with Na_2CO_3, you determine that the concentration of sodium carbonate in the diluted solution is 0.0890 M. Calculate the percentage of Na_2CO_3 in the original batch of material. The molar mass of Na_2CO_3 is 105.99 g. (Hint: Determine the number of moles in the original solution and convert to mass of Na_2CO_3.)

384. A student wants to prepare 0.600 L of a stock solution of copper(II) chloride so that 20.0 mL of the stock solution diluted by adding 130.0 mL of water will yield a 0.250 M solution. What mass of $CuCl_2$ should be used to make the stock solution?

385. You have a bottle containing a 2.15 M $BaCl_2$ solution. You must tell other students how to dilute this solution to get various volumes of a 0.65 M $BaCl_2$ solution. By what factor will you tell them to dilute the stock solution? In other words, when a student removes any volume, V, of the stock solution, how many times V of water should be added to dilute to 0.65 M?

386. You have a bottle containing an 18.2% solution of strontium nitrate (density = 1.02 g/mL).
a. What mass of strontium nitrate is dissolved in 80.0 mL of this solution?
b. How many moles of strontium nitrate are dissolved in 80.0 mL of the solution?
c. If 80.0 mL of this solution is diluted with 420.0 mL of water, what is the molarity of the solution?

Colligative Properties: Chap. 13, Sec. 2

387. Determine the freezing point of a solution of 60.0 g of glucose, $C_6H_{12}O_6$, dissolved in 80.0 g of water.

388. What is the freezing point of a solution of 645 g of urea, H_2NCONH_2, dissolved in 980. g of water?

389. What is the expected boiling point of a brine solution containing 30.00 g of KBr dissolved in 100.00 g of water?

390. What is the expected boiling point of a $CaCl_2$ solution containing 385 g of $CaCl_2$ dissolved in 1.230×10^3 g of water?

391. A solution of 0.827 g of an unknown non-electrolyte compound in 2.500 g of water has a freezing point of −10.18°C. Calculate the molar mass of the compound.

392. A 0.171 g sample of an unknown organic compound is dissolved in ether. The solution has a total mass of 2.470 g. The boiling point of the solution is found to be 36.43°C. What is the molar mass of the organic compound?

Mixed Review

In each of the following problems, assume that the solute is a nonelectrolyte unless otherwise stated.

393. Calculate the freezing point and boiling point of a solution of 383 g of glucose dissolved in 400. g of water.

394. Determine the boiling point of a solution of 72.4 g of glycerol dissolved in 122.5 g of water.

395. What is the boiling point of a solution of 30.20 g of ethylene glycol, $HOCH_2CH_2OH$, in 88.40 g of phenol?

396. What mass of ethanol, CH_3CH_2OH, should be dissolved in 450. g of water to obtain a freezing point of −4.5°C?

397. Calculate the molar mass of a nonelectrolyte that lowers the freezing point of 25.00 g of water to −3.9°C when 4.27 g of the substance is dissolved in the water.

398. What is the freezing point of a solution of 1.17 g of 1-naphthol, $C_{10}H_8O$, dissolved in 2.00 mL of benzene at 20°C? The density of benzene at 20°C is 0.876 g/mL. K_f for benzene is −5.12°C/m, and benzene's normal freezing point is 5.53°C.

399. The boiling point of a solution containing 10.44 g of an unknown nonelectrolyte in 50.00 g of acetic acid is 159.2°C. What is the molar mass of the solute?

400. A 0.0355 g sample of an unknown molecular compound is dissolved in 1.000 g of liquid camphor at 200.0°C. Upon cooling, the camphor freezes at 157.7°C. Calculate the molar mass of the unknown compound.

401. Determine the boiling point of a solution of 22.5 g of fructose, $C_6H_{12}O_6$, in 294 g of phenol.

402. Ethylene glycol, $HOCH_2CH_2OH$, is effective as an antifreeze, but it also raises the boiling temperature of automobile coolant, which helps prevent loss of coolant when the weather is hot.
a. What is the freezing point of a 50.0% solution of ethylene glycol in water?
b. What is the boiling point of the same 50.0% solution?

403. The value of K_f for cyclohexane is −20.0°C/m, and its normal freezing point is 6.6°C. A mass of 1.604 g of a waxy solid dissolved in 10.000 g of cyclohexane results in a freezing point of −4.4°C. Calculate the molar mass of the solid.

404. What is the expected freezing point of an aqueous solution of 2.62 kg of nitric acid, HNO_3, in a solution with a total mass of 5.91 kg? Assume that the nitric acid is completely ionized.

405. An unknown organic compound is mixed with 0.5190 g of naphthalene crystals to give a mixture having a total mass of 0.5959 g. The mixture is heated until the naphthalene melts and the unknown substance dissolves. Upon cooling, the solution freezes at a temperature of 74.8°C. What is the molar mass of the unknown compound?

406. What is the boiling point of a solution of 8.69 g of the electrolyte sodium acetate, $NaCH_3COO$, dissolved in 15.00 g of water?

407. What is the expected freezing point of a solution of 110.5 g of H_2SO_4 in 225 g of water? Assume sulfuric acid completely dissociates in water.

408. A compound called pyrene has the empirical formula C_8H_5. When 4.04 g of pyrene is dissolved in 10.00 g of benzene, the boiling point of the solution is 85.1°C. Calculate the molar mass of pyrene and determine its molecular formula. The molal boiling-point constant for benzene is 2.53°C/m. Its normal boiling point is 80.1°C.

409. What mass of $CaCl_2$, when dissolved in 100.00 g of water, gives an expected freezing point of −5.0°C; $CaCl_2$ is ionic? What mass of glucose would give the same result?

410. A compound has the empirical formula CH_2O. When 0.0866 g is dissolved in 1.000 g of ether, the solution's boiling point is 36.5°C. Determine the molecular formula of this substance.

411. What is the freezing point of a 28.6% (by mass) aqueous solution of HCl? Assume the HCl is 100% ionized.

412. What mass of ethylene glycol, $HOCH_2CH_2OH$, must be dissolved in 4.510 kg of water to result in a freezing point of − 18.0°C? What is the boiling point of the same solution?

413. A water solution containing 2.00 g of an unknown molecular substance dissolved in 10.00 g of water has a freezing point of − 4.0°C.
- **a.** Calculate the molality of the solution.
- **b.** When 2.00 g of the substance is dissolved in acetone instead of in water, the boiling point of the solution is 58.9°C. The normal boiling point of acetone is 56.00°C, and its K_b is 1.71°C/*m*. Calculate the molality of the solution from this data.

414. A chemist wants to prepare a solution with a freezing point of − 22.0°C and has 100.00 g of glycerol on hand. What mass of water should the chemist mix with the glycerol?

415. An unknown carbohydrate compound has the empirical formula CH_2O. A solution consisting of 0.515 g of the carbohydrate dissolved in 1.717 g of acetic acid freezes at 8.8°C. What is the molar mass of the carbohydrate? What is its molecular formula?

416. An unknown organic compound has the empirical formula C_2H_2O. A solution of 3.775 g of the unknown compound dissolved in 12.00 g of water is cooled until it freezes at a temperature of − 4.72°C. Determine the molar mass and the molecular formula of the compound.

pH: Chap. 15, Sec. 1

417. The hydroxide ion concentration of an aqueous solution is 6.4×10^{-5} M. What is the hydronium ion concentration?

418. Calculate the H_3O^+ and OH^- concentrations in a 7.50×10^{-4} M solution of HNO_3, a strong acid.

419. Determine the pH of a 0.001 18 M solution of HBr.

420.
- **a.** What is the pH of a solution that has a hydronium ion concentration of 1.0 M?
- **b.** What is the pH of a 2.0 M solution of HCl, assuming the acid remains 100% ionized?
- **c.** What is the theoretical pH of a 10. M solution of HCl?

421. What is the pH of a solution with the following hydroxide ion concentrations?
- **a.** 1×10^{-5} M
- **b.** 5×10^{-8} M
- **c.** 2.90×10^{-11} M

422. What are the pOH and hydroxide ion concentration of a solution with a pH of 8.92?

423. What are the pOH values of solutions with the following hydronium ion concentrations?
- **a.** 2.51×10^{-13} M
- **b.** 4.3×10^{-3} M
- **c.** 9.1×10^{-6} M
- **d.** 0.070 M

424. A solution is prepared by dissolving 3.50 g of sodium hydroxide in water and adding water until the total volume of the solution is 2.50 L. What are the OH^- and H_3O^+ concentrations?

425. If 1.00 L of a potassium hydroxide solution with a pH of 12.90 is diluted to 2.00 L, what is the pH of the resulting solution?

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426. Calculate the H_3O^+ and OH^- concentrations in the following solutions. Each is either a strong acid or a strong base.
- **a.** 0.05 M sodium hydroxide
- **b.** 0.0025 M sulfuric acid
- **c.** 0.013 M lithium hydroxide
- **d.** 0.150 M nitric acid
- **e.** 0.0200 M calcium hydroxide
- **f.** 0.390 M perchloric acid

427. What is the pH of each solution in item 426?

428. Calculate $[H_3O^+]$ and $[OH^-]$ in a 0.160 M solution of potassium hydroxide. Assume that the solute is 100% dissociated at this concentration.

429. The pH of an aqueous solution of NaOH is 12.9. What is the molarity of the solution?

430. What is the pH of a 0.001 25 M HBr solution? If 175 mL of this solution is diluted to a total volume of 3.00 L, what is the pH of the diluted solution?

431. What is the pH of a 0.0001 M solution of NaOH? What is the pH of a 0.0005 M solution of NaOH?

432. A solution is prepared using 15.0 mL of 1.0 M HCl and 20.0 mL of 0.50 M HNO_3. The final volume of the solution is 1.25 L. Answer the following questions:
- **a.** What are the $[H_3O^+]$ and $[OH^-]$ in the final solution?
- **b.** What is the pH of the final solution?

433. A container is labeled 500.0 mL of 0.001 57 M nitric acid solution. A chemist finds that the container was not sealed and that some evaporation has taken place. The volume of solution is now 447.0 mL.
- **a.** What was the original pH of the solution?
- **b.** What is the pH of the solution now?

434. Calculate the hydroxide ion concentration in an aqueous solution that has a 0.000 35 M hydronium ion concentration.

435. A solution of sodium hydroxide has a pH of 12.14. If 50.00 mL of the solution is diluted to 2.000 L with water, what is the pH of the diluted solution?

436. An acetic acid solution has a pH of 4.0. What are the $[H_3O^+]$ and $[OH^-]$ in this solution?

437. What is the pH of a 0.000 460 M solution of $Ca(OH)_2$?

438. A solution of strontium hydroxide with a pH of 11.4 is to be prepared. What mass of strontium hydroxide would be required to make 1.00 L of this solution?

439. A solution of NH_3 has a pH of 11.00. What are the concentrations of hydronium and hydroxide ions in this solution?

440. Acetic acid does not completely ionize in solution. Percent ionization of a substance dissolved in water is equal to the moles of ions produced as a percentage of the moles of ions that would be produced if the substance were completely ionized. Calculate the percent ionization of acetic acid in the following solutions.
a. 1.0 M acetic acid solution with a pH of 2.40
b. 0.10 M acetic acid solution with a pH of 2.90
c. 0.010 M acetic acid solution with a pH of 3.40

441. Calculate the pH of a solution that contains 5.00 g of HNO_3 in 2.00 L of solution.

442. A solution of HCl has a pH of 1.50. Determine the pH of the solutions made in each of the following ways.
a. 1.00 mL of the solution is diluted to 1000. mL with water.
b. 25.00 mL is diluted to 200. mL with distilled water.
c. 18.83 mL of the solution is diluted to 4.000 L with distilled water.
d. 1.50 L is diluted to 20.0 kL with distilled water.

443. An aqueous solution contains 10 000 times more hydronium ions than hydroxide ions. What is the concentration of each ion?

444. A potassium hydroxide solution has a pH of 12.90. Enough acid is added to react with half of the OH^- ions present. What is the pH of the resulting solution? Assume that the products of the neutralization have no effect on pH and that the amount of additional water produced is negligible.

445. A hydrochloric acid solution has a pH of 1.70. What is the $[H_3O^+]$ in this solution? Considering that HCl is a strong acid, what is the HCl concentration of the solution?

446. What is the molarity of a solution of the strong base $Ca(OH)_2$ in a solution that has a pH of 10.80?

447. You have a 1.00 M solution of the strong acid, HCl. What is the pH of this solution? You need a solution of pH 4.00. To what volume would you dilute 1.00 L of the HCl solution to get this pH? To what volume would you dilute 1.00 L of the pH 4.00 solution to get a solution of pH 6.00? To what volume would you dilute 1.00 L of the pH 4.00 solution to get a solution of pH 8.00?

448. A solution of chloric acid, $HClO_3$, a strong acid, has a pH of 1.28. How many moles of NaOH would be required to react completely with the $HClO_3$ in 1.00 L of the solution? What mass of NaOH is required?

449. A solution of the weak base NH_3 has a pH of 11.90. How many moles of HCl would have to be added to 1.00 L of the ammonia to react with all of the OH^- ions present at pH 11.90?

450. The pH of a citric acid solution is 3.15. What are the $[H_3O^+]$ and $[OH^-]$ in this solution?

Titrations: Chap. 15, Sec. 2

In each of the following problems, the acids and bases react in a mole ratio of 1 mol base : 1 mol acid.

451. A student titrates a 20.00 mL sample of a solution of HBr with unknown molarity. The titration requires 20.05 mL of a 0.1819 M solution of NaOH. What is the molarity of the HBr solution?

452. Vinegar can be assayed to determine its acetic acid content. Determine the molarity of acetic acid in a 15.00 mL sample of vinegar that requires 22.70 mL of a 0.550 M solution of NaOH to reach the equivalence point.

453. A 20.00 mL sample of a solution of $Sr(OH)_2$ is titrated to the equivalence point with 43.03 mL of 0.1159 M HCl. What is the molarity of the $Sr(OH)_2$ solution?

454. A 35.00 mL sample of ammonia solution is titrated to the equivalence point with 54.95 mL of a 0.400 M sulfuric acid solution. What is the molarity of the ammonia solution?

In the problems below, assume that impurities are not acidic or basic and that they do not react in an acid-base titration.

455. A supply of glacial acetic acid has absorbed water from the air. It must be assayed to determine the actual percentage of acetic acid. 2.000 g of the acid is diluted to 100.00 mL, and 20.00 mL is titrated with a solution of sodium hydroxide. The base solution has a concentration of 0.218 M, and 28.25 mL is used in the titration. Calculate the percentage of acetic acid in the original sample. Write the titration equation to get the mole ratio.

456. A shipment of crude sodium carbonate must be assayed for its Na_2CO_3 content. You receive a small jar containing a sample from the shipment and weigh out 9.709 g into a flask, where it is dissolved in water and diluted to 1.0000 L with distilled water. A 10.00 mL sample is taken from the flask and titrated to the equivalence point with 16.90 mL of a 0.1022 M HCl solution. Determine the percentage of Na_2CO_3 in the sample. Write the titration equation to get the mole ratio.

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457. A 50.00 mL sample of a potassium hydroxide is titrated with a 0.8186 M HCl solution. The titration requires 27.87 mL of the HCl solution to reach the equivalence point. What is the molarity of the KOH solution?

458. A 15.00 mL sample of acetic acid is titrated with 34.13 mL of 0.9940 M NaOH. Determine the molarity of the acetic acid.

459. A 12.00 mL sample of an ammonia solution is titrated with 1.499 M HNO_3 solution. A total of 19.48 mL of acid is required to reach the equivalence point. What is the molarity of the ammonia solution?

460. A certain acid and base react in a 1:1 ratio.

a. If the acid and base solutions are of equal concentration, what volume of acid will titrate a 20.00 mL sample of the base?

b. If the acid is twice as concentrated as the base, what volume of acid will be required to titrate 20.00 mL of the base?

c. How much acid will be required if the base is four times as concentrated as the acid, and 20.00 mL of base is used?

461. A 10.00 mL sample of a solution of hydrofluoric acid, HF, is diluted to 500.00 mL. A 20.00 mL sample of the diluted solution requires 13.51 mL of a 0.1500 M NaOH solution to be titrated to the equivalence point. What is the molarity of the original HF solution?

462. A solution of oxalic acid, a diprotic acid, is used to titrate a 16.22 mL sample of a 0.5030 M KOH solution. If the titration requires 18.41 mL of the oxalic acid solution, what is its molarity?

463. A H_2SO_4 solution of unknown molarity is titrated with a 1.209 M NaOH solution. The titration requires 42.27 mL of the NaOH solution to reach the equivalent point with 25.00 mL of the H_2SO_4 solution. What is the molarity of the acid solution?

464. Potassium hydrogen phthalate, $KHC_8H_4O_4$, is a solid acidic substance that reacts in a 1:1 mole ratio with bases that have one hydroxide ion. Suppose that 0.7025 g of potassium hydrogen phthalate is titrated to the equivalence point by 20.18 mL of a KOH solution. What is the molarity of the KOH solution?

465. A solution of citric acid, a triprotic acid, is titrated with a sodium hydroxide solution. A 20.00 mL sample of the citric acid solution requires 17.03 mL of a 2.025 M solution of NaOH to reach the equivalence point. What is the molarity of the acid solution?

466. A flask contains 41.04 mL of a solution of potassium hydroxide. The solution is titrated and reaches an equivalence point when 21.65 mL of a 0.6515 M solution of HNO_3 is added. Calculate the molarity of the base solution.

467. A bottle is labeled 2.00 M H_2SO_4. You decide to titrate a 20.00 mL sample with a 1.85 M NaOH solution. What volume of NaOH solution would you expect to use if the label is correct?

468. What volume of a 0.5200 M solution of H_2SO_4 would be needed to titrate 100.00 mL of a 0.1225 M solution of $Sr(OH)_2$?

469. A sample of a crude grade of KOH is sent to the lab to be tested for KOH content. A 4.005 g sample is dissolved and diluted to 200.00 mL with water. A 25.00 mL sample of the solution is titrated with a 0.4388 M HCl solution and requires 19.93 mL to reach the equivalence point. How many moles of KOH were in the 4.005 g sample? What mass of KOH is this? What is the percent of KOH in the crude material?

470. What mass of magnesium hydroxide would be required for the magnesium hydroxide to react to the equivalence point with 558 mL of 3.18 M hydrochloric acid?

471. An ammonia solution of unknown concentration is titrated with a solution of hydrochloric acid. The HCl solution is 1.25 M, and 5.19 mL are required to titrate 12.61 mL of the ammonia solution. What is the molarity of the ammonia solution?

472. What volume of 2.811 M oxalic acid solution is needed to react to the equivalence point with a 5.090 g sample of material that is 92.10% NaOH? Oxalic acid is a diprotic acid.

473. Standard solutions of accurately known concentration are available in most laboratories. These solutions are used to titrate other solutions to determine their concentrations. Once the concentration of the other solutions are accurately known, they may be used to titrate solutions of unknowns.

The molarity of a solution of HCl is determined by titrating the solution with an accurately known solution of $Ba(OH)_2$, which has a molar concentration of 0.1529 M. A volume of 43.09 mL of the $Ba(OH)_2$ solution titrates 26.06 mL of the acid solution. The acid solution is in turn used to titrate 15.00 mL of a solution of rubidium hydroxide. The titration requires 27.05 mL of the acid.

a. What is the molarity of the HCl solution?

b. What is the molarity of the RbOH solution?

474. A truck containing 2800 kg of a 6.0 M hydrochloric acid has been in an accident and is in danger of spilling its load. What mass of $Ca(OH)_2$ should be sent to the scene in order to neutralize all of the acid in case the tank bursts? The density of the 6.0 M HCl solution is 1.10 g/mL.

475. A 1.00 mL sample of a fairly concentrated nitric acid solution is diluted to 200.00 mL. A 10.00 mL sample of the diluted solution requires 23.94 mL of a 0.0177 M solution of $Ba(OH)_2$ to be titrated to the equivalence point. Determine the molarity of the original nitric acid solution.

476. What volume of 4.494 M H_2SO_4 solution would be required to react to the equivalence point with 7.2280 g of $LiOH(s)$?

Thermochemistry: Chap. 16, Sec. 1

477. Calculate the reaction enthalpy for the following reaction:

$$5CO_2(g) + Si_3N_4(s) \rightarrow 3SiO(s) + 2N_2O(g) + 5CO(g)$$

Use the following equations and data:

(1) $CO(g) + SiO_2(s) \rightarrow SiO(g) + CO_2(g)$

(2) $8CO_2(g) + Si_3N_4(s) \rightarrow 3SiO_2(s) + 2N_2O(g) + 8CO(g)$

$\Delta H_{\text{reaction 1}} = +520.9$ kJ

$\Delta H_{\text{reaction 2}} = +461.05$ kJ

Determine ΔH for each of the following three reactions.

478. The following reaction is used to make CaO from limestone:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

479. The following reaction represents the oxidation of FeO to Fe_2O_3:

$$2FeO(s) + O_2(g) \rightarrow Fe_2O_3(s)$$

480. The following reaction of ammonia and hydrogen fluoride produces ammonium fluoride:

$$NH_3(g) + HF(g) \rightarrow NH_4F(s)$$

481. Calculate the free energy change, ΔG, for the combustion of hydrogen sulfide according to the following chemical equation. Assume reactants and products are at 25°C:

$$H_2S(g) + O_2(g) \rightarrow H_2O(l) + SO_2(g)$$

$\Delta H_{reaction} = -562.1$ kJ/mol

$\Delta S_{reaction} = -0.09278$ kJ/mol•K

482. Calculate the free energy change for the decomposition of sodium chlorate. Assume reactants and products are at 25°C:

$$NaClO_3(s) \rightarrow NaCl(s) + O_2(g)$$

$\Delta H_{reaction} = -19.1$ kJ/mol

$\Delta S_{reaction} = -0.1768$ kJ/mol•K

483. Calculate the free energy change for the combustion of 1 mol of ethane. Assume reactants and products are at 25°C:

$$C_2H_6(g) + O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

$\Delta H_{reaction} = -1561$ kJ/mol

$\Delta S_{reaction} = -0.4084$ kJ/mol•K

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484. Calculate ΔH for the reaction of fluorine with water:

$$F_2(g) + H_2O(l) \rightarrow 2HF(g) + O_2(g)$$

485. Calculate ΔH for the reaction of calcium oxide and sulfur trioxide:

$$CaO(s) + SO_3(g) \rightarrow CaSO_4(s)$$

Use the following equations and data:

$$H_2O(l) + SO_3(g) \rightarrow H_2SO_4(l)$$

$\Delta H = -132.5$ kJ/mol

$$H_2SO_4(l) + Ca(s) \rightarrow CaSO_4(s) + H_2(g)$$

$\Delta H = -602.5$ kJ/mol

$$Ca(s) + O_2(g) \rightarrow CaO(s)$$

$\Delta H = -634.9$ kJ/mol

$$H_2(g) + O_2(g) \rightarrow H_2O(l)$$

$\Delta H = -285.8$ kJ/mol

486. Calculate ΔH for the reaction of sodium oxide with sulfur dioxide:

$$Na_2O(s) + SO_2(g) \rightarrow Na_2SO_3(s)$$

487. Use enthalpies of combustion to calculate ΔH for the oxidation of 1-butanol to make butanoic acid:

$$C_4H_9OH(l) + O_2(g) \rightarrow C_3H_7COOH(l) + H_2O(l)$$

Combustion of butanol:

$$C_4H_9OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$$

$\Delta H_c = -2675.9$ kJ/mol

Combustion of butanoic acid:

$$C_3H_7COOH(l) + 5O_2(g) \rightarrow 4CO_2(g) + 4H_2O(l)$$

$\Delta H_c = -2183.6$ kJ/mol

488. Determine the free energy change for the reduction of CuO with hydrogen. Products and reactants are at 25°C.

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$$

$\Delta H = -128.5$ kJ/mol
$\Delta S = -70.1$ J/mol•K

489. Calculate the enthalpy change at 25°C for the reaction of sodium iodide and chlorine. Use only the data given.

$$NaI(s) + Cl_2(g) \rightarrow NaCl(s) + I_2(l)$$

$\Delta S = -79.9$ J/mol•K
$\Delta G = -98.0$ kJ/mol

490. The element bromine can be produced by the reaction of hydrogen bromide and manganese(IV) oxide:

$$4HBr(g) + MnO_2(s) \rightarrow MnBr_2(s) + 2H_2O(l) + Br_2(l)$$

ΔH for the reaction is -291.3 kJ/mol at 25°C. Use this value and the following values of ΔH_f^0 to calculate ΔH_f^0 of $MnBr_2(s)$.

$\Delta H^0_{f\,HBr} = -36.29$ kJ/mol
$\Delta H^0_{f\,MnO_2} = -520.0$ kJ/mol
$\Delta H^0_{f\,H_2O} = -285.8$ kJ/mol
$\Delta H^0_{f\,Br_2} = 0.00$ kJ/mol

491. Calculate the change in entropy, ΔS, at 25°C for the reaction of calcium carbide with water to produce acetylene gas:

$$CaC_2(s) + 2H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(s)$$

$\Delta G = -147.7$ kJ/mol
$\Delta H = -125.6$ kJ/mol

492. Calculate the free energy change for the explosive decomposition of ammonium nitrate at 25°C. Note that H_2O is a gas in this reaction:

$$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$$

$\Delta S = 446.4$ J/mol•K

493. In locations where natural gas, which is mostly methane, is not available, many people burn propane, which is delivered by truck and stored in a tank under pressure.

a. Write the chemical equations for the complete combustion of 1 mol of methane, CH_4, and 1 mol of propane, C_3H_8.

b. Calculate the enthalpy change for each reaction to determine the amount of energy as heat evolved by burning 1 mol of each fuel.

c. Using the molar enthalpies of combustion you calculated, determine the energy output per kilogram of each fuel. Which fuel yields more energy per unit mass?

494. The hydration of acetylene to form acetaldehyde is shown in the following equation:

$$C_2H_2(g) + H_2O(l) \rightarrow CH_3CHO(l)$$

Use enthalpies of combustion for C_2H_2 and CH_3CHO to compute the enthalpy of the above reaction.

$$C_2H_2(g) + 2O_2(g) \rightarrow 2CO_2(g) + H_2O(l)$$

$\Delta H = -1299.6$ kJ/mol

$$CH_3CHO(l) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$$

$\Delta H = -1166.9$ kJ/mol

495. Calculate the enthalpy for the combustion of decane. ΔH_f^0 for liquid decane is -300.9 kJ/mol.

$$C_{10}H_{22}(l) + 15O_2(g) \rightarrow 10CO_2(g) + 11H_2O(l)$$

496. Find the enthalpy of the reaction of magnesium oxide with hydrogen chloride:

$$MgO(s) + 2HCl(g) \rightarrow MgCl_2(s) + H_2O(l)$$

Use the following equations and data.

$$Mg(s) + 2HCl(g) \rightarrow MgCl_2(s) + H_2(g)$$

$\Delta H = -456.9$ kJ/mol

$$Mg(s) + O_2(g) \rightarrow MgO(s)$$

$\Delta H = -601.6$ kJ/mol

$$H_2O(l) \rightarrow H_2(g) + O_2(g)$$

$\Delta H = +285.8$ kJ/mol

497. What is the free energy change for the following reaction at 25°C?

$$2NaOH(s) + 2Na(s) \xrightarrow{\Delta} 2\,Na_2O(s) + H_2(g)$$

$\Delta S = 10.6$ J/mol•K $\quad \Delta H^0_{f\,NaOH} = -425.9$ kJ/mol

498. The following equation represents the reaction between gaseous HCl and gaseous ammonia to form solid ammonium chloride:

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

Calculate the entropy change in J/mol•K for the reaction of hydrogen chloride and ammonia at 25°C using the following data and the table following item 500.

$\Delta G = -91.2$ kJ/mol

499. The production of steel from iron involves the removal of many impurities in the iron ore. The following equations show some of the purifying reactions. Calculate the enthalpy for each reaction. Use the table following item 500 and the data given below.

a. $3C(s) + Fe_2O_3(s) \rightarrow 3CO(g) + 2Fe(s)$
$\Delta H^0_{f\,CO(g)} = -110.53$ kJ/mol

b. $3Mn(s) + Fe_2O_3(s) \rightarrow 3MnO(s) + 2Fe(s)$
$\Delta H^0_{f\,MnO(s)} = -384.9$ kJ/mol

c. $12P(s) + 10Fe_2O_3(s) \rightarrow 3P_4O_{10}(s) + 20Fe(s)$
$\Delta H^0_{f\,P_4O_{10}(s)} = -3009.9$ kJ/mol

d. $3Si(s) + 2Fe_2O_3(s) \rightarrow 3SiO_2(s) + 4Fe(s)$
$\Delta H^0_{f\,SiO_2(s)} = -910.9$ kJ/mol

e. $3S(s) + 2Fe_2O_3(s) \rightarrow 3SO_2(g) + 4Fe(s)$

For problems 498–499

Substance	ΔH_f^0 *(kj/mol)*	*Substance*	ΔH_f^0 *(kj/mol)*
$NH_3(g)$	−45.9	$HF(g)$	−273.3
$NH_4Cl(s)$	−314.4	$H_2O(g)$	−241.82
$NH_4F(s)$	−125	$H_2O(l)$	−285.8
$NH_4NO_3(s)$	−365.56	$H_2O_2(l)$	−187.8
$Br_2(l)$	0.00	$H_2SO_4(l)$	−813.989
$CaCO_3(s)$	−1207.6	$FeO(s)$	−825.5
$CaO(s)$	−634.9	$Fe_2O_3(s)$	−1118.4
$CH_4(g)$	−74.9	$MnO_2(s)$	−520.0
$C_3H_8(g)$	−104.7	$N_2O(g)$	+82.1
$CO_2(g)$	−393.5	$O_2(g)$	0.00
$F_2(g)$	0.00	$Na_2O(s)$	−414.2
$H_2(g)$	0.00	$Na_2SO_3(s)$	−1101
$HBr(g)$	−36.29	$SO_2(g)$	−296.8
$HCl(g)$	−92.3	$SO_3(g)$	−395.7

Equilibrium: Chap. 18, Sec. 1

500. Calculate the equilibrium constants for the following hypothetical reactions. Assume that all components of the reactions are gaseous.

a. $A \rightleftarrows C + D$

At equilibrium, the concentration of A is 2.24×10^{-2} M and the concentrations of both C and D are 6.41×10^{-3} M.

b. $A + B \rightleftarrows C + D$

At equilibrium, the concentrations of both A and B are 3.23×10^{-5} M and the concentrations of both C and D are 1.27×10^{-2} M.

c. $A + B \rightleftarrows 2C$

At equilibrium, the concentrations of both A and B are 7.02×10^{-3} M and the concentration of C is 2.16×10^{-2} M.

d. $2A \rightleftarrows 2C + D$

At equilibrium, the concentration of A is 6.59×10^{-4} M. The concentration of C is 4.06×10^{-3} M, and the concentration of D is 2.03×10^{-3} M.

e. $A + B \rightleftarrows C + D + E$

At equilibrium, the concentrations of both A and B are 3.73×10^{-4} M and the concentrations of C, D, and E are 9.35×10^{-4} M.

f. $2A + B \rightleftarrows 2C$

At equilibrium, the concentration of A is 5.50×10^{-3} M, the concentration of B is 2.25×10^{-3}, and the concentration of C is 1.02×10^{-2} M.

501. Calculate the concentration of product D in the following hypothetical reaction:

$$2A(g) \rightleftarrows 2C(g) + D(g)$$

At equilibrium, the concentration of A is 1.88×10^{-1} M, the concentration of C is 6.56 M, and the equilibrium constant is 2.403×10^2.

502. At a temperature of 700 K, the equilibrium constant is 3.164×10^3 for the following reaction system for the hydrogenation of ethene, C_2H_4, to ethane, C_2H_6:

$$C_2H_4(g) + H_2(g) \rightleftarrows C_2H_6(g)$$

What will be the equilibrium concentration of ethene if the concentration of H_2 is 0.0619 M and the concentration of C_2H_6 is 1.055 M?

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503. Using the reaction $A + 2B \rightleftarrows C + 2D$, determine the equilibrium constant if the following equilibrium concentrations are found. All components are gases.

[A] = 0.0567 M

[B] = 0.1171 M

[C] = 0.000 3378 M

[D] = 0.000 6756 M

504. In the reaction $2A \rightleftarrows 2C + 2D$, determine the equilibrium constant when the following equilibrium concentrations are found. All components are gases.

[A] = 0.1077 M

[C] = 0.000 4104 M

[D] = 0.000 4104 M

505. Calculate the equilibrium constant for the following reaction. Note the phases of the components.

$$2A(g) + B(s) \rightleftarrows C(g) + D(g)$$

The equilibrium concentrations of the components are

[A] = 0.0922 M

[C] = 4.11×10^{-4} M

[D] = 8.22×10^{-4} M

506. The equilibrium constant of the following reaction for the decomposition of phosgene at 25°C is 4.282×10^{-2}.

$$COCl_2(g) \rightleftarrows CO(g) + Cl_2(g)$$

a. What is the concentration of $COCl_2$ when the concentrations of both CO and Cl_2 are 5.90×10^{-3} M?

b. When the equilibrium concentration of $COCl_2$ is 0.003 70 M, what are the concentrations of CO and Cl_2? Assume the concentrations are equal.

507. Consider the following hypothetical reaction.

$$A(g) + B(s) \rightleftarrows C(g) + D(s)$$

a. If $K = 1$ for this reaction at 500 K, what can you say about the concentrations of A and C at equilibrium?

b. If raising the temperature of the reaction results in an equilibrium with a higher concentration of C than A, how will the value of K change?

508. The following reaction occurs when steam is passed over hot carbon. The mixture of gases it generates is called *water gas* and is useful as an industrial fuel and as a source of hydrogen for the production of ammonia.

$$C(s) + H_2O(g) \rightleftarrows CO(g) + H_2(g)$$

The equilibrium constant for this reaction is 4.251×10^{-2} at 800 K. If the equilibrium concentration of $H_2O(g)$ is 0.1990 M, what concentrations of CO and H_2 would you expect to find?

509. When nitrogen monoxide gas comes in contact with air, it oxidizes to the brown gas nitrogen dioxide according to the following equation:

$$2NO(g) + O_2(g) \rightleftarrows 2NO_2(g)$$

a. The equilibrium constant for this reaction at 500 K is 1.671×10^4. What concentration of NO_2 is present at equilibrium if [NO] = 6.200×10^{-2} M and $[O_2] = 8.305 \times 10^{-3}$ M?

b. At 1000 K, the equilibrium constant, K, for the same reaction is 1.315×10^{-2}. What will be the concentration of NO_2 at 1000 K given the same concentrations of NO and O_2 as were in (a)?

510. Consider the following hypothetical reaction, for which $K = 1$ at 300 K:

$$A(g) + B(g) \rightleftarrows 2C(g)$$

a. If the reaction begins with equal concentrations of A and B and a zero concentration of C, what can you say about the relative concentrations of the components at equilibrium?

b. Additional C is introduced at equilibrium, and the temperature remains constant. When equilibrium is restored, how will the concentrations of all components have changed? How will K have changed?

511. The equilibrium constant for the following reaction of hydrogen gas and bromine gas at 25°C is 5.628×10^{18}:

$$H_2(g) + Br_2(g) \rightleftarrows 2HBr(g)$$

a. Write the equilibrium expression for this reaction.

b. Assume that equimolar amounts of H_2 and Br_2 were present at the beginning. Calculate the equilibrium concentration of H_2 if the concentration of HBr is 0.500 M.

c. If equal amounts of H_2 and Br_2 react, which reaction component will be present in the greatest concentration at equilibrium? Explain your reasoning.

512. The following reaction reaches an equilibrium state:

$$N_2F_4(g) \rightleftarrows 2NF_2(g)$$

At equilibrium at 25°C the concentration of N_2F_4 is found to be 0.9989 M and the concentration of NF_2 is 1.131×10^{-3} M. Calculate the equilibrium constant of the reaction.

513. The equilibrium between dinitrogen tetroxide and nitrogen dioxide is represented by the following equation:

$$N_2O_4(g) \rightleftarrows 2NO_2(g)$$

A student places a mixture of the two gases into a closed gas tube and allows the reaction to reach equilibrium at 25°C. At equilibrium, the concentration of N_2O_4 is found to be 5.95×10^{-1} M and the concentration of NO_2 is found to be 5.24×10^{-2} M. What is the equilibrium constant of the reaction?

514. Consider the following equilibrium system:

$$NaCN(s) + HCl(g) \rightleftarrows HCN(g) + NaCl(s)$$

a. Write a complete expression for the equilibrium constant of this system.

b. The equilibrium constant for this reaction is 2.405×10^6. What is the concentration of HCl remaining when the concentration of HCN is 0.8959 M?

515. The following reaction is used in the industrial production of hydrogen gas:

$$CH_4(g) + H_2O(g) \rightleftarrows CO(g) + 3H_2(g)$$

The equilibrium constant of this reaction at 298 K (25°C) is 3.896×10^{-27}, but at 1100 K the constant is 3.112×10^2.

a. What do these equilibrium constants tell you about the progress of the reaction at the two temperatures?

b. Suppose the reaction mixture is sampled at 1100 K and found to contain 1.56 M of hydrogen, 3.70×10^{-2} M of methane, and 8.27×10^{-1} M of gaseous H_2O. What concentration of carbon monoxide would you expect to find?

516. Dinitrogen tetroxide, N_2O_4, is soluble in cyclohexane, a common nonpolar solvent. While in solution, N_2O_4 can break down into NO_2 according to the following equation:

$$N_2O_4(\mathit{cyclohexane}) \rightleftarrows NO_2(\mathit{cyclohexane})$$

At 20°C, the following concentrations were observed for this equilibrium reaction:

$[N_2O_4] = 2.55 \times 10^{-3}$ M

$[NO_2] = 10.4 \times 10^{-3}$ M

What is the value of the equilibrium constant for this reaction? Note: the chemical equation must be balanced first.

517. The reaction given in item 516 also occurs when the dinitrogen tetroxide and nitrogen dioxide are dissolved in carbon tetrachloride, CCl_4, another nonpolar solvent.

$$N_2O_4(CCl_4) \rightleftarrows NO_2(CCl_4)$$

The following experimental data were obtained at 20°C:

$[N_2O_4] = 2.67 \times 10^{-3}$ M

$[NO_2] = 10.2 \times 10^{-3}$ M

Calculate the value of the equilibrium constant for this reaction occurring in carbon tetrachloride.

Equilibrium of Acids and Bases K_a and K_b: Chap. 18, Sec. 3

518. At 25°C, a 0.025 M solution of formic acid, HCOOH, is found to have a hydronium ion concentration of 2.03×10^{-3} M. Calculate the ionization constant of formic acid.

519. The pH of a 0.400 M solution of iodic acid, HIO_3, is 0.726 at 25°C. What is the K_a at this temperature?

520. The pH of a 0.150 M solution of hypochlorous acid, HClO, is found to be 4.55 at 25°C. Calculate the K_a for HClO at this temperature.

521. The compound propylamine, $CH_3CH_2CH_2NH_2$, is a weak base. At equilibrium, a 0.039 M solution of propylamine has an OH^- concentration of 3.74×10^{-3} M. Calculate the pH of this solution and K_b for propylamine.

522. The K_a of nitrous acid is 4.6×10^{-4} at 25°C. Calculate the $[H_3O^+]$ of a 0.0450 M nitrous acid solution.

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523. Hydrazoic acid, HN_3, is a weak acid. The $[H_3O^+]$ of a 0.102 M solution of hydrazoic acid is 1.39×10^{-3} M. Determine the pH of this solution, and calculate K_a at 25°C for HN_3.

524. Bromoacetic acid, $BrCH_2COOH$, is a moderately weak acid. A 0.200 M solution of bromoacetic acid has a H_3O^+ concentration of 0.0192 M. Determine the pH of this solution and the K_a of bromoacetic acid at 25°C.

525. A base, B, dissociates in water according to the following equation:

$$B + H_2O \rightleftarrows BH^+ + OH^-$$

Complete the following table for base solutions with the characteristics given.

Initial [B]	*[B] at Equilibrium*	*$[OH^-]$*	*K_b*	*$[H_3O^+]$*	*pH*
a. 0.400 M	NA	2.70×10^{-4} M	?	? M	?
b. 0.005 50 M	? M	8.45×10^{-4} M	?	NA	?
c. 0.0350 M	? M	? M	?	? M	11.29
d. ? M	0.006 28 M	0.000 92 M	?	NA	?

526. The solubility of benzoic acid, C_6H_5COOH, in water at 25°C is 2.9 g/L. The pH of this saturated solution is 2.92. Determine K_a at 25°C for benzoic acid. (Hint: first calculate the initial concentration of benzoic acid.)

527. A 0.006 50 M solution of ethanolamine, $H_2NCH_2CH_2OH$, has a pH of 10.64 at 25°C. Calculate the K_b of ethanolamine. What concentration of undissociated ethanolamine remains at equilibrium?

528. The weak acid hydrogen selenide, H_2Se, has two hydrogen atoms that can form hydronium ions. The second ionization is so small that the concentration of the resulting H_3O^+ is insignificant. If the $[H_3O^+]$ of a 0.060 M solution of H_2Se is 2.72×10^{-3} M at 25°C, what is the K_a of the first ionization?

529. Pyridine, C_5H_5N, is a very weak base. Its K_b at 25°C is 1.78×10^{-9}. Calculate the $[OH^-]$ and pH of a 0.140 M solution. Assume that the concentration of pyridine at equilibrium is equal to its initial concentration because so little pyridine is dissociated.

530. A solution of a monoprotic acid, HA, at equilibrium is found to have a 0.0208 M concentration of nonionized acid. The pH of the acid solution is 2.17. Calculate the initial acid concentration and K_a for this acid.

531. Pyruvic acid, $CH_3COCOOH$, is an important intermediate in the metabolism of carbohydrates in the cells of the body. A solution made by dissolving 438 mg of pyruvic acid in 10.00 mL of water is found to have a pH of 1.34 at 25°C. Calculate K_a for pyruvic acid.

532. The $[H_3O^+]$ of a solution of acetoacetic acid, CH_3COCH_2COOH, is 4.38×10^{-3} M at 25°C. The concentration of nonionized acid is 0.0731 M at equilibrium. Calculate K_a for acetoacetic acid at 25°C.

533. The K_a of 2-chloropropanoic acid, $CH_3CHClCOOH$, is 1.48×10^{-3}. Calculate the $[H_3O^+]$ and the pH of a 0.116 M solution of 2-chloropropionic acid. Let $x = [H_3O^+]$. The degree of ionization of the acid is too large to ignore. If your set up is correct, you will have a quadratic equation to solve.

534. Sulfuric acid ionizes in two steps in water solution. For the first ionization shown in the following equation, the K_a is so large that in moderately dilute solution the ionization can be considered 100%.

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$$

The second ionization is fairly strong, and $K_a = 1.3 \times 10^{-2}$:

$$HSO_4^- + H_2O \rightleftarrows H_3O^+ + SO_4^{2-}$$

Calculate the total $[H_3O^+]$ and pH of a 0.0788 M H_2SO_4 solution. Hint: If the first ionization is 100%, what will $[HSO_4^-]$ and $[H_3O^+]$ be? Remember to account for the already existing concentration of H_3O^+ in the second ionization. Let $x = [SO_4^{2-}]$.

535. The hydronium ion concentration of a 0.100 M solution of cyanic acid, HOCN, is found to be 5.74×10^{-3} M at 25°C. Calculate the ionization constant of cyanic acid. What is the pH of this solution?

536. A solution of hydrogen cyanide, HCN, has a 0.025 M concentration. The cyanide ion concentration is found to be 3.16×10^{-6} M.

a. What is the hydronium ion concentration of this solution?
b. What is the pH of this solution?
c. What is the concentration of nonionized HCN in the solution? Be sure to use the correct number of significant figures.
d. Calculate the ionization constant of HCN.
e. How would you characterize the strength of HCN as an acid?
f. Determine the $[H_3O^+]$ for a 0.085 M solution of HCN.

537. A 1.20 M solution of dichloroacetic acid, CCl_2HCOOH, at 25°C has a hydronium ion concentration of 0.182 M.

a. What is the pH of this solution?
b. What is the K_a of dichloroacetic acid at 25°C?
c. What is the concentration of nonionized dichloroacetic acid in this solution?
d. What can you say about the strength of dichloroacetic acid?

538. Phenol, C_6H_5OH, is a very weak acid. The pH of a 0.215 M solution of phenol at 25°C is found to be 5.61. Calculate the K_a for phenol.

539. A solution of the simplest amino acid, glycine (NH_2CH_2COOH), is prepared by dissolving 3.75 g in 250.0 mL of water at 25°C. The pH of this solution is found to be 0.890.

a. Calculate the molarity of the glycine solution.
b. Calculate the K_a for glycine.

540. Trimethylamine, $(CH_3)_3N$, dissociates in water the same way that NH_3 does—by accepting a proton from a water molecule. The $[OH^-]$ of a 0.0750 M solution of trimethylamine at 25°C is 2.32×10^{-3} M. Calculate the pH of this solution and the K_b of trimethylamine.

541. Dimethylamine, $(CH_3)_2NH$, is a weak base similar to the trimethylamine in item 540. A 5.00×10^{-3} M solution of dimethylamine has a pH of 11.20 at 25°C. Calculate the K_b of dimethylamine. Compare this K_b with the K_b for trimethylamine that you calculated in item 540. Which substance is the stronger base?

542. Hydrazine dissociates in water solution according to the following equations:

$$H_2NNH_2 + H_2O(l) \rightleftarrows H_2NNH_3^+(aq) + OH^-(aq)$$
$$H_2NNH_3^+(aq) + H_2O(l) \rightleftarrows H_3NNH_3^{2+}(aq) + OH^-(aq)$$

The K_b of this second dissociation is 8.9×10^{-16}, so it contributes almost no hydroxide ions in solution and can be ignored here.

a. The pH of a 0.120 M solution of hydrazine at 25°C is 10.50. Calculate K_b for the first ionization of hydrazine. Assume that the original concentration of H_2NNH_2 does not change.
b. Make the same assumption as you did in (a) and calculate the $[OH^-]$ of a 0.020 M solution.
c. Calculate the pH of the solution in (b).

Equilibrium of Salts, K_{sp}: Chap. 18, Sec. 4

543. Silver bromate, $AgBrO_3$, is slightly soluble in water. A saturated solution is found to contain 0.276 g $AgBrO_3$ dissolved in 150.0 mL of water. Calculate K_{sp} for silver bromate.

544. 2.50 L of a saturated solution of calcium fluoride leaves a residue of 0.0427 g of CaF_2 when evaporated to dryness. Calculate the K_{sp} of CaF_2.

545. The K_{sp} of calcium sulfate, $CaSO_4$, is 9.1×10^{-6}. What is the molar concentration of $CaSO_4$ in a saturated solution?

546. A salt has the formula X_2Y, and its K_{sp} is 4.25×10^{-7}.

a. What is the molarity of a saturated solution of the salt?
b. What is the molarity of a solution of AZ if its K_{sp} is the same value?

In each of the following problems, include the calculated ion product with your answer.

547. Will a precipitate of $Ca(OH)_2$ form when 320. mL of a 0.046 M solution of NaOH mixes with 400. mL of a 0.085 M $CaCl_2$ solution? K_{sp} of $Ca(OH)_2$ is 5.5×10^{-6}.

548. 20.00 mL of a 0.077 M solution of silver nitrate, $AgNO_3$, is mixed with 30.00 mL of a 0.043 M solution of sodium acetate, $NaC_2H_3O_2$. Does a precipitate form? The K_{sp} of $AgC_2H_3O_2$ is 2.5×10^{-3}.

549. If you mix 100. mL of 0.036 M $Pb(C_2H_3O_2)_2$ with 50. mL of 0.074 M NaCl, will a precipitate of $PbCl_2$ form? The K_{sp} of $PbCl_2$ is 1.9×10^{-4}.

550. If 20.00 mL of a 0.0090 M solution of $(NH_4)_2S$ is mixed with 120.00 mL of a 0.0082 M solution of $Al(NO_3)_3$, does a precipitate form? The K_{sp} of Al_2S_3 is 2.00×10^{-7}.

Mixed Review

551. The molar concentration of a saturated calcium chromate, $CaCrO_4$, solution is 0.010 M at 25°C. What is the K_{sp} of calcium chromate?

552. A 10.00 mL sample of a saturated lead selenate solution is found to contain 0.00136 g of dissolved $PbSeO_4$ at 25°C. Determine the K_{sp} of lead selenate.

553. A 22.50 mL sample of a saturated copper(I) thiocyanate, CuSCN, solution at 25°C is found to have a 4.0×10^{-6} M concentration.
a. Determine the K_{sp} of CuSCN.
b. What mass of CuSCN would be dissolved in 1.0×10^3 L of solution?

554. A saturated solution of silver dichromate, $Ag_2Cr_2O_7$, has a concentration of 3.684×10^{-3} M. Calculate the K_{sp} of silver dichromate.

555. The K_{sp} of barium sulfite, $BaSO_3$, at 25°C is 8.0×10^{-7}.
a. What is the molar concentration of a saturated solution of $BaSO_3$?
b. What mass of $BaSO_3$ would dissolve in 500. mL of water?

556. The K_{sp} of lead(II) chloride at 25°C is 1.9×10^{-4}. What is the molar concentration of a saturated solution at 25°C?

557. The K_{sp} of barium carbonate at 25°C is 1.2×10^{-8}.
a. What is the molar concentration of a saturated solution of $BaCO_3$ at 25°C?
b. What volume of water would be needed to dissolve 0.10 g of barium carbonate?

558. The K_{sp} of $SrSO_4$ is 3.2×10^{-7} at 25°C.
a. What is the molar concentration of a saturated $SrSO_4$ solution?
b. If 20.0 L of a saturated solution of $SrSO_4$ were evaporated to dryness, what mass of $SrSO_4$ would remain?

559. The K_{sp} of strontium sulfite, $SrSO_3$, is 4.0×10^{-8} at 25°C. If 1.0000 g of $SrSO_3$ is stirred in 5.0 L of water until the solution is saturated and then filtered, what mass of $SrSO_3$ would remain?

560. The K_{sp} of manganese(II) arsenate is 1.9×10^{-11} at 25°C. What is the molar concentration of $Mn_3(AsO_4)_2$ in a saturated solution? Note that five ions are produced from the dissociation of $Mn_3(AsO_4)_2$.

561. Suppose that 30.0 mL of a 0.0050 M solution of $Sr(NO_3)_2$ is mixed with 20.0 mL of a 0.010 M solution of K_2SO_4 at 25°C. The K_{sp} of $SrSO_4$ is 3.2×10^{-7}.
a. What is the ion product of the ions that can potentially form a precipitate?
b. Does a precipitate form?

562. Lead(II) bromide, $PbBr_2$, is slightly soluble in water. Its K_{sp} is 6.3×10^{-6} at 25°C. Suppose that 120. mL of a 0.0035 M solution of $MgBr_2$ is mixed with 180. mL of a 0.0024 M $Pb(C_2H_3O_2)_2$ solution at 25°C.
a. What is the ion product of Br^- and Pb^{2+} in the mixed solution?
b. Does a precipitate form?

563. The K_{sp} of $Mg(OH)_2$ at 25°C is 1.5×10^{-11}.
a. Write the equilibrium equation for the dissociation of $Mg(OH)_2$.
b. What volume of water would be required to dissolve 0.10 g of $Mg(OH)_2$?
c. Considering that magnesium hydroxide is essentially insoluble, why is it possible to titrate a suspension of $Mg(OH)_2$ to an equivalence point with a strong acid such as HCl?

564. Lithium carbonate is somewhat soluble in water; its K_{sp} at 25°C is 2.51×10^{-2}.
a. What is the molar concentration of a saturated Li_2CO_3 solution?
b. What mass of Li_2CO_3 would you dissolve in order to make 3440 mL of saturated solution?

565. A 50.00 mL sample of a saturated solution of barium hydroxide, $Ba(OH)_2$, is titrated to the equivalence point by 31.61 mL of a 0.3417 M solution of HCl. Determine the K_{sp} of $Ba(OH)_2$.

566. Calculate the K_{sp} for salts represented by QR that dissociate into two ions, Q^+ and R^-, in each of the following solutions:
a. saturated solution of QR is 1.0 M
b. saturated solution of QR is 0.50 M
c. saturated solution of QR is 0.1 M
d. saturated solution of QR is 0.001 M

567. Suppose that salts QR, X_2Y, KL_2, A_3Z, and D_2E_3 form saturated solutions that are 0.02 M in concentration. Calculate K_{sp} for each of these salts.

568. The K_{sp} at 25°C of silver bromide is 5.0×10^{-13}. What is the molar concentration of a saturated AgBr solution? What mass of silver bromide would dissolve in 10.0 L of saturated solution at 25°C?

569. The K_{sp} at 25°C for calcium hydroxide is 5.5×10^{-6}.
a. Calculate the molarity of a saturated $Ca(OH)_2$ solution.
b. What is the OH^- concentration of this solution?
c. What is the pH of the saturated solution?

570. The K_{sp} of magnesium carbonate is 3.5×10^{-8} at 25°C. What mass of $MgCO_3$ would dissolve in 4.00 L of water at 25°C?

Redox Equations: Chap. 19, Sec. 2

Reactions in Acidic Solution

Balance the following redox equations. Assume that all reactions take place in an acid environment where H^+ and H_2O are readily available.

571. $Fe + SnCl_4 \rightarrow FeCl_3 + SnCl_2$

572. $H_2O_2 + FeSO_4 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + H_2O$

573. $CuS + HNO_3 \rightarrow Cu(NO_3)_2 + NO + S + H_2O$

574. $K_2Cr_2O_7 + HI \rightarrow CrI_3 + KI + I_2 + H_2O$

Reactions in Basic Solution

Balance the following redox equations. Assume that all reactions take place in a basic environment where OH^- and H_2O are readily available.

575. $CO_2 + NH_2OH \rightarrow CO + N_2 + H_2O$

576. $Bi(OH)_3 + K_2SnO_2 \rightarrow Bi + K_2SnO_3$
(Both of the potassium-tin-oxygen compounds dissociate into potassium ions and tin-oxygen ions.)

Mixed Review

Balance each of the following redox equations. Unless stated otherwise, assume that the reaction occurs in acidic solution.

577. $Mg + N_2 \rightarrow Mg_3N_2$

578. $SO_2 + Br_2 + H_2O \rightarrow HBr + H_2SO_4$

579. $H_2S + Cl_2 \rightarrow S + HCl$

580. $PbO_2 + HBr \rightarrow PbBr_2 + Br_2 + H_2O$

581. $S + HNO_3 \rightarrow NO_2 + H_2SO_4 + H_2O$

582. $NaIO_3 + N_2H_4 + HCl \rightarrow N_2 + NaICl_2 + H_2O$ (N_2H_4 is hydrazine; do not separate it into ions.)

583. $MnO_2 + H_2O_2 + HCl \rightarrow MnCl_2 + O_2 + H_2O$

584. $AsH_3 + NaClO_3 \rightarrow H_3AsO_4 + NaCl$ (AsH_3 is arsine, the arsenic analogue of ammonia, NH_3.)

585. $K_2Cr_2O_7 + H_2C_2O_4 + HCl \rightarrow CrCl_3 + CO_2 + KCl + H_2O$ ($H_2C_2O_4$ is oxalic acid; it can be treated as $2H^+ + C_2O_4^{2-}$.)

586. $Hg(NO_3)_2 \xrightarrow{\text{heat}} HgO + NO_2 + O_2$ (The reaction is not in solution.)

587. $HAuCl_4 + N_2H_4 \rightarrow Au + N_2 + HCl$ ($HAuCl_4$ can be considered as $H^+ + AuCl_4^-$.)

588. $Sb_2(SO_4)_3 + KMnO_4 + H_2O \rightarrow H_3SbO_4 + K_2SO_4 + MnSO_4 + H_2SO_4$

589. $Mn(NO_3)_2 + NaBiO_3 + HNO_3 \rightarrow Bi(NO_3)_2 + HMnO_4 + NaNO_3 + H_2O$

590. $H_3AsO_4 + Zn + HCl \rightarrow AsH_3 + ZnCl_2 + H_2O$

591. $KClO_3 + HCl \rightarrow Cl_2 + H_2O + KCl$

592. The same reactants as in item 591 can combine in the following way when more $KClO_3$ is present. Balance the equation.

$$KClO_3 + HCl \rightarrow Cl_2 + ClO_2 + H_2O + KCl$$

593. $MnCl_3 + H_2O \rightarrow MnCl_2 + MnO_2 + HCl$

594. $NaOH + H_2O + Al \rightarrow NaAl(OH)_4 + H_2$ in basic solution

595. $Br_2 + Ca(OH)_2 \rightarrow CaBr_2 + Ca(BrO_3)_2 + H_2O$ in basic solution

596. $N_2O + NaClO + NaOH \rightarrow NaCl + NaNO_2 + H_2O$ in basic solution

597. Balance the following reaction, which can be used to prepare bromine in the laboratory:

$$HBr + MnO_2 \rightarrow MnBr_2 + H_2O + Br_2$$

598. The following reaction occurs when gold is dissolved in *aqua regia*. Balance the equation.

$$Au + HCl + HNO_3 \rightarrow HAuCl_4 + NO + H_2O$$

Electrochemistry: Chap. 20, Sec. 2

Use the reduction potentials in the table on page 915 to determine whether the following reactions are spontaneous as written. Report the E^0_{cell} for the reactions.

599. $Cu^{2+} + Fe \rightarrow Fe^{2+} + Cu$

600. $Pb^{2+} + Fe^{2+} \rightarrow Fe^{3+} + Pb$

601. $Mn^{2+} + 4H_2O + Sn^{2+} \rightarrow MnO_4^- + 8H^+ + Sn$

602. $MnO_4^{2-} + Cl_2 \rightarrow MnO_4^- + 2Cl^-$

603. $Hg_2^{2+} + 2MnO_4^{2-} \rightarrow 2Hg + 2MnO_4^-$

604. $2Li^+ + Pb \rightarrow 2Li + Pb^{2+}$

605. $Br_2 + 2Cl^- \rightarrow 2Br^- + Cl_2$

606. $S + 2I^- \rightarrow S^{2-} + I_2$

If a cell is constructed in which the following pairs of reactions are possible, what would be the cathode reaction, the anode reaction, and the overall cell voltage?

607. $Ca^{2+} + 2e^- \rightleftarrows Ca$
$Fe^{3+} + 3e^- \rightleftarrows Fe$

608. $Ag^+ + e^- \rightleftarrows Ag$
$S + 2H^+ + 2e^- \rightleftarrows H_2S$

609. $Fe^{3+} + e^- \rightleftarrows Fe^{2+}$
$Sn^{2+} + 2e^- \rightleftarrows Sn$

610. $Cu^{2+} + 2e^- \rightleftarrows Cu$
$Au^{3+} + 3e^- \rightleftarrows Au$

Mixed Review

Use reduction potentials to determine whether the reactions in the following 10 problems are spontaneous.

611. $Ba + Sn^{2+} \rightarrow Ba^{2+} + Sn$

612. $Ni + Hg^{2+} \rightarrow Ni^{2+} + Hg$

613. $2Cr^{3+} + 7\,H_2O + 6Fe^{3+} \rightarrow Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+}$

614. $Cl_2 + Sn \rightarrow 2Cl^- + Sn^{2+}$

615. $Al + 3Ag^+ \rightarrow Al^{3+} + 3Ag$

616. $Hg_2^{2+} + S^{2-} \rightarrow 2Hg + S$

617. $Ba + 2Ag^+ \rightarrow Ba^{2+} + 2Ag$

618. $2I^- + Ca^{2+} \rightarrow I_2 + Ca$

619. $Zn + 2MnO_4^- \rightarrow Zn^{2+} + 2MnO_4^{2-}$

620. $2Cr^{3+} + 3Mg^{2+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+ + 3Mg$

In the following problems, you are given a pair of reduction half-reactions. If a cell were constructed in which the pairs of half-reactions were possible, what would be the balanced equation for the overall cell reaction that would occur? Write the half-reactions that occur at the cathode and anode, and calculate the cell voltage.

621. $Cl_2 + 2e^- \rightleftarrows 2Cl^-$
$Ni^{2+} + 2e^- \rightleftarrows Ni$

622. $Fe^{3+} + 3e^- \rightleftarrows Fe$
$Hg^{2+} + 2e^- \rightleftarrows Hg$

623. $MnO_4^- + e^- \rightleftarrows MnO_4^{2-}$
$Al^{3+} + 3e^- \rightleftarrows Al$

624. $MnO_4^- + 8H^+ + 5e^- \rightleftarrows Mn^{2+} + 4H_2O$
$S + 2H^+ + 2e^- \rightleftarrows H_2S$

625. $Ca^{2+} + 2e^- \rightleftarrows Ca$
$Li^+ + e^- \rightleftarrows Li$

626. $Br_2 + 2e^- \rightleftarrows 2Br^-$
$MnO_4^- + 8H^+ + 5e^- \rightleftarrows Mn^{2+} + 4H_2O$

627. $Sn^{2+} + 2e^- \rightleftarrows Sn$
$Fe^{3+} + e^- \rightleftarrows Fe^{2+}$

628. $Zn^{2+} + 2e^- \rightleftarrows Zn$
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftarrows 2Cr^{3+} + 7H_2O$

629. $Ba^{2+} + 2e^- \rightleftarrows Ba$
$Ca^{2+} + 2e^- \rightleftarrows Ca$

630. $Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$
$Cd^{2+} + 2e^- \rightleftarrows Cd$

For problems 599–606

Reduction Half-reaction	Standard Electrode Potential, E^0 (in volts)	Reduction Half-reaction	Standard Electrode Potential, E^0 (in volts)
$MnO_4^- + 8H^+ + 5e^- \rightleftarrows Mn^{2+} + 4H_2O$	+1.50	$Fe^{3+} + 3e^- \rightleftarrows Fe$	−0.04
$Au^{3+} + 3e^- \rightleftarrows Au$	+1.50	$Pb^{2+} + 2e^- \rightleftarrows Pb$	−0.13
$Cl_2 + 2e^- \rightleftarrows 2Cl^-$	+1.36	$Sn^{2+} + 2e^- \rightleftarrows Sn$	−0.14
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftarrows 2Cr^{3+} + 7H_2O$	+1.23	$Ni^{2+} + 2e^- \rightleftarrows Ni$	−0.26
$MnO_2 + 4H^+ + 2e^- \rightleftarrows Mn^{2+} + 2H_2O$	+1.22	$Cd^{2+} + 2e^- \rightleftarrows Cd$	−0.40
$Br_2 + 2e^- \rightleftarrows 2Br^-$	+1.07	$Fe^{2+} + 2e^- \rightleftarrows Fe$	−0.45
$Hg^{2+} + 2e^- \rightleftarrows Hg$	+0.85	$S + 2e^- \rightleftarrows S^{2-}$	−0.48
$Ag^+ + e^- \rightleftarrows Ag$	+0.80	$Zn^{2+} + 2e^- \rightleftarrows Zn$	−0.76
$Hg_2^{2+} + 2e^- \rightleftarrows 2Hg$	+0.80	$Al^{3+} + 3e^- \rightleftarrows Al$	−1.66
$Fe^{3+} + e^- \rightleftarrows Fe^{2+}$	+0.77	$Mg^{2+} + 2e^- \rightleftarrows Mg$	−2.37
$MnO_4^- + e^- \rightleftarrows MnO_4^{2-}$	+0.56	$Na^+ + e^- \rightleftarrows Na$	−2.71
$I_2 + 2e^- \rightleftarrows 2I^-$	+0.54	$Ca^{2+} + 2e^- \rightleftarrows Ca$	−2.87
$Cu^{2+} + 2e^- \rightleftarrows Cu$	+0.34	$Ba^{2+} + 2e^- \rightleftarrows Ba$	−2.91
$S + 2H^+(aq) + 2e^- \rightleftarrows H_2S(aq)$	+0.14	$K^+ + e^- \rightleftarrows K$	−2.93
$2H^+(aq) + 2e^- \rightleftarrows H_2$	0.00	$Li^+ + e^- \rightleftarrows Li$	−3.04

APPENDIX E

Selected Answers

Matter and Change

Math Tutor Practice

1. a. 5 significant figures
   b. 4 significant figures
2. a. 4.21 g/cm^3
   b. 16.5 g

Measurements and Calculations

Practice Problems A

1. 2.75 g/cm^3
2. 1.14 g
3. 5.60 mL

Practice Problems B

1. 1645 cm, 0.01645 km
2. 0.000 014 g

Practice Problems C

1. −17%
2. 2.7%

Practice Problems D

1. a. 5
   b. 6
   c. 4
   d. 1
   e. 5
   f. 6
2. a. 7000 cm
   b. 7000. cm
   c. 7000.00 cm

Practice Problems E

1. 2.156 g
2. 85.6 cm
3. 1.00 μm2
4. 440 g

Practice Problems F

1. 9.69 mL
2. 1.67 g/cm^3
3. 5.12×10^{11} mm
4. 5.2×10^3 s

Math Tutor Practice

1. a. 7.45×10^{-5} g
   b. 5.984102×10^6 nm
2. a. -9.11×10^3
   b. 8.25×10^{-2}

Atoms: The Building Blocks of Matter

Practice Problems A

1. 35 protons, 35 electrons, 45 neutrons
2. $^{13}_{6}C$
3. phosphorus-30

Practice Problems B

1. 126 g Fe
2. 14.7 g K
3. 0.310 g Na
4. 957 g Ni

Practice Problems C

1. 0.125 mol Ca
2. 1.83×10^{-7} mol Au
3. 8.18×10^{-3}

Practice Problems D

1. 2.49×10^{-12} mol Pb
2. 4.2×10^{-21} mol Sn
3. 1.66×10^{24} atoms Al

Practice Problems E

1. 7.3×10^{-7} g Ni
2. 7.51×10^{22} atoms S
3. 66 g Au

Math Tutor Practice

1. a. 2.25 g
   b. 59 300 L
2. a. 7.2×10^1 μg
   b. 3.98×10^3 km

Arrangement of Electrons in Atoms

Practice Problems A

1. 7, 7, $\underset{1s}{\uparrow\downarrow}$ $\underset{2s}{\uparrow\downarrow}$ $\underbrace{\uparrow \;\; \uparrow \;\; \uparrow}_{2p}$
2. 9, 2

Practice Problems B

1. a. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^5$, [Kr]$4d^{10}5s^25p^5$, 46
   b. 27, 26, 1
2. a. [Kr]$4d^{10}5s^25p^2$, 2
   b. 10, germanium

3. **a.** $1s^22s^22p^63s^23p^63d^54s^2$
   b. manganese
4. **a.** 9, $1s^22s^22p^63s^23p^6$
   b. argon

Practice Problems C

1. **a.** $1s^22s^22p^63s^23p^63d^{10}4s^24p^6 4d^{10}5s^25p^66s^2$, $[Xe]6s^2$
   b. Be, Mg, Ca, Sr
2. **a.** $[Xe]4f^{14}5d^{10}6s^1$
   b. Au, Cs, Pt

Math Tutor Practice

1. 85.47 amu
2. 28.1 amu

The Periodic Law

Practice Problems A

1. Group 1, fifth period, *s* block
2. **a.** ns^2
   b. $1s^22s^22p^63s^23p^64s^2$
   c. Ca, $[Ar]4s^2$

Practice Problems B

1. fourth period, *d* block, Group 10
2. $4d^{10}5s^2$

Practice Problems C

1. **a.** $3s^23p^5$
   b. chlorine, nonmetal
2. **a.** fourth period, *p* block, Group 15
   b. arsenic, metalloid

Practice Problems D

1. **a.** *p* block, second period, Group 17, halogens, fluorine, nonmetal, high reactivity
   b. *d* block, fourth period, Group 11, transition elements, copper, metal, low reactivity

Practice Problems E

1. Li; F
2. All of the elements are in Group 2. Of the four, barium has the highest atomic number and is farthest down the group. Therefore, barium has the largest atomic radius because atomic radii increase down a group.
3. All of the elements are in Period 3. Of the four, silicon has the largest atomic number and therefore is the farthest to the right on the periodic table. Therefore, silicon has the smallest atomic radius because atomic radii decrease from *left to right* across a period.

Practice Problems F

1. **a.** Q is in the *p* block, R is in the *s* block, T is in the *p* block, and X is in the *p* block.
   b. Q and R, and X and T are in the same period. Q and T are in the same group.
   c. Q would have the highest ionization energy, and R would have the lowest.
   d. R
   e. R

Practice Problems G

1. **a.** All are in the *p* block. E, J, and M are in the same period, and E, G, and L are in the same group.
   b. E should have the highest electron affinity; E, G, and L are most likely to form 1– ions; E should have the highest electronegativity.
   c. The ionic radius would be larger.
   d. E, G, and L

Math Tutor Practice

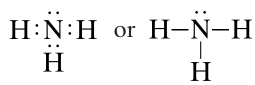
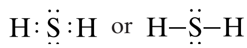
1. **a.** $1s^22s^22p^63s^23p^1$
   b. $1s^22s^22p^6$
   c. $1s^22s^22p^63s^23p^63d^{10}4s^24p^6 4d^{10}5s^25p^2$
   d. $1s^22s^22p^63s^23p^64s^1$
2. **a.** $[Ne]3s^23p^2$
   b. $[Kr]5s^1$
   c. $[Kr]4d^{10}5s^25p^3$
   d. $[Ar]3d^{10}4s^24p^3$

Chemical Bonding

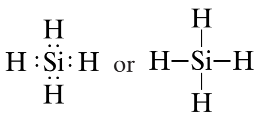
Practice Problems A

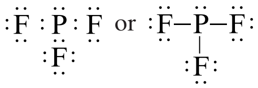
See table below.

Practice Problems C

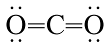
1. 
2. 

Bonding between chlorine and	Electronegativity difference	Bond type	More-negative atom
calcium	3.0 – 1.0 = 2.0	ionic	chlorine
oxygen	3.5 – 3.0 = 0.5	polar-covalent	oxygen
bromine	3.0 – 2.8 = 0.2	nonpolar-covalent	chlorine

3. 

4. 

Practice Problems D

1. 
2. $H-C\equiv N:$

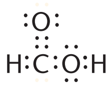
Practice Problems E

1. **a.** linear
   b. tetrahedral
   c. tetrahedral

Practice Problems F

1. **a.** bent or angular
   b. trigonal-pyramidal

Math Tutor Practice

1. **a.** 
   b. 
2. **a.** 
   b. 

Chemical Formulas and Chemical Compounds

Practice Problems A

1. **a.** KI
   b. $MgCl_2$
   c. Na_2S
   d. Al_2S_3
   e. AlN
2. **a.** silver chloride
   b. zinc oxide
   c. calcium bromide
   d. strontium fluoride
   e. barium oxide
   f. calcium chloride

Practice Problems B

1. **a.** $CuBr_2$, copper(II) bromide
   b. FeO, iron(II) oxide
   c. $PbCl_2$, lead(II) chloride
   d. HgS, mercury(II) sulfide
   e. SnF_2, tin(II) fluoride
   f. Fe_2O_3, iron(III) oxide
2. **a.** copper(II) oxide
   b. cobalt(III) fluoride
   c. tin(IV) iodide
   d. iron(II) sulfide

Practice Problems C

1. **a.** NaI
   b. $CaCl_2$
   c. K_2S
   d. $LiNO_3$
   e. $CuSO_4$
   f. Na_2CO_3
   g. $Ca(NO_2)_2$
   h. $KClO_4$
2. **a.** silver oxide
   b. calcium hydroxide
   c. potassium chlorate
   d. ammonium hydroxide
   e. iron(III) chromate
   f. potassium hypochlorite

Practice Problems D

1. **a.** sulfur trioxide
   b. iodine trichloride
   c. phosphorus pentabromide
2. **a.** CI_4
   b. PCl_3
   c. N_2O_3

Practice Problem E

a. +1, –1
b. +4, –1
c. +3, –1
d. +4, –2
e. +1, +5, –2
f. +1, –1
g. +5, –2
h. +1, +5, –2
i. +5, –2
j. +2, –1

Practice Problem F

a. 98.09 amu
b. 164.10 amu
c. 94.97 amu
d. 95.21 amu

Practice Problems G

1. **a.** 2 mol Al, 3 mol S
   b. 1 mol Na, 1 mol N, 3 mol O
   c. 1 mol Ba, 2 mol O, 2 mol H
2. **a.** 150.17 g/mol
   b. 85.00 g/mol
   c. 171.35 g/mol

Practice Problems I

1. **a.** 0.0499 mol
   b. 61 mol
2. **a.** 1.53×10^{23} molecules
   b. 2.20×10^{23} molecules
3. 1170 g

Practice Problems K

1. **a.** 74.51% Pb, 25.49% Cl
   b. 52.55% Ba, 10.72% N, 36.73% O
2. 43.85% H_2O
3. 96.0 g O; 6.00 mol O

Practice Problems M

1. FeS
2. $K_2Cr_2O_7$
3. $CaBr_2$

Practice Problems N

1. C_6H_6
2. H_2O_2

Math Tutor Practice

1. 43.38% Na, 11.33% C, 45.29% O
2. 61.13% I

Chemical Equations and Reactions

Practice Problems B

1. **a.** calcium + sulfur $\longrightarrow$ calcium sulfide; $8Ca(s) + S_8(s) \longrightarrow 8CaS(s)$
   b. hydrogen + fluorine $\longrightarrow$ hydrogen fluoride; $H_2(g) + F_2(g) \longrightarrow 2HF(g)$
   c. aluminum + zinc chloride $\longrightarrow$ zinc + aluminum chloride; $2Al(s) + 3ZnCl_2(aq) \longrightarrow 3Zn(s) + 2AlCl_3(aq)$
2. **a.** Liquid carbon disulfide reacts with oxygen gas to produce carbon dioxide gas and sulfur dioxide gas.
   b. Aqueous solutions of sodium chloride and silver nitrate react to produce aqueous sodium nitrate and a precipitate of silver chloride.
3. $N_2H_4(l) + O_2(g) \longrightarrow N_2(g) + 2H_2O(l)$

Practice Problems C

1. **a.** Word: magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen
   Formula: $Mg(s) + HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
   Balanced: $Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
   b. Word: nitric acid + magnesium hydroxide $\longrightarrow$ magnesium nitrate + water
   Formula: $HNO_3(aq) + Mg(OH)_2(s) \longrightarrow Mg(NO_3)_2(aq) + H_2O(l)$
   Balanced: $2HNO_3(aq) + Mg(OH)_2(s) \longrightarrow Mg(NO_3)_2(aq) + 2H_2O(l)$
2. $Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$

Practice Problems E

1. **a.** $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$
   b. $Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$
   c. $Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g)$

Practice Problems F

1. **a.**no
   b. no
   c. yes; $Cd(s) + 2HBr(aq) \longrightarrow CdBr_2(aq) + H_2(g)$
   d. yes; $Mg(s) + 2H_2O(g) \longrightarrow Mg(OH)_2(aq) + H_2(g)$
2. Pb
3. Mn

Math Tutor Practice

1. $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$
2. **a.** $2KI(aq) + Cl_2(g) \longrightarrow 2KCl(aq) + I_2(s)$
   b. $2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

Stoichiometry

Practice Problems A

1. 4 mol NH_3
2. 10. mol $KClO_3$

Practice Problems C

1. 80.6 g MgO
2. 300 g $C_6H_{12}O_6$

Practice Problems D

1. 7.81 mol HgO
2. 7.81 mol Hg

Practice Problems E

1. **a.** 60.0 g NH_4NO_3
   b. 27.0 g H_2O
2. 339 g Ag
3. 2.6 kg Al

Practice Problems F

1. **a.** H_2O_2
   b. 0.500 mol N_2H_4
   c. 0.250 mol N_2, 1.00 mol H_2O

Practice Problems G

1. **a.** Zn
   b. 0.75 mol S_8 remains
   c. 2.00 mol ZnS
2. **a.** carbon
   b. 2.40 mol H_2 and 2.40 mol CO
   c. 4.85 g H_2 and 67.2 g CO

Practice Problems H

1. 79.7%
2. 3.70 g Cu

Math Tutor Practice

1. 24.48 mol SO_3
2. 30.75 g O_2

States of Matter

Practice Problems A

1. 169 kJ
2. 2.19×10^5 g

Math Tutor Practice

1. 11.65 kJ/mol
2. 74.7 kJ

Gases

Practice Problems A

1. 177 kPa, 1330 mm Hg
2. 7.37×10^6 Pa

Practice Problems B

1. 760.0 torr

Practice Problems C

1. 1000 mL He

Practice Problems D

1. 941 mL
2. 91°C

Practice Problems E

1. 1.30 atm
2. 1.29 atm
3. 219°C

Practice Problems F

1. 26.3 mL
2. 3.94×10^5 Pa; or 394 kPa

Practice Problems G

1. 159 L N_2
2. 0.629 mol H_2

Practice Problems H

1. 9.10 L H_2
2. 0.313 L O_2
3. 236 L NO

Practice Problems I

1. 2.01 atm
2. 3.98 atm

Practice Problems J

1. CO_2 will effuse about 0.9 times as fast as HCl
2. 160 g/mol
3. about 235 m/s

Math Tutor Practice

1. $T_2 = \frac{P_2 T_1}{P_1}$
2. 694 mL

Solutions

Practice Problems A–C

1. 0.282 M KI
2. 0.0750 mol
3. 0.834 L

Practice Problems D

1. 22.0 *m* acetone
2. 3.13 g CH_3OH

Math Tutor Practice

1. 0.700 M Na_2SO_4
2. 0.4758 M $Cd(NO_3)_2$

Ions in Aqueous Solutions and Colligative Properties

Practice Problems A

1. **a.** $NH_4Cl(s) \xrightarrow{H_2O} NH_4^+(aq) + Cl^-(aq)$; 1 mol MH_4^+, 1 mol Cl^-, 2 mol ions

   b. $Na_2S(s) \xrightarrow{H_2O} 2Na^+(aq) + S^{2-}(aq)$; 2 mol Na^+, 1 mol S^{2-}, 3 mol ions

   c. $Ba(NO_3)_2(s) \xrightarrow{H_2O} Ba^+(aq) + 2NO_3^-(aq)$; 0.5 mol Ba^{2+}, 1 mol NO_3^-, 1.5 mol ions

Practice Problems B

1. Yes; $Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$
2. No
3. Yes; Na^+ and Cl^-; $Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$
4. $Ni_2(aq) + S^{2-}(aq) \longrightarrow NiS(s)$

Practice Problems C and D

1. –0.426°C
2. 0.175 *m*
3. –118.1°C
4. **a.** –9.0°C

   b. 4.8 *m*

Practice Problems E

1. 0.15°C
2. 102.7°C
3. 2.0 *m*
4. **a.** 0.75°C

   b. 1.5 *m*

Practice Problems F

1. –7.4°C
2. 2.6°C
3. 0.054 *m* NaCl

Math Tutor Practice

1. –4.77°C
2. 106.3°C

Acids and Bases

Math Tutor Practice

1. Formula equation: $CuSO_4(aq) + Na_2S(aq) \longrightarrow Na_2SO_4(aq) + CuS(s)$
   Full ionic equation: $Cu^{2+}(aq) + SO_4^{2-}(aq) + 2Na^+(aq) + S^{2-}(aq) \longrightarrow 2Na^+(aq) + SO_4^{2-}(aq) + CuS(s)$
   Net ionic equation: $Cu^{2+}(aq) + S^{2-}(aq) \longrightarrow CuS(s)$
2. Full ionic equation: $Cd^{2+}(aq) + 2Cl^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \longrightarrow 2Na^+(aq) + 2Cl^-(aq) + CdCO_3(s)$
   Net ionic equation: $Cd^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow CdCO_3(s)$

Acid-Base Titration and pH

Practice Problems A

1. $[H_3O^+] = 1 \times 10^{-4}$ M; $[OH^-] = 1 \times 10^{-10}$ M
2. $[H_3O^+] = 1.0 \times 10^{-3}$ M; $[OH^-] = 1.0 \times 10^{-11}$ M
3. $[H_3O^+] = 3.3 \times 10^{-13}$ M; $[OH^-] = 3.0 \times 10^{-2}$ M
4. $[H_3O^+] = 5.0 \times 10^{-11}$ M; $[OH^-] = 2.0 \times 10^{-4}$ M

Practice Problems B

1. **a.** pH = 3.0
   b. pH = 5.00
   c. pH = 10.0
   d. pH = 12.00

Practice Problems C

1. pH = 3.17
2. pH = 1.60
3. pH = 5.60
4. pH = 12.60

Practice Problems E

1. $[H_3O^+] = 1 \times 10^{-5}$ M
2. $[H_3O^+] = 1 \times 10^{-12}$ M
3. $[H_3O^+] = 3.2 \times 10^{-2}$ M; $[OH^-] = 3.2 \times 10^{-13}$ M
4. $[H_3O^+] = 2.1 \times 10^{-4}$ M

Practice Problems F

1. 0.157 M CH_3COOH
2. 0.0128 M H_2SO_4

Math Tutor Practice

1. pH = 3.07
2. 8.9×10^{-5} M OH^-

Reaction Energy

Practice Problems A

1. 0.14 J/(g•K)
2. 329 K

Practice Problems B

1. −890.8 kJ
2. 2 kJ

Practice Problems C

1. −125.4 kJ
2. +66.4 kJ
3. −296.1 kJ

Practice Problems D

1. above 333 K

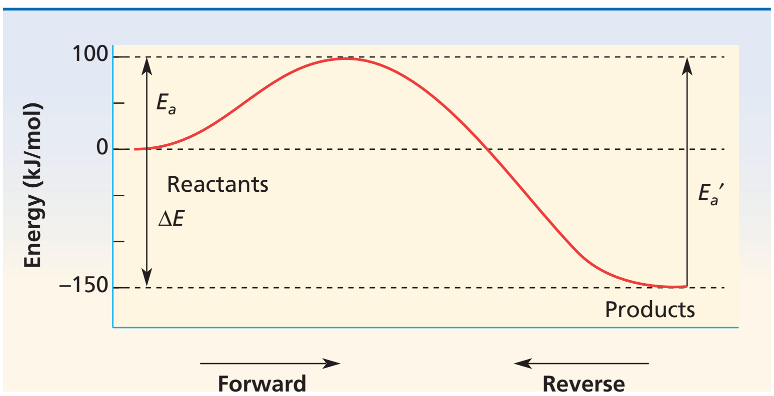
Math Tutor Practice

1. $\Delta H^0 = -396.0$ kJ
2. $\Delta H^0 = -441.8$ kJ

Reaction Kinetics

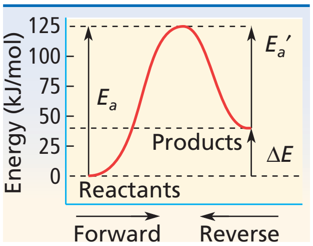
Practice Problems A

1. **a.** See figure below
   $\Delta E_{forward} = -150$ kJ/mol
   $\Delta E_{reverse} = +150$ kJ/mol
   $E_a = 100$ kJ/mol
   $E_a' = 250$ kJ/mol



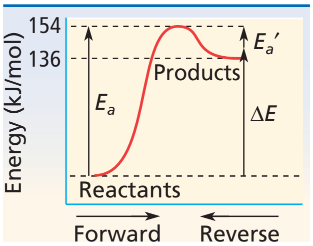
b. exothermic; The energy of the reactants is greater than the energy of the products.

2. a.



b. $\Delta E_{forward} = 39$ kJ/mol
$\Delta E_{reverse} = -39$ kJ/mol
c. endothermic; The energy of the products is greater that the energy of the products

3. a.



b. E_a (reverse) = 18 kJ/mol

Practice Problems B

1. rate = $k[A]^2$
2. 27

Practice Problems E

1. $R = k[L][M]^2$
2. $R = k[NO_2]^2$

Math Tutor Practice

1. $R = k[O_2][NO]_2$
2. $R = k[H_2]$; Students should observe that changing the concentration of C_2H_2 has no effect on the rate. The rate depends on only the concentration of hydrogen.

Chemical Equilibrium

Practice Problems A

1. 0.286
2. 4.9×10^{-3}
3. 4.36

Practice Problems B

1. 1.9×10^{-4}
2. 1.6×10^{-5}

Practice Problems C

1. 8.9×10^{-14} mol/L
2. 5.7×10^{-4} mol/L

Practice Problems D

1. AgBr precipitates.
2. $PbCl_2$ does *not* precipitate.

Math Tutor Practice

1. **a.** $K = \dfrac{[AB_2]}{[A][B]^2}$

   b. $K = \dfrac{[D_2][E_2]^2}{[DE_2]^2}$
2. $K = 2.6 \times 10^{-9}$

Oxidation-Reduction Reactions

Practice Problems A

1. $Cu + 2H_2SO_4 \longrightarrow CuSO_4 + SO_2 + 2H_2O$
2. $8HNO_3 + 6KI \longrightarrow 6KNO_3 + 3I_2 + 2NO + 4H_2O$

Math Tutor Practice

1. $2MnO_2 + NaClO_3 + 2NaOH \longrightarrow 2NaMnO_4 + NaCl + H_2O$
2. $N_2O + 2KClO + 2KOH \longrightarrow 2KCl + 2KNO_2 + H_2O$

Electrochemistry

Practice Problems A

1. **a.** $Cr_2O_7^{2-} + 14H^+ + 3Ni \longrightarrow 2Cr^{3+} + 3Ni^{2+} + 7H_2O$;
   $E^0 = 1.33 - (-0.23) = 1.56$ V

   b. $2Fe^{3+} + H_2 \longrightarrow 2Fe^{2+} + 2H^+$;
   $E^0 = 0.77 - 0.0 = 0.77$ V

Math Tutor Practice

1. $E^0 = 1.82$ V
2. $E^0 = 1.20$ V

Nuclear Chemistry

Practice Problems A

1. ${}^{253}_{99}Es + {}^{4}_{2}He \longrightarrow {}^{1}_{0}n + {}^{256}_{101}Md$
2. ${}^{142}_{61}Pm + {}^{0}_{-1}e \longrightarrow {}^{142}_{60}Nd$

Practice Problems B

1. 0.25 mg
2. 6396 years
3. 7.648 days
4. 0.00977 mg
5. 4.46×10^9 years

Math Tutor Practice

1. 1.4×10^{-6} g chromium-51
2. 8 half-lives or 420 000 years (expressed with 2 significant figures)

Organic Chemistry

Practice Problems A

1. methylbutane
2. 3-ethyl-4-methylhexane

Practice Problems B

1. 2-hexene
2. 2-methyl-2-butene *or* methyl-2-butene
3. 2-methyl-3-hexene
4. 2,5-dimethyl-2,5-heptadiene

Math Tutor Practice

1. CH_4N_2O
2. C_3H_5Br
3. CH_2O

Biological Chemistry

Math Tutor Practice

1. leucine–histidine–aspartic acid–tyrosine–asparagine–tryptophan
2. leucine–threonine–glycine; The codon UGA is a stop codon, so no more amino acids will be added.

GLOSSARY

A

absolute zero the temperature at which molecular energy is at a minimum (0 K on the Kelvin scale or –273.15 ºC on the Celsius scale) (371)

accuracy a description of how close a measurement is to the true value of the quantity measured (44)

acid-base indicator a substance that changes in color depending on the pH of the solution that the substance is in (511)

acid ionization constant the term K_a (605)

actinide any of the series of heavy radioactive elements that extends from thorium (atomic number 90) through lawrencium (atomic number 103) on the periodic table (136)

activated complex a molecule in an unstable state intermediate to the reactants and the products in the chemical reaction (565)

activation energy the minimum amount of energy required to start a chemical reaction (564)

activity series a series of elements that have similar properties and that are arranged in descending order of chemical activity; examples of activity series include metals and halogens (285)

actual yield the measured amount of a product of a reaction (317)

addition reaction a reaction in which an atom or molecule is added to an unsaturated molecule (735)

adensosine diphosphate (ADP) an organic molecule that is involved in energy metabolism; composed of a nitrogenous base, a sugar, and two phosphate groups (767)

adenosine triphosphate (ATP) an organic molecule that acts as the main energy source for cell processes; composed of a nitrogenous base, a sugar, and three phosphate groups (766)

alcohol an organic compound that contains one or more hydroxyl groups attached to carbon atoms (731)

aldehyde an organic compound that contains the carbonyl group, —CHO (733)

alkali metal one of the elements of Group 1 of the periodic table (lithium, sodium, potassium, rubidium, cesium, and francium) (142)

alkaline-earth metal one of the elements of Group 2 of the periodic table (beryllium, magnesium, calcium, strontium, barium, and radium) (142)

alkane a hydrocarbon characterized by a straight or branched carbon chain that contains only single bonds (716)

alkene a hydrocarbon that contains one or more double bonds (724)

alkyl group a group of atoms that forms when one hydrogen atom is removed from an alkane molecule (719)

alkyl halide a compound formed from an alkyl group and a halogen (fluorine, chlorine, bromine, or iodine) (732)

alkyne a hydrocarbon that contains one or more triple bonds (727)

alpha particle a positively charged atom that is released in the disintegration of radioactive elements and that consists of two protons and two neutrons (686)

amine an organic compound that can be considered to be a derivative of ammonia (733)

amino acid any one of 20 different organic molecules that contain a carboxyl and an amino group and that combine to form proteins (756)

amorphous solid a solid in which the particles are not arranged with periodicity or order (338)

amphoteric describes a substance, such as water, that has the properties of an acid and the properties of a base (485)

anabolism the metabolic synthesis of proteins, fats, and other large biomolecules from smaller molecules; requires energy in the form of ATP (769)

angular momentum quantum number the quantum number that indicates the shape of an orbital (107)

anion an ion that has a negative charge (159)

anode the electrode on whose surface oxidation takes place; anions migrate toward the anode, and electrons leave the system from the anode (656)

aromatic hydrocarbon a member of the class of hydrocarbons (of which benzene is the first member) that consists of assemblages of cyclic conjugated carbon atoms and that is characterized by large resonance energies (729)

Arrhenius acid a substance that increases the concentration of hydronium ions in aqueous solution (473)

Arrhenius base a substance that increases the concentration of hydroxide ions in aqueous solution (473)

artificial transmutation the transformation of atoms of one element into atoms of another element as a result of a nuclear reaction, such as bombardment with neutrons (691)

atmosphere of pressure the pressure of Earth's atmosphere at sea level; exactly equivalent to 760 mm Hg (364)

atom the smallest unit of an element that maintains the chemical properties of that element (6, 72)

atomic mass unit a unit of mass that describes the mass of an atom or molecule; it is exactly 1/12 of the mass of a carbon atom with mass number 12 (abbreviation, amu) (80)

atomic number the number of protons in the nucleus of an atom; the atomic number is the same for all atoms of an element (77)

atomic radius one-half of the distance between the center of identical atoms that are not bonded together (150)

Aufbau principle the principle that states that the structure of each successive element is obtained by adding one proton to the nucleus of the atom and one electron to the lowest-energy orbital that is available (111)

autotroph an organism that produces its own nutrients from inorganic substances or from the environment instead of consuming other organisms (766)

average atomic mass the weighted average of the masses of all naturally occurring isotopes of an element (81)

Avogadro's law the law that states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules (379)

Avogadro's number 6.02×10^{23}, the number of atoms or molecules in 1 mol (83)

B

barometer an instrument that measures atmospheric pressure (363)

benzene the simplest aromatic hydrocarbon (729)

beta particle a charged electron emitted during certain types of radioactive decay, such as beta decay (686)

binary acid an acid that does not contain oxygen, such as hydrofluoric acid (468)

binary compound a compound composed of two different elements (222)

boiling the conversion of a liquid to a vapor within the liquid as well as at the surface of the liquid at a specific temperature and pressure; occurs when the vapor pressure of the liquid equals the atmospheric pressure (344)

boiling point the temperature and pressure at which a liquid and a gas are in equilibrium (344)

boiling-point elevation the difference between the boiling point of a liquid in pure state and the boiling point of the liquid in solution; the increase depends on the amount of solute particles present (450)

bond energy the energy required to break the bonds in 1 mol of a chemical compound (181)

Boyle's law the law that states that for a fixed amount of gas at a constant temperature, the volume of the gas increases as the pressure of the gas decreases and the volume of the gas decreases as the pressure of the gas increases (370)

Brønsted-Lowry acid a substance that donates a proton to another substance (478)

Brønsted-Lowry acid-base reaction the transfer of protons from one reactant (the acid) to another (the base) (479)

Brønsted-Lowry base a substance that accepts a proton (479)

buffered solution a solution that can resist changes in pH when an acid or a base is added to it; a buffer (606)

C

calorimeter a device used to measure the energy as heat absorbed or released in a chemical or physical change (531)

capillary action the attraction of the surface of a liquid to the surface of a solid, which causes the liquid to rise or fall (335)

carbohydrate any organic compound that is made of carbon, hydrogen, and oxygen and that provides nutrients to the cells of living things (751)

carboxylic acid an organic acid that contains the carboxyl functional group (734)

catabolism the chemical decomposition of complex biological substances, such as carbohydrates, proteins, and glycogen, accompanied by the release of energy (768)

catalysis the acceleration of a chemical reaction by a catalyst (570)

catalyst a substance that changes the rate of a chemical reaction without being consumed or changed significantly (570)

catenation the binding of an element to itself to form chains or rings (712)

cathode the electrode on whose surface reduction takes place (656)

cation an ion that has a positive charge (159)

chain reaction a reaction in which the material that starts the reaction is also one of the products and can start another reaction (697)

change of state the change of a substance from one physical state to another (8)

Charles's law the law that states that for a fixed amount of gas at a constant pressure, the volume of the gas increases as the temperature of the gas increases and the volume of the gas decreases as the temperature of the gas decreases (372)

chemical any substance that has a defined composition (4)

chemical bond the attractive force that holds atoms or ions together (175)

chemical change a change that occurs when one or more substances change into entirely new substances with different properties (9)

chemical equation a representation of a chemical reaction that uses symbols to show the relationship between the reactants and the products (261)

chemical equilibrium a state of balance in which the rate of a forward reaction equals the rate of the reverse reaction and the concentrations of products and reactants remain unchanged (590)

chemical equilibrium expression the equation for the equilibrium constant, K_{eq} (592)

chemical formula a combination of chemical symbols and numbers to represent a substance (178)

chemical kinetics the area of chemistry that is the study of reaction rates and reaction mechanisms (568)

chemical property a property of matter that describes a substance's ability to participate in chemical reactions (8)

chemical reaction the process by which one or more substances change to produce one or more different substances (9)

chemistry the scientific study of the composition, structure, and properties of matter and the changes that matter undergoes (3)

clone an organism that is produced by asexual reproduction and that is genetically identical to its parent; to make a genetic duplicate (775)

coefficient a small whole number that appears as a factor in front of a formula in a chemical equation (263)

colligative property a property that is determined by the number of particles present in a system but that is independent of the properties of the particles themselves (446)

collision theory the theory that states that the number of new compounds formed in a chemical reaction is equal to the number of molecules that collide, multiplied by a factor that corrects for low-energy collisions (562)

colloid a mixture consisting of tiny particles that are intermediate in size between those in solutions and those in suspensions and that are suspended in a liquid, solid, or gas (403)

combined gas law the relationship between the pressure, volume, and temperature of a fixed amount of gas (374)

combustion reaction the oxidation reaction of an element or compound, in which energy as heat is released (283)

common-ion effect the phenomenon in which the addition of an ion common to two solutes brings about precipitation or reduces ionization (603)

composition stoichiometry calculations involving the mass relationships of elements in compounds (299)

compound a substance made up of atoms of two or more different elements joined by chemical bonds (7)

concentration the amount of a particular substance in a given quantity of a mixture, solution, or ore (418)

condensation the change of state from a gas to a liquid (342)

condensation reaction a chemical reaction in which two or more molecules combine to produce water or another simple molecule (736, 752)

conjugate acid an acid that forms when a base gains a proton (483)

conjugate base a base that forms when an acid loses a proton (483)

continuous spectrum the uninterrupted broad band of all colors (wavelengths) emitted by incandescent solids (100)

control rod a neutron-absorbing rod that helps control a nuclear reaction by limiting the number of free neutrons (698)

conversion factor a ratio that is derived from the equality of two different units and that can be used to convert from one unit to the other (40)

copolymer a polymer made from two different monomers (737)

covalent bond a bond formed when atoms share one or more pairs of electrons (175)

critical mass the minimum mass of a fissionable isotope that provides the number of neutrons needed to sustain a chain reaction (698)

critical point the temperature and pressure at which the gas and liquid states of a substance become identical and form one phase (347)

critical pressure the lowest pressure at which a substance can exist as a liquid at the critical temperature (348)

critical temperature the temperature above which a substance cannot exist in the liquid state (347)

crystal a solid whose atoms, ions, or molecules are arranged in a regular, repeating pattern (338)

crystal structure the arrangement of atoms, ions, or molecules in a regular way to form a crystal (339)

crystalline solid a solid that consists of crystals (338)

cycloalkane a saturated carbon chain that forms a loop or a ring (718)

D

Dalton's law of partial pressures the law that states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of the component gases (365)

daughter nuclide a nuclide produced by the radioactive decay of another nuclide (690)

decay series a series of radioactive nuclides produced by successive radioactive decay until a stable nuclide is reached (690)

decomposition reaction a reaction in which a single compound breaks down to form two or more simpler substances (279)

denature to change irreversibly the structure or shape—and thus the solubility and other properties—of a protein by heating, shaking, or treating the protein with acid, alkali, or other species (764)

density the ratio of the mass of a substance to the volume of the substance; often expressed as grams per cubic centimeter for solids and liquids and as grams per liter for gases (38)

deposition the change of state from a gas directly to a solid (346)

derived unit a unit of measure that is a combination of other measurements (36)

diffusion the movement of particles from regions of higher density to regions of lower density (331)

dimensional analysis a mathematical technique for studying dimensions of physical quantities (40)

dipole a molecule or a part of a molecule that contains both positively and negatively charged regions (204)

diprotic acid an acid that has two ionizable hydrogen atoms in each molecule, such as sulfuric acid (480)

direct proportion the relationship between two variables whose ratio is a constant value (55)

disaccharide a sugar formed from two monosaccharides (752)

disproportionation the process by which a substance is transformed into two or more dissimilar substances, usually by simultaneous oxidation and reduction (645)

dissociation the separating of a molecule into simpler molecules, atoms, radicals, or ions (435)

DNA replication the process of making a copy of DNA (772)

double-displacement reaction a reaction in which a gas, a solid precipitate, or a molecular compound forms from the apparent exchange of atoms or ions between two compounds (282)

ductility the ability of a substance to be hammered thin or drawn out into a wire (196)

E

effervescence a bubbling of a liquid caused by the rapid escape of a gas rather than by boiling (413)

effusion the passage of a gas under pressure through a tiny opening (332)

elastic collision a collision between ideally elastic bodies in which the final and initial kinetic energies are the same (329)

electrochemistry the branch of chemistry that is the study of the relationship between electric forces and chemical reactions (655)

electrode a conductor used to establish electrical contact with a nonmetallic part of a circuit, such as an electrolyte (656)

electrode potential the difference in potential between an electrode and its solution (662)

electrolysis the process in which an electric current is used to produce a chemical reaction, such as the decomposition of water (279, 670)

electrolyte a substance that dissolves in water to give a solution that conducts an electric current (405)

electrolytic cell an electrochemical device in which electrolysis takes place when an electric current is in the device (667)

electromagnetic radiation the radiation associated with an electric and magnetic field; it varies periodically and travels at the speed of light (97)

electromagnetic spectrum all of the frequencies or wavelengths of electromagnetic radiation (97)

electron affinity the energy needed to remove an electron from a negative ion to form a neutral atom or molecule (157)

electron capture the process in which an inner orbital electron is captured by the nucleus of the atom that contains the electron (687)

electron configuration the arrangement of electrons in an atom (111)

electron-dot notation an electron-configuration notation in which only the valence electrons of an atom of the a particular element are shown, indicated by dots placed around the element's symbol (184)

electronegativity a measure of the ability of an atom in a chemical compound to attract electrons (161)

electroplating the electrolytic process of plating or coating an object with a metal (668)

element a substance that cannot be separated or broken down into simpler substances by chemical means; all atoms of an element have the same atomic number (6)

elimination reaction a reaction in which a simple molecule, such as water or ammonia, is removed and a new compound is produced (737)

emission-line spectrum a diagram or graph that indicates the degree to which a substance emits radiant energy with respect to wavelength (100)

empirical formula a chemical formula that shows the composition of a compound in terms of the relative numbers and kinds of atoms in the simplest ratio (245)

end point the point in a titration at which a marked color change takes place (516)

enthalpy change the amount of energy released or absorbed as heat by a system during a process at constant pressure (534)

enthalpy of combustion the energy released as heat by the complete combustion of a specific amount of a substance at constant pressure or constant volume (539)

enthalpy of reaction the amount of energy released or absorbed as heat during a chemical reaction (534)

enthalpy of solution the amount of energy released or absorbed as heat when a specific amount of solute dissolves in a solvent (416)

entropy a measure of the randomness or disorder of a system (547)

enzyme a type of protein that speeds up metabolic reactions in plants and animals without being permanently changed or destroyed (763)

equilibrium in chemistry, the state in which a chemical process and the reverse chemical process occur at the same rate such that the concentrations of reactants and products do not change; in physics, the state in which the net force on an object is zero (342)

equilibrium constant a number that relates the concentrations of starting materials and products of a reversible chemical reaction to one another at a given temperature (592)

equilibrium vapor pressure the vapor pressure of a system at equilibrium (343)

equivalence point the point at which the two solutions used in a titration are present in chemically equivalent amounts (516)

ester an organic compound formed by combining an organic acid with an alcohol such that water is eliminated (734)

ether an organic compound in which two carbon atoms bond to the same oxygen atom (732)

evaporation the change of state from a liquid to a gas (335)

excess reactant the substance that is not used up completely in a reaction (312)

excited state a state in which an atom has more energy than it does at its ground state (100)

extensive property a property that depends on the extent or size of a system (7)

F

family a vertical column of the periodic table (17)

fatty acid an organic acid that is contained in lipids, such as fats or oils (754)

film badge a device that measures the approximate amount of radiation received in a given period of time by people who work with radiation (694)

fluid a nonsolid state of matter in which the atoms or molecules are free to move past each other, as in a gas or liquid (333)

formula equation a representation of the reactants and products of a chemical reaction by their symbols or formulas (264)

formula mass the sum of the average atomic masses of all atoms represented in the formula of any molecule, formula unit, or ion (237)

formula unit the collection of atoms corresponding to an ionic compound's formula such that the molar mass of the compound is the same as the mass of 1 mol of formula units (190)

free energy the energy in a system that is available for work; a system's capacity to do useful work (548)

free-energy change the difference between the change in enthalpy, ΔH, and the product of the Kelvin temperature and the entropy change, which is defined as $T\Delta S$, at a constant pressure and temperature (548)

freezing the change of state in which a liquid becomes a solid as energy as heat is removed (336)

freezing point the temperature at which a solid and liquid are in equilibrium at 1 atm pressure; the temperature at which a liquid substance freezes (345)

freezing-point depression the difference between the freezing points of a pure solvent and a solution, which is directly proportional to the amount of solute present (448)

frequency the number of cycles or vibrations per unit of time; *also* the number of waves produced in a given amount of time (98)

functional group the portion of a molecule that is active in a chemical reaction and that determines the properties of many organic compounds (730)

G

gamma ray the high-energy photon emitted by a nucleus during fission and radioactive decay (687)

gas a form of matter that does not have a definite volume or shape (8)

Gay-Lussac's law the law that states that the volume occupied by a gas at a constant pressure is directly proportional to the absolute temperature (373)

Gay-Lussac's law of combining volumes of gases the law that states that the volumes of gases involved in a chemical change can be represented by a ratio of small whole numbers (378)

Geiger-Müller counter an instrument that detects and measures the intensity of radiation by counting the number of electric pulses that pass between the anode and the cathode in a tube filled with gas (694)

geometric isomer a compound that exists in two or more geometrically different configurations (714)

Graham's law of effusion the law that states that the rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses (387)

ground state the lowest energy state of a quantized system (100)

group a vertical column of elements in the periodic table; elements in a group share chemical properties (17)

H

half-cell a single electrode immersed in a solution of its ions (656)

half-life the time required for half of a sample of a radioactive isotope to break down by radioactive decay to form a daughter isotope (688)

half-reaction the part of a reaction that involves only oxidation or reduction (633)

halogen one of the elements of Group 17 (fluorine, chlorine, bromine, iodine, and astatine); halogens combine with most metals to form salts (147)

heat the energy transferred between objects that are at different temperatures; energy is always transferred from higher-temperature objects to lower-temperature objects until thermal equilibrium is reached (532)

Heisenberg uncertainty principle the principle that states that determining both the position and velocity of an electron or any other particle simultaneously is impossible (105)

Henry's law the law that states that at constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas on the surface of the liquid (413)

Hess's law the overall enthalpy change in a reaction is equal to the sum of the enthalpy changes for the individual steps in the process (539)

heterogeneous composed of dissimilar components (12)

heterogeneous catalyst a catalyst that is in a different phase from the phase of the reactants (570)

heterogeneous reaction a reaction in which the reactants are in two different phases (568)

heterotroph an organism that obtains organic food molecules by eating other organisms or their byproducts and that cannot synthesize organic compounds from inorganic materials (766)

homogeneous describes something that has a uniform structure or composition throughout (12)

homogeneous catalyst a catalyst that is in the same phase as the reactants are (570)

homogeneous reaction a reaction in which all of the reactants and products are in the same phase (562)

Hund's rule the rule that states that for an atom in the ground state, the number of unpaired electrons is the maximum possible and these unpaired electrons have the same spin (112)

hybrid orbitals orbitals that have the properties to explain the geometry of chemical bonds between atoms (202)

hybridization the mixing of two or more atomic orbitals of the same atom to produce new orbitals; hybridization represents the mixing of higher- and lower-energy orbitals to form orbitals of intermediate energy (201)

hydration the strong affinity of water molecules for particles of dissolved or suspended substances that causes electrolytic dissociation (411)

hydrocarbon an organic compound composed only of carbon and hydrogen (712)

hydrogen bond the intermolecular force occurring when a hydrogen atom that is bonded to a highly electronegative atom of one molecule is attracted to two unshared electrons of another molecule (206)

hydrolysis a chemical reaction between water and another substance to form two or more new substances; a reaction between water and a salt to create an acid or a base (608, 752)

hydronium ion an ion consisting of a proton combined with a molecule of water; H_3O^+ (441)

hypothesis an explanation that is based on prior scientific research or observations and that can be tested (30)

I

ideal gas an imaginary gas whose particles are infinitely small and do not interact with each other (329)

ideal gas constant the proportionality constant that appears in the equation of state for 1 mol of an ideal gas; R = 0.082 057 84 L • atm/mol • K (384)

ideal gas law the law that states the mathematical relationship of pressure (P), volume (V), temperature (T), the gas constant (R), and the number of moles of a gas (n); $PV = nRT$ (383)

immiscible describes two or more liquids that do not mix with each other (412)

intensive property a property that does not depend on the amount of matter present, such as pressure, temperature, or density (7)

intermediate a substance that forms in a middle stage of a chemical reaction and is considered a stepping stone between the parent substance and the final product (562)

inverse proportion the relationship between two variables whose product is constant (56)

ion an atom, radical, or molecule that has gained or lost one or more electrons and has a negative or positive charge (153)

ionic bond a force that attracts electrons from one atom to another, which transforms a neutral atom into an ion (175)

ionic compound a compound composed of ions bound together by electrostatic attraction (190)

ionization the process of adding or removing electrons from an atom or molecule, which gives the atom or molecule a net charge (153, 441)

ionization energy the energy required to remove an electron from an atom or ion (abbreviation, IE) (153)

isomer one of two or more compounds that have the same chemical composition but different structures (712)

isotope an atom that has the same number of protons (or the same atomic number) as other atoms of the same element do but that has a different number of neutrons (and thus a different atomic mass) (78)

J

joule the unit used to express energy; equivalent to the amount of work done by a force of 1 N acting through a distance of 1 m in the direction of the force (abbreviation, J) (531)

K

ketone an organic compound in which a carbonyl group is attached to two alkyl groups; obtained by the oxidation of secondary alcohols (733)

kinetic-molecular theory a theory that explains that the behavior of physical systems depends on the combined actions of the molecules constituting the system (329)

L

lanthanide a member of the rare-earth series of elements, whose atomic numbers range from 58 (cerium) to 71 (lutetium) (136)

lattice energy the energy associated with constructing a crystal lattice relative to the energy of all constituent atoms separated by infinite distances (192)

law of conservation of mass the law that states that mass cannot be created or destroyed in ordinary chemical and physical changes (68)

law of definite proportions the law that states that a chemical compound always contains the same elements in exactly the same proportions by weight or mass (68)

law of multiple proportions the law that states that when two elements combine to form two or more compounds, the mass of one element that combines with a given mass of the other is in the ratio of small whole numbers (68)

Lewis acid an atom, ion, or molecule that accepts a pair of electrons (481)

Lewis acid-base reaction the formation of one or more covalent bonds between an electron-pair donor and an electron-pair acceptor (482)

Lewis base an atom, ion, or molecule that donates a pair of electrons (482)

Lewis structure a structural formula in which electrons are represented by dots; dot pairs or dashes between two atomic symbols represent pairs in covalent bonds (185)

limiting reactant the substance that controls the quantity of product that can form in a chemical reaction (312)

lipid a type of biochemical that does not dissolve in water, including fats and steroids; lipids store energy and make up cell membranes (754)

liquid the state of matter that has a definite volume but not a definite shape (8)

London dispersion force the intermolecular attraction resulting from the uneven distribution of electrons and the creation of temporary dipoles (207)

M

magic numbers the numbers (2, 8, 20, 28, 50, 82, and 126) that represent the number of particles in an extra stable atomic nucleus that has completed shells of protons and neutrons (683)

magnetic quantum number the quantum number that corresponds to the alignment of the angular momentum component with a magnetic field (108)

main-group element an element in the *s*-block or *p*-block of the periodic table (146)

malleability the ability of a substance to be hammered or beaten into a sheet (196)

mass a measure of the amount of matter in an object (6)

mass defect the difference between the mass of an atom and the sum of the masses of the atom's protons, neutrons, and electrons (681)

mass number the sum of the numbers of protons and neutrons that make up the nucleus of an atom (78)

matter anything that has mass and takes up space (6)

melting the change of state in which a solid becomes a liquid by adding energy as heat or changing pressure (338)

melting point the temperature and pressure at which a solid becomes a liquid (338)

metabolism the sum of all chemical processes that occur in an organism (766)

metal an element that is shiny and that conducts heat and electricity well (18)

metallic bond a bond formed by the attraction between positively charged metal ions and the electrons around them (195)

metalloid an element that has properties of both metals and nonmetals; sometimes referred to as a semiconductor (19)

millimeters of mercury a unit of pressure (364)

miscible describes two or more liquids that can dissolve into each other in various proportions (412)

mixture a combination of two or more substances that are not chemically combined (11)

model a pattern, plan, representation, or description designed to show the structure or workings of an object, system, or concept (31)

moderator a material that slows the velocity of neutrons so that they may be absorbed by the nuclei (698)

molal boiling-point constant a quantity calculated to represent the boiling-point elevation of a 1-molal solution of a nonvolatile, nonelectrolyte solution (450)

molal freezing-point constant a quantity calculated to represent the freezing-point depression of a 1-molal solution of a nonvolatile, nonelectrolyte solute (448)

molality the concentration of a solution expressed in moles of solute per kilogram of solvent (422)

molar enthalpy of formation the amount of energy as heat resulting from the formation of 1 mol of a substance at constant pressure (537)

molar enthalpy of fusion the amount of energy as heat required to change 1 mol of a substance from solid to liquid at constant temperature and pressure (346)

molar enthalpy of vaporization the amount of energy as heat required to evaporate 1 mol of a liquid at constant pressure and temperature (345)

molar mass the mass in grams of 1 mol of a substance (83)

molarity a concentration unit of a solution expressed as moles of solute dissolved per liter of solution (418)

mole the SI base unit used to measure the amount of a substance whose number of particles is the same as the number of atoms of carbon in exactly 12 g of carbon-12 (83)

mole ratio a conversion factor that relates the amounts in moles of any two substances involved in a chemical reaction (300)

molecular compound a chemical compound whose simplest units are molecules (178)

molecular formula a chemical formula that shows the number and kinds of atoms in a molecule, but not the arrangement of the atoms (178)

molecule a group of atoms that are held together by chemical forces; a molecule is the smallest unit of matter that can exist by itself and retain all of a substance's chemical properties (178)

monatomic ion an ion formed from a single atom (220)

monomer a simple molecule that can combine with other like or unlike molecules to make a polymer (737)

monoprotic acid an acid that can donate only one proton to a base (479)

monosaccharide a simple sugar that is the basic subunit of a carbohydrate (751)

multiple bond a bond in which the atoms share more than one pair of electrons, such as a double bond or a triple bond (187)

N

natural gas a mixture of gaseous hydrocarbons located under the surface of Earth, often near petroleum deposits; used as a fuel (723)

net ionic equation an equation that includes only those compounds and ions that undergo a chemical change in a reaction in an aqueous solution (439)

neutralization the reaction of the ions that characterize acids (hydronium ions) and the ions that characterize bases (hydroxide ions) to form water molecules and a salt (489)

newton the SI unit for force; the force that will increase the speed of a 1 kg mass by 1 m/s each second that the force is applied (abbreviation, N) (362)

noble gas one of the elements of Group 18 of the periodic table (helium, neon, argon, krypton, xenon, and radon); noble gases are unreactive (117)

noble-gas configuration an outer main energy level fully occupied, in most cases, by eight electrons (118)

nomenclature a naming system (222)

nonelectrolyte a liquid or solid substance or mixture that does not allow an electric current (406)

nonmetal an element that conducts heat and electricity poorly and that does not form positive ions in an electrolytic solution (19)

nonpolar covalent bond a covalent bond in which the bonding electrons are equally attracted to both bonded atoms (176)

nonvolatile substance a substance that has little tendency to become a gas under existing conditions (446)

nuclear binding energy the energy released when a nucleus is formed from nucleons (682)

nuclear fission the splitting of the nucleus of a large atom into two or more fragments; releases additional neutrons and energy (697)

nuclear forces the interaction that binds protons and neutrons, protons and protons, and neutrons and neutrons together in a nucleus (76)

nuclear fusion the combination of the nuclei of small atoms to form a larger nucleus; releases energy (699)

nuclear power plant a facility that uses heat from nuclear reactors to produce electrical energy (698)

nuclear radiation the particles that are released from the nucleus during radioactive decay, such as neutrons, electrons, and photons (685)

nuclear reaction a reaction that affects the nucleus of an atom (684)

nuclear reactor a device that uses controlled nuclear reactions to produce energy or nuclides (698)

nuclear shell model a model which represents nucleons as existing in different energy levels, or shells, in the nucleus (683)

nuclear waste waste that contains radioisotopes (696)

nucleic acid an organic compound, either RNA or DNA, whose molecules are made up of one or two chains of nucleotides and carry genetic information (770)

nucleon a proton or neutron (681)

nuclide an atom that is identified by the number of protons and neutrons in its nucleus (79, 681)

O

orbital a region in an atom where there is a high probability of finding electrons (106)

order in chemistry, a classification of chemical reactions that depends on the number of molecules that appear to enter into the reaction (572)

organic compound a covalently bonded compound that contains carbon, excluding carbonates and oxides (711)

osmosis the diffusion of water or another solvent from a more dilute solution (of a solute) to a more concentrated solution (of the solute) through a membrane that is permeable to the solvent (452)

osmotic pressure the external pressure that must be applied to stop osmosis (452)

oxidation a reaction that removes one or more electrons from a substance such that the substance's valence or oxidation state increases (632)

oxidation number the number of electrons that must be added to or removed from an atom in a combined state to convert the atom into the elemental form (232)

oxidation state the condition of an atom expressed by the number of electrons that the atom needs to reach its elemental form (232)

oxidation-reduction reaction any chemical change in which one species is oxidized (loses electrons) and another species is reduced (gains electrons); also called *redox reaction* (633)

oxidized describes an element that has lost electrons and that has increased its oxidation number (632)

oxidizing agent the substance that gains electrons in an oxidation-reduction reaction and that is reduced (642)

oxyacid an acid that is a compound of hydrogen, oxygen, and a third element, usually a nonmetal (469)

oxyanion a polyatomic ion that contains oxygen (225)

P

parent nuclide a radionuclide that yields a specific daughter nuclide as a later member of a radioactive series (690)

partial pressure the pressure of each gas in a mixture (365)

pascal the SI unit of pressure; equal to the force of 1 N exerted over an area of 1 m^2 (abbreviation, Pa) (364)

Pauli exclusion principle the principle that states that two particles of a certain class cannot be in exactly the same energy state (112)

percentage composition the percentage by mass of each element in a compound (243)

percentage error a figure that is calculated by subtracting the accepted value from the experimental value, dividing the difference by the accepted value, and then multiplying by 100 (45)

percentage yield the ratio of the actual yield to the theoretical yield, multiplied by 100 (317)

period in chemistry, a horizontal row of elements in the periodic table (17)

periodic law the law that states that the repeating chemical and physical properties of elements change periodically with the atomic numbers of the elements (135)

periodic table an arrangement of the elements in order of their atomic numbers such that elements with similar properties fall in the same column, or group (135)

petroleum a liquid mixture of complex hydrocarbon compounds; used widely as a fuel source (723)

pH a value that is used to express the acidity or alkalinity (basicity) of a system; each whole number on the scale indicates a tenfold change in acidity; a pH of 7 is neutral, a pH of less than 7 is acidic, and a pH of greater than 7 is basic (503)

pH meter a device used to determine the pH of a solution by measuring the voltage between the two electrodes that are placed in the solution (512)

phase in chemistry, one of the four states or conditions in which a substance can exist: solid, liquid, gas, or plasma; a part of matter that is uniform (342)

phase diagram a graph of the relationship between the physical state of a substance and the temperature and pressure of the substance (347)

photoelectric effect the emission of electrons from a material when light of certain frequencies shines on the surface of the material (99)

photon a unit or quantum of light; a particle of electromagnetic radiation that has zero rest mass and carries a quantum of energy (100)

physical change a change of matter from one form to another without a change in chemical properties (7)

physical property a characteristic of a substance that does not involve a chemical change, such as density, color, or hardness (7)

plasma in physical science, a state of matter that starts as a gas and then becomes ionized; it consists of free-moving ions and electrons, it takes on an electric charge, and its properties differ from those of a solid, liquid, or gas (8)

pOH the negative of the common logarithm of the hydroxide ion concentration of a solution (503)

polar describes a molecule in which the positive and negative charges are separated (176)

polar covalent bond a covalent bond in which a pair of electrons shared by two atoms is held more closely by one atom (176)

polyatomic ion an ion made of two or more atoms (194)

polymer a large molecule that is formed by more than five monomers, or small units (737)

polyprotic acid an acid that can donate more than one proton per molecule (479)

polysaccharide one of the carbohydrates made up of long chains of simple sugars; polysaccharides include starch, cellulose, and glycogen (753)

positron a particle that has the same mass and spin as an electron but that has a positive charge (686)

precipitate a solid that is produced as a result of a chemical reaction in solution (262)

precision the exactness of a measurement (44)

pressure the amount of force exerted per unit area of a surface (361)

primary standard a highly purified solid compound used to check the concentration of a known solution in a titration (517)

principal quantum number the quantum number that indicates the energy and orbital of an electron in an atom (107)

product a substance that forms in a chemical reaction (9)

protein an organic compound that is made of one or more chains of amino acids and that is a principal component of all cells (757)

pure substance a sample of matter, either a single element or a single compound, that has definite chemical and physical properties (13)

Q

quantity something that has magnitude, size, or amount (33)

quantum the basic unit of electromagnetic energy; it characterizes the wave properties of electrons (99)

quantum number a number that specifies certain properties of electrons (107)

quantum theory the study of the structure and behavior of the atom and of subatomic particles from the view that all energy comes in tiny, indivisible bundles (105)

R

radioactive dating the process by which the approximate age of an object is determined based on the amount of certain radioactive nuclides present (695)

radioactive decay the disintegration of an unstable atomic nucleus into one or more different nuclides, accompanied by the emission of radiation, the nuclear capture or ejection of electrons, or fission (685)

radioactive nuclide a nuclide that contains isotopes that decay and that emit radiation (685)

radioactive tracer a radioactive material that is added to a substance so that its distribution can be detected later (695)

rate law the expression that shows how the rate of formation of product depends on the concentration of all species other than the solvent that take part in a reaction (572)

rate-determining step in a multistep chemical reaction, the step that has the lowest velocity, which determines the rate of the overall reaction (576)

reactant a substance or molecule that participates in a chemical reaction (9)

reaction mechanism the way in which a chemical reaction takes place; expressed in a series of chemical equations (561)

reaction rate the rate at which a chemical reaction takes place; measured by the rate of formation of the product or the rate of disappearance of the reactants (568)

reaction stoichiometry calculations involving the mass relationships between reactants and products in a chemical reaction (299)

real gas a gas that does not behave completely like a hypothetical ideal gas because of the interactions between the gas molecules (332)

redox reaction [see *oxidation-reduction reaction*] (633)

reduced describes a substance that has gained electrons, lost an oxygen atom, or gained a hydrogen atom (633)

reducing agent a substance that has the potential to reduce another substance (642)

reduction a chemical change in which electrons are gained, either by the removal of oxygen, the addition of hydrogen, or the addition of electrons (633)

reduction potential the decrease in voltage that takes place when a positive ion becomes less positive or neutral or when a neutral atom becomes negative ion (662)

rem the quantity of ionizing radiation that does as much damage to human tissue as 1 roentgen of high-voltage X rays does (693)

resonance the bonding in molecules or ions that cannot be correctly represented by a single Lewis structure (189)

reversible reaction a chemical reaction in which the products re-form the original reactants (266, 589)

roentgen a unit of radiation dose of X rays or gamma rays that is equal to the amount of radiation that will produce 2.58×10^{-4} of ions per kilogram of air at atmospheric pressure (693)

S

salt an ionic compound that forms when a metal atom or a positive radical replaces the hydrogen of an acid (231, 489)

saponification a chemical reaction in which esters of fatty acids react with a strong base to produce glycerol and a fatty acid salt; the process that is used to make soap (754)

saturated hydrocarbon an organic compound formed only by carbon and hydrogen linked by single bonds (716)

saturated solution a solution that cannot dissolve any more solute under the given conditions (409)

scientific method a series of steps followed to solve problems, including collecting data, formulating a hypothesis, testing the hypothesis, and stating conclusions (29)

scientific notation a method of expressing a quantity as a number multiplied by 10 to the appropriate power (50)

scintillation counter an instrument that converts scintillating light into an electrical signal for detecting and measuring radiation (694)

self-ionization of water a process in which two water molecules produce a hydronium ion and a hydroxide ion by transfer of a proton (499)

semipermeable membrane a membrane that permits the passage of only certain molecules (452)

shielding a radiation-absorbing material that is used to decrease radiation leakage from nuclear reactors (698)

SI Le **S**ystème **I**nternational d'Unités, or the International System of Units, which is the measurement system that is accepted worldwide (33)

significant figure a prescribed decimal place that determines the amount of rounding off to be done based on the precision of the measurement (46)

single bond a covalent bond in which two atoms share one pair of electrons (185)

single-displacement reaction a reaction in which one element or radical takes the place of another element or radical in a compound (281)

solid the state of matter in which the volume and shape of a substance are fixed (8)

solubility the ability of one substance to dissolve in another at a given temperature and pressure; expressed in terms of the amount of solute that will dissolve in a given amount of solvent to produce a saturated solution (410)

solubility product constant the equilibrium constant for a solid that is in equilibrium with the solid's dissolved ions (613)

soluble capable of dissolving in a particular solvent (401)

solute in a solution, the substance that dissolves in the solvent (402)

solution a homogeneous mixture of two or more substances uniformly dispersed throughout a single phase (402)

solution equilibrium the physical state in which the opposing processes of dissolution and crystallization of a solute occur at equal rates (408)

solvated describes a solute molecule that is surrounded by solvent molecules (415)

solvent in a solution, the substance in which the solute dissolves (402)

specific heat the quantity of heat required to raise a unit mass of homogeneous material 1 K or 1ºC in a specified way given constant pressure and volume (532)

spectator ions ions that are present in a solution in which a reaction is taking place but that do not participate in the reaction (439)

spin quantum number the quantum number that describes the intrinsic angular momentum of a particle (110)

standard electrode potential the potential developed by a metal or other material immersed in an electrolyte solution relative to the potential of the hydrogen electrode, which is set at zero (663)

standard solution a solution of known concentration, expressed in terms of the amount of solute in a given amount of solvent or solution (517)

standard temperature and pressure for a gas, the temperature of 0ºC and the pressure 1.00 atm (364)

strong acid an acid that ionizes completely in a solvent (474)

strong electrolyte a compound that completely or largely dissociates in an aqueous solution, such as soluble mineral salts (442)

structural formula a formula that indicates the location of the atoms, groups, or ions relative to one another in a molecule and that indicates the number and location of chemical bonds (185, 712)

structural isomers two or more compounds that have the same number and kinds of atoms and the same molecular weight but that differ in the order in which the atoms are attached to one another (713)

sublimation the process in which a solid changes directly into a gas (the term is sometimes also used for the reverse process) (346)

substitution reaction a reaction in which one or more atoms replace another atom or group of atoms in a molecule (735)

supercooled liquid a liquid that is cooled below its normal freezing point without solidifying (338)

supersaturated solution a solution that holds more dissolved solute than is required to reach equilibrium at a given temperature (409)

surface tension the force that acts on the surface of a liquid and that tends to minimize the area of the surface (335)

suspension a mixture in which particles of a material are more or less evenly dispersed throughout a liquid or gas (403)

synthesis reaction a reaction in which two or more substances combine to form a new compound (276)

system a set of particles or interacting components considered to be a distinct physical entity for the purpose of study (29)

T

temperature a measure of how hot (or cold) something is; specifically, a measure of the average kinetic energy of the particles in an object (531)

theoretical yield the maximum amount of product that can be produced from a given amount of reactant (317)

theory an explanation for some phenomenon that is based on observation, experimentation, and reasoning (31)

thermochemical equation an equation that includes the quantity of energy as heat released or absorbed during the reaction as written (535)

thermochemistry the branch of chemistry that is the study of the energy changes that accompany chemical reactions and changes of state (531)

titration a method to determine the concentration of a substance in solution by adding a solution of known volume and concentration until the reaction is completed, which is usually indicated by a change in color (515)

transition element one of the metals that can use the inner shell before using the outer shell to bond (144)

transition interval the range in concentration over which a variation in a chemical indicator can be observed (512)

transmutation the transformation of atoms of one element into atoms of a different element as a result of a nuclear reaction (684)

transuranium element a synthetic element whose an atomic number is greater than that of uranium (atomic number 92) (692)

triple point the temperature and pressure conditions at which the solid, liquid, and gaseous phases of a substance coexist at equilibrium (347)

triprotic acid an acid that has three ionizable protons per molecule, such as phosphoric acid (480)

U

unit cell the smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire lattice (339)

unsaturated hydrocarbon a hydrocarbon that has available valence bonds, usually from double or triple bonds with carbon (724)

unsaturated solution a solution that contains less solute than a saturated solution does and that is able to dissolve additional solute (409)

V

valence electron an electron that is found in the outermost shell of an atom and that determines the atom's chemical properties (160)

vaporization the process by which a liquid or solid changes to a gas (335)

volatile liquid a liquid that evaporates readily or at a low temperature (343)

voltaic cell a primary cell that consists of two electrodes made of different metals immersed in an electrolyte; used to generate voltage (658)

volume a measure of the size of a body or region in three-dimensional space (37)

VSEPR theory a theory that predicts some molecular shapes based on the idea that pairs of valence electrons surrounding an atom repel each other (197)

W

wavelength the distance from any point on a wave to an identical point on the next wave (97)

weak acid an acid that releases few hydrogen ions in aqueous solution (474)

weak electrolyte a compound that dissociates only to a small extent in aqueous solution (443)

weight a measure of the gravitational force exerted on an object; its value can change with the location of the object in the universe (35)

word equation an equation in which the reactants and products in a chemical reaction are represented by words (263)

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