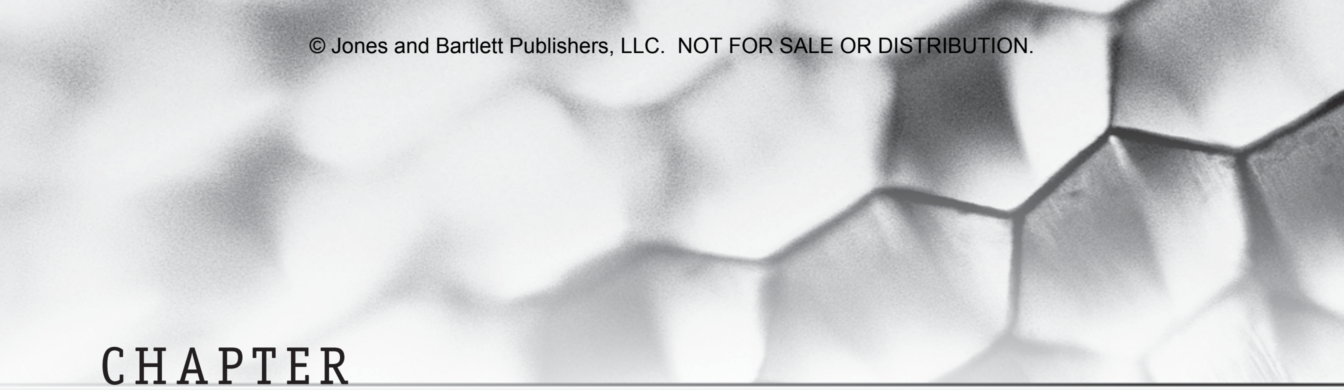
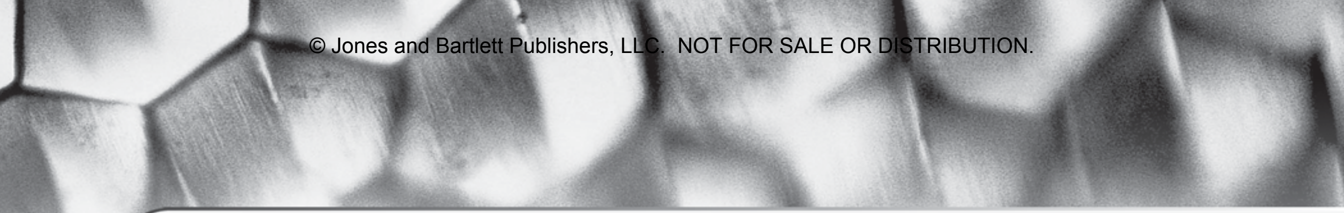


CHAPTER

1

CHAPTER OUTLINE

- **The Formation of the Universe**
  - Galaxies and Stars
  - The Planets in Our Solar System
  - The Sun
- **Differentiation of the Earth into Layers**
  - Heating of the Earth
  - The Core
  - The Mantle
  - The Crust
  - Relative Abundance of the Elements in the Earth
- **Formation of the Oceans and the Atmosphere**
- **Rocks and Minerals**
- **Rocks as Natural Resources**
- **Ores and Metals**
  - Iron
  - Aluminum
  - Copper
  - Strategic Metals
- **Mineral Reserves**
- **The Origin of Life on Earth**
- **The Uniqueness of the Earth**
- **The Environment**
- **Ecosystems**
  - Producers and Consumers
- **The Flow of Energy Through Ecosystems**
  - What Is Energy?
  - Energy Transformations
  - Food Chains and Trophic Levels
  - Energy and Biomass



Planet Earth: Rocks, Life, and History

- **Concentration Units**
  - Molarity and Molar Solutions
  - Parts per Million
  - Parts per Billion
- **Nutrient Cycles**
  - The Carbon Cycle
  - The Nitrogen Cycle
  - The Oxygen Cycle
  - The Phosphorus Cycle
  - Nature's Cycles in Balance
- **Additional Sources of Information**
- **Keywords**
- **Questions and Problems**

To understand how our **environment** works, we must first look back billions of years at the time when the Earth was born and see how it evolved into the life-supporting planet that we inhabit today. In this chapter, we consider the formation of the universe, including the origin of the galaxies, the stars, and our own planet, Earth. We look at how the oceans, the atmosphere, and the rocky surface on which we live were formed; examine the Earth's mineral resources; and discuss the ways in which society uses them. We see how life developed on Earth and how all living organisms interact with their physical surroundings and with each other, how all of these interactions are intertwined, and how a continuing flow of energy through all of its parts fuels the entire system.

The Formation of the Universe

If we gaze at the sky on a clear night, away from the lights of any city, we see myriads of stars. All of the stars that we see are a part of our galaxy, the **Milky Way**. This pinwheel-shaped body, which is made up of clouds of gas and cosmic dust and billions and billions of stars, includes our solar system: the sun and its nine orbiting planets. What we see is only a minute fraction of the entire **universe**. Beyond the Milky Way, extending into space for distances beyond our comprehension, are countless other galaxies. It was probably only when we humans first ventured into space in the 1960s that we began to appreciate the smallness and insignificance of our planet in relationship to the universe as a whole. The first photographs of the Earth taken from the moon showed us our planet suspended in the black vastness of space (Figure 1.1).

According to the most recent research, the universe began between 12- to 13.5-billion years ago. Although differences of opinion still exist, many scientists believe that all of the matter in the universe was once compressed into an infinitesimally small and infinitely dense mass that exploded with tremendous force. This explosion of unimaginable proportions—appropriately called the **big bang**—generated enormous amounts of light, heat, and energy, and released the cosmic matter from which the galaxies and stars were eventually formed. The universe began expanding in all directions and, according to most astronomers, has been expanding ever since.

Figure 1.1 The Earth seen from the surface of the moon.

Galaxies and Stars

As the universe expanded, it cooled very, very slowly, and cosmic matter gradually condensed to form the first galaxies. Atoms of hydrogen—the simplest and lightest of all of the elements—formed in the swirling clouds of condensing matter. Over billions of years, the galaxies gave birth to the early stars, which generated sufficient heat to cause hydrogen atoms to fuse (join) to form atoms of helium, the second lightest of the elements. The energy released during these **fusion reactions** initiated further fusion reactions, in which all 90 of the remaining naturally occurring elements found on Earth were formed. In the universe as a whole, 90% of all atoms are hydrogen, and 9% are helium, whereas the remaining 1% are atoms of all of the other elements. Scientists believe that subsequent explosions of the early stars scattered the elements and that our sun was born from the debris of one of these explosions. The sun, which to us appears very bright, is an average-sized star that is located toward the edge of the Milky Way.

The Planets in Our Solar System

Scientists still do not know with any certainty how the planets in our **solar system** developed (a solar system is a group of planets that revolve around a star), but it is generally believed that they began to form approximately 5-billion years ago from hot, mainly gaseous matter rotating about the sun. With time, the matter slowly cooled, and solid particles condensed from the gases. The particles gradually coalesced into clumps of matter. Larger clumps had stronger gravity and gradually drew in and retained additional particles, eventually forming the eight planets that revolve around the sun: Mercury, Venus, Earth, Mars, Jupiter, Saturn, Uranus, Neptune, along with the dwarf planet, Pluto (Figure 1.2).

The four planets closest to the sun—Mercury, Venus, Earth, and Mars—are called **terrestrial planets** and are small and dense. The more distant **giant planets**—Jupiter, Saturn, Uranus, and Neptune—are much larger and are of lower density than the terrestrial planets.

The Earth and the other terrestrial planets formed close to the sun and were so hot that lighter, easily evaporated materials could not condense and were swept away. Only substances with extremely high boiling points, such as metals and minerals, condensed on these planets. Mercury, the planet closest to the sun and therefore the hottest, is composed mainly of iron. On the Earth, which formed at a somewhat lower temperature, silicates and other metals besides iron were able to condense. (Silicates are minerals that are formed from the elements silicon, oxygen, and a variety of metals.) The larger planets, with their greater mass and thus a stronger gravitational pull, retained gases—mostly hydrogen and helium—in the atmospheres surrounding them. Some important features of the planets as they exist today are listed in Table 1.1.

The Sun

The sun is the ultimate source of energy for life on Earth. It makes up 99.9% of the mass of the solar system, and its diameter is approximately 110 times as great as that of the Earth. Scientists estimate that temperatures near the center of this immense rotating sphere of extremely hot gases reach almost 15,000,000°C (27,000,000°F). Fusion reactions occur at these incredibly high temperatures, continually releasing tremendous amounts of energy that are continually in the form of heat and light. These fusion reactions have

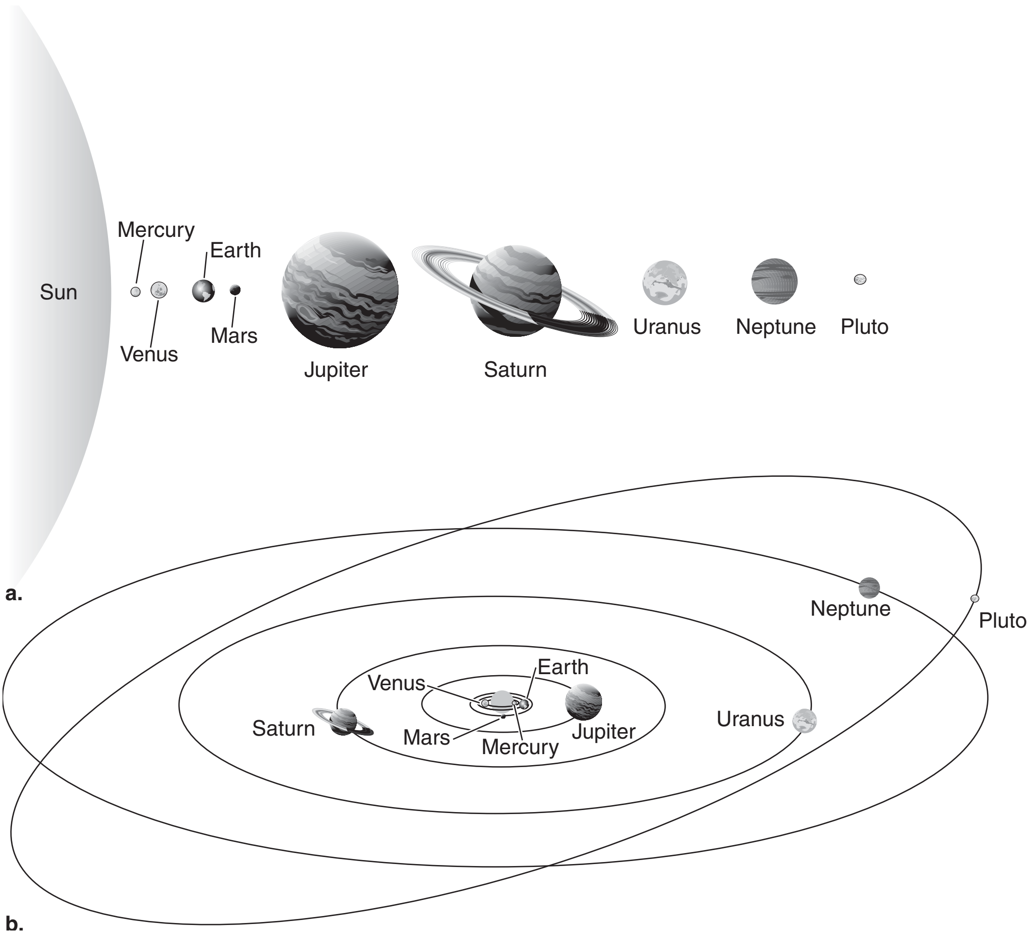


Figure 1.2 The solar system. (a) The relative sizes of the planets. (b) The planets in their orbits around the sun.

allowed the sun to shine brightly for billions of years and will allow it to continue for billions more.

■ Differentiation of the Earth into Layers

Exactly how the Earth evolved to its present state is not known, but Earth scientists believe that when the Earth was first formed approximately 4.7-billion years ago. It was homogeneous in composition: a dense, rocky sphere with no water on its surface and no atmosphere. Then, over time, the interior of the sphere gradually grew hotter, and the Earth became differentiated into layers, with each layer having a different chemical composition. This crucial period in the development of the Earth led to the formation of its magnetic field, atmosphere, oceans, and continents, and—ultimately—to life.

Table 1.1

Important Features of the Planets in Our Solar System

Planet	Diameter (km)	Diameter (mi)	Mass (Earth = 1)	Density (water = 1)	Gravity (Earth = 1)	Time for One Rotation on Axis (Earth days or hours)	Time for One Revolution around Sun (Earth years)	Distance from Sun (million km)	Distance from Sun (million mi)	Composition of Atmosphere
Terrestrial										
Mercury	4835	3004	0.055	5.69	0.38	59 days	0.24	57.7	36.8	None
Venus	12,194	7577	0.815	5.16	0.89	243 days	0.62	107	66.9	CO_2
Earth	12,756	7926	1.00	5.52	1.00	1.00 days	1.0	149	92.6	N_2, O_2
Mars	6760	4200	0.108	3.89	0.38	1.03 days	1.9	226	141	CO_2, N_2, Ar
Giant										
Jupiter	141,600	87,986	318	1.25	2.64	9.83 hours	12	775	482	H_2, He
Saturn	120,800	75,061	95.1	0.62	1.17	10.23 hours	29	1421	883	H_2, He
Uranus	47,100	29,266	14.5	1.60	1.03	23.00 hours	84	2861	1777	H_2, He, CH_4
Neptune	44,600	27,713	17.0	2.21	1.50	22.00 hours	165	4485	2787	H_2, He, CH_4

Heating of the Earth

Three factors are believed to have caused the Earth to heat. First, the cosmic particles that collided and clumped to form the Earth were drawn inward by the pull of gravity. As more particles collided with the developing planet, heat was released. Some of this heat was retained within the Earth; this heat gradually built up as increasing amounts of material accumulated.

As the Earth grew, material in the center was compressed by the weight of new material that struck the surface and was retained. Some of the energy that was expended in compression was converted to heat and caused a further rise in the temperature within the Earth.

The third and very significant factor in the warming of the Earth was the decay of radioactive elements within the interior that released energy in the form of heat. The atoms in radioactive elements are unstable and disintegrate spontaneously, emitting atomic particles and energy. In this process, which continues today, the radioactive elements are converted into atoms of other elements (discussed in Chapter 10). Only a very small percentage of naturally occurring elements has atoms that disintegrate in this way, and the heat generated with each disintegration is extremely small. Nonetheless, Earth scientists have calculated that the retention of this heat within the Earth over billions of years (together with the heat released as new material accumulated and was compressed) would have been sufficient to raise the temperature of the material at the center of the Earth to the point where it became molten.

It seems probable that this critical temperature was reached approximately 1-billion years after the Earth was born. Metallic iron, which melts at 1535°C (2795°F) and makes up over 30% of the mass of the Earth, began to melt. This heavy molten iron, together with some molten nickel, sank to the center of the Earth. As the molten iron sank, less dense material was displaced and rose toward the surface. As a result, the Earth ceased to be homogeneous and eventually became differentiated into three distinct layers: the **core**, the **mantle**, and the **crust** (Figure 1.3).

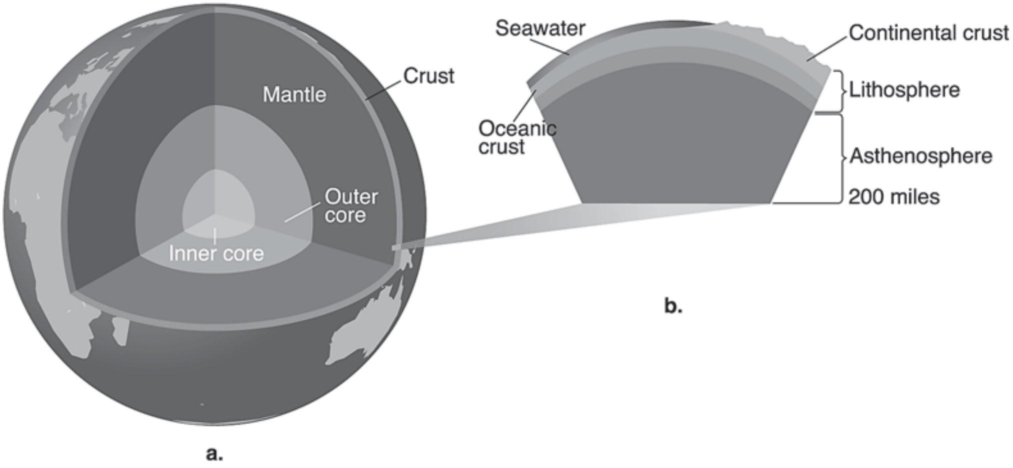


Figure 1.3 The structure of the Earth: (a) The Earth is differentiated into three distinct layers called the core, the mantle, and the crust. (b) The lithosphere, which comprises the continental and oceanic crust together with the solid upper part of the mantle, rests on the partially molten asthenosphere.

The Core

The Earth's core, which extends 3500 km (2200 miles) from the planet's center, is believed to be composed of iron and small amounts of nickel. The metals are thought to be in solid form in the inner core and molten in the surrounding outer core (Figure 1.3a). Because the core is inaccessible to us, there is no way to prove that it consists primarily of iron, but considerable indirect evidence supports this view. For example, analysis of light emitted by the sun and stars has revealed that iron is the most abundant metal in the universe, and most of the meteorites that have landed on the Earth from outer space are composed of iron. Furthermore, analysis of waves generated by Earthquakes has shown that the core is very dense, and iron is the densest metal found in any quantity on Earth.

The Mantle

The Earth's mantle, which lies between the core and the crust, is approximately 2900 km (1800 miles) thick (Figure 1.3b). The relatively thin upper part of the mantle is solid and rigid, but the layer below it—called the **asthenosphere**—although essentially solid, is able to flow extremely slowly, like a very thick, viscous liquid. In the deep mantle, below the asthenosphere, the rock is believed to be rigid.

The Crust

Above the mantle is the crust, which forms the thin outer skin of the Earth (Figure 1.3b). The crust is thicker beneath the continents than beneath the oceans. Its thickness ranges from 6 km (4 miles) under the oceans to 70 km (45 miles) under mountainous regions. Although the crust makes up a very small part of the Earth as a whole, we gather from it practically all of the resources that sustain our way of life.

Together the crust and the solid upper part of the mantle make up the relatively cool and rigid **lithosphere**, which floats on the hotter, partially molten asthenosphere. The boundary between the lithosphere and the asthenosphere is not caused by a difference in the chemical composition of their rocks but by a change in the physical properties of the rocks that occurs as temperature and pressure increase with depth.

Relative Abundance of the Elements in the Earth

By mass, the four most abundant elements in the Earth are iron, oxygen, silicon, and magnesium, which together account for approximately 93% of the Earth's mass (Figure 1.4a). Nickel, sulfur, calcium, and aluminum make up another 6.5%. The remaining 0.5% or so of the Earth's mass is made of the other 84 naturally occurring elements.

Primarily because most of the iron sank to the center of the Earth during the period of differentiation, the relative abundance of the elements in the crust differs greatly from that in the Earth as a whole (Figure 1.4b). Seventy-four percent of the crust consists of oxygen and silicon, whereas aluminum, iron, magnesium calcium, potassium, and sodium together account for 25%.

It might have been expected that as the Earth became differentiated into layers, the elements would have been distributed strictly according to mass, with the heavier elements falling to the Earth's center and the lighter ones rising to the surface. This distribution did not occur, however, because some elements combined with other elements to form compounds, and the melting points and densities of the compounds (rather than those of the elements from which they were formed) primarily determined how the elements were distributed in the Earth. For example, silicon, oxygen, and various metals combined to form

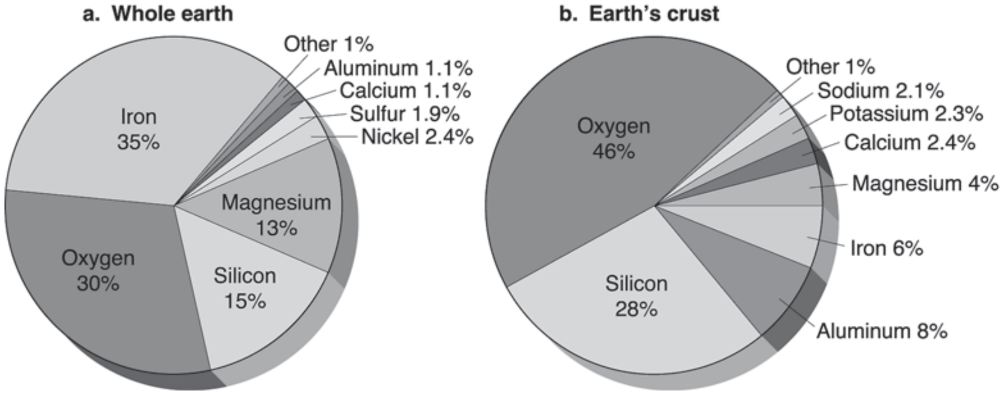


Figure 1.4 The relative abundance (by mass) of elements in the whole Earth and in the Earth's crust. Because of the differentiation that occurred early in the Earth's history, the percentage of iron in the crust (b) is lower than that in the whole Earth (a), and the percentages of aluminum, silicon, and oxygen (the elements that combine to form silicates) are higher.

silicates, which are relatively light compounds that melt at relatively low temperatures. When the Earth's interior was hot, these silicates rose to the surface. They are the most abundant minerals in the Earth's crust.

As a result of the chemical changes that occurred during the period of differentiation, the distribution of the elements on the Earth is very uneven. The relative abundance in the Earth's crust of the economically valuable elements is shown in Table 1.2. Of these, only four—aluminum, iron, magnesium, and potassium—are present in amounts greater than 1% of the total mass of the crust. It is fortunate for us that as the result of geologic processes that have been occurring for millions of years, the less abundant (but valuable) elements such as gold and silver are concentrated in specific regions of the world. If these elements had been distributed evenly throughout the Earth's crust, their concentrations would be too low to make their extraction technically or economically feasible.

Formation of the Oceans and the Atmosphere

It is generally accepted that there was no water on the Earth's surface for millions of years after the planet was formed. Then, as the interior of the Earth heated up, minerals below the Earth's surface became molten. The molten material rose to the surface, and oxygen (O) and hydrogen (H) atoms that were chemically bound to certain minerals escaped explosively into the atmosphere as clouds of water (H_2O) vapor. In these tremendous volcanic eruptions, which were widespread and numerous, carbon dioxide (CO_2) and other gases were also released from the planet's interior (Figure 1.5). The lighter gases escaped into space, but the heavier ones, including water vapor and carbon dioxide, were held by gravity and formed as a thick blanket of clouds surrounding the Earth. In time, as the Earth's surface cooled, the water vapor condensed, and the clouds released their moisture. For the first time, rain fell on the Earth. During the next several million years, volcanoes continued to erupt, and the oceans filled with water as more rain fell.

Table 1.2

The Relative Abundance of the Economically Valuable Elements in the Earth's Crust

Name	Chemical Symbol	Abundance in Crust (% by mass)
Aluminum	Al	8.00
Iron	Fe	5.80
Magnesium	Mg	2.77
Potassium	K	1.68
Titanium	Ti	0.86
Hydrogen	H	0.14
Phosphorus	P	0.101
Fluorine	F	0.0460
Sulfur	S	0.030
Chlorine	Cl	0.019
Chromium	Cr	0.0096
Zinc	Zn	0.0082
Nickel	Ni	0.0072
Copper	Cu	0.0058
Cobalt	Co	0.0028
Lead	Pb	0.00010
Arsenic	As	0.00020
Tin	Sn	0.00015
Uranium	U	0.00016
Tungsten	W	0.00010
Silver	Ag	0.000008
Mercury	Hg	0.000002
Platinum	Pt	0.0000005
Gold	Au	0.0000002

Source: Adapted from F. Press and R. Siever, *Earth,* 3rd ed. (New York: W. H. Freeman, 1982), p. 553.

The Earth's first atmosphere was quite different from the one that surrounds the Earth today. Volcanic eruptions continued to occur long after the Earth's surface had cooled to the point where water vapor began to condense to form the oceans. Evidence suggests that in addition to water vapor and carbon dioxide, the enormous volumes of gases emitted were mostly nitrogen, with smaller amounts of carbon monoxide, hydrogen, and hydrogen chloride—the same gases that erupting volcanoes emit today. Hydrogen gas, being very light, was lost into space, but the Earth's gravitational pull held other gases near the surface.

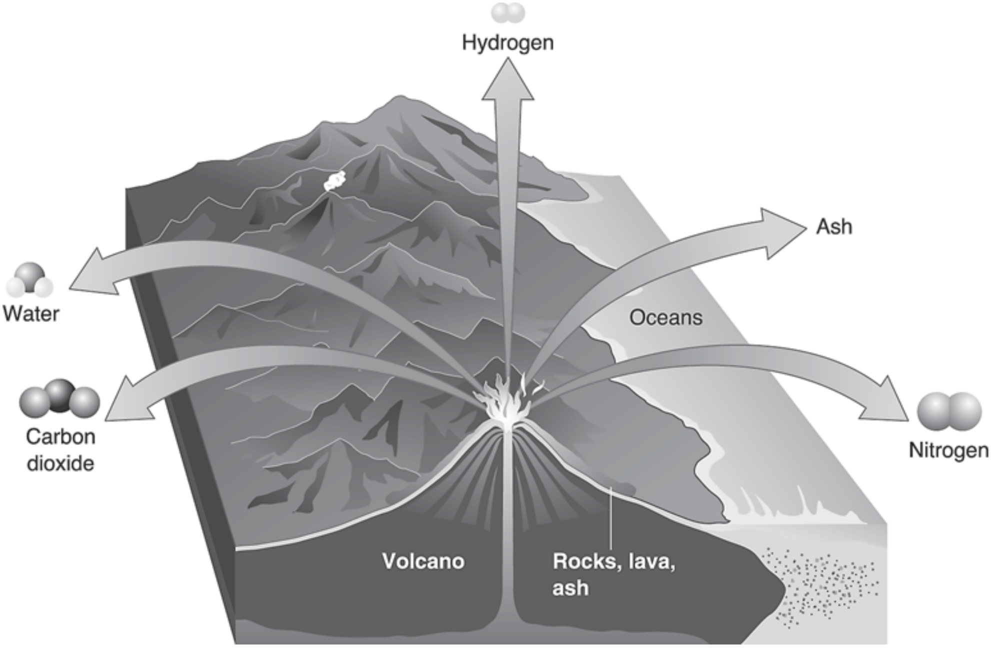


Figure 1.5 In volcanic eruptions, huge quantities of rocks, lava, ash, and gases (mainly water vapor, carbon dioxide, and nitrogen) are ejected. Early in the Earth's history, when volcanic eruptions were frequent and widespread, water vapor released to the atmosphere condensed and fell to Earth as rain and the oceans were formed; carbon dioxide and nitrogen became the main constituents of the early atmosphere.

After millions of years of volcanic activity, the atmosphere was rich in nitrogen and carbon dioxide but was completely devoid of oxygen. Today, the Earth's atmosphere is still rich in nitrogen (78%), but only 0.03% of the atmosphere is carbon dioxide, whereas oxygen accounts for 21%. There were two main ways in which the excess carbon dioxide was removed: First, when rain began to fall on the Earth, very large quantities of carbon dioxide dissolved in the oceans that were formed, and much of it combined with calcium in the water to form limestone (calcium carbonate). Second, approximately 3-billion years ago, the first primitive blue-green algae, or **cyanobacteria**, developed in shallow waters. Similar to the more advanced plants on Earth today, these organisms used energy from the sun to convert carbon dioxide and water into simple carbohydrates and oxygen by the process of **photosynthesis**. Oxygen escaped from the water and for the first time entered the atmosphere.

As cyanobacteria multiplied, increasingly larger amounts of carbon dioxide were removed from the atmosphere and replaced with oxygen. Eventually, when photosynthesis had been going on for millions of years, the Earth's atmosphere attained its present composition. Although photosynthesis was important in reducing the carbon dioxide content of the atmosphere, dissolution in the oceans followed by the formation of limestone was the major factor in its removal.

Since the beginning of the Industrial Revolution, humans have been pouring unprecedented amounts of carbon dioxide back into the atmosphere by burning carbon-containing fuels such as coal and petroleum. There is growing concern that the continuing increase in atmospheric carbon dioxide may lead to a rise in the Earth's temperature, which could have catastrophic consequences for our planet (discussed in Chapters 2 and 3).

■ Rocks and Minerals

The rocks that make up the hard surface of the Earth, the lithosphere, are composed of one or more substances called **minerals.** A mineral is a naturally occurring, usually crystalline, substance that has a definite composition or a restricted range of composition; it may be either an element or a compound.

Minerals form the **inorganic** part of the Earth's crust; materials derived from the decayed remains of plants and animals make up the **organic** part. The terms *inorganic* and *organic* were introduced in the 18th century to distinguish between compounds derived from nonliving matter and compounds derived from plant and animal sources. At that time, chemists believed that the complex compounds that make up living matter—such as carbohydrates and proteins—could be produced only by living organisms. Once it was discovered that these compounds, which all contain carbon, could be synthesized in the laboratory, the definition of organic compounds was broadened to include all nonmineral compounds of carbon.

Some minerals have been known and used since ancient times. There is evidence that very early in human history flint and obsidian (a volcanic glass) were shaped to make weapons and primitive knives, and clay was formed into pottery vessels and bricks. Gold, silver, copper, and brightly colored minerals such as jade and amethyst were fashioned into jewelry and other objects, and pigments were made from red and black iron oxides.

Over 2500 distinct minerals have now been identified, but only a few of them are distributed widely over the Earth's surface. Many of the more valuable minerals are found in only a few limited regions of the world, where they became concentrated as a result of the upheaval and subsidence of crust materials and other rock-forming processes that have gone on for millions of years. The study of the composition of rocks has been an important factor in helping to explain how the Earth was formed.

■ Rocks as Natural Resources

Rocks and minerals are **natural resources**, which are anything taken from the physical environment to meet the needs of society. Such resources may be **renewable** or **nonrenewable.** Resources such as soil, natural vegetation, fresh water, and wildlife are all renewable; if not depleted too rapidly, they are replaced in natural recycling processes. Rocks and minerals—as well as oil, natural gas, and coal—are nonrenewable. They are present in the Earth in fixed amounts and are not replaced as they are used.

Rocks and minerals are quarried and used widely, often in modified form, in the construction and chemical industries and for making ceramics and many other products.

■ Ores and Metals

Minerals in which a particular metallic element occurs in a sufficiently high concentration to make mining and extracting it economically feasible are termed **ores.** Silicates, although very abundant, are seldom used as ores because extraction of their metallic elements involves high costs and technologic difficulties. Instead, most metals are extracted from sulfide, oxide, carbonate, chloride, and phosphate ores, which are found concentrated in a relatively small number of regions of the world. We now consider briefly the sources and uses of a few of the metals that are essential for our modern way of life.

Iron

Iron (Fe), the fourth most abundant element in the Earth's crust (see Figure 1.4), is the metal that industrialized nations use in the greatest quantity. The ores from which it is extracted usually contain a mixture of two iron oxides: hematite (Fe_2O_3) and magnetite (Fe_3O_4). Nearly all of the iron extracted from ores is used to manufacture **steel**, an alloy of iron with a small amount of carbon. The percentage of carbon determines the properties of steel. Low-carbon steel (less than 0.25% carbon) is relatively soft and suitable for making cans and wire. High-carbon steel (up to 1.5% carbon) is very hard and strong and is used for making tools and surgical instruments. Steels with a variety of properties and uses are made by alloying iron with small amounts of other metals.

Although world reserves of iron are still large—and the United States has an abundant supply within its own borders—the demand for iron ore and steel continues to rise, and known sources of ores must eventually run out.

Aluminum

Aluminum (Al) is the second most highly used metal in industrialized nations and is the third most abundant element in the Earth's crust (see Figure 1.4). Nearly all naturally occurring aluminum is a component of complex silicates, but currently, there is no economically viable way to extract it from these silicates. Instead, the source of practically all aluminum is **bauxite**, an ore rich in aluminum oxide that is found in quantity in only a few places in the world. Very large deposits of bauxite are found in Jamaica and Australia.

Aluminum is a light, strong metal that is used primarily for making beverage cans. In the building industry, it is used to make doors, windows, and siding; it is ideal for cooking utensils and many household appliances. Because it is a good conductor of electricity, aluminum is used extensively for high-voltage transmission lines. When alloyed with magnesium, it forms a light but strong material that is vital to the manufacture of airplane bodies.

Aluminum corrodes less easily than does iron; this feature is an important advantage for certain purposes such as the construction of homes, but it also means that aluminum cans are very slow to degrade. Tin cans (which are made of steel and coated with tin) eventually break down completely, but discarded aluminum cans remain in the environment for a very long time.

Aluminum oxides that include traces of certain metal impurities are valuable as gemstones. Rubies are crystalline aluminum oxide that is colored red with traces of chromium. Sapphires, which occur in various shades of yellow, green, and blue, owe their color to traces of nickel, magnesium, cobalt, iron, or titanium.

Copper

Another valuable and extensively used metal is copper (Cu), which today is obtained from low-grade copper sulfide ores. Because the copper content of these ores is 1% or less, the cost of obtaining the pure metal is high. Valuable by-products of copper production are gold and silver, which are frequently present in very small quantities in the original ores. Increasingly, pure copper is being obtained by recycling copper-containing materials.

Copper is an excellent conductor of electricity and is used extensively for electrical wiring. It is also used for plumbing fixtures and as a constituent of alloys. Brass is an alloy of copper and zinc, and bronze is an alloy of copper and tin. Copper and copper alloys are also used for coinage.

Strategic Metals

Because metal ores are very unevenly distributed in the Earth's crust, many countries must depend on imports for their supplies. Within the United States, for instance, there are no deposits of ores of many **strategic metals**, which are metals that are essential for industry and defense. Large reserves of chromium (Cr), manganese (Mn), and the platinum (Pt) group of metals (platinum [Pt], palladium [Pd], rhodium [Rh], iridium [Ir], osmium [Os], and ruthenium [Ru]), which are needed for the manufacture of specialty steels, heat-resistant alloys, industrial catalysts, and parts for automobiles and aircraft, are found in only South Africa and the former Soviet Union.

■ Mineral Reserves

The world's human population continues to grow at an ever-increasing rate, and with it grows the demand not only for food but also for material goods. To meet these material needs, metals and minerals (nonrenewable resources) are being consumed at a tremendous rate, one that is bound to rise as the developing nations become increasingly industrialized.

The industrialized nations currently consume a disproportionate amount of the Earth's mineral reserves. The North American continent, for example, has less than 10% of the world's population, but it consumes almost 75% of the world's production of aluminum. The same disproportionate usage rate holds true for many other metals. At this rate of consumption, supplies of many important metals will be severely depleted by the end of the first quarter of the 21st century.

It is possible that new mineral deposits will be discovered in the future. However, because geologists have already thoroughly explored most of the Earth's surface, it is unlikely that significant quantities of ores will be found. One source of minerals that has not yet been fully explored is the ocean floor. Areas where a new floor is being formed have been shown to be rich sources of manganese and polymetallic sulfides. As deep-sea mining technology develops, sites on the ocean floor may provide a much needed source of metals.

One of the most important ways to conserve our mineral supplies is to recover metals by recycling. More and more communities are collecting aluminum beverage cans (and also glass containers and paper) for recycling, which saves energy as well as conserves natural resources. Approximately half as much energy is required to make new aluminum cans from old cans as is needed to make them from bauxite.

■ The Origin of Life on Earth

Thus far in this chapter we have considered only the nonliving part of the Earth. We now turn our attention to the living creatures that inhabit the planet. It is generally agreed that life on Earth began between 3.5- and 4.0-billion years ago, but exactly how it began will probably never be fully understood. The early atmosphere is thought to have consisted mainly of CO_2 and nitrogen (N_2), with smaller amounts of ammonia (NH_3) and methane (CH_4), which would have dissolved to some extent in the early oceans. Some scientists believe that life began in tidal pools or lagoons where evaporation would have concentrated dissolved chemicals, making it possible for them to combine to form simple amino acids. **Amino acids**, which are basic building blocks of living tissues, might then have joined together to form simple proteins, and further reactions in the chemical "soup" of the ocean pools or lagoons could have produced other compounds that are essential for life. The essential compounds produced in this way might then have gradually clumped to form larger masses. Membranes formed around these masses, separating them from the surrounding environment. The organic matter gradually acquired the characteristics of living cells.

Other researchers believe that life is more likely to have begun near volcanic vents on the ocean floor where there was heat and protection from destructive ultraviolet (UV) radiation. Still others believe that the first living organisms did not arise on Earth but came from outer space in meteorites or interplanetary dust.

No matter how the first one-celled organisms were formed, the environment was devoid of oxygen. These early **anaerobic bacteria** (bacteria that require an oxygen-free environment) flourished until the development of oxygen-producing cyanobacteria. Oxygen was lethal to the anaerobic bacteria and, except in a few specialized locations, they gradually died.

■ The Uniqueness of the Earth

The Earth is unique, as it is the only planet in our solar system that developed an environment that is capable of supporting life as we know it. The position of the Earth relative to the sun made possible the formation of the atmosphere and the oceans, which together maintain the temperature on the Earth's surface within a very narrow range: a range that extends approximately from the freezing point of water (0°C, 32°F) to the boiling point of water (100°C, 212°F). If the Earth had formed a little closer to the sun, it would have been too hot to support life; if it were a little farther away, it would have been too cold.

The size of our planet is another important factor. If the Earth were much smaller, the pull of gravity would be too weak to hold the atmosphere around the Earth. Without an atmosphere, we would be exposed to life-destroying amounts of UV radiation from the sun. If the Earth were much larger, the atmosphere would be thicker and would contain more kinds of gases, possibly some that are poisonous.

■ The Environment

When speaking of the environment, we are referring to all of the factors, both living and nonliving, that in any way affect living organisms on Earth. The living, or **biotic**, factors include plants, animals, fungi, and bacteria. The nonliving, or **abiotic**, factors include physical and chemical components such as temperature, rainfall, nutrient supplies, and sunlight.

■ Ecosystems

For purposes of study and the sake of simplicity, it is useful to subdivide the environment, which comprises the entire Earth, into small functional units called **ecosystems.** An ecosystem consists of all of the different organisms living within a finite geographic region and their nonliving surroundings. It may be a forest, desert, grassland, marsh, or just a pond or a field (Figure 1.6). Interrelationships between the organisms and the surroundings are such that an ecosystem is usually self-contained and self-sustaining.

Producers and Consumers

Ecosystems are sustained by the energy that flows through them. The biotic part of any ecosystem can be divided into producers of energy and consumers of energy. Green plants and cyanobacteria (blue-green algae) are the **producers**; they are able to manufacture all of their own food. By means of photosynthesis, they absorb light energy from the sun and use it to convert H_2O and CO_2 from the air into the simple carbohydrate glucose ($C_6H_{12}O_6$). At the same time, O_2 is released to the atmosphere. By further reactions between glucose and chemicals obtained from water and soil, plants manufacture all of the complex materials that they need (Figure 1.7). The plant world includes trees, bushes, flowers, grasses, mosses, and algae.

Consumers are unable to harness energy from the sun to manufacture their own food and must consume plants or other creatures to obtain the nutrients and energy that they need. Consumers can be divided into four main groups according to their food source: herbivores, carnivores, omnivores, and decomposers. **Herbivores** feed directly on producers. Examples of herbivores are deer, cows, mice, and grasshoppers. **Carnivores** eat other animals and include spiders, frogs, hawks, and all cats (lions, tigers, and domestic cats). The animals that carnivores eat may be herbivores, carnivores, or omnivores. **Omnivores** are creatures, including rats, raccoons, bears, and most humans, that feed on both plants and animals.

Decomposers feed on **detritus**, the freshly dead or partly decomposed remains of plants and animals. Decomposers include bacteria, fungi, Earthworms, and many insects. Decomposers perform the very useful task of breaking down complex organic compounds in dead plants and animals into simpler chemicals and returning them to the soil for the producers to reuse. In this way, many nutrients are endlessly recycled through an ecosystem.

Figure 1.6 An ecosystem is a group of plants and animals interacting with one another and their surroundings. It may cover a small area, such as a pond.

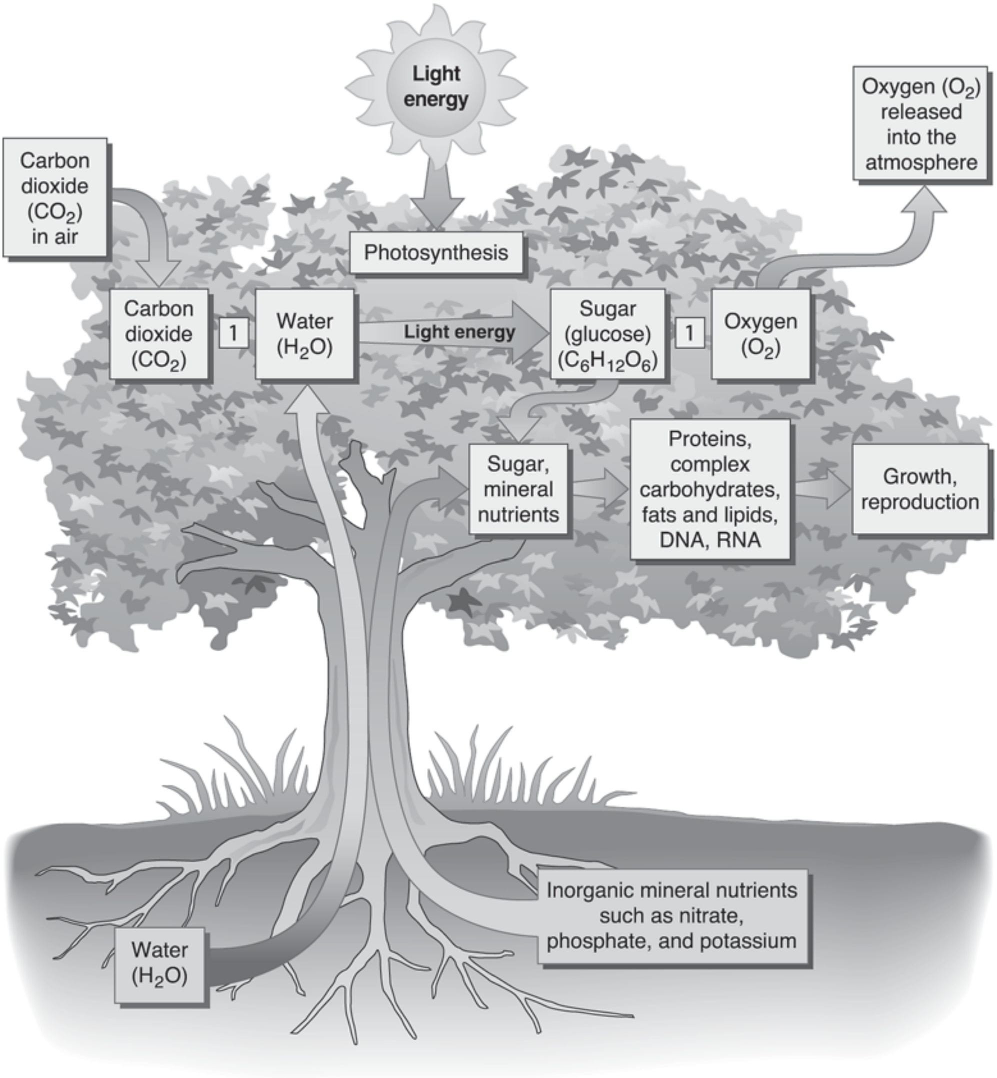


Figure 1.7 Green plants such as this tree are producers. In the process of photosynthesis, they use light energy from the sun to convert carbon dioxide and water to glucose and oxygen. The oxygen is released to the atmosphere; the glucose, together with mineral nutrients from the soil, is used to produce the complex organic compounds that make up plant tissues.

■ The Flow of Energy Through Ecosystems

All of the activities that go on in an ecosystem require **energy**. Without a constant flow of energy from the sun to producers and then to consumers, an ecosystem would not be able to maintain itself. Before you can understand these energy relationships, you need to have some appreciation of what is meant by *energy*. We all know that it takes enormous amounts of energy to maintain our industrialized society. Energy is needed to run automobiles, heat and cool buildings, supply light, and grow food. We receive energy from the food that we eat and use it when we perform any activity. Energy exists in many forms, including light, heat, electrical energy, nuclear energy, and chemical energy. The ultimate source of energy for our planet is the sun. But what exactly is energy?

What Is Energy?

Energy is usually defined as the ability to do work or bring about change. **Work** is done whenever any form of matter is moved over a distance. Everything that goes on in the universe involves work, in which one form of energy is transformed into one or more other forms of energy. Energy is the capacity to make something happen.

All forms of energy can be classified as either kinetic or potential. **Kinetic energy** is energy of motion. A moving car, wind, swiftly flowing water, and a falling rock all have kinetic energy and are capable of doing work. Wind, for example, can turn a windmill, and water rushing from a dam (Figure 1.8a) can turn a turbine to produce electricity. **Potential energy** is stored energy, which is converted to kinetic energy when it is released. Water held behind a dam (Figure 1.8b) or a rock poised at the edge of a cliff has potential energy because of its position. When the water is released or the rock falls off the cliff, the potential energy is converted to kinetic energy. Potential energy is also stored in chemical compounds, such as those present in food and gasoline. When food is digested, chemical bonds are broken, and energy that the body needs to function is released. Similarly, combustion of gasoline in a car engine releases energy to set the car in motion.

Energy Transformations

In the universe, energy transformations occur continuously. Stars convert nuclear energy into light and heat. Plants convert light energy from the sun into chemical energy in the bonds within sugar molecules, and animals convert the chemical energy in sugars into energy of motion. None of these transformations is 100% efficient; in every case, a part of the energy is converted to some useless form of energy, usually heat. In any transformation, however, no new energy is created, and no energy is destroyed. This is the **first law of thermodynamics**

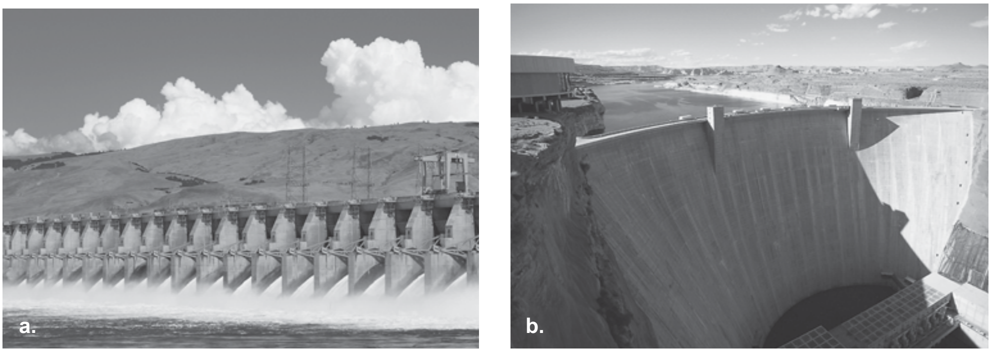


Figure 1.8 (a) The kinetic energy of water rushing from a dam can be harnessed to do useful work. (b) The potential energy of the water in the lake behind the dam is converted to kinetic energy as it is released from the dam.

(also known as the *law of conservation of energy*): Energy can be neither created nor destroyed; it can only be transformed from one form to another.

The efficiency of any energy-transfer process is defined as the percentage of the total energy that is transformed into some useful form of energy. For example, the efficiency of an incandescent electric light bulb is very low. Only approximately 5% of the electrical input is converted into light energy; the remainder, as anyone who has touched a lighted bulb knows, is converted to heat (Figure 1.9). If the light bulb is in its normal surroundings in a home or office, there is no way for the heat energy to do useful work or be converted into some useful form of energy. It is essentially lost.

The loss of useful energy is summed in the **second law of thermodynamics**: In every energy transformation, some energy is always lost in the form of heat energy that thereafter is unavailable to do useful work. This statement means that all systems tend to run "downhill." In other words, whenever any work is done, high-quality energy is converted into lower quality energy. All of the life-sustaining processes that go on in the human body and in all other living organisms follow this pattern of energy flow.

Although energy is never destroyed, the fact that energy is lost as heat in all transformations means that, unlike many material resources, energy cannot be recycled. This fact has important implications for our society, which is dependent on so many energy-inefficient machines. The gasoline engine, for instance, is only 10% efficient.

Food Chains and Trophic Levels

Green plants (producers) are the only organisms that can take energy from the sun and use the process of photosynthesis to store some of that energy in chemical bonds in sugars, starches, and other large molecules. When an animal eats a plant, the sugars and other chemical substances in the plant are broken down in chemical reactions in the animal's body. Bonds that had connected the atoms in the plant's molecules are broken, and energy is released. The energy is used to power the many activities that enable the animal to grow and survive.

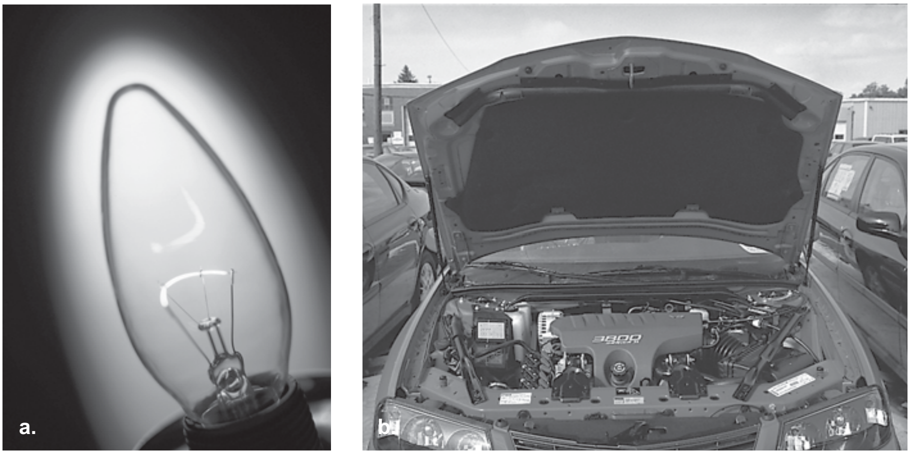


Figure 1.9 (a) When electrical energy passes through the filament of a light bulb, only about 5% of it is converted to light; the rest is lost as heat. (b) An automobile engine converts about 10% of the chemical energy in gasoline to mechanical energy that can be used to drive the vehicle; 90% is lost as heat.

Any ecosystem has innumerable feeding pathways, or **food chains**, through which energy flows. In one typical food chain, grasshoppers eat green leaves. Frogs eat grasshoppers, and fish eat frogs (Figure 1.10a). Each step in the chain is called a **trophic level.** In this example, there are four trophic levels: green leaves (plants) at the first trophic level, grasshoppers (herbivores) at the second trophic level, frogs (carnivores eating herbivores) at the third trophic level, and fish (carnivores eating carnivores) at the top of this food chain, at the fourth trophic level. A human eating a carrot would be at the second trophic level; a human eating beef that has been raised on corn would be at the third trophic level. Organisms at any one trophic level are dependent on the organisms at the level below them for their energy needs. Ultimately, all animals—including humans—are dependent on producers for their existence.

Energy and Biomass

At each trophic level in a food chain, organisms use the energy at that level to maintain their own life processes. Inevitably, as a result of the second law of thermodynamics, some energy is lost to the surroundings as heat. It is estimated that in going from one trophic level to the next, about approximately 90% of the energy that was present at the lower level is lost (Figure 1.10b). Thus, in any ecosystem, the energy at the second trophic layer (herbivores) is only approximately 10% of the energy at the first trophic level (producers). The energy at the third trophic level (carnivores) is a mere 1% of that at the first trophic level (producers). This progression has an important implication for humans. It means that it is much more efficient to eat grain than to eat beef that has been fed on grain. It also means that to support a given mass of herbivores requires a mass of producers that is 10 times as large. To support a given mass

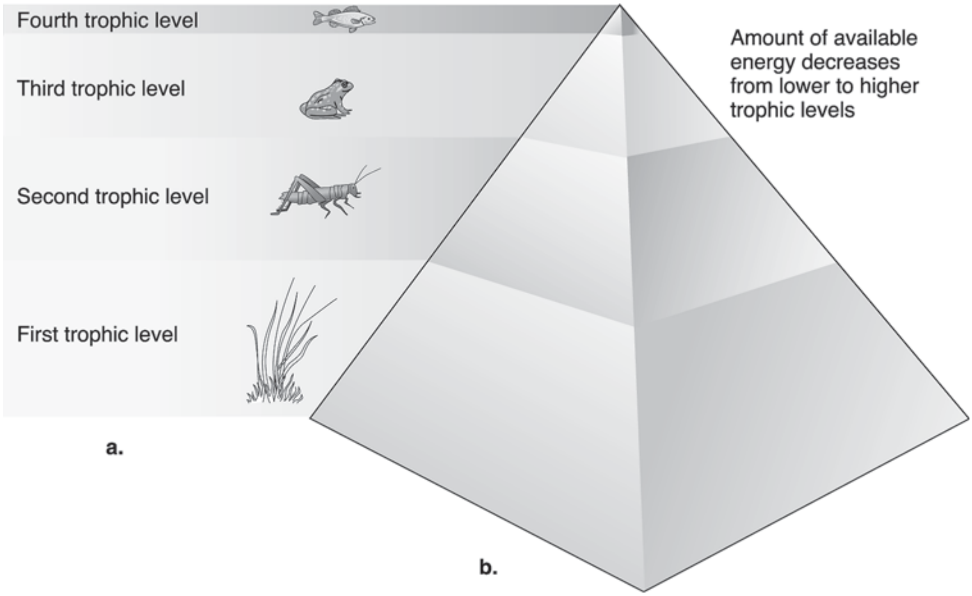


Figure 1.10 (a) A typical food chain with four trophic levels. (b) As energy passes to a higher trophic level in a food chain, approximately 90% of the useful energy is lost. High trophic levels contain less energy and fewer organisms than lower levels.

of carnivores requires a mass of producers that is 100 times as large. Because of these mass requirements, food chains rarely go beyond four trophic levels.

■ Concentration Units

When describing an environmental process, it is important to know how much of each participating chemical is present. Because the chemicals of interest are often present in a mixture of solids, liquids, or gases, chemists need to know the exact **concentration** of each participating chemical before they can assess the process. In addition, government agencies, such as the Environmental Protection Agency (EPA) that regulates the release of pollutants, often establish regulations with rules that limit the concentration of pollutants in our air, water, and soil.

These mixtures can be considered solutions. A *solution* is a homogeneous mixture of two or more substances. The minor species in a solution is called the *solute,* and the major species the *solvent*. In water, that which is dissolved is the solute, and water is the solvent. There are various ways of expressing concentration; we consider molarity, parts per million (ppm) and parts per billion (ppb).

Molarity and Molar Solutions

For aqueous solutions, the unit of concentration that is most often used by chemists is **molarity:** the number of moles of solute per liter of solution.

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

A 1-M solution contains the molecular (or formula) mass of the solute dissolved in 1 L of solution. Molarity can also be expressed as millimoles of solute per milliliter of solution (mmol/mL).

Sea water, for instance, contains 1.06% of Na^+ (percentage wt/vol, weight of Na per 100 mL of solution). Thus, to convert 1.06% Na^+ to molarity of Na^+, do the following:

Determine how many grams of Na^+ would be present in 1 L of sea water.

$$1.06\% = \frac{1.06 \text{ g Na}^+}{100 \text{ mL}} \quad \text{so to find how much in 1 L multiply by 10}$$

$$1.06\% = \frac{10.6 \text{ g Na}^+}{1000 \text{ mL}}$$

Determine how many moles are present in 10.6 g of Na^+.

$$\text{Moles Na}^+ = \frac{10.6 \text{ g}}{23.0 \text{ g/mol}} = 0.46 \text{ moles}$$

Thus, seawater has a molarity (M) of Na^+ of 0.46 M.

For elements or chemicals that are present at low concentration, it is often more convenient to express the concentration in millimolar (millimoles [10^{-3}]/L), micromolar (micromoles [10^{-6}]/L), or nanomolar (nanomoles [10^{-9}]/L) units.

Parts per Million

The unit ppm is a convenient way to describe very dilute solutions. It is frequently used for stating concentrations of pollutants in water. Consider drinking water in which the concentration of dissolved lead is 1 ppm. This means there is one part of lead in every 1 million parts of water. Parts can be expressed in any unit of mass (e.g., ounces, tons, micrograms), but the same unit must be used for both solute and solvent. We use grams.

$$1\ \text{ppm} = \frac{1\ \text{g of solute}}{1\ \text{million g of water}}$$

Because it is more convenient to measure liquids by volume rather than by mass, we change the mass of water to a volume of water:

1 g of pure water has a volume of 1 mL (density = 1.00 g/mL)

Therefore,

$$1\ \text{ppm} = \frac{1\ \text{g of solute}}{1\ \text{million}\ (1{,}000{,}000)\ \text{mL of water}}$$

We change milliliters to liters:

$$1\ \text{ppm} = \frac{1000\ \text{mL}}{1\ \text{L}} \times \frac{1\ \text{g}}{1{,}000{,}000\ \text{mL}}$$

$$= \frac{1\ \text{g}}{1000\ \text{L}}$$

We change grams to milligrams:

$$1\ \text{ppm} = \frac{1000\ \text{mg}}{1\ \text{g}} \times \frac{1\ \text{g}}{1000\ \text{L}}$$

Therefore,

$$1\ \text{ppm} = 1\ \text{mg per L}$$

Each liter of the drinking water contains 1 mg of lead.

To interconvert ppm concentration and molarity, determine how many moles of solute are present. Using the solution that contains 1 ppm lead:

$$1\ \text{ppm} = \frac{1\ \text{mg}}{1\ \text{L}}$$

To convert to moles, divide 1 mg by lead's gram atomic weight.

$$\frac{1\ \text{mg}}{207.2\ \text{g/mol}} = 0.00482\ \text{moles}$$

$$1\ \text{ppm lead} = 0.00482\ \text{mol/L} = 0.00482\ \text{M} = 4.82 \times 10^{-3}\ \text{M}$$

2ND REVISE

Parts per Billion

For certain solutions, particularly water samples containing minute traces of contaminants, it is often more convenient to express concentration in ppb rather than ppm.

$$1 \text{ ppb} = \frac{1 \text{ ppm}}{1000}$$

$$= \frac{1 \text{ mg}}{1 \text{ L}} \times \frac{1}{1000}$$

We change milligrams to micrograms as follows:

$$1 \text{ ppb} = \frac{1000\ \mu\text{g}}{1 \text{ mg}} \times \frac{1 \text{ mg}}{1 \text{ L}} \times \frac{1}{1000}$$

Therefore, 1 ppb = 1 μg/L.

EXAMPLE 1.1

The U.S. EPA set a limit for the concentration of lead in drinking water at 15 ppb. A laboratory finds the concentration of lead in a sample taken from a water fountain to be 18 μg/100 mL. Is this above or below the EPA limit? By how much?

Solution

Step 1. Given the concentration of lead in the sample

$$\text{concentration} = \frac{18\ \mu\text{g}}{100 \text{ mL}}$$

Step 2. To convert the concentration to μg/L (ppb), multiply the top and bottom by 10 to make the denominator 1000 mL (1 L)

$$\text{concentration} = \frac{18\ \mu\text{g}}{100 \text{ mL}} \times \frac{10}{10}$$

$$= \frac{180\ \mu\text{g}}{1000 \text{ mL}}$$

$$= \frac{180\ \mu\text{g}}{1 \text{ L}}$$

Remember, 1 ppb = 1 μg/L. Therefore,

$$\frac{180\ \mu\text{g}}{1 \text{ L}} = 180 \text{ ppb}$$

Step 3. This concentration is above the EPA limit of 15 ppb. How much above?

$$180 - 15 = 165 \text{ ppb}$$

The concentration (180 ppb) is above the EPA limit by 165 ppb.

The assumption that we made in these calculations is that the density of water is 1.00 g/mL. This is true in freshwater but not in seawater. The density of seawater is 1.025 g/mL. If measurements were being taken in the Chesapeake Bay, where Northern samples are freshwater and Southern samples are seawater, an error in the measured concentration would be introduced because of the difference in density between seawater and freshwater. It would be better to use a more unambiguous unit than mg/L. It is more preferable to describe the concentration as mass solute per mass solution. This means that the ppm concentration would be expressed as mg solute per kg solution and the ppb concentration as g solute per kg solution.

When expressing the concentration of solid samples, such as soil, sludge, rocks, and hazardous materials in ppm concentration, it is best to use mg/kg as the unit of measure. When expressing the concentration in ppb, it is better to use μg/kg.

Nutrient Cycles

To survive, a community of plants and animals in an ecosystem requires a constant supply of both energy and **nutrients.** The energy that sustains the system is not recycled. It flows endlessly from producers to consumers, entering as light from the sun and leaving as waste heat that cannot be reused. Nutrients, however, are continually recycled and reused. When living organisms die, their tissues are broken down, and vital chemicals are returned to the soil, water, and atmosphere.

Analysis of tissues from living organisms shows that more than 95% of the mass of the tissues is made from just 6 of the Earth's 92 naturally occurring elements: carbon (C), hydrogen (H), oxygen (O), nitrogen (N), sulfur (S), and phosphorus (P). These six elements are the building blocks for the manufacture of carbohydrates, proteins, and fats. These compounds, along with water, make up almost the entire mass of all living organisms. Plants are composed primarily of carbohydrates, whereas animals are composed primarily of proteins.

Small amounts of other elements are also required in order for plants and animals to survive and thrive. Iron (Fe), magnesium (Mg), and calcium (Ca) together make up most of the remaining 5% of the mass of living organisms. In many animals, iron is bound to hemoglobin, the protein in the blood that supplies oxygen to all parts of the body. In green plants, magnesium is bound to chlorophyll, the protein that absorbs light from the sun and that is a vital part of the process of photosynthesis. Animals with skeletons need calcium as well as phosphorus to make bones and cartilage.

Trace amounts of approximately 16 other elements are also required. Copper (Cu) and zinc (Zn), for example, are essential components of certain enzymes, specialized proteins that facilitate many vital chemical reactions within the animals' bodies.

Plants obtain the essential elements from the soil and the atmosphere. Animals obtain them from their food: plants and other animals. Consider the cycles by which supplies of oxygen, carbon, and nitrogen are constantly renewed.

The Carbon Cycle

The **carbon cycle** is illustrated in Figure 1.11. The major sources of carbon for our planet are the carbon dioxide gas in the atmosphere and the carbon dioxide dissolved in the oceans. Enormous quantities of carbon are also present in rocks, tied up in carbonates such

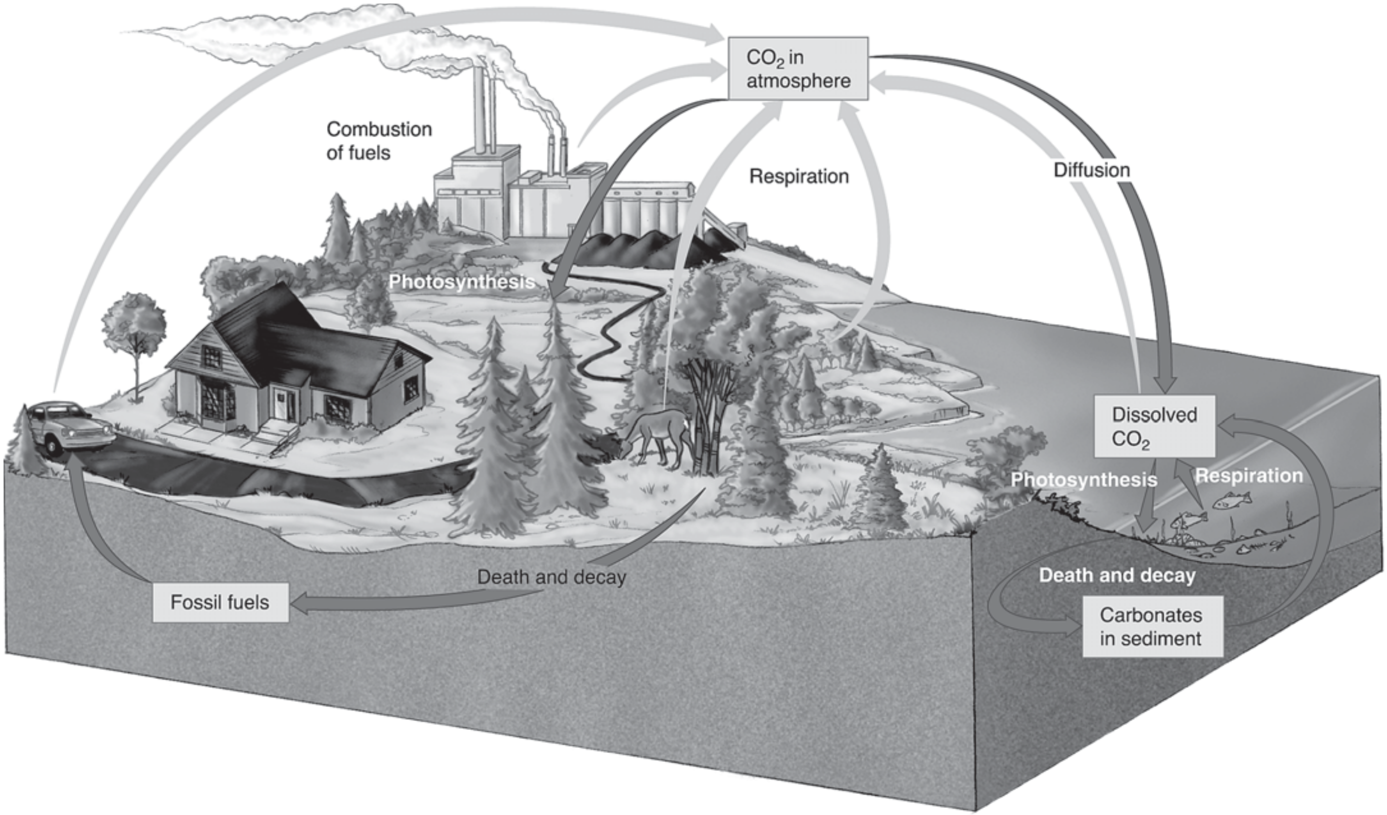


Figure 1.11 The carbon cycle: Atmospheric carbon dioxide is consumed by green plants in photosynthesis. Respiration in animals and plants and combustion of fossil fuels return carbon dioxide to the atmosphere. Carbon also cycles through water; dissolved carbon dioxide reacts with minerals and water to form carbonates, which are deposited in sediments.

as limestone, but this source recycles so slowly that it is not available to plants and animals for their daily needs.

Atmospheric carbon dioxide, although it makes up only 0.036% (360 ppm) of the atmosphere by volume, is the starting material on which all living organisms depend. In the process of photosynthesis, carbon dioxide is taken into the leaves of green plants, where it combines with water to form sugar (glucose) and oxygen, as shown in this equation:

$$\text{Solar energy} + \underset{\text{carbon dioxide}}{6\,CO_2} + \underset{\text{water}}{6\,H_2O} \xrightarrow{\text{chlorophyll}} \underset{\text{glucose}}{C_6H_{12}O_6} + \underset{\text{oxygen}}{6\,O_2}$$

Plants on land obtain the needed water from the soil; aquatic plants obtain it from their surroundings. The water is absorbed through the plant's roots and is then transported to the leaves. Part of the sugar that is formed is stored in the leaves, and part is converted into the complex carbohydrates and other large molecules that make up plant tissues. The oxygen is released to the atmosphere.

Photosynthesis is a very complex process that is still not fully understood. It involves many chemical reactions in which the green pigment chlorophyll plays an important role. We consider only the overall result of photosynthesis, as summarized by the previous equation.

When an animal such as a deer or rabbit eats a green plant, the carbohydrates in the plant are digested and broken into simple sugars, including glucose. Glucose is absorbed into the bloodstream and carried to the cells of the animal's body, where in the process of **respiration**, it reacts with oxygen in the blood. Carbon dioxide and water are formed, and energy is released.

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

Part of the energy released in respiration is used to power the many activities that go on in living cells, and part is lost as heat.

Notice that the overall reaction for respiration is the same as the overall reaction for photosynthesis written backward. However, although cellular respiration is essentially the reverse of photosynthesis, the complex intermediate steps that are involved in the two processes are very different.

Plants engage in both photosynthesis and respiration. During the day, photosynthesis is the dominant process. At night, when there is no sunlight, respiration is dominant. Decomposers also play a part in the carbon cycle. They feed on the dead remains of plants and animals and via respiration release carbon dioxide and water to the atmosphere.

Approximately 300-million years ago, huge quantities of dead and decaying plant and animal remains became buried deeply under sediments before they could be completely broken down. Over time, the remains were compressed, and chemical reactions gradually transformed them into the fossil fuels: coal, oil, and natural gas. When these fuels are burned to release the chemical energy stored within their molecules' bonds, oxygen from the atmosphere is used to convert their carbon into carbon dioxide.

$$C + O_2 \rightarrow CO_2 + \text{energy}$$

The combustion of the fossil fuels that powers our industrialized society therefore forms an integral part of the carbon cycle. Humans also intervene in the carbon cycle when they cut down more trees than they replace, thereby decreasing the amount of carbon dioxide that otherwise would be taken from the atmosphere and converted to nutrients. The climatic implications of these two activities, both of which tend to increase carbon dioxide concentration in the atmosphere, are discussed in Chapter 3.

Another important part of the carbon cycle is the continual exchange of carbon dioxide between the atmosphere and the oceans, a process that is important in maintaining the carbon dioxide concentration of the atmosphere at a constant level. The amount of carbon dioxide that dissolves in the oceans depends mainly on the temperature of the ocean water and the relative concentrations of carbon dioxide in the atmosphere and the water. When the temperature falls or when the carbon dioxide concentration in the water becomes relatively low, more atmospheric carbon dioxide dissolves. Only the surface layer of the ocean is at equilibrium with the atmosphere. A very small portion of the dissolved carbon dioxide reacts with chemicals in the water, such as calcium, to form carbonates. The carbonates, primarily limestone ($CaCO_3$), are insoluble (do not dissolve) in water and settle on the ocean floor.

Rock formation and weathering are other aspects of the carbon cycle. Sedimentary rocks such as limestone and dolomite ($CaMg[CO_3]_2$) were formed millions of years ago from the skeletal remains of coral and other marine creatures that were rich in calcium carbonates. When chemically weathered by rain with a slight natural acidity, limestone rocks very gradually dissolve, releasing carbon dioxide into the atmosphere. This recycling of carbon through rock is a very slow process.

The **solubility product** is the equilibrium constant for the reaction in which a solid salt dissolves to give its constituent ions in solution. The solution is saturated; that is, some of the undissolved solute is in contact with the solution. For example, when limestone (calcium carbonate) is placed in water, the dissociation process is described with this equation:

$$CaCO_3 \rightleftharpoons Ca^{2+} + CO_3^{\ 2-}$$

The equilibrium constant, K for this reaction, can be written as follows:

$$K = \frac{[Ca^{2+}][CO_3^{\ 2-}]}{[CaCO_3(s)]}$$

remember that [ ] means "molar concentration"
M in mol/L or mmol/mL

The solid $CaCO_3$ is not dissolved. It is considered to be in a standard state and is omitted from the equilibrium constant. The solubility product equilibrium constant, K_{sp} for this reaction, is defined as follows:

$$K_{sp} = [Ca^{2+}][CO_3^{\ 2-}]$$

This equation describes the equilibrium concentration of the calcium and carbonate ions in solution. The K_{sp} for various solid compounds are listed in Appendix A.

If calcium carbonate ($CaCO_3$) is placed in water, the equilibrium concentration of calcium ion in solution is calculated as follows:

The K_{sp} of $CaCO_3$ is 4.9×10^{-9}

This equilibrium shows that every $CaCO_3$ that dissolves produces one Ca^{2+} and one CO_3^{2-}.

$$\left[Ca^{2+}\right] = \left[CO_3^{2-}\right]$$

Therefore,

$$K_{sp} = \left[Ca^{2+}\right]^2 = 4.9 \times 10^{-9}$$

Take the square root of both sides.

$$\left[Ca^{2+}\right] = 6.8 \times 10^{-5}\ M$$

This is a molar concentration, M, expressed as mol/L or mmol/mL.

The mass of dissolved calcium ion is (40.08 g/mol) (6.8×10^{-5} mol/L) = 2.8×10^{-3} g/L or 2.8 mg/L. This corresponds to 2.8 ppm (mg/L). The Ksp indicates that the equilibrium concentration of calcium ion should be 2.8 ppm. This means that if a limestone rock is in contact with pure water (carbon dioxide free), it will dissolve until the concentration of calcium ion reaches 2.8 ppm. Once this concentration is attained, dissolution will stop and the limestone will no longer dissolve, even though it is in contact with the solution. This calculation ignores the fact that dissolved carbon dioxide can react with calcium carbonate.

Calcium ion is the most common metal ion in rivers and lakes. When measured in lakes, the concentration is much higher than the 2.8-ppm concentration that the Ksp indicates. In natural waters, Ca^{2+} is produced by the dissolution of calcium carbonate by the action of dissolved CO_2 to produce 2 moles of bicarbonate (HCO_3^-) for each mole of Ca^{2+}.

$$CaCO_3(s) + CO_2(aq) + H_2O \qquad Ca^{2+} + 2\,HCO_3^-$$

If the pH is near neutral, most of the product is bicarbonate, not CO_3^{2-} or H_2CO_3. Measurements of Ca^{2+} and HCO_3^- in many rivers confirm that there is one calcium ion for every two bicarbonate ion. This process raises the $[Ca^{2+}]$ to 20 ppm (for an average atmospheric carbon dioxide concentration). As can be seen in Figure 1.12, the concentration of calcium and bicarbonate in river water has been measured across the world. Rivers such as the Congo, the Mississippi, and the Danube all show a relationship of $2[Ca^{2+}] = [HCO_3^-]$ and appear to be saturated with calcium carbonate. Rivers such as the Nile and the Amazon for which $2[Ca^{2+}]$ is less than $[HCO_3^-]$ are not saturated with calcium carbonate.

In principle, it may be possible to use the Ksp of a solid to estimate the solubility of a solid. In practice, the solubility of the solid is affected by the presence of other solutes. As was shown previously here, the presence of an acid can greatly increase the solid solubility. The presence of complexing agents or common ions also dramatically affects the solubility of solids.

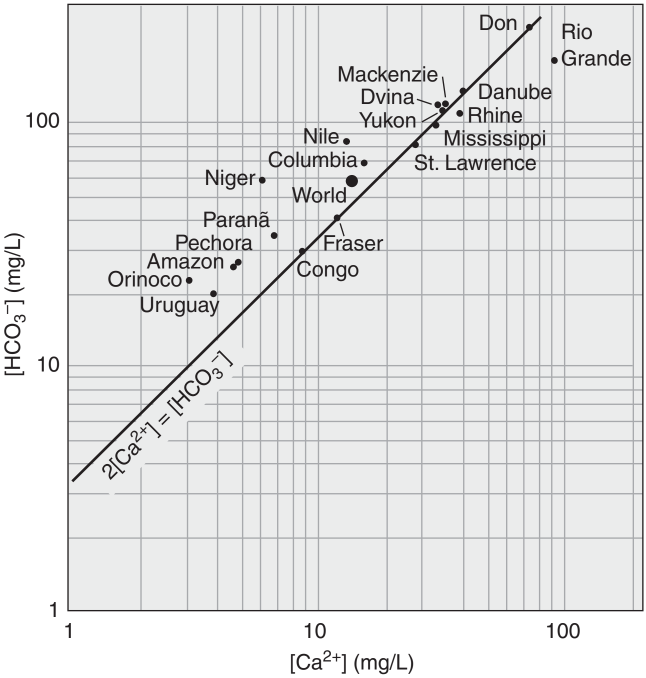


Figure 1.12 Concentrations of bicarbonate and calcium in many rivers conform to the mass balance for the reaction: $[HCO_3^-] = 2[Ca^{2+}]$. Data from W. Stumm and J. J. Morgan, *Aquatic Chemistry*, 3rd ed. (New York: Wiley Interscience, 1996), p. 189, and H. D. Holland, *The Chemistry of the Atmosphere and Oceans*. (New York: Wiley Interscience, 1978).

The Nitrogen Cycle

The **nitrogen cycle** is illustrated in Figure 1.13. Nitrogen is an essential component of proteins and of the genetic material that makes up DNA (deoxyribonucleic acid) and is a constant supply that is vital for all living organisms. Although 7% of the Earth's atmosphere is composed of nitrogen gas (N_2), plants and animals cannot use this nitrogen directly. Atmospheric nitrogen must be converted to other nitrogen compounds before it can be absorbed through the roots of plants. This change is achieved by **nitrogen fixation**, a process that is carried out by specialized bacteria that have the ability to transform atmospheric nitrogen into ammonia (NH_3). Some nitrogen-fixing bacteria live in soil, whereas others live in nodules on the roots of leguminous plants such as peas, beans, clover, and alfalfa. The ammonia produced in root nodules is converted into a variety of nitrogen compounds that are then transported through the plant as needed.

Another means by which atmospheric nitrogen is converted to a usable form is lightning. The electric discharges in lightning cause nitrogen and oxygen in the atmosphere to combine and form oxides of nitrogen, which in turn react with water in the atmosphere to form nitric acid (HNO_3). The nitric acid, which reaches the Earth's surface dissolved in rainwater, reacts with materials in soil and water to form nitrates (NO_3^-) that are directly absorbed through plant roots. Compared with biological fixation, lightning accounts for only a small fraction of the usable nitrogen in soil.

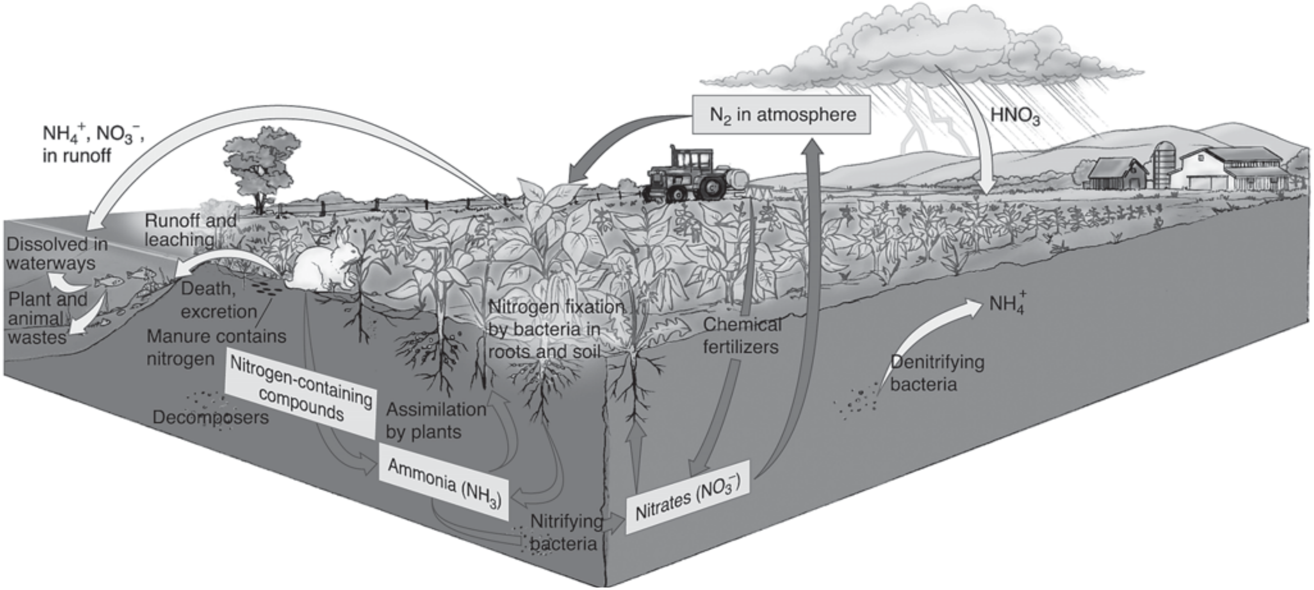


Figure 1.13 The nitrogen cycle: In nitrogen fixation, specialized bacteria convert atmospheric nitrogen to ammonia and nitrates, which plants absorb through their roots. Some nitrogen gas is fixed by lightening. Animals obtain the nitrogen they need to make tissues from plants. The wastes animals produce and their dead bodies return nitrogen to the soil in forms plants can use. In denitrification, bacteria convert nitrates in soil back to nitrogen gas.

Although some trees and grasses can absorb ammonia produced by nitrogen-fixing bacteria directly from the soil, most plants can use only nitrogen that is in the form of nitrates. The transformation of ammonia into nitrates is carried out by specialized soil bacteria in the process known as **nitrification.** A soil pH of 6.5 to 8 is optimum, and the reaction rate decreases when the pH is below 6. The two-step nitrification reaction is as follows:

$$2\,NH_4^+ + 3\,O_2 + 2\,H_2O \rightarrow 2\,NO_2^- + 4\,H_3O^+$$

$$2\,NO_2^- + O_2 + H_2O \rightarrow 2\,NO_3^-$$

When these two reactions are added together, they yield this overall reaction:

$$NH_4^+ + 2\,O_2 + H_2O \rightarrow NO_3^- + 2\,H_3O^+$$

Plants convert ammonia or nitrates that they get from soil or root nodules into proteins and other essential nitrogen-containing compounds. Animals get their essential nitrogen supplies by eating plants. When plants and animals die and decompose, the nitrogen-containing compounds in their tissues are broken down by decomposers; ammonia is eventually formed and returned to the soil. Nitrogen is also returned to the soil in animal wastes. Both urine and feces have a high content of nitrogen-containing compounds. In this way, nitrogen is continually cycled through food chains.

Not all of the ammonia and nitrates that are formed in soil by the processes just described become available for plants. Both ammonia and nitrates are very soluble in water. As rainwater percolates downward through the ground, these compounds are leached out of topsoil. They are often carried away in runoff into nearby streams, rivers, and lakes, where they are recycled through aquatic food chains. Another process that removes nitrates from the soil is **denitrification**, in which bacteria carry out a series of reactions that convert nitrates back to nitrogen gas. One of the denitrification reactions can be written as follows:

$$4\,NO_3^- + 5\,CH_2O + 4\,H_3O^+ \rightarrow 2\,N_2 + 5\,CO_2 + 11\,H_2O$$

In a few locations around the world, nitrates have accumulated in large mineral deposits. In Chile, for example, as mountain streams originating in the Andes Mountains flowed across dry, hot desert toward the sea over thousand of years, much of the water evaporated, leaving behind huge deposits of sodium nitrate.

In the natural environment, a balance is maintained between the amount of nitrogen removed from the atmosphere and the amount returned. However, because most soils contain an insufficient amount of nitrogen for maximum plant growth, farmers frequently apply synthetic inorganic fertilizers containing ammonia and nitrates. As a result of runoff from fertilized farmland, the extra nitrogen-containing compounds reaching rivers and lakes may upset the natural balance, sometimes with damaging consequences for the environment.

An alternative to the use of synthetic fertilizers for replenishing nitrogen in the soil is to plant a nitrogen-fixing crop, such as clover, and plow it back into the soil. Another method is to spread manure and allow the natural soil bacteria to degrade it and release the nitrogen-containing compounds that plants can absorb.

The Oxygen Cycle

Oxygen is all around us. As O_2, it makes up 21% of the atmosphere and is a component of all of the important organic compounds in living organisms. Oxygen is a very reactive element that combines readily with many other elements. It is a component of CO_2, nitrate (NO_2^-), and phosphate (PO_4^{3-}), and thus is an integral part of the recycling of carbon, nitrogen, and phosphorus. Oxygen is also a constituent of most rocks and minerals, including silicates, limestone ($CaCO_3$), and iron ores (Fe_2O_3, Fe_3O_4). The **oxygen cycle** is very complex and is interconnected with many other cycles. Here we briefly consider only some of the more important pathways in the cycle.

Photosynthesis and respiration are the basis of both the carbon cycle and the oxygen cycle. Oxygen is released during photosynthesis and is consumed during respiration. The oxygen and carbon cycles are also interconnected when coal, wood, or any other organic materials are burned. During burning, oxygen is consumed, and carbon dioxide is released.

Another part of the oxygen cycle is the constant exchange of oxygen between the atmosphere and bodies of water, especially oceans. Oxygen dissolved at the surface of water is carried to deeper levels by currents. Dissolved oxygen is essential for fish and other aquatic life.

The Phosphorus Cycle

The **phosphorus cycle** is illustrated in Figure 1.14. Phosphorus is a component of many important biological compounds, including DNA and enzymes that play an essential role in

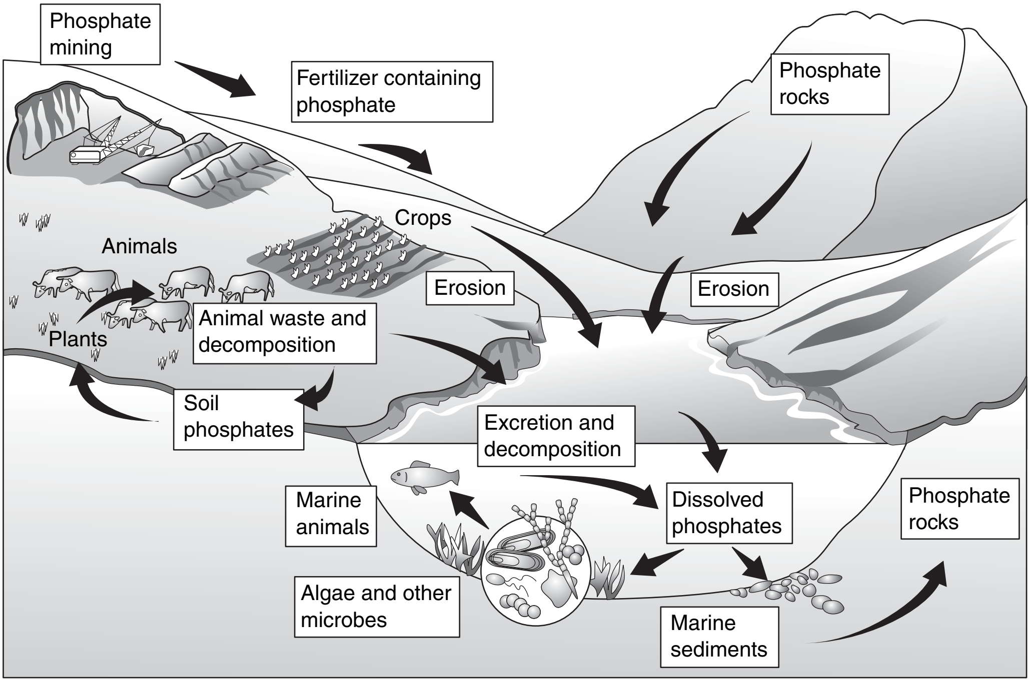


Figure 1.14 The phosphorus cycle. The main reservoir of phosphorus is phosphate rock. Phosphate released to the soil during weathering is absorbed through plant roots. Animals obtain phosphorus form plants and return it to soil in wastes and dead tissue. Phosphates leached from the soil enter waterways and are carried to the oceans.

the transfer of energy in living cells. It is a major constituent of cell membranes and is present in high concentration in bones, teeth, and shells.

Unlike the carbon and nitrogen cycles, the phosphorus cycle does not include an atmospheric phase, and the ultimate source of phosphorus for plants and animals is rock, nearly all of which contains small amounts of phosphorus, mostly in the form of phosphates (PO_4^{3-}). Two examples are apatite, $Ca_5(PO_4)_3(F)$, and vivianite, $Fe_3^{2+}(PO_4)_2 \cdot 8H_2O$. The phosphorus cycle is primarily a sedimentary cycle. As rocks weather, phosphates are very slowly dissolved and released to the soil. The crustal average abundance of apatite is 1.2 mg/g.

Plants absorb dissolved phosphate directly from the soil through their roots and transport it to their leaves, where it is incorporated into large biological molecules. Animals get their phosphorus by eating plants. Not all of the phosphate formed from rock by weathering becomes available to plants. Some of it becomes tightly bound to elements such as aluminum, iron, and calcium in compounds that are not very soluble in water.

In water, phosphorus solubility is largely controlled by pH. The phosphate ion has three forms: $H_2PO_4^-$, HPO_4^{2-}, and PO_4^{3-}.

$$H_3PO_4 \xrightarrow{pK_a\ 2.15} H_2PO_4^- + H^+ \xrightarrow{pK_a\ 7.2} HPO_4^{2-} + H^+ \xrightarrow{pK_a\ 12.3} PO_4^{3-} + H^+$$

At a pH of less than 4, insoluble iron and aluminum phosphates are formed. At a pH of greater than 8, almost insoluble calcium phosphates are formed. Iron phosphate, $Fe_3^{2+}(PO_4)_2 \cdot 8H_2O$, has a solubility product constant (K_{sp}) of 1×10^{-36}, and calcium phosphate, $Ca_3(PO_4)_2$, has a K_{sp} of 1×10^{-24}. At a pH of between 4 and 8, phosphorous has its maximum solubility.

Renewal of phosphorus through rock cycling is an extremely slow process; the main route for plants is through the recycling of the phosphorus in dead and decomposing organisms. When plants and animals die, microorganisms break down the organic phosphorus compounds in the dead tissues into inorganic phosphates, which immediately become available to plants. The average phosphorous content of plants is 0.1%.

Soil is often poor in the types of phosphates that plants can absorb readily, and on agricultural land, synthetic fertilizers rich in available phosphate are used to replenish it. As was the case with nitrogen-containing fertilizers, runoff of large quantities of phosphate-containing fertilizers can have serious effects on aquatic ecosystems.

An important natural source of phosphates is **guano**. Droppings from thousands of sea birds that breed on islands off the west coast of South America have built up into huge deposits of phosphate-rich material called guano that because of low rainfall are not washed away. Large guano deposits also have been found in Arizona and New Mexico in dry caves where thousands of bats gather. These guano deposits are mined and are an important source of phosphorus for the manufacture of fertilizers. Also important for the manufacture of fertilizer are deposits of phosphate that formed in certain regions of the world millions of years ago from the skeletal remains of sea creatures.

Phosphorus is continually recycled through plants and animals, but inevitably, some is lost by leaching and erosion of soil into streams and rivers. Human activities such as mining and farming often accelerate this loss. When phosphates reach the oceans, they react with other chemicals in sea water, and most are converted to insoluble phosphates that sink to the ocean floor and, for all practical purposes, are permanently lost. It can take millions of years for phosphate lost to the oceans to be replenished by weathering of rock. Humans add to the problem of phosphorus loss by disposing of large amounts of their wastes—which contain significant quantities of phosphorus—into waterways instead of recycling them and returning them to the land.

Reserves of phosphorus in high-grade phosphate ores are estimated to be sufficient for several hundred years. We do not have an immediate problem, but we would be wise to conserve the reserves of phosphorus that we still have and reduce phosphorus loss as far as possible.

Nature's Cycles in Balance

Life on Earth depends on a continual supply of energy from the sun and a continual recycling of materials. As you have learned here, the numerous and complex processes that sustain life are interlinked and interdependent. Because all aspects of an ecosystem are so completely interwoven, human intervention in any part of the system can easily lead to widespread disturbances of the natural balance in the environment.

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■ Keywords

abiotic
amino acids
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carbon cycle
carnivores
concentration
consumers
core
crust
cyanobacteria
decomposers
denitrification
detritus
ecosystems
energy
environment
first law of thermodynamics
food chains
fusion reactions
giant planets
guano
herbivores
inorganic
kinetic energy
lithosphere
mantle
Milky Way
minerals

molarity
natural resource
nitrification
nitrogen cycle
nitrogen fixation
nonrenewable resource
nutrients
omnivores
ores
organic
oxygen cycle
phosphorus cycle
photosynthesis
potential energy
ppb (parts per billion)
ppm (parts per million)
producers
renewable resource
respiration
second law of thermodynamics
solar system
solubility product
steel
strategic metals
terrestrial planets
trophic level
universe
work (energy)

■ Questions and Problems

1. What percentage of the mass of the solar system is accounted for by the mass of the sun?
    a. Less than 1%
    b. Approximately 10%
    c. Nearly 30%
    d. More than 99%
2. Why does the large number of hydrogen atoms in the universe suggest that other elements were built from hydrogen rather than by larger elements breaking into smaller ones?
3. Make a diagram showing a cross-section of the Earth and the three main layers into which the Earth can be divided. What element predominates at the center of the Earth?
4. By weight, which are the four most abundant elements in the whole Earth? Explain why elements are not uniformly distributed throughout the Earth.
5. By weight, which are the four most abundant elements in the Earth's crust?
6. How did water form on the Earth's surface?
7. Why does Earth's sister planet, Venus, not support life?
8. The early Earth is thought to have been devoid of oxygen. How did oxygen first appear in the Earth's atmosphere?
9. What are anaerobic bacteria?
10. If aluminum cans and steel cans are thrown into the city dump, which will degrade first? Briefly explain your answer.
11. Name the mineral from which most of the world's aluminum is extracted. From what countries does the United States obtain its supply of aluminum?
12. Give three uses for the metal copper.
13. Making an aluminum can from recycled aluminum rather than starting with aluminum ore produces an energy savings of
    a. 10%
    b. 50%
    c. 75%
14. Give three uses for the metal aluminum.

15. What valuable metals are by-products of copper production?

16. What is the purpose of adding carbon to steel?

17. What is meant by a strategic metal?

18. Significant quantities of which strategic metal are found in
   a. South America
   b. Africa
   c. Europe
   d. Australia

19. Discuss each of the following claims:
   a. New discoveries of mineral deposits will provide all we need in the future.
   b. Our children will have sufficient mineral resources.
   c. The ocean will supply all our mineral needs in the future.
   d. Recycling scarce mineral resources will guarantee a continued ample supply of those minerals.

20. What is meant by biotic and abiotic? Give examples of each.

21. Living organisms can be divided into producers and consumers. Describe how producers make their food and how consumers get the food that they need.

22. Consumers can be divided into groups according to their food source. Name the main groups and give two examples of each.

23. Organisms that consume the remains of dead plants and animals are necessary for a balanced ecosystem. Explain their purpose.

24. Define what is meant by energy. Name four forms of energy.

25. State the first law of thermodynamics. Give an example of the conversion of one form of energy to another.

26. State the second law of thermodynamics. Give an example that illustrates this law.

27. Of what type is the unusable energy that is released during almost all energy transformations?

28. If 10,000 units of energy are available to organisms at the first trophic level of a food chain, how many units of energy will be available to organisms that occupy the third trophic level in the food chain?

29. Why do most food chains contain only three or four trophic levels?

30. Why are people with limited food resources often herbivores?

31. Each time energy flows from a tropic level to the one above it, approximately what percentage of useful energy is lost?

32. What four elements are most abundant in the bodies of plants and animals? Name three other elements that are essential for life.

33. What is the molar concentration (molarity) of a 1.0-L solution that contains 156 g of nitric acid, HNO_3, or another 1.0-L solution that contains 7.8 g?

34. Seawater contains the following dissolved ions. Convert the concentration of each from ppm to molarity.
   a. 19,000 ppm Cl^-
   b. 2600 ppm SO_4^{2-}
   c. 1300 ppm Mg^{2+}

35. Seawater contains 2.7 g of NaCl per 100 mL.
   a. What is the concentration of NaCl expressed as a percentage (wt/vol)?
   b. What is the molarity of NaCl in seawater?

36. The molarity of $MgCl_2$ in seawater is 0.54.
   a. How many grams of $MgCl_2$ are there in every 100 mL of seawater?
   b. What is the concentration of $MgCl_2$ expressed as a percentage (wt/vol)?
   c. How many mmol/mL of $MgCl_2$ does seawater contain?

37. Express 0.000056 M as
   a. A millimolar concentration
   b. A micromolar concentration
   c. A nanomolar concentration

38. What is the molarity of a 20.0-L solution that contains 40 g of sodium hydroxide (NaOH)?

39. Waste discharge water from a paper mill is sampled for dioxin, a toxic substance that is formed in the bleaching of paper pulp. Analysis of the waste water from the stream shows a dioxin concentration of 0.010 μg/1.0 mL. Express this concentration in ppb. The EPA standard for dioxin is 1 ppb. Is this waste water in violation of the EPA standard?

40. Which is the higher concentration in each pair?
   a. 100 ppb or 0.05 ppm?
   b. 500 ppb or 250 ppm?
   c. 10 ppb or 1 ppm?

41. If a 100 g water sample contains 1.5 mg of arsenic, what is the concentration of arsenic in the sample in ppm?

42. A laboratory has measured the amount of mercury in a sewage sludge sample to be 0.0005 g per kilogram. What is the concentration of mercury in ppm? In ppb?

43. An ore sample contains 0.0000068 g of gold. What is the concentration of gold in ppm and ppb?

44. Calculate the following:
   a. The molarity of a sample of freshwater that contains 10-ppm NH_4^+
   b. The ppm concentration of calcium in a freshwater sample that contains 5.4×10^{-6} M $CaCO_3$

45. Calculate the following:
   a. The molarity of a sample of freshwater that contains 10 ppm NO_3^-
   b. The ppm concentration of Cu in a freshwater sample that contains 5.4×10^{-4} M $CuCl_2$
   c. The ppm concentration of Cl in a freshwater sample that contains 5.4×10^{-4} M $CuCl_2$

46. What is the main source of carbon for living things? Describe the process by which plants use this source of carbon to make their food.

47. Make a diagram that illustrates one of the following nutrient cycles:
   a. carbon
   b. nitrogen

48. What part do decomposers play in the carbon cycle?

49. What part does each of the following play in the carbon cycle?
   a. The oceans
   b. Limestone
   c. Crude oil

50. What is the name of the class of compounds formed, in addition to oxygen, as products in the photosynthesis reaction?

51. Name the process by which the nitrogen of the atmosphere is converted into a form that can be used by plants.

a. Where are the bacteria that bring about this process found?
b. What chemical compound is produced by these bacteria?

52. Explain the process by which lightning can convert atmospheric nitrogen to a form that plants can use.

53. Why do nitrates not occur as large deposits in many places in the crust of the Earth?

54. List two ways streams and rivers receive excessive amounts of nitrogen-containing compounds because of the activities of people.

55. What is meant by *nitrification*? How is the nitrogen in animal waste recycled by nature?

56. Describe two ways by which a farmer might replenish the nitrogen in her fields without using inorganic fertilizers.

57. What two parts of the oxygen cycle are intertwined with the carbon cycle?

58. List these three salts in order of decreasing solubility.

CuBr	$K_{sp} = 5.2 \times 10^{-9}$
CuCl	$K_{sp} = 1.2 \times 10^{-6}$
CuI	$K_{sp} = 1.1 \times 10^{-12}$

59. Which of the following lead salts will produce the highest concentration of dissolved lead in water? Assume 1 mol of each is placed in 1-L of water.

Lead sulfate	$PbSO_4$	$K_{sp} = 1.6 \times 10^{-8}$
Lead chromate	$PbCrO_4$	$K_{sp} = 1.8 \times 10^{-14}$
Lead sulfide	PbS	$K_{sp} = 7 \times 10^{-28}$

60. Silver ion Ag^+ is an effective disinfectant for the water in swimming pools when it is kept at a concentration of 10 to 100 ppb (ng/L). A silver ion concentration higher than 300 ppb in drinking water, however, is considered unhealthy. One manufacturer of pool disinfectants provides a slightly soluble silver salt in pellet form that releases an equilibrium concentration of Ag^+ to the pool water, which is effective as a disinfectant but is not a danger to the swimmer's health. Which of the following silver salts would be the best candidate for this task?

AgCl	$K_{sp} = 1.8 \times 10^{-10}$
AgBr	$K_{sp} = 5.0 \times 10^{-13}$
AgI	$K_{sp} = 8.3 \times 10^{-17}$

61. What is the primary phosphate species at the following pH?

pH = 4, pH = 10, pH = 13

62. The concentration of arsenic in the soil around an abandoned pest control business is found to be 0.02 mg per 100 g of soil. Express the concentration in ppm and ppb.