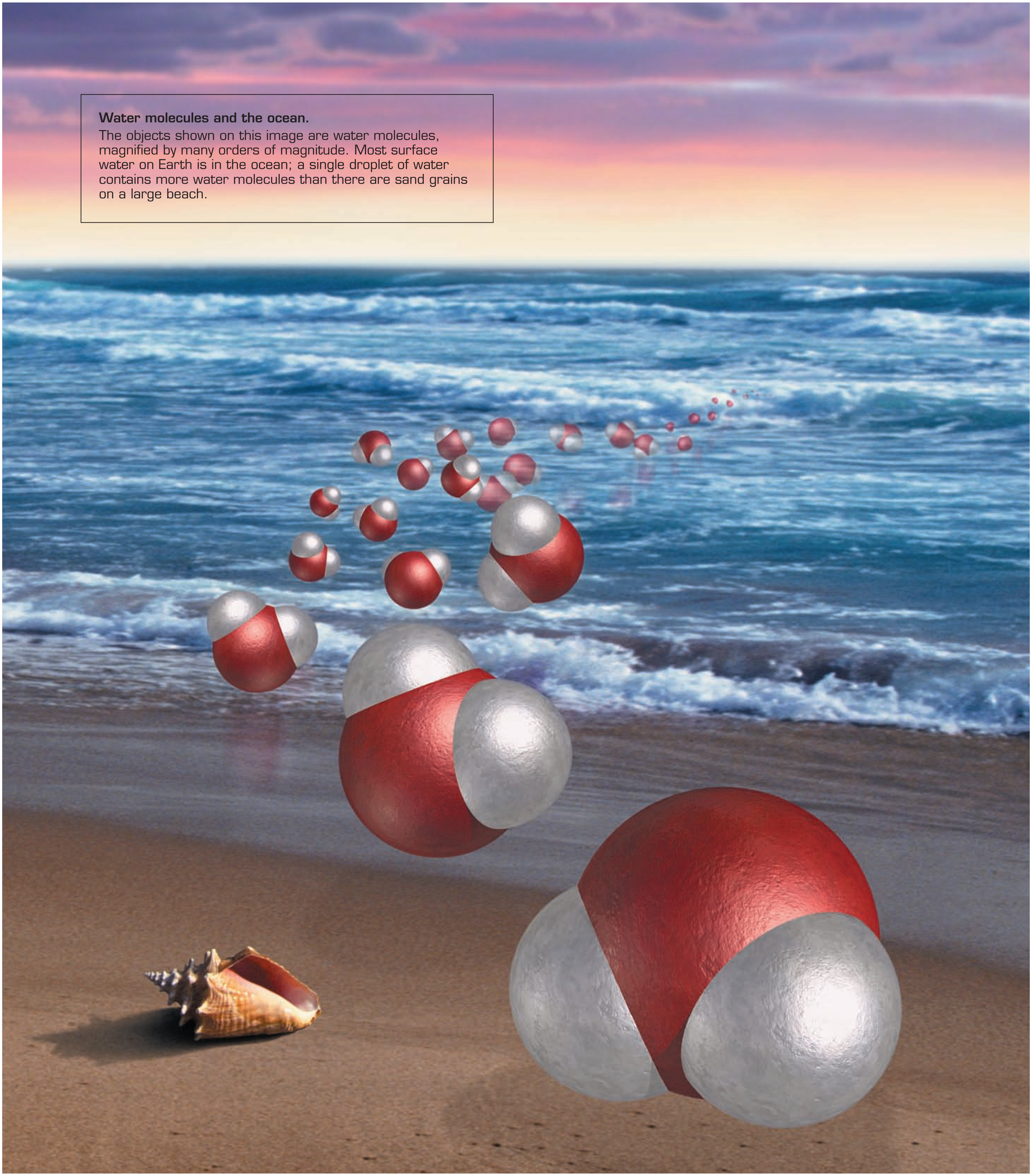

Water molecules and the ocean.
The objects shown on this image are water molecules, magnified by many orders of magnitude. Most surface water on Earth is in the ocean; a single droplet of water contains more water molecules than there are sand grains on a large beach.

> *"Chemistry . . . is one of the broadest branches of science, if for no other reason that, when we think about it, everything is chemistry."*
>
> *—Luciano Caglioti,*
> *The Two Faces of Chemistry (1985)*

5

WATER AND SEAWATER

CHAPTER AT A GLANCE

- The water molecule (H_2O) has an unusual bend in its geometry, with its two hydrogen atoms on the same side of the oxygen atom, giving water many of its unique dissolving and thermal properties.
- Seawater is mostly water molecules but includes a small amount of dissolved substances that gives seawater its characteristic salinity; various surface processes cause salinity to vary.
- The ocean is layered based on density; both salinity and especially temperature affect seawater density.

Why are temperature extremes found at places far from the ocean, while those areas close to the ocean rarely experience severe temperature variations? The mild climates found in coastal regions are made possible by the unique thermal properties of water. These and other properties of water, which stem from the arrangement of its atoms and how its molecules stick together, give water the ability to store vast quantities of heat and to dissolve almost everything.

Water is so common we often take it for granted, yet it is one of the most peculiar substances on Earth. For example, almost every other liquid contracts as it approaches its freezing point, but water actually *expands* as it freezes. Thus water stays at the surface as it starts to freeze, and ice floats—a rare property shared by very few other substances. If its nature were otherwise, all temperate-zone lakes, ponds, rivers, and even oceans would eventually freeze solid from the bottom up, and life as we know it could not exist. Instead, a floating skin of ice cocoons life in the liquid water beneath a layer of insulation, enabling it to persist under the frozen surface.

The chemical properties of water are also essential for sustaining all forms of life. In fact, the primary component of all living organisms is water. The water content of organisms, for instance, ranges from about 65% (humans) to 95% (most plants) to 99% in some jellyfish. Water is the ideal medium to have within our bodies because it facilitates chemical reactions. Our blood, which serves to transport nutrients and remove wastes within our bodies, is 83% water. The very presence of water on our planet makes life possible, and its remarkable properties make our planet livable.

5.1 Why Does Water Have Such Unusual Chemical Properties?

To understand why water has such unusual properties, let's examine its chemical structure.

Atomic Structure

Atoms (*a* = not, *tomos* = cut) are the basic building blocks of all matter. Every physical substance in our world—chairs, tables, books, people, the air we breathe—is composed of atoms. An atom resembles a microscopic sphere (Figure 5.1) and was originally thought to be the smallest form of matter. Additional study has revealed that atoms are composed of even smaller particles, called subatomic particles.[1] As shown in Figure 5.1, the **nucleus** (*nucleos* = a little nut) of an atom is composed

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[1] It has been discovered that subatomic particles themselves are composed of a variety of even smaller particles such as *quarks*, *leptons*, and *bosons*.

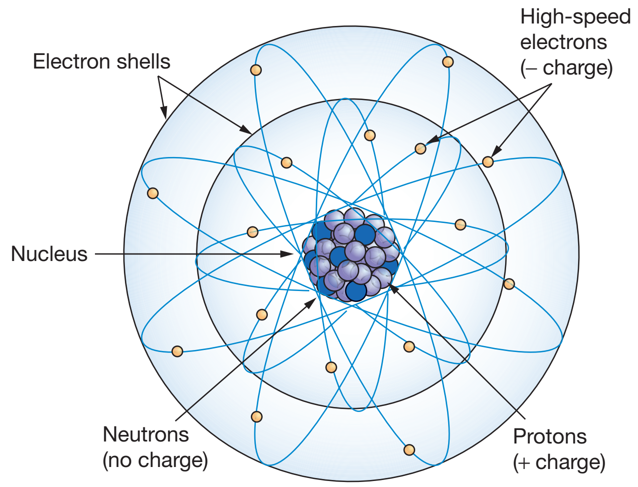


FIGURE 5.1 Simplified model of an atom. An atom consists of a central nucleus composed of protons and neutrons that is encircled by electrons.

of **protons** (*protos* = first) and **neutrons** (*neutr* = neutral) which are bound together by strong forces. Protons have a positive electrical charge, whereas neutrons have no electrical charge. Both protons and neutrons have about the same mass, which is extremely small. Surrounding the nucleus are particles called **electrons** (*electro* = electricity) which have about $^1\!/_{2000}$ the mass of either protons or neutrons. Electrical attraction between positively charged protons and negatively charged electrons holds electrons in layers or shells around the nucleus.

The overall electrical charge of most atoms is balanced because each atom contains an equal number of protons and electrons. An oxygen atom, for example, has eight protons and eight electrons. Most oxygen atoms also have eight neutrons, which do not affect the overall electrical charge because neutrons are electrically neutral. The number of protons is what distinguishes atoms of the 118 known chemical elements from one another. For example, an oxygen atom (and only an oxygen atom) has eight protons. Similarly, a hydrogen atom (and only a hydrogen atom) has one proton, a helium atom has two protons, and so on (for more details, see Appendix IV, "A Chemical Background: Why Water has 2 H's and 1 O"). In some cases, an atom will lose or gain one or more electrons and thus have an overall electrical charge. These atoms are called **ions** (*ienai* = to go).

The Water Molecule

A **molecule** (*molecula* = a mass) is a group of two or more atoms held together by mutually shared electrons. It is the smallest form of a substance that can exist yet still retain the original properties of that substance. When atoms combine with other atoms to form molecules, they share or trade electrons and establish chemical bonds. For instance, the chemical formula for water—H_2O—indicates that a water molecule is composed of two hydrogen atoms chemically bonded to one oxygen atom.

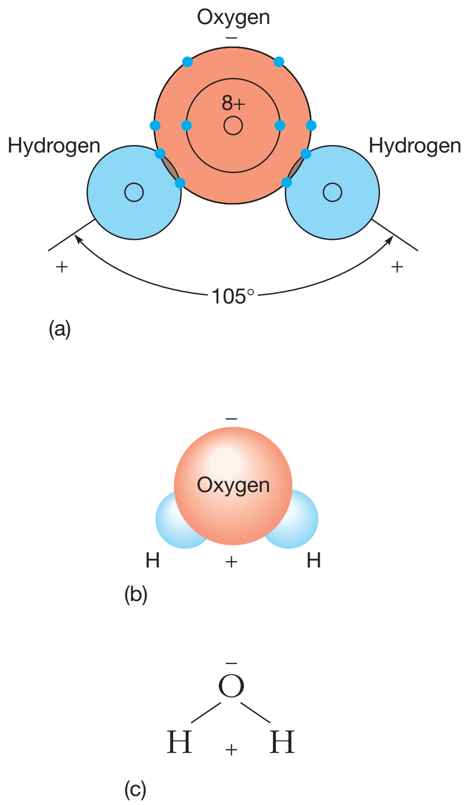


FIGURE 5.2 The water molecule. **(a)** Geometry of a water molecule. The oxygen end of the molecule is negatively charged, and the hydrogen regions exhibit a positive charge. Covalent bonds occur between the oxygen and the two hydrogen atoms. **(b)** A three-dimensional representation of the water molecule. **(c)** The water molecule represented by letters (*H* = hydrogen, *O* = oxygen).

GEOMETRY Atoms can be represented as spheres of various sizes, and the more electrons the atom contains, the larger the sphere. It turns out that an oxygen atom (with eight electrons) is about twice the size of a hydrogen atom (with one electron). A water molecule consists of a central oxygen atom covalently bonded to the two hydrogen atoms, which are separated by an angle of about 105 degrees (Figure 5.2a). The **covalent** (*co* = with, *valere* = to be strong) **bonds** in a water molecule are due to the sharing of electrons between oxygen and each hydrogen atom. They are relatively strong chemical bonds, so a lot of energy is needed to break them.

Figure 5.2b shows a water molecule in a more compact representation, and in Figure 5.2c letter symbols are used to represent the atoms in water (*O* for oxygen, *H* for hydrogen). Instead of water's atoms being in a straight line, *both hydrogen atoms are on the same side of the oxygen atom*. This curious bend in the geometry of the water molecule is the underlying cause of most of the unique properties of water.

POLARITY The bent geometry of the water molecule gives a slight overall negative charge to the side of the oxygen atom and a slight overall positive charge to the side of the hydrogen atoms (Figure 5.2a). This slight separation of charges gives the entire molecule an electrical **polarity** (*polus* = pole, *ity* = having the quality of) so water molecules are **dipolar** (*di* = two, *polus* =pole) other common dipolar objects are flashlight batteries, car batteries, and bar magnets. Although the electrical charges are weak, water molecules behave as if they contain a tiny bar magnet.

INTERCONNECTIONS OF MOLECULES If you've ever experimented with bar magnets, you know they have polarity and orient themselves relative to one another such that the positive end of one bar magnet is attracted to the negative

STUDENTS SOMETIMES ASK . . .

Why does a water molecule have the unusual shape that it does?

Based on simple symmetry considerations and charge separations, a water molecule should have its two hydrogen atoms on opposite sides of the oxygen atom, thus producing a linear shape like many other molecules. But water's odd shape where *both* hydrogen atoms are on the same side of the oxygen atom stems from the fact that oxygen has four bonding sites, which are evenly spaced around the oxygen atom. No matter which two bonding sites are occupied by hydrogen atoms, it results in the curious bend of the water molecule.

end of another. Water molecules have polarity, too, so they orient themselves relative to one another. In water, the positively charged hydrogen area of one water molecule interacts with the negatively charged oxygen end of an adjacent water molecule, forming a **hydrogen bond** (Figure 5.3). The hydrogen bonds between water molecules are much weaker than the covalent bonds that hold individual water molecules together. In essence, weaker hydrogen bonds form *between* adjacent water molecules and stronger covalent bonds occur *within* water molecules.

Even though hydrogen bonds are weaker than covalent bonds, they are strong enough to cause water molecules to stick to one another and exhibit **cohesion** (*cohaesus* = a clinging together). The cohesive properties of water cause it to "bead up" on a waxed surface, such as a freshly waxed car. They also give water its **surface tension**. Water's surface has a thin "skin" that allows a glass to be filled just above the brim without spilling any of the water. Surface tension results from the formation of hydrogen bonds between the outermost layer of water molecules and the underlying molecules. Water's ability to form hydrogen bonds causes it to have the highest surface tension of any liquid except the element mercury.[2]

WATER: THE UNIVERSAL SOLVENT Water molecules stick not only to other water molecules, but also to other polar chemical compounds. In doing so, water molecules can reduce the attraction between ions of opposite charges by as much as 80 times. For instance, ordinary table salt—sodium[3] chloride, NaCl—consists of an alternating array of positively charged sodium ions and negatively charged chloride ions (Figure 5.4a). The **electrostatic** (*electro* = electricity, *stasis* = standing) **attraction** between oppositely charged ions produces an **ionic** (*ienai* = to go) **bond**. When solid NaCl is placed in water, the electrostatic attraction (ionic bonding) between the sodium and chloride ions is reduced by 80 times. This, in turn, makes it much easier for the sodium ions and chloride ions to separate. When the ions separate, the positively charged sodium ions become attracted to the negative ends of the water molecules, the negatively charged chloride ions become attracted to the positive ends of the water molecules (Figure 5.4b), and the salt is dissolved in water. The process by which water molecules completely surround ions is called *hydration* (*hydra* = water, *ation* = action or process).

Because water molecules interact with other water molecules and other polar molecules, water is able to dissolve nearly everything.[4] Given enough time, water

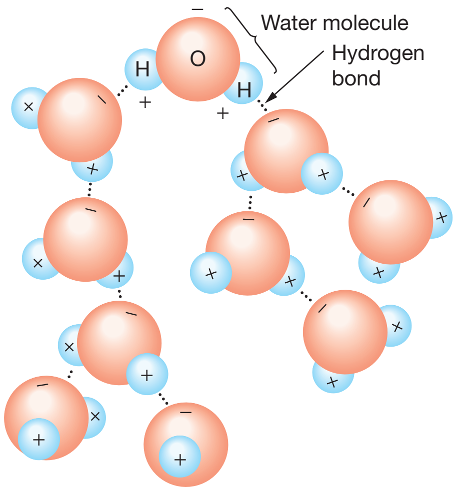
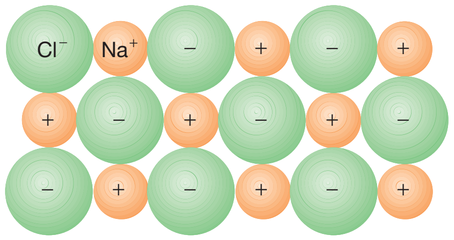


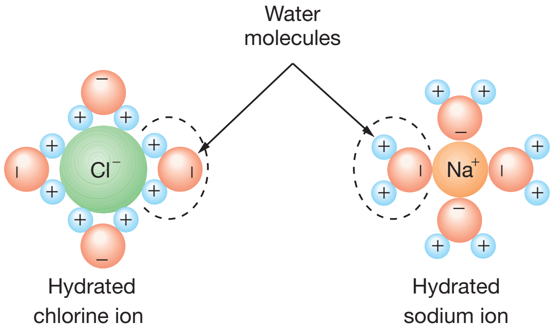
FIGURE 5.3 Hydrogen bonding in water. Dashed lines indicate locations of hydrogen bonds, which occur between water molecules.

KEY CONCEPT

A water molecule has a bend in its geometry, with the two hydrogen atoms on the same side of the oxygen atom. This property gives water its polarity and ability to form hydrogen bonds.



(a) Sodium chloride, solid crystal structure



(b) Sodium chloride, in solution

FIGURE 5.4 Water as a solvent. (a) Atomic structure of table salt, which is composed of sodium chloride (Na^+ = sodium ion, Cl^- = chlorine ion). **(b)** As sodium chloride is dissolved, the positively charged ends of water molecules are attracted to the negatively charged Cl^- ion, while the negatively charged ends are attracted to the positively charged Na^+ ion.

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[2] Mercury is the only metal that is a liquid at normal surface temperatures. Although it is commonly used in thermometers, it is also quite toxic.

[3] Sodium is represented by the letters Na because the Latin term for sodium is *natrium*.

[4] If water is such a good solvent, why doesn't oil dissolve in water? As you might have guessed, the chemical structure of oil is remarkably nonpolar. With no positive or negative ends to attract the polar water molecule, oil will not dissolve in water.

STUDENTS SOMETIMES ASK...

How can it be that water—a liquid at room temperature—can be created by combining hydrogen and oxygen—two gases at room temperature?

It is true that combining two parts hydrogen gas with one part oxygen gas produces liquid water. This can be accomplished as a chemistry experiment, although care should be taken because much energy is released during the reaction (don't try this at home!). Oftentimes, when combining two elements, the product has very different properties than the pure substances. For instance, combining elemental sodium (Na), a highly reactive metal, with pure chlorine (Cl_2) a toxic nerve gas, produces cubes of harmless table salt (NaCl). This is what most people find amazing about chemistry.

can dissolve more substances and in greater quantity than any other known substance. This is why water is called "the universal solvent." It is also why the ocean contains so much dissolved material—an estimated 50 quadrillion tons (50 million billion tons) of salt—which makes seawater taste "salty."

5.2 What Other Important Properties Does Water Possess?

Water's other important properties include its thermal properties (such as water's freezing and boiling points, heat capacity, and latent heats) and its density.

Water's Thermal Properties

Water exists on Earth as a solid, a liquid, and a gas and has the ability to store and release great amounts of heat. Water's thermal properties influence the world's heat budget and are in part responsible for the development of tropical cyclones, worldwide wind belts, and ocean surface currents.

HEAT, TEMPERATURE, AND CHANGES OF STATE Matter around us is usually in one of the three common states: solid, liquid, or gas.[5] What must happen to change the state of a compound? The attractive forces between molecules or ions in the substance must be overcome if the state of the substance is to be changed from solid to liquid or from liquid to gas. These attractive forces include hydrogen bonds and van der Waals forces. The **van der Waals forces**—named for Dutch physicist Johannes Diderik van der Waals (1837–1923)—are relatively weak interactions that become significant only when molecules are very close together, as in the solid and liquid states (but not the gaseous state). Energy must be added to the molecules or ions so they can move fast enough to overcome these attractions.

What form of energy changes the state of matter? Very simply, adding or removing heat is what causes a substance to change its state of matter. For instance, adding heat to ice cubes causes them to melt and removing heat from water causes ice to form. Before proceeding, let's clarify the difference between heat and temperature:

- **Heat** is the *energy transfer from one body to another due to a difference in temperature*. Heat is proportional to the energy level of moving molecules and thus is the total internal energy—both **kinetic** (*kinetos* = moving) **energy** and **potential** (*potentia* = power) **energy**—transferred from one body to another. For example, water can exist as a solid, liquid, or gas depending on the amount of heat added. Heat may be generated by combustion (a chemical reaction commonly called "burning"), through other chemical reactions, by friction, or from radioactivity; it can be transferred by conduction, by convection, or by radiation. A **calorie** (*calor* = heat) is the amount of heat required to raise the temperature of 1 gram of water[6] by 1 degree centigrade. The familiar "calories" used to measure the energy content of foods is actually a *kilocalorie*, or 1000 calories. Although the metric unit for thermal energy is the *joule*, calories are directly tied to some of water's thermal properties, as will be discussed in the next section.

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[5] *Plasma* is widely recognized as a fourth state of matter distinct from solids, liquids, and normal gases. Plasma is a gaseous substance in which atoms have been ionized—that is to say, stripped of electrons. Plasma television monitors take advantage of the fact that plasmas are strongly influenced by electric currents.

[6] One gram (0.035 ounce) of water is equal to about 10 drops.

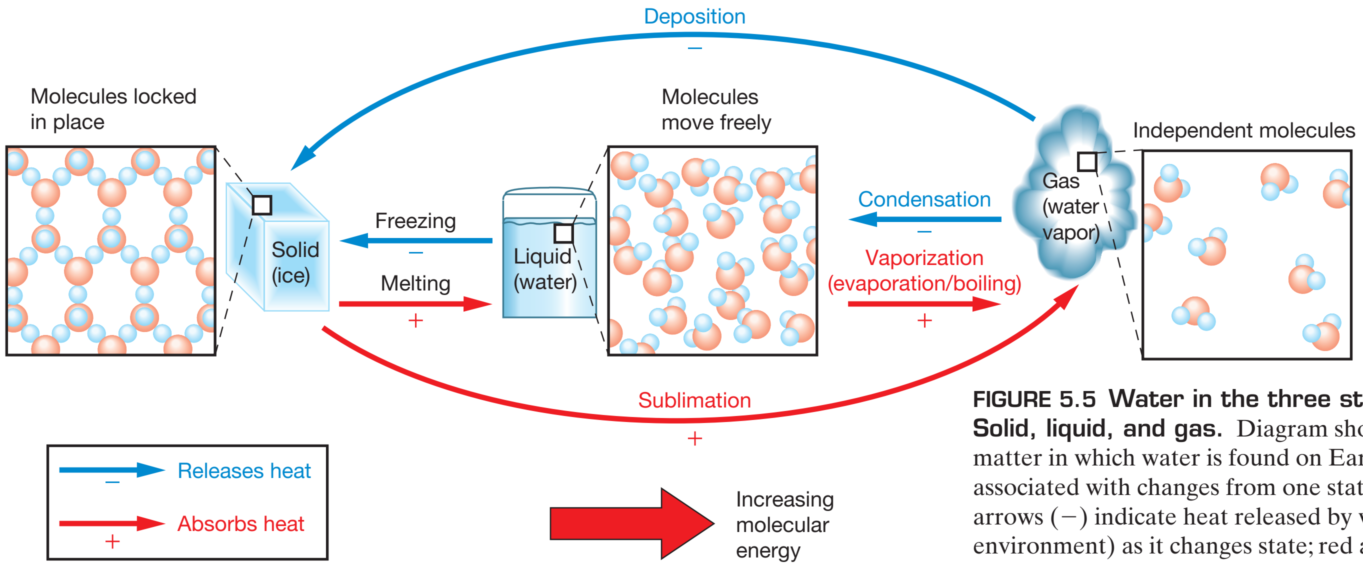

FIGURE 5.5 Water in the three states of matter: Solid, liquid, and gas. Diagram showing the three states of matter in which water is found on Earth and the processes associated with changes from one state to another. Blue arrows (−) indicate heat released by water (which warms the environment) as it changes state; red arrows (+) indicate heat absorbed by water (which cools the environment).

- **Temperature** is the *direct measure of the average kinetic energy of the molecules that make up a substance*. The greater the temperature, the greater the kinetic energy of the substance. Temperature changes when heat energy is added to or removed from a substance. Temperature is usually measured in degrees centigrade (°C) or degrees Fahrenheit (°F).

Figure 5.5 shows water molecules in the solid, liquid, and gaseous states. In the *solid state* (ice), water has a rigid structure and does not normally flow over short time scales. Intermolecular bonds are constantly being broken and reformed, but the molecules remain firmly attached. That is, the molecules vibrate with energy but remain in relatively fixed positions. As a result, solids do not conform to the shape of their container.

In the *liquid state* (water), water molecules still interact with each other, but they have enough kinetic energy to flow past each other and take the shape of their container. Intermolecular bonds are being formed and broken at a much greater rate than in the solid state.

In the *gaseous state* (water **vapor**), water molecules no longer interact with one another except during random collisions. Water vapor molecules flow very freely, filling the volume of whatever container they are placed in.

WATER'S FREEZING AND BOILING POINTS If enough heat energy is added to a solid, it melts to a liquid. The temperature at which melting occurs is the substance's **melting point**. If enough heat energy is removed from a liquid, it freezes to a solid. The temperature at which freezing occurs is the substance's **freezing point**, which is the same temperature as the melting point (Figure 5.5). For pure water, melting and freezing occur at 0°C (32°F).[7]

If enough heat energy is added to a liquid, it converts to a gas. The temperature at which boiling occurs is the substance's **boiling point**. If enough heat energy is removed from a gas, it **condenses** to a liquid. The highest temperature at which condensation occurs is the substance's **condensation point**, which is the same temperature as the boiling point (Figure 5.5). For pure water, boiling and condensation occur at 100°C (212°F).

Both the freezing and boiling points of water are unusually high compared to other similar substances. As shown in Figure 5.6, if water followed the pattern of

web animation

Phase Changes of Water

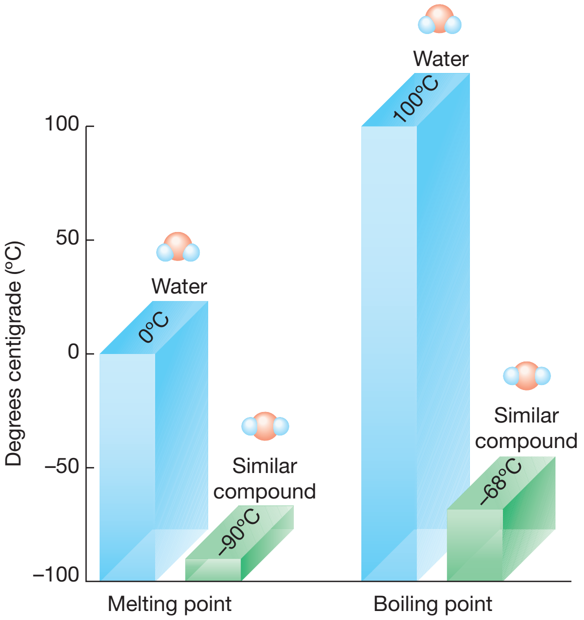

FIGURE 5.6 Comparison of melting and boiling points of water with similar chemical compounds. Bar graph showing the melting and boiling points of water compared to the melting and boiling points of similar chemical compounds. Water would have properties like those of similar chemical compounds if it did not have its unique geometry and resulting polarity.

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[7] All melting/freezing/boiling points discussed in this chapter assume a standard sea level pressure of 1 atmosphere (14.7 pounds per square inch).

other chemical compounds with molecules of similar mass, it should melt at −90°C (−130°F) and boil at −68°C (−90°F). If that were the case, all water on Earth would be in the gaseous state. Instead, water melts and boils at the relatively high temperatures of 0°C (32°F) and 100°C (212°F),[8] respectively, because additional heat energy is required to overcome its hydrogen bonds and van der Waals forces. Thus, if not for the unusual geometry and resulting polarity of the water molecule, all water on Earth would be boiled away and life as we know it would not exist.

WATER'S HEAT CAPACITY AND SPECIFIC HEAT **Heat capacity** is *the amount of heat energy required to raise the temperature of a substance by 1 degree centigrade*. Substances that have high heat capacity can absorb (or lose) large quantities of heat with only a small change in temperature. Conversely, substances that change temperature rapidly when heat is applied—such as oil or metals—have lower heat capacity.

The heat capacity per unit mass of a body, called *specific heat capacity* or more simply **specific heat**, is used to more directly compare the heat capacity of substances. For example, as shown in Figure 5.7, pure water has a high specific heat capacity that is exactly 1 calorie per gram,[9] whereas other common substances have much lower specific heats. Note that metals such as iron and copper—which heat up rapidly when heat is applied—have capacity values that are about 10 times lower than water.

Why does water have such high heat capacity? The reason is because it takes more energy to increase the kinetic energy of hydrogen-bonded water molecules than it does for substances in which the dominant intermolecular interaction is the much weaker van der Waals force. As a result, water gains or loses much more heat than other common substances while undergoing an equal temperature change. In addition, water resists any change in temperature, as you may have observed when heating a large pot of water. When heat is applied to the pot, which is made of metal that has a low heat capacity, the pot heats up quickly. The water *inside* the pot, however, takes a long time to heat up (hence, the tale that a watched pot never boils but an unwatched pot boils over!). Making the water boil takes even more heat because all the hydrogen bonds must be broken. The exceptional capacity of water to absorb large quantities of heat helps explain why water is used in home heating, industrial and automobile cooling systems, and home cooking applications.

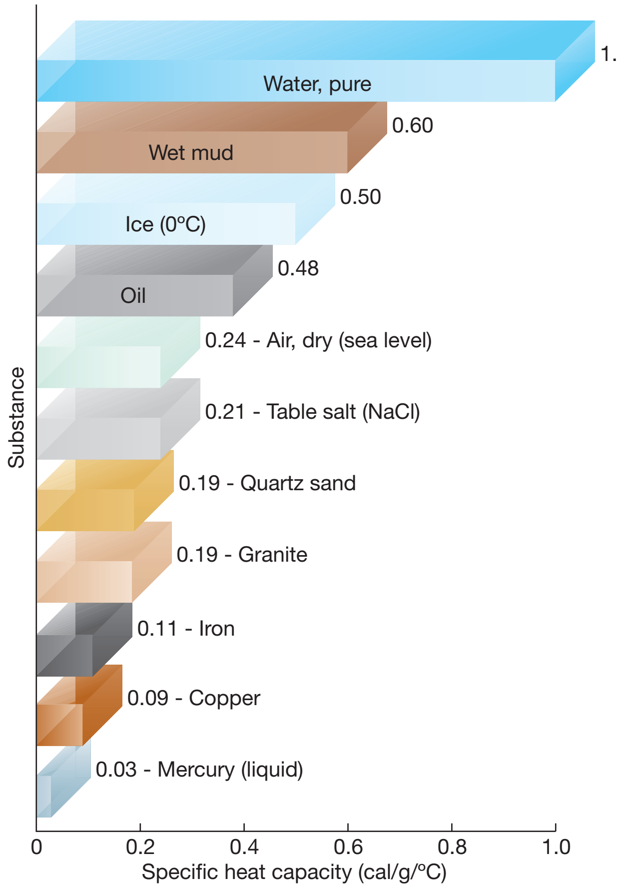

FIGURE 5.7 Specific heat capacity of common substances. Graph showing the specific heat capacity of common substances at 20°C. Note that water has a very high specific heat capacity, which means that it takes a lot of energy to increase water temperature.

WATER'S LATENT HEATS When water undergoes a change of state—that is, when ice melts or water freezes, or when water boils or water vapor condenses—a large amount of heat is absorbed or released. The amount of heat absorbed or released is due to water's high latent (*latent* = hidden) heats and is closely related to water's unusually high heat capacity. As water evaporates from your skin, it cools your body by absorbing heat (this is why sweating cools your body). Conversely, if you ever have been scalded by water vapor—steam—you know that steam releases an enormous amount of latent heat when it condenses to a liquid.

Latent Heat of Melting The graph in Figure 5.8 shows how latent heat affects the amount of energy needed to increase

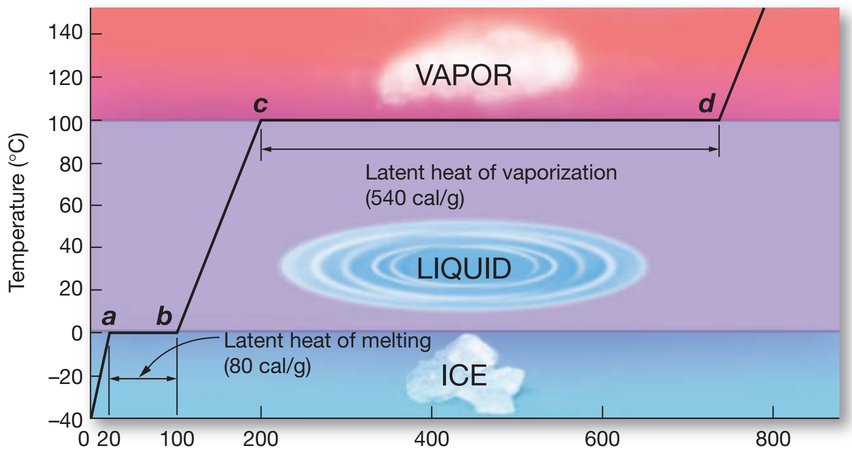

FIGURE 5.8 Latent heats and changes of state of water. The latent heat of melting (80 calories per gram) is much less than the latent heat of vaporization (540 calories per gram). See text for description of points *a*, *b*, *c*, and *d*.

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[8] Note that the temperature scale *centigrade* (*centi* = a hundred, *grad* = step) is based on 100 even divisions between the melting and boiling points of pure water. It is also called the Celsius scale after its founder (see Appendix I, "Metric and English Units Compared").

[9] Note that the specific heat capacity of water is used as the unit of heat quantity, the calorie. Thus, water is the standard against which the specific heats of other substances are compared.

water temperature and change the state of water. Beginning with 1 gram of ice (*lower left*), the addition of 20 calories of heat raises the temperature of the ice by 40 degrees, from −40°C to 0°C (point *a* on the graph). The temperature remains at 0°C (32°F) even though more heat is being added, as shown by the plateau on the graph between points *a* and *b*. The temperature of the water does not change until 80 more calories of heat energy have been added. The **latent heat of melting** is the energy needed to break the intermolecular bonds that hold water molecules rigidly in place in ice crystals. The temperature remains unchanged until most of the bonds are broken and the mixture of ice and water has changed completely to 1 gram of water.

After the change from ice to liquid water has occurred at 0°C, additional heat raises the water temperature between points *b* and *c* in Figure 5.8. As it does, it takes 1 calorie of heat to raise the temperature of the gram of water 1°C (or 1.8°F). Therefore, another 100 calories must be added before the gram of water reaches the boiling point of 100°C (212°F). So far, a total of 200 calories has been added to reach point *c*.

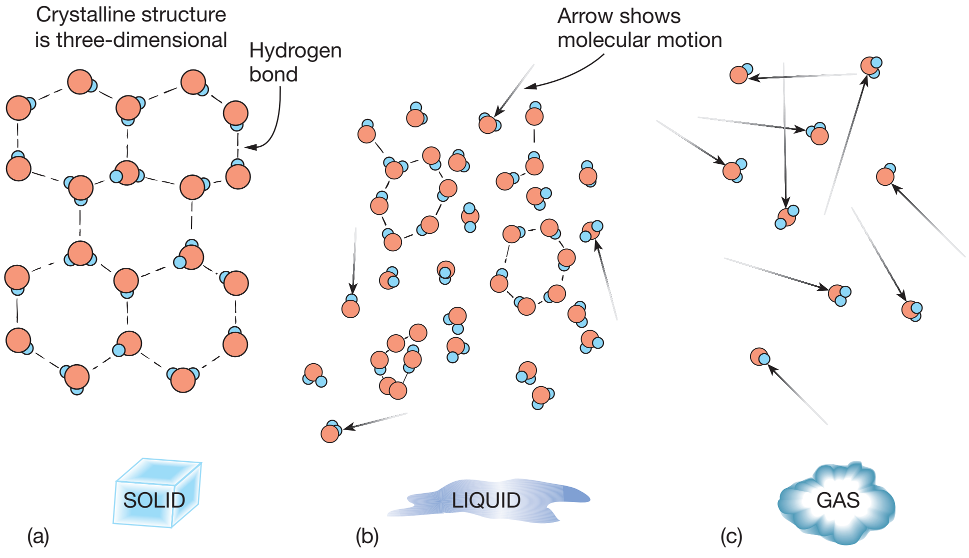

FIGURE 5.9 Hydrogen bonds in H_2O and the three states of matter. **(a)** In the solid state, water exists as ice, in which there are hydrogen bonds between all water molecules. **(b)** In the liquid state, there are some hydrogen bonds. **(c)** In the gaseous state, there are no hydrogen bonds and the water molecules are moving rapidly and independently.

Latent Heat of Vaporization The graph in Figure 5.8 flattens out again at 100°C, between points *c* and *d*. This plateau represents the **latent heat of vaporization**, which is 540 calories per gram for water. This is the amount of heat that must be added to 1 gram of a substance at its boiling point to break the intermolecular bonds and complete the change of state from liquid to vapor (gas).

The drawings in Figure 5.9, which show the structure of water molecules in the solid, liquid, and gaseous states, help explain why the latent heat of vaporization is so much greater than the latent heat of melting. To go from a solid to a liquid, just enough hydrogen bonds must be broken to allow water molecules to slide past one another. To go from a liquid to a gas, however, all of the hydrogen bonds must be completely broken so that individual water molecules can move about freely.

Latent Heat of Evaporation Sea surface temperatures average 20°C (68°F) or less. How, then, does liquid water convert to vapor at the surface of the ocean? the conversion of a liquid to a gas below the boiling point is called **evaporation**. At ocean surface temperatures, individual molecules converted from the liquid to the gaseous state have less energy than do water molecules at 100°C. To gain the additional energy necessary to break free of the surrounding ocean water molecules, an individual molecule must capture heat energy from its neighbors. In other words, the molecules left behind have lost heat energy to those that evaporate, which explains the cooling effect of evaporation.

It takes more than 540 calories of heat to produce 1 gram of water vapor from the ocean surface at temperatures less than 100°C. At 20°C (68°F), for instance, the **latent heat of evaporation** is 585 calories per gram. More heat is required because more hydrogen bonds must be broken. At higher temperatures, liquid water has fewer hydrogen bonds because the molecules are vibrating and jostling about more.

Latent Heat of Condensation When water vapor is cooled sufficiently, it condenses to a liquid and releases its **latent heat of condensation** into the surrounding air. On a small scale, the heat released is enough to cook food; this is how a “steamer” works. On a large scale, the heat released is sufficient to power large thunderstorms and even hurricanes (see Chapter 6, “Air–Sea Interaction”).

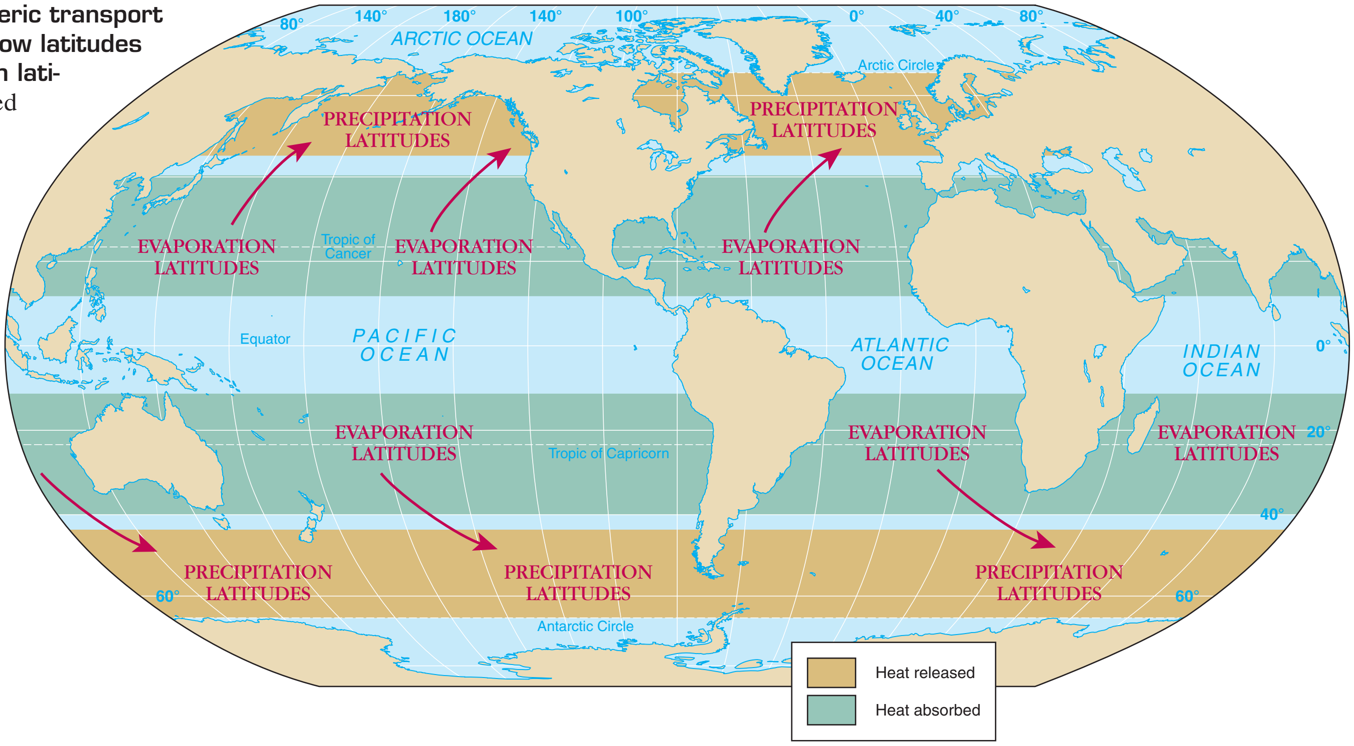

FIGURE 5.10 Atmospheric transport of surplus heat from low latitudes into heat-deficient high latitudes. The heat removed from the tropical ocean (*evaporation latitudes*) is carried toward the poles (*red arrows*) and is released at higher latitudes through precipitation (*precipitation latitudes*), thus moderating Earth's climate.

Latent Heat of Freezing Heat is also released when water freezes. The amount of heat released when water freezes is the same amount that was absorbed when the water was melted in the first place. Thus, the **latent heat of freezing** is identical to the latent heat of melting. Similarly, the latent heats of vaporization and condensation are identical.

GLOBAL THERMOSTATIC EFFECTS The **thermostatic** (*thermos* = heat, *stasis* = standing) **effects** of water are those properties that act to moderate changes in temperature, which in turn affect Earth's climate. For example, the huge amount of heat energy exchanged in the evaporation–condensation cycle helps make life possible on Earth. The Sun radiates energy to Earth, where some is stored in the oceans. Evaporation removes this heat energy from the oceans and carries it high into the atmosphere. In the cooler upper atmosphere, water vapor condenses into clouds, which are the basis of **precipitation** (mostly rain and snow) that releases latent heat of condensation. The map in Figure 5.10 shows how this cycle of evaporation and condensation removes huge amounts of heat energy from the low-latitude oceans and adds huge amounts of heat energy to the heat-deficient higher latitudes. In addition, the heat released when sea ice forms further moderates Earth's high-latitude regions.

The exchange of latent heat between ocean and atmosphere is very efficient. For every gram of water that condenses in cooler latitudes, the amount of heat released to warm these regions equals the amount of heat removed from the tropical ocean when that gram of water was evaporated initially. The end result is that the thermal properties of water have prevented wide variations in Earth's temperature, thus moderating Earth's climate. Because rapid change is the enemy of all life, our planet's moderated climate is one of the main reasons life exists on Earth.

Another thermostatic effect of the ocean can be seen in Figure 5.11, which shows the temperature difference between day and night. The map shows that in the ocean, there is only a small difference in temperature between day and night while the land experiences a much greater variation. This difference between ocean and land is due to the higher heat capacity of water, which gives it the ability to absorb the daily gains and minimize the daily losses of heat energy much more

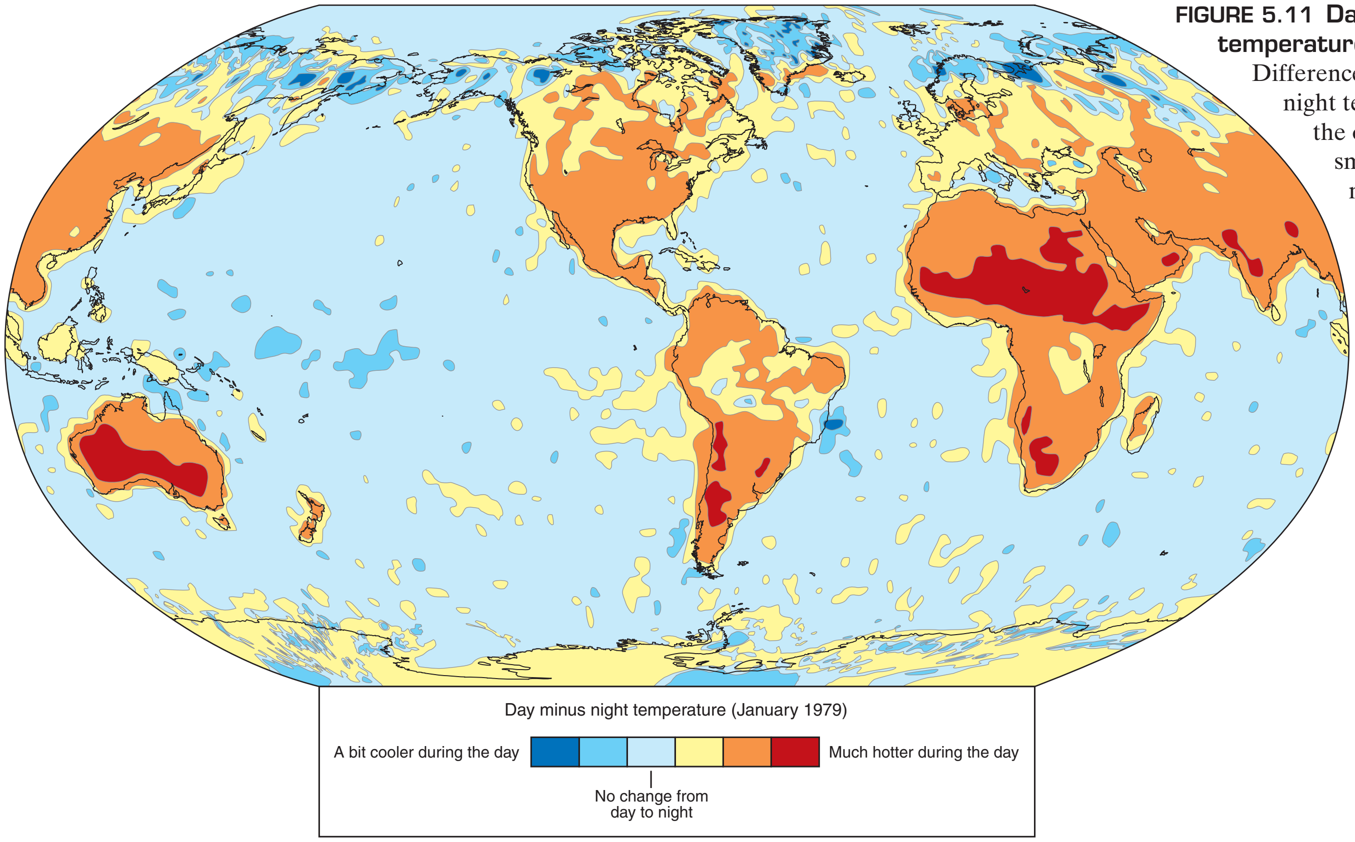

FIGURE 5.11 Day minus night temperature difference. Differences between day and night temperature show that the ocean experiences only a small change while continental areas experience a much larger change. Data based on satellite measurement of the average difference in temperature from 2 P.M. to 2 A.M. during January 1979.

easily than land materials. The term **marine effect** describes locations that experience the moderating influences of the ocean, usually along coastlines or islands. **Continental effect**—a condition of *continentality*—refers to areas less affected by the sea and therefore having a greater range of temperature differences (both daily and yearly).

KEY CONCEPT

Water's unique thermal properties include water's latent heats and high heat capacity, which redistribute heat on Earth and have moderated Earth's climate.

Water Density

Recall from Chapter 1 that density is mass per unit volume and can be thought of as *how heavy something is for its size*. Ultimately, density is related to how tightly the molecules or ions of a substance are packed together. Typical units of density are grams per cubic centimeter (g/cm^3). Pure water, for example, has a density of 1.0 g/cm^3. Temperature, salinity, and pressure all affect water density.

The density of most substances increases as temperature decreases. For example, cold air sinks and warm air rises because cold air is denser than warm air. Density increases as temperature decreases because the molecules lose energy and slow down, so the same number of molecules occupy less space. This shrinkage caused by cold temperatures, called **thermal contraction**, also occurs in water, but only to a certain point. As pure water cools to 4°C (39°F), its density increases. From 4°C down to 0°C (32°F), however, its density *decreases*. In other words, water stops contracting and actually expands, which is highly unusual among Earth's many substances. The result is that ice is less dense than liquid water, so ice floats on water. For most other substances, the solid state is denser than the liquid state, so the solid sinks.

Why is ice less dense than water? Figure 5.12 shows how molecular packing changes as water approaches its freezing point. From points *a* to *c* in the figure, the temperature decreases from 20°C (68°F) to 4°C (39°F) and the density increases from 0.9982 g/cm^3 to 1.0000 g/cm^3. Density increases because the

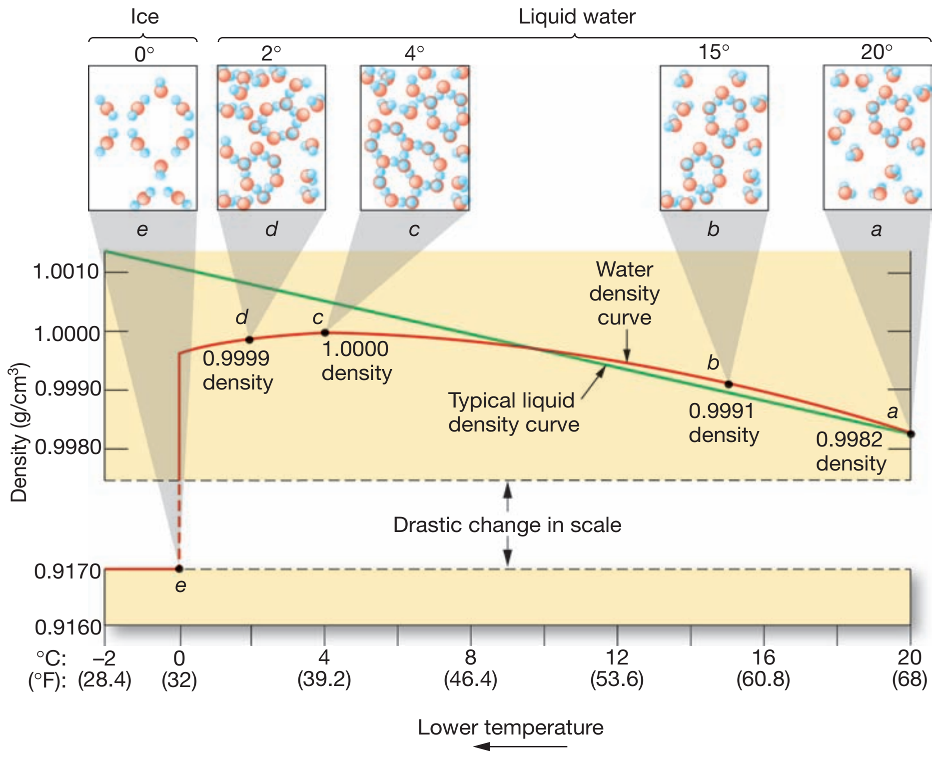

FIGURE 5.12 Water density as a function of temperature and the formation of ice. The density of freshwater (*red curve*) as it freezes (*right to left*); the density of a typical liquid is also shown (*green curve*). Water reaches its maximum density at 4°C, but below that water becomes less dense as ice begins to form. At 0°C, ice forms, its crystal structure expands dramatically, and density decreases. As a result, ice floats.

amount of thermal motion decreases, so the water molecules occupy less volume. As a result, the window at point *c* contains more water molecules than the windows at points *a* or *b*. When the temperature is lowered below 4°C (39°F), the overall volume increases again because water molecules begin to line up to form ice crystals. Ice crystals are bulky, open, six-sided structures in which water molecules are widely spaced. Their characteristic hexagonal shape (Figure 5.13) mimics the hexagonal molecular structure resulting from hydrogen bonding between water molecules (see Figure 5.9a). By the time water fully freezes (point *e*), the density of the ice is much less than that of water at 4°C (39°F), the temperature at which water achieves its maximum density.

When water freezes, its volume increases by about 9%. Anyone who has put a beverage in a freezer for "just a few minutes" to cool it down and inadvertently forgotten about it has experienced the volume increase associated with water's expansion as it freezes—usually resulting in a burst beverage container (Figure 5.14). The force exerted when ice expands is powerful enough to break apart rocks, split pavement on roads and sidewalks, and crack water pipes.

Increasing the pressure or adding dissolved substances decreases the temperature of maximum density for freshwater because the formation of bulky ice crystals is inhibited. Increasing pressure increases the number of water molecules in a given volume and inhibits the number of ice crystals that can be created. Increasing amounts of dissolved substances inhibits the formation of hydrogen bonds, which further restricts the number of ice crystals that can form. To produce ice crystals equal in volume to those that could be produced at 4°C (39°F) in freshwater, more energy must be removed, causing a reduction in the temperature of maximum density.

Dissolved solids reduce the freezing point of water, too. It's one of the reasons why most seawater never freezes, except near Earth's frigid poles (and even then, only at the surface). It's also why salt is spread on roads and sidewalks during the winter in cold climates. The salt lowers the freezing point of water, allowing ice-free roads and sidewalks at temperatures that are several degrees below freezing.

Web Table 5.1 summarizes the physical and biological significance of the unusual properties of seawater.

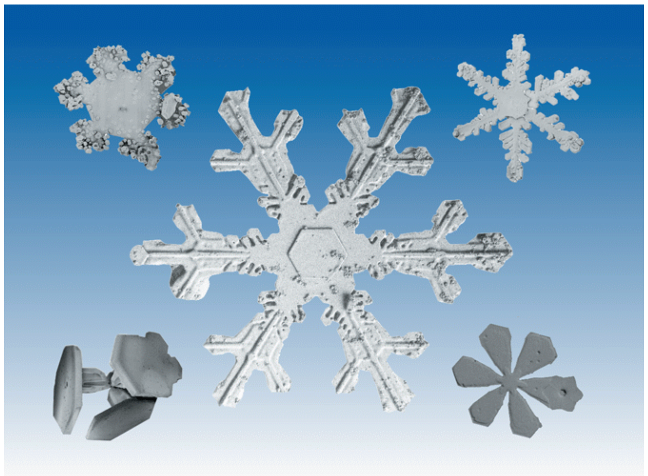

FIGURE 5.13 Snowflakes. Scanning electron photomicrograph of actual snowflakes magnified about 500 times. Hexagonal snowflakes indicate the internal structure of water molecules held together by hydrogen bonds.

5.3 How Salty Is Seawater?

What is the difference between pure water and seawater? One of the most obvious differences is that seawater contains dissolved substances that give it a distinctly salty taste. These dissolved substances are not simply sodium chloride (table salt)—they include various other salts, metals, and dissolved gases. The oceans, in fact, contain enough salt to cover the entire planet with a layer more than 150 meters (500 feet) thick (about the height of a 50-story skyscraper). Unfortunately, the salt content of seawater makes it unsuitable for drinking or irrigating most crops and causes it to be highly corrosive to many materials.

Salinity

Salinity (*salinus* = salt) is the total amount of solid material dissolved in water including dissolved gases (because even gases become solids at low enough temperatures) but excluding dissolved organic substances. Salinity does *not* include fine particles being held in suspension (turbidity) or solid material in contact with water because these materials are not dissolved. Salinity is the ratio of the mass of dissolved substances to the mass of the water sample.

The salinity of seawater is typically about 3.5%, about 220 times saltier than freshwater. Seawater with a salinity of 3.5% indicates that it also contains 96.5% pure water, as shown in Figure 5.15. Because seawater is mostly pure water, its physical properties are very similar to those of pure water, with only slight variations.

Figure 5.15 and Table 5.1 show that the elements chlorine, sodium, sulfur (as the sulfate ion), magnesium, calcium, and potassium account for over 99% of the dissolved solids in seawater. More than 80 other chemical elements have been identified in seawater—most in extremely small amounts—and probably all of Earth's naturally occurring elements exist in the sea. Remarkably, trace amounts of dissolved components in seawater are vital for human survival (Box 5.1).

Salinity is often expressed in **parts per thousand (‰)**. For example, as 1% is one part in 100, 1‰ is one part in 1000. When converting from percent to parts per thousand, the decimal is simply moved one place to the right. For instance, typical seawater salinity of 3.5% is the same as 35‰. Advantages of expressing salinity in parts per thousand are that decimals are often avoided and values convert directly to grams of salt per kilogram of seawater. For example, 35‰ seawater has 35 grams of salt in every 1000 grams of seawater.[10]

FIGURE 5.14 Glass bottle shattered by frozen water. This glass bottle was filled with water, sealed, and put into a freezer. As water freezes, it expands by 9% as it forms hydrogen bonds and forms an open lattice structure, which increased the pressure and caused the bottle to fracture.

Determining Salinity

Early methods of determining seawater salinity involved evaporating a carefully weighed amount of seawater and weighing the salts that precipitated from it. However, the accuracy of this time-consuming method is limited because some water can remain bonded to salts that precipitate and some substances can evaporate along with the water.

WEB VIDEO
Seawater Evaporating
(Time Lapse)

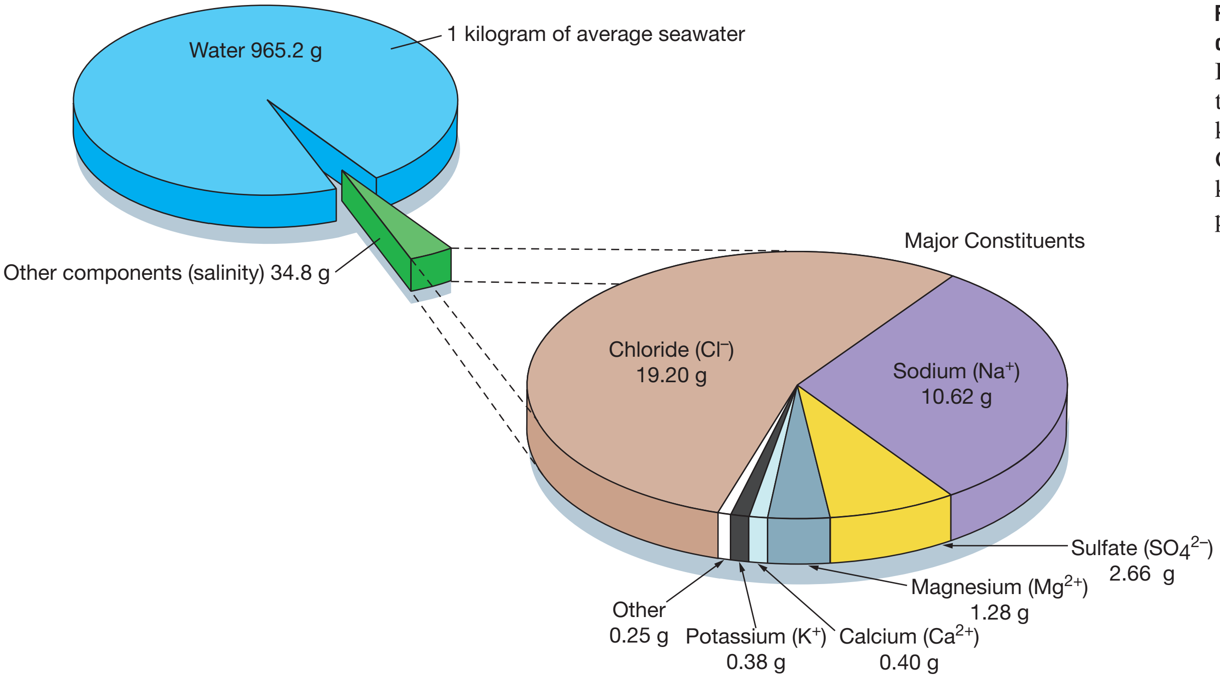

FIGURE 5.15 Major dissolved components in seawater. Diagrammatic representation of the most abundant components in a kilogram of 35‰ salinity seawater. Constituents are listed in grams per kilogram, which is equivalent to parts per thousand (‰).

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[10] Note that the units "parts per thousand" are effectively parts per thousand by weight. Salinity values, however, lack units because the salinity of a water sample is determined as the *ratio* of the electrical conductivity of the sample to the electrical conductivity of a standard. Thus, salinity values are sometimes reported in *p.s.u.*, or *practical salinity units*, which are equivalent to parts per thousand.

TABLE 5.1 SELECTED DISSOLVED MATERIALS IN 35‰ SEAWATER

1. Major Constituents (in parts per thousand, ‰)

Constituent	Concentration (‰)	Ratio of constituent/total salts (%)
Chloride (Cl^-)	19.2	55.04
Sodium (Na^+)	10.6	30.61
Sulfate (SO_4^{2-})	2.7	7.68
Magnesium (Mg^{2+})	1.3	3.69
Calcium (Ca^{2+})	0.40	1.16
Potassium (K^+)	0.38	1.10
Total	**34.58‰**	**99.28%**

2. Minor Constituents (in parts per million, ppm[a])

Gases		Nutrients		Others	
Constituent	Concentration (ppm)	Constituent	Concentration (ppm)	Constituent	Concentration (ppm)
Carbon dioxide (CO_2)	90	Silicon (Si)	3.0	Bromide (Br^-)	65.0
Nitrogen (N_2)	14	Nitrogen (N)	0.5	Carbon (C)	28.0
Oxygen (O_2)	6	Phosphorus (P)	0.07	Strontium (Sr)	8.0
		Iron (Fe)	0.002	Boron (B)	4.6

3. Trace Constituents (in parts per billion, ppb[b])

Constituent	Concentration (ppb)	Constituent	Concentration (ppb)	Constituent	Concentration (ppb)
Lithium (Li)	185	Zinc (Zn)	10	Lead (Pb)	0.03
Rubidium (Rb)	120	Aluminum (Al)	2	Mercury (Hg)	0.03
Iodine (I)	60	Manganese (Mn)	2	Gold (Au)	0.005

[a] Note that 1000 ppm = 1‰.

[b] Note that 1000 ppb = 1 ppm.

Another way to measure salinity is to use the *principle of constant proportions*, which was firmly established by chemist William Dittmar when he analyzed the water samples collected during the *Challenger* Expedition (Box 5.2). The **principle of constant proportions** states that the major dissolved constituents responsible for the salinity of seawater occur nearly everywhere in the ocean in the exact same proportions, independent of salinity. The ocean, therefore, is well mixed. When salinity changes, moreover, the salts don't leave (or enter) the

5.1 OCEANS AND PEOPLE

HOW TO AVOID GOITERS

FIGURE 5A A woman with goiters.

The nutritional label on containers of salt usually proclaims "this product contains iodine, a necessary nutrient." Why is iodine necessary in our diets? It turns out that if a person's diet contains an insufficient amount of iodine, a potentially life-threatening disease called **goiters** (*guttur* = throat) may result (Figure 5A).

Iodine is used by the thyroid gland, which is a butterfly-shaped organ located in the neck in front of and on either side of the trachea (windpipe). The thyroid gland manufactures hormones that regulate cellular metabolism essential for mental development and physical growth. If people lack iodine in their diet, their thyroid glands cannot function properly. Often, this results in the enlargement or swelling of the thyroid gland. Severe symptoms include dry skin, loss of hair, puffy face, weakness of muscles, weight increase, diminished vigor, mental sluggishness, and a large nodular growth on the neck called a goiter. If proper steps are not taken to correct this disease, it can lead to cancer. Iodine ingested regularly often begins to reverse the effects. In advanced stages, surgery to remove the goiter or exposure to radioactivity is the only course of action.

How can you avoid goiters? Fortunately, goiters can be prevented with a diet that contains just *trace amounts* of iodine. Where can you get iodine in your diet? All products from the sea contain trace amounts of iodine because iodine is one of the many elements dissolved in seawater. Sea salt, seafood, seaweed, and other sea products contain plenty of iodine to help prevent goiters. Although goiters are rarely a problem in developed nations like the United States, goiters pose a serious health hazard in many underdeveloped nations, especially those far from the sea. In the United States, however, many people get too much iodine in their diet, leading to the overproduction of hormones by the thyroid gland. That's why most stores that sell iodized salt also carry noniodized salt for those people who have a *hyperthyroid* (*hyper* = excessive, *thyroid* = the thyroid gland) condition and must restrict their intake of iodine.

ocean, but water molecules do. Seawater has *constancy of composition*, so the concentration of a single major constituent can be measured to determine the total salinity of a given water sample. The constituent that occurs in the greatest abundance and is the easiest to measure accurately is the chloride ion, Cl^-. The weight of this ion in a water sample is its **chlorinity**.

In any sample of ocean water worldwide, the chloride ion accounts for 55.04% of the total proportion of dissolved solids (Figure 5.15 and Table 5.1). Therefore, by measuring only the chloride ion concentration, the total salinity of a seawater sample can be determined using the following relationship:

$$\text{Salinity (‰)} = 1.80655 \times \text{chlorinity (‰)}^{*} \tag{5.1}$$

For example, the average chlorinity of the ocean is 19.2‰, so the average salinity is 1.80655 × 19.2‰ which rounds to 34.7‰. In other words, on average there are 34.7 parts of dissolved material in every 1000 parts of seawater.

Standard seawater consists of ocean water analyzed for chloride ion content to the nearest ten-thousandth of a part per thousand by the Institute of Oceanographic Services in Wormly, England. It is then sealed in small glass vials called *ampules* and sent to laboratories throughout the world for use as a reference standard in calibrating analytical equipment.

Seawater salinity can be measured very accurately with modern oceanographic instruments such as a **salinometer** (*salinus* = salt, *meter* = measure). Most

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*The number 1.80655 comes from dividing 1 by 0.5504 (the chloride ion's proportion in seawater of 55.04%). However, if you actually divide this, you will get 1.81686, which is different from the original value by 0.57%. Empirically, oceanographers found that seawater's constancy of composition is an approximation and have agreed to use 1.80655 because it more accurately represents the total salinity of seawater.

5.2 HISTORICAL FEATURE

THE HMS *CHALLENGER* EXPEDITION: BIRTH OF OCEANOGRAPHY

Oceanography as a scientific discipline began in 1872 with the **HMS *Challenger*** expedition, the first large-scale voyage with the express purpose of studying the ocean for scientific purposes. What inspired such a voyage? One of the principal scientific disputes in the mid 1800s was initiated by one of the most influential biologists of his time, Edwin Forbes (see Web Box 15.1), who asserted that life below about 550 meters (1800 feet) was impossible because of high pressure and lack of light. Could life exist in the deep ocean? If so, what were the physical and chemical conditions there? What was the nature of sea floor deposits?

In 1871, the Royal Society of England recommended that funds be raised for an expedition to investigate the distribution of life in the sea as well as the physical and chemical conditions of the water column from the surface to the sea floor. The British government agreed to sponsor such an expedition and in 1872, a reserve warship was refitted to support scientific studies and renamed HMS *Challenger* (Figure 5B, *inset*). It contained a staff of six scientists under the direction of Charles Wyville Thomson, a professor of natural history at Scotland's University of Edinburgh.

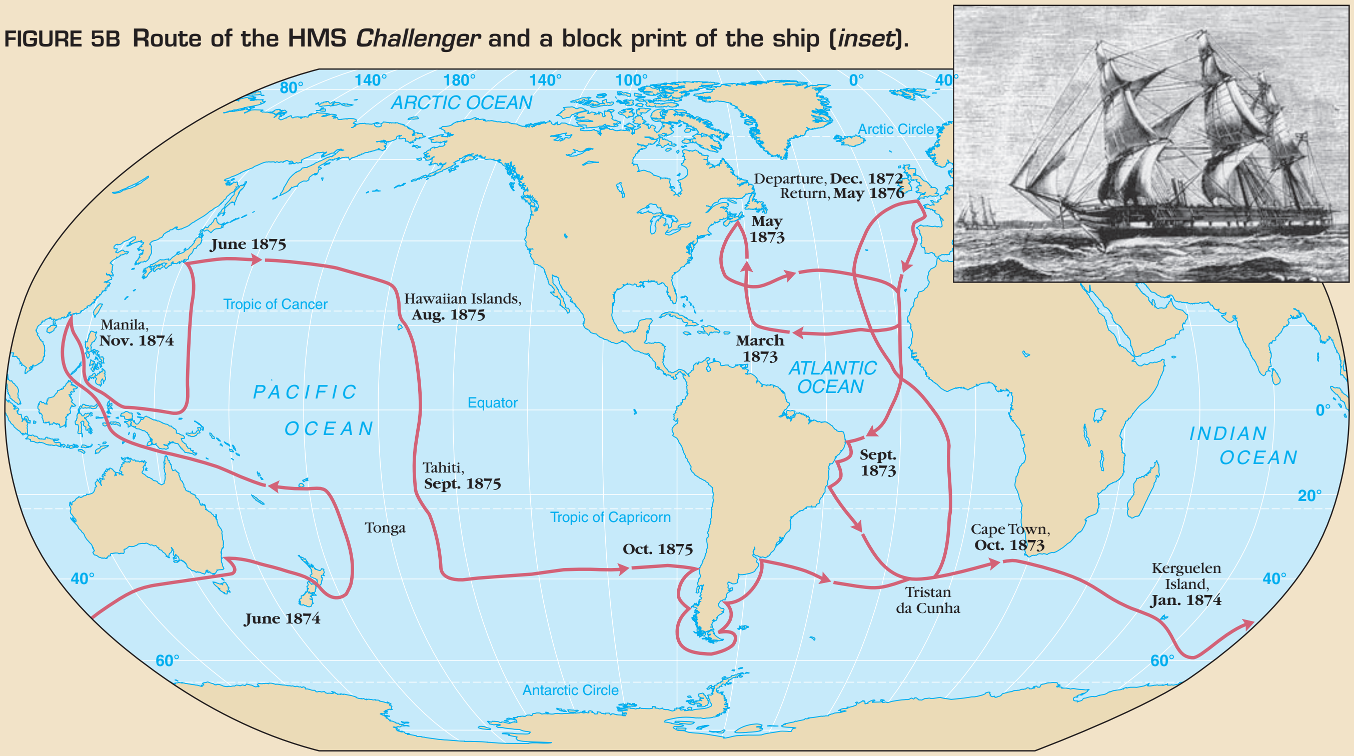
From time to time during the course of the voyage, the ship would stop to measure the water depth using a sounding line, the bottom temperature with newly developed thermometers that could withstand the high pressure at depth, and atmospheric and meteorologic conditions. In addition, a sample of the bottom water was collected, and the bottom sediment was dredged. Other measurements included trawling the bottom for life using a net, collecting organisms at the surface, determining temperature at various depths, gathering samples of seawater from certain depths, and recording surface and deep-water currents.

Challenger returned in May 1876, after circumnavigating the globe for nearly three and a half years (Figure 5B). During the 127,500-kilometer (79,200-mile) voyage, the scientists performed 492 deep-sea soundings, dredged the bottom 133 times, trawled the open water 151 times, took 263 water temperature readings, and collected water samples from as deep as 1830 meters (6000 feet). As with most other oceanographic expeditions, the real work of analyzing the data was just beginning. In fact, it took nearly 20 years to compile the expedition results into 50 volumes.

Major accomplishments of the voyage included verifying the existence of life at all ocean depths (thus proving Forbes wrong), classifying 4717 new marine species, measuring a then-record water depth of 8185 meters (26,850 feet) in the Mariana Trench in the western North Pacific Ocean, demonstrating that the ocean floor was not flat but had significant relief, and discovering manganese nodules.

Pioneering work on the chemistry of the oceans was completed, too. Analysis by chemist William Dittmar of 77 ocean water samples collected during the *Challenger* expedition revealed that the oceans had a remarkably consistent chemical composition, even down to minor dissolved substances. Not only were the ratios between various salts constant at the surface from ocean to ocean, but they were also constant at depth. This relationship is the basis of what is now known as the *principle of constant proportions* or *Forchammer's principle*, which has contributed greatly to the understanding of ocean salinity.

FIGURE 5B **Route of the HMS *Challenger* and a block print of the ship (*inset*).**



salinometers measure seawater's *electrical conductivity* (the ability of a substance to transmit electric current), which increases as more substances are dissolved in water (Figure 5.16). Salinometers can determine salinity to resolutions of better than 0.003‰.

Comparing Pure Water and Seawater

Table 5.2 compares various properties of pure water and seawater. Because seawater is 96.5% water, most of its physical properties are very similar to those of pure water. For instance, the color of pure water and seawater is identical.

The dissolved substances in seawater, however, give it slightly different yet important physical properties as compared to pure water. For example, recall that dissolved substances interfere with pure water changing state. The freezing points and boiling points in Table 5.2 show that dissolved substances decrease the freezing point and increase the boiling point of water. Thus, seawater freezes at a temperature 1.9°C (3.4°F) lower than pure water. Similarly, seawater boils at a temperature 0.6°C (1.1°F) higher than pure water. As a result, the salts in seawater extend the range of temperatures in which water is a liquid. The same principle applies to antifreeze used in automobile radiators. Antifreeze lowers the freezing point of the water in a radiator and increases the boiling point, thus extending the range over which the water remains in the liquid state. Antifreeze, therefore, protects your radiator from freezing in the winter *and* from boiling over in the summer.

Other important properties of seawater (such as pH and density) are discussed in later parts of this chapter.

TABLE 5.2 COMPARISON OF SELECTED PROPERTIES OF PURE WATER AND SEAWATER

Property	Pure water	35‰ Seawater
Color (light transmission)		
• Small quantities of water	Clear (high transparency)	Same as for pure water
• Large quantities of water	Blue-green because water molecules scatter blue and green wavelengths best	Same as for pure water
Odor	Odorless	Distinctly marine
Taste	Tasteless	Distinctly salty
pH	7.0 (neutral)	Surface waters range = 8.0–8.3; average = 8.1 (slightly alkaline)
Density at 4°C (39°F)	1.000 g/cm^3	1.028 g/cm^3
Freezing point	0°C (32°F)	−1.9°C (28.6°F)
Boiling point	100°C (212°F)	100.6°C (213.1°F)

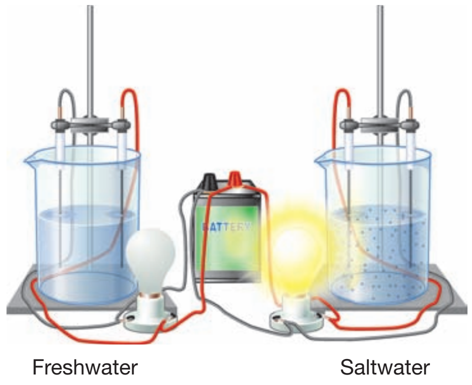


FIGURE 5.16 Salinity affects water conductivity. Increasing the amount of dissolved substances increases the conductivity of the water. A light bulb with bare electrodes shows that the higher the salinity, the more electricity is transmitted and the brighter the bulb will be lit.

KEY CONCEPT

Seawater salinity can be measured by salinometer and averages 35‰. The dissolved components in seawater give it different yet important physical properties as compared to pure water.

STUDENTS SOMETIMES ASK...

What is the strategy behind adding salt to a pot of water when making pasta? Does it make the water boil faster?

Adding salt to water will not make the water boil faster. It will, however, make the water boil at a slightly higher temperature because *dissolved substances raise* its boiling point (and *lower* its freezing point; see Table 5.2). Thus, the pasta will cook in slightly less time. In addition, the salt adds flavoring, so the pasta may taste better, too. Be sure to add the salt *after* the water has come to a boil, though, or it will take longer to reach a boil. This is a wonderful use of chemical principles—helping you to cook better!

STUDENTS SOMETIMES ASK...

I've seen the labels on electric cords warning against using electrical appliances close to water. Are these warnings because water's polarity allows electricity to be transmitted through it?

Yes and no. Water molecules are polar, so you might assume that water is a good conductor of electricity. Pure water is a very poor conductor, however, because water molecules are neutral overall and will not move toward the negatively or positively charged pole in an electrical system. If an electrical appliance is dropped into a tub of absolutely pure water, the water molecules will transmit no electricity. Instead, the water molecules will simply orient their positively charged hydrogen ends toward the negative pole of the appliance and their negatively charged oxygen ends toward its positive pole, which tends to neutralize the electric field. Interestingly, it is the dissolved substances that transmit electrical current through water (see Figure 5.16). Even slight amounts, such as those in tap water, allow electricity to be transmitted. That's why there are warning labels on the electric cords of household appliances that are commonly used in the bathroom, such as blow dryers, electric razors, and heaters. That's also why it is recommended to stay out of any water—including a bathtub or shower—during a lightning storm!

5.4 Why Does Seawater Salinity Vary?

Using salinometers and other techniques, oceanographers have determined that salinity varies from place to place in the oceans.

Salinity Variations

In the open ocean far from land, salinity varies between about 33 and 38‰. In coastal areas, salinity variations can be extreme. In the Baltic Sea, for example, salinity averages only 10‰ because physical conditions create **brackish** (*brak* = salt, *ish* = somewhat) water. Brackish water is produced in areas where freshwater (from rivers and high rainfall) and seawater mix. In the Red Sea, on the other hand, salinity averages 42‰ because physical conditions produce **hypersaline** (*hyper* = excessive, *salinus* = salt) water. Hypersaline water is typical of seas and inland bodies of water that experience high evaporation rates and limited open-ocean circulation.

Some of the most hypersaline water in the world is found in inland lakes, which are often called seas because they are so salty. The Great Salt Lake in Utah, for example, has a salinity of 280‰, and the Dead Sea on the border of Israel and Jordan has a salinity of 330‰. The water in the Dead Sea, therefore, contains 33% dissolved solids and is almost *10 times saltier than seawater*. As a result, hypersaline waters are so dense and buoyant that one can easily float (Figure 5.17), even with arms and legs sticking up above water level! Hypersaline waters also taste much saltier than seawater.

Salinity of seawater in coastal areas also varies seasonally. For example, the salinity of seawater off Miami Beach, Florida, varies from about 34.8‰ in October to 36.4‰ in May and June when evaporation is high. Offshore of Astoria, Oregon, seawater salinity is always relatively low because of the vast freshwater input from the Columbia River. Here, surface water salinity can be as low as 0.3‰ in April and May (when the Columbia River is at its maximum flow rate) and 2.6‰ in October (the dry season).

Other types of water have much lower salinity. Tap water, for instance, has salinity somewhere below 0.8‰, and good-tasting tap water is usually below 0.6‰. Salinity of premium bottled water is on the order of 0.3‰, with the salinity often displayed prominently on its label, usually as total dissolved solids (TDS) in units of parts per million (ppm), where 1000 ppm equals 1‰.

Processes Affecting Seawater Salinity

Processes affecting seawater salinity change either the amount of water (H_2O molecules) or the amount of dissolved substances within the water. Adding more water, for instance, dilutes the dissolved component and lowers the salinity of the sample. Conversely, removing water increases salinity. Changing the salinity in these ways does not affect the *amount* or the *composition* of the dissolved components, which remain in constant proportions. Let's first examine processes that affect the amount of water in seawater before turning our attention to processes that influence dissolved components.

PROCESSES THAT DECREASE SEAWATER SALINITY Table 5.3 summarizes the processes affecting seawater salinity. Precipitation, **runoff** (stream discharge), melting icebergs, and melting sea ice *decrease* seawater salinity by adding more freshwater to the ocean. Precipitation is the way atmospheric water returns to Earth as rain, snow, sleet, and hail. Worldwide, about three-quarters of all precipitation falls directly back into the ocean and one-quarter falls onto land. Precipitation falling directly into the oceans adds freshwater, reducing seawater salinity.

FIGURE 5.17 High-salinity water of the Dead Sea allows swimmers to easily float. The Dead Sea, which has 330‰ salinity (almost 10 times the salinity of seawater), has high density. As a result, it also has high buoyancy that allows swimmers to float easily.

Most of the precipitation that falls on land returns to the oceans indirectly as stream runoff. Even though this water dissolves minerals on land, the runoff is relatively pure water, as shown in Table 5.4. Runoff, therefore, adds mostly water to the ocean, causing seawater salinity to decrease.

Icebergs are chunks of ice that have broken free (*calved*) from a glacier when it flows into an ocean or marginal sea and begins to melt. Glacial ice originates as snowfall in high mountain areas, so icebergs are composed of freshwater. When icebergs melt in the ocean, they add freshwater, which is another way in which seawater salinity is reduced.

Sea ice forms when ocean water freezes in high-latitude regions and is composed primarily of freshwater. When warmer temperatures return to high-latitude regions in the summer, sea ice melts in the ocean, adding mostly freshwater with a small amount of salt to the ocean. Seawater salinity, therefore, is decreased.

PROCESSES THAT INCREASE SEAWATER SALINITY The formation of sea ice and evaporation *increase* seawater salinity by removing water from the ocean (Table 5.3). Sea ice forms when seawater freezes. Depending on the salinity of seawater and the rate of ice formation, about 30% of the dissolved components in seawater are retained in sea ice. This means that 35‰ seawater creates sea ice with about 10‰ salinity (30% of 35‰ is 10‰). Consequently, the formation of sea ice removes mostly freshwater from seawater, increasing the salinity of the remaining unfrozen water. High-salinity water also has a high density, so it sinks below the surface.

Recall that evaporation is the conversion of water molecules from the liquid state to the vapor state at temperatures below the boiling point. Evaporation removes water from the ocean, leaving its dissolved substances behind. Evaporation, therefore, increases seawater salinity.

THE HYDROLOGIC CYCLE Figure 5.18 shows how the **hydrologic** (*hydro* = water, *logos* = study of) **cycle** relates the processes that affect seawater salinity. These processes recycle water among the ocean, the atmosphere, and the continents, so water is in continual motion between the different components

STUDENTS SOMETIMES ASK...

What would happen to a person if he or she drank seawater?

It depends on the quantity. The salinity of seawater is about four times greater than that of your body fluids. In your body, seawater causes your internal membranes to lose water through *osmosis* (*osmos* = to push) which transports water molecules from higher concentrations (the normal body chemistry of your internal fluids) to areas of lower concentrations (your digestive tract containing seawater). Thus, your natural body fluids would move into your digestive tract and eventually be expelled, causing dehydration.

Don't worry too much if you've inadvertently swallowed some seawater. As a nutritional drink, seawater provides seven important nutrients and contains no fat, cholesterol, or calories. Some people even claim that drinking a small amount of seawater daily gives them good health! However, beware of microbial contaminants in seawater such as viruses and bacteria that can often exist in great quantities.

KEY CONCEPT

Various surface processes either decrease seawater salinity (precipitation, runoff, icebergs melting, or sea ice melting) or increase seawater salinity (sea ice forming and evaporation).

TABLE 5.3 PROCESSES AFFECTING SEAWATER SALINITY

Process	How accomplished	Adds or removes	Effect on salt in seawater	Effect on H_2O in seawater	Salinity increase or decrease?	Source of freshwater from the sea?
Precipitation	Rain, sleet, hail, or snow falls directly on the ocean	Adds very fresh water	None	More H_2O	Decrease	N/A
Runoff	Streams carry water to the ocean	Adds mostly fresh water	Negligible addition of salt	More H_2O	Decrease	N/A
Icebergs melting	Glacial ice calves into the ocean and melts	Adds very fresh water	None	More H_2O	Decrease	Yes, icebergs from Antarctic have been towed to South America
Sea ice melting	Sea ice melts in the ocean	Adds mostly fresh water and some salt	Adds a small amount of salt	More H_2O	Decrease	Yes, sea ice can be melted and is better than drinking seawater
Sea ice forming	Seawater freezes in cold ocean areas	Removes mostly fresh water	30% of salts in seawater are retained in ice	Less H_2O	Increase	Yes, through multiple freezings, called *freeze separation*
Evaporation	Seawater evaporates in hot climates	Removes very pure water	None (essentially all salts are left behind)	Less H_2O	Increase	Yes, through evaporation of seawater and condensation of water vapor, called *distillation*

(*reservoirs*) of the hydrologic cycle. Earth's water supply exists in the following proportions:

- 97.2% in the world ocean
- 2.15% frozen in glaciers and ice caps
- 0.62% in groundwater and soil moisture
- 0.02% in streams and lakes
- 0.001% as water vapor in the atmosphere

In addition, Figure 5.18 shows the average yearly amounts of transfer or *flux* of water between various reservoirs.

Earth's Water and the Hydrologic Cycle

Dissolved Components Added to and Removed from Seawater

Seawater salinity is a function of the amount of dissolved components in seawater. Interestingly, dissolved substances do not remain in the ocean forever. Instead, they are cycled into and out of seawater by the processes shown in Figure 5.19. These processes include stream runoff in which streams dissolve ions from continental rocks and carry them to the sea, and volcanic eruptions, both on the land and on the sea floor. Other sources include the atmosphere (which contributes gases) and biologic interactions.

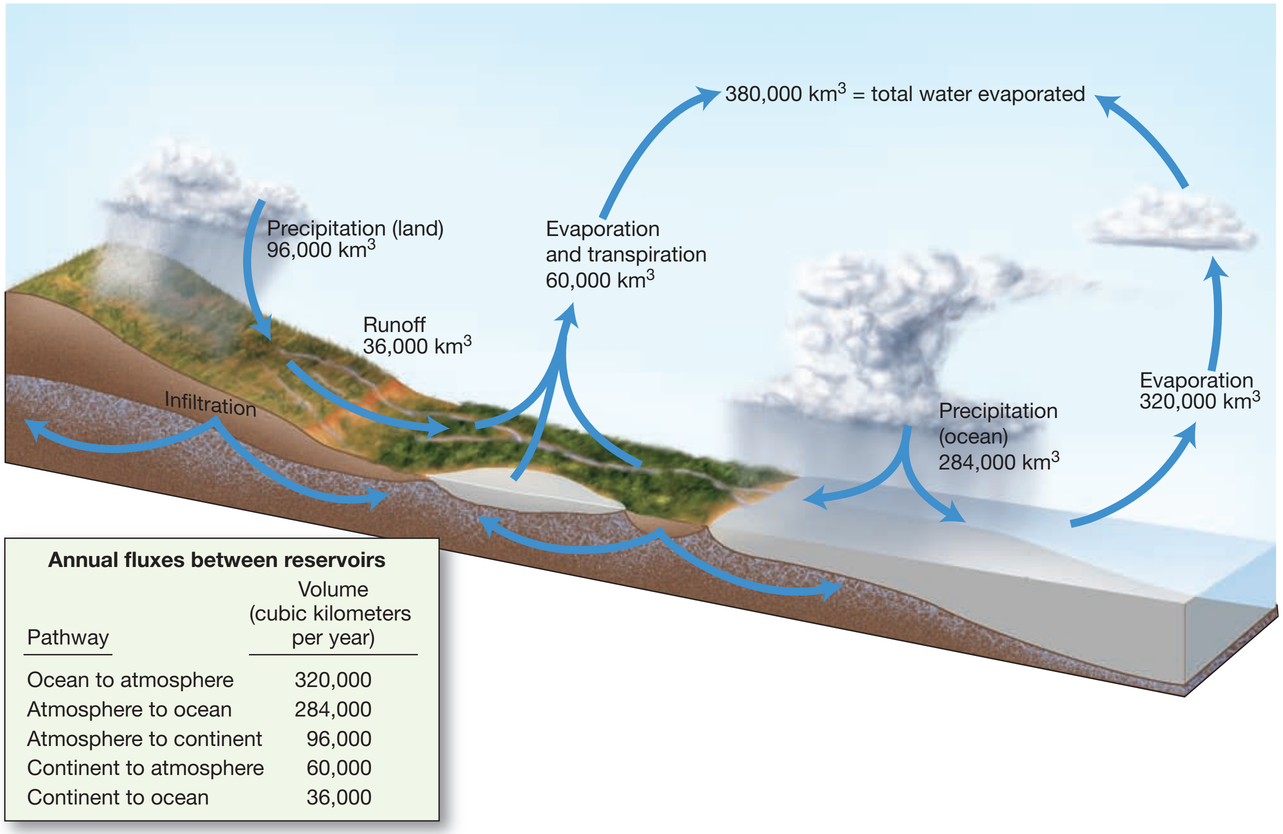
Stream runoff is the primary method by which dissolved substances are added to the oceans. Table 5.4 compares the major components dissolved in stream water with those in seawater. It shows that streams have far lower salinity and a vastly different composition of dissolved substances than seawater. For example, bicarbonate ion (HCO_3^-) is the most abundant dissolved constituent in stream water yet is found in only trace amounts in seawater. Conversely, the most abundant dissolved component in seawater is the chloride ion (Cl^-) which exists in very small concentrations in streams.

If stream water is the main source of dissolved substances in seawater, why do the components of the two not match each other more closely? One of the reasons is that some dissolved substances stay in the ocean and accumulate over time. **Residence time** is the average length of time that a substance resides in the ocean. Long residence times lead to higher concentrations of the dissolved substance. The sodium ion (Na^+) for instance, has a residence time of 260 million years, and, as a result, has a high concentration in the ocean. Other elements such as aluminum have a residence time of only 100 years and occur in seawater in much lower concentrations.

STUDENTS SOMETIMES ASK...

You mentioned that when seawater freezes, it produces ice with about 10‰ salinity. Once that ice melts, can a person drink it with no ill effects?

Early Arctic explorers found out the answer to your question by necessity. Some of these explorers who traveled by ship in high-latitude regions became inadvertently or purposely entrapped by sea ice (see, for example, Web Box 7.1, which describes the voyage of the *Fram*). Lacking other water sources, they used melted sea ice. Although newly formed sea ice contains little salt, it does trap a significant amount of brine (drops of salty water). Depending on the rate of freezing, newly formed ice may have a total salinity from 4 to 15‰. The more rapidly it forms, the more brine it captures and the higher the salinity. Melted sea ice with salinity this high doesn't taste very good, and it still causes dehydration, but not as quickly as drinking 35‰ seawater does. Over time, however, the brine will trickle down through the coarse structure of the sea ice, so its salinity decreases. By the time it is a year old, sea ice normally becomes relatively pure. Drinking melted sea ice enabled these early explorers to survive.

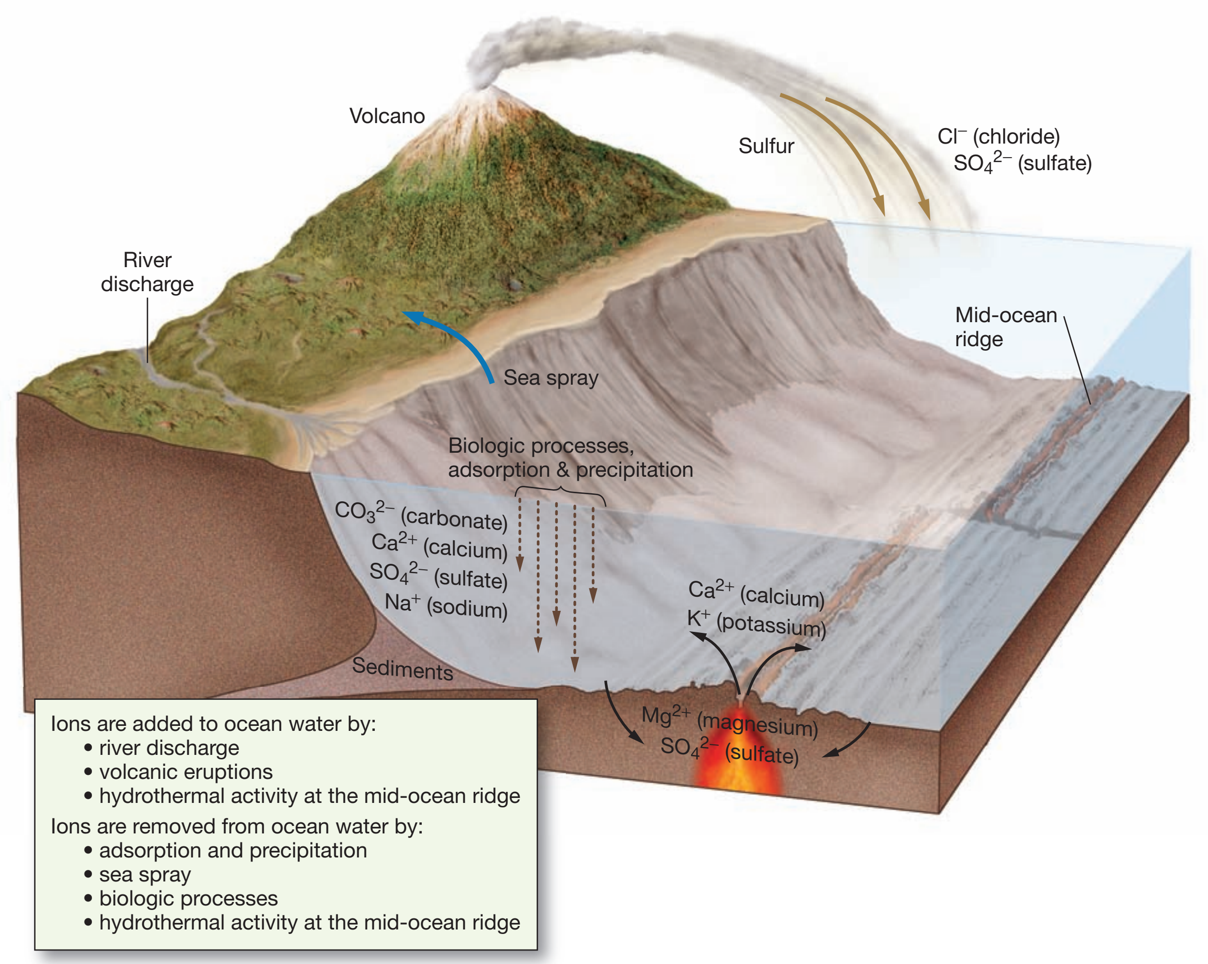


Annual fluxes between reservoirs

Pathway	Volume (cubic kilometers per year)
Ocean to atmosphere	320,000
Atmosphere to ocean	284,000
Atmosphere to continent	96,000
Continent to atmosphere	60,000
Continent to ocean	36,000

FIGURE 5.18 The hydrologic cycle. All water is in continual motion between the various components (reservoirs) of the hydrologic cycle. Volumes are Earth's average yearly amounts in cubic kilometers; table shows average yearly flux between reservoirs; ice not shown.

FIGURE 5.19 The cycling of dissolved components in seawater. Dissolved components are added to seawater primarily by river discharge and volcanic eruptions, while they are removed by adsorption, precipitation, ion entrapment in sea spray, and marine organisms that produce hard parts. Chemical reactions at the mid-ocean ridge add and remove various dissolved components.



Ions are added to ocean water by:
- river discharge
- volcanic eruptions
- hydrothermal activity at the mid-ocean ridge

Ions are removed from ocean water by:
- adsorption and precipitation
- sea spray
- biologic processes
- hydrothermal activity at the mid-ocean ridge

TABLE 5.4 COMPARISON OF MAJOR DISSOLVED COMPONENTS IN STREAMS WITH THOSE IN SEAWATER

Constituent	Concentration in streams (parts per million)	Concentration in seawater (parts per million)
Bicarbonate ion (HCO_3^-)	58.4	trace
Calcium ion (Ca^{2+})	15.0	400
Silicate (SiO_2)	13.1	3
Sulfate ion (SO_4^{2-})	11.2	2700
Chloride ion (Cl^-)	7.8	19,200
Sodium ion (Na^+)	6.3	10,600
Magnesium ion (Mg^{2+})	4.1	1300
Potassium ion (K^+)	2.3	380
Total (parts per million)	119.2 ppm	34,793 ppm
Total (‰)	0.1192‰	34.8‰

Are the oceans becoming saltier through time? This might seem logical since new dissolved components are constantly being added to the oceans and because most salts have long residence times. However, analysis of ancient marine organisms and sea floor sediments suggests that the oceans have not increased in salinity over time. This must be because the rate at which an element is added to the ocean equals the rate at which it is removed, so the average *amounts* of various elements remain constant (this is called a *steady-state* condition).

Materials added to the oceans are counteracted by several processes that cycle dissolved substances out of seawater. When waves break at sea, for example, sea spray releases tiny salt particles into the atmosphere where they may be blown over land before being washed back to Earth. The amount of material leaving the ocean in this way is enormous: According to a recent study, as much as 3.3 billion metric tons (3.6 billion short tons) of salt as sea spray enter the atmosphere each year. Another example is the infiltration of seawater along mid-ocean ridges near hydrothermal vents (see Figure 5.17), which incorporates magnesium and sulfate ions into sea floor mineral deposits. In fact, chemical studies of seawater indicate that the *entire volume of ocean water* is recycled through this hydrothermal circulation system at the mid-ocean ridge approximately every 3 million years. As a result, the chemical exchange between ocean water and the basaltic crust has a major influence on the composition of ocean water.

Dissolved substances are also removed from seawater in other ways. Calcium, carbonate, sulfate, sodium, and silicon are deposited in ocean sediments within the shells of dead microscopic organisms and animal feces. Vast amounts of dissolved substances can be removed when inland arms of seas dry up, leaving salt deposits called *evaporites* (such as those beneath the Mediterranean Sea; see Web Box 4.1). In addition, ions dissolved in ocean water are removed by adsorption (physical attachment) to the surfaces of sinking clay and biologic particles.

5.5 Is Seawater Acidic or Basic?

An **acid** is a compound that releases hydrogen ions (H^+) when dissolved in water. The resulting solution is said to be *acidic*. A strong acid readily and completely releases hydrogen ions when dissolved in water. An **alkaline** or a **base** is a compound that releases hydroxide ions (OH^-) when dissolved in water. The resulting solution is said to be *alkaline* or *basic*. A strong base readily and completely releases hydroxide ions when dissolved in water.

Both hydrogen ions and hydroxide ions are present in extremely small amounts at all times in water because water molecules dissociate and reform. Chemically, this is represented by the equation:

$$H_2O \underset{\text{reform}}{\overset{\text{dissociate}}{\rightleftarrows}} H^+ + OH^- \tag{5.2}$$

Note that if the hydrogen ions and hydroxide ions in a solution are due only to the dissociation of water molecules, they are always found in equal concentrations, and the solution is consequently neutral.

When substances dissociate in water, they can make the solution acidic or basic. For example, if hydrochloric acid (HCl) is added to water, the resulting solution will be acidic because there will be a large excess of hydrogen ions from the dissociation of the HCl molecules. Conversely, if a base such as baking soda (sodium bicarbonate, $NaHCO_3$) is added to water, the resulting solution will be basic because there will be an excess of hydroxide ions (OH^-) from the dissociation of $NaHCO_3$ molecules.

The pH Scale

Figure 5.20 shows the **pH** (potential of hydrogen) **scale**, which is a measure of the hydrogen ion concentration of a solution. Values for pH range from 0 (strongly acidic) to 14 (strongly alkaline or basic) and the pH of a **neutral** solution such as pure water is 7.0. A decrease of one pH unit corresponds to a 10-fold increase in the concentration of hydrogen ions, making the water more acidic, whereas a change of one unit upward corresponds to a 10-fold decrease, making the water more alkaline.

Ocean surface waters have a pH that averages about 8.1 and ranges from about 8.0 to 8.3, so seawater is slightly alkaline. Water in the ocean combines with carbon dioxide to form a weak acid, called carbonic acid (H_2CO_3) which dissociates and releases hydrogen ions (H^+):

$$H_2O + CO_2 \longrightarrow H_2CO_3 \longrightarrow H^+ + HCO_3^- \qquad (5.3)$$

This reaction would seem to make the ocean slightly acidic. Carbonic acid, however, keeps the ocean slightly alkaline through the process of *buffering*.

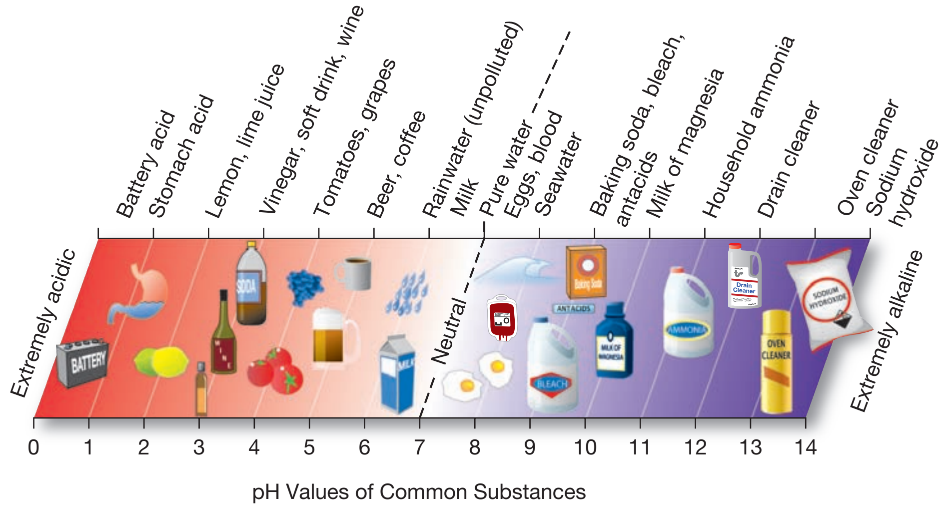


FIGURE 5.20 The pH scale. The pH scale ranges from 0 (highly acidic) to 14 (highly alkaline). A pH of 7 is neutral; the pH values of common substances are also shown.

KEY CONCEPT

Reactions involving carbonate chemicals serve to buffer the ocean and help maintain its average pH at 8.1 (slightly alkaline or basic).

The Carbonate Buffering System

The chemical reactions in Figure 5.21 show that carbon dioxide (CO_2) combines with water (H_2O) to form carbonic acid (H_2CO_3). Carbonic acid can then lose a hydrogen ion (H^+) to form the negatively charged bicarbonate ion (HCO_3^-). The bicarbonate ion can lose its hydrogen ion, too, though it does so less readily than carbonic acid. When the bicarbonate ion loses its hydrogen ion, it forms the double-charged negative carbonate ion (CO_3^{2-}) some of which combines with calcium ions to form calcium carbonate ($CaCO_3$). Some of the calcium carbonate is precipitated by various inorganic and organic means and then it sinks and cycles back into the ocean by dissolving at depth.

The equations below Figure 5.21 show how these chemical reactions involving carbonate minimize changes in the pH of the ocean in a process called **buffering**. Buffering protects the ocean from getting too acidic or too basic, similar to how buffered aspirin protects sensitive stomachs. For example, if the pH of the ocean increases (becomes too basic), it causes H_2CO_3 to release H^+ and pH drops. Conversely, if the pH of the ocean decreases (becomes too acidic), HCO_3^- combines with H^+ to remove it, causing pH to rise. In this way, buffering prevents large swings of ocean water pH and allows the ocean to stay within a limited range of pH values. Recently, however, increasing amounts of carbon dioxide from human emissions are beginning to enter the ocean and change the ocean's pH by making it more acidic. For more details on this, see Chapter 16, "The Oceans and Climate Change."

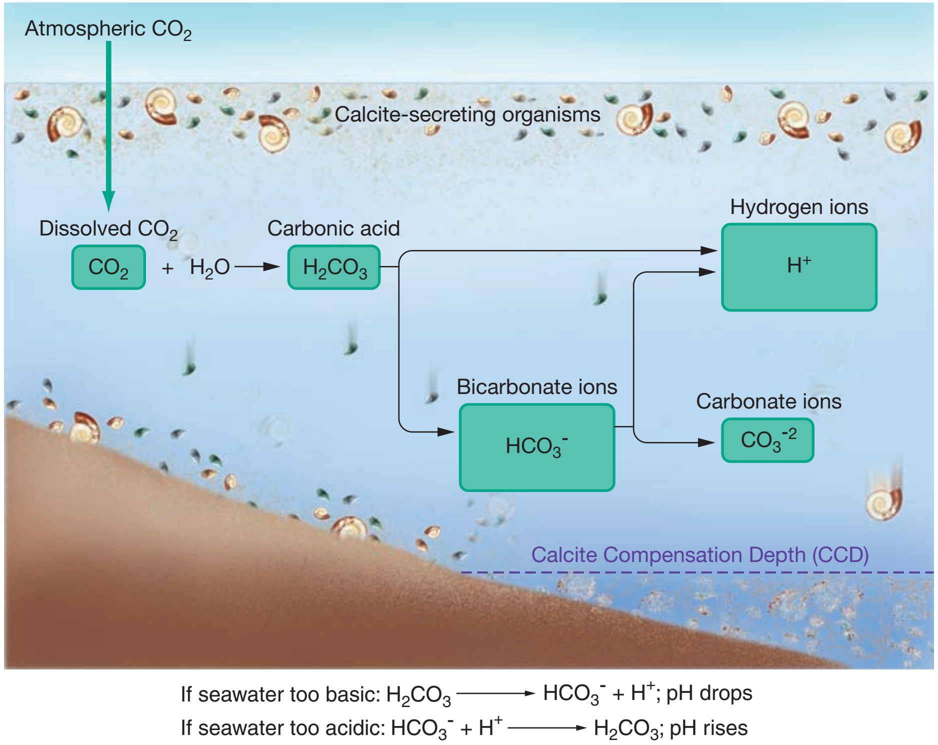


FIGURE 5.21 The carbonate buffering system. Atmospheric carbon dioxide (CO_2) enters the ocean and undergoes chemical reactions. If seawater is too basic, chemical reactions occur that release H^+ into seawater and lower pH. If seawater is too acidic, chemical reactions occur that remove H^+ from seawater and cause pH to rise. Thus, buffering keeps the pH of seawater constant.

Deep-ocean water contains more carbon dioxide than surface water because deep water is cold and has the ability to dissolve more gases. Also, the higher pressures of the deep ocean further aid the dissolution of gases in seawater. Because carbon dioxide combines with water to form carbonic acid, why isn't the cold water of the deep ocean highly acidic? When microscopic marine organisms that make their shells out of calcium carbonate (calcite) die and sink into the deep ocean, they neutralize the acid through buffering. In essence, these organisms act as an "antacid" for the deep ocean analogous to the way commercial antacids use calcium carbonate to neutralize excess stomach acid. As explained in Chapter 4, these shells are readily dissolved below the calcite (calcium carbonate) compensation depth (CCD).

5.6 How Does Seawater Salinity Vary at the Surface and With Depth?

Average seawater salinity is 35‰, but it varies significantly from place to place at the surface and also with depth.

Surface Salinity Variation

Figure 5.22 shows how salinity varies at the surface with latitude. The red curve shows temperature, which decreases at high latitudes and increases near the equator. The green curve shows salinity, which is lowest at high latitudes, highest at the Tropics of Cancer and Capricorn, and dips near the equator.

Why does surface salinity vary in the pattern shown in Figure 5.22? At high latitudes, abundant precipitation and runoff and the melting of freshwater icebergs all decrease salinity. In addition, cool temperatures limit the amount of evaporation that takes place (which would increase salinity). The formation and melting of sea ice balance each other out in the course of a year and are not a factor in changes in salinity.

The pattern of Earth's atmospheric circulation (see Chapter 6, "Air-Sea Interaction") causes warm dry air to descend at lower latitudes near the Tropics of Cancer and Capricorn, so evaporation rates are high and salinity increases. In addition, little precipitation and runoff occur to decrease salinity. As a result, the regions near the Tropics of Cancer and Capricorn are the continental *and* maritime deserts of the world.

STUDENTS SOMETIMES ASK...

If water molecules are so good at dissolving almost everything, then why does pure water have a neutral pH of 7.0?

Indeed, pure water's neutral pH might seem surprising in light of its tremendous ability to dissolve substances. Intuitively, it seems like water should be acidic and thus have a low pH. However, pH measures the amount of hydrogen ions (H^+) in solution, not the ability of a substance to dissolve by forming hydrogen bonds (as water molecules do).

STUDENTS SOMETIMES ASK...

Why do carbonated beverages burn my throat when I drink them?

When carbon dioxide gas (CO_2) dissolves in water (H_2O), its molecules often cling to water molecules and form carbonic acid (H_2CO_3). Carbonic acid is a weak acid, an acid in which most molecules are intact at any given moment. However, some of those molecules naturally break apart and exist as two fragments: a negatively charged $H_2CO_3^-$ ion and a positively charged H^+ ion. The H^+ ions are responsible for acidity—the higher their concentration in a solution, the more acidic that solution. The presence of carbonic acid in carbonated water makes that water acidic—the more carbonated, the more acidic. What you're feeling when you drink a carbonated beverage is the moderate acidity of that beverage irritating your throat.

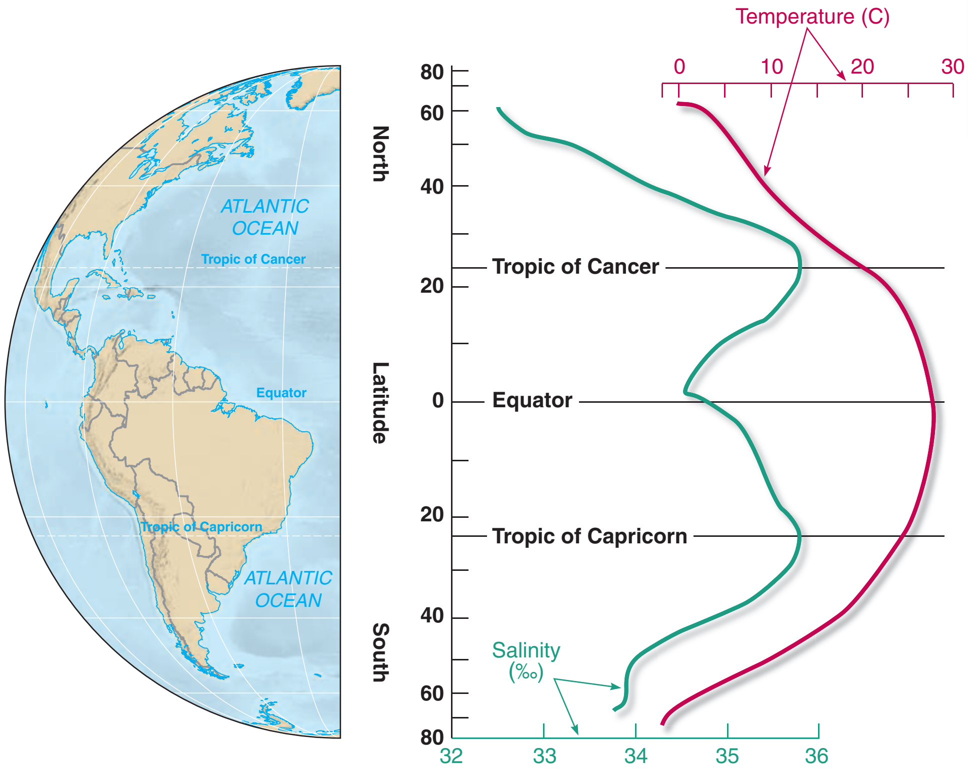


FIGURE 5.22 Surface salinity variation. Sea surface temperature (*red curve*) is lowest at the poles and highest at the equator. Surface seawater salinity (*green curve*) is lowest at the poles, peaks at the Tropics of Cancer and Capricorn, and dips near the equator. The presence of large amounts of runoff from land in far northern latitudes causes salinity to be lower there as compared to equivalent latitudes in the Southern Hemisphere.

Temperatures are warm near the equator, so evaporation rates are high enough to increase salinity. Increased precipitation and runoff partially offsets the high salinity, though. For example, daily rain showers are common along the equator, adding water to the ocean and lowering its salinity.

The map in Figure 5.23 shows how ocean surface salinity varies worldwide. Notice how the overall pattern matches the graph in Figure 5.22. The Atlantic Ocean's higher overall salinity is caused by its proximity to land, which experiences continental effect. This causes high rates of evaporation in the narrow Atlantic, particularly along the Tropics of Cancer and Capricorn.

Salinity Variation with Depth

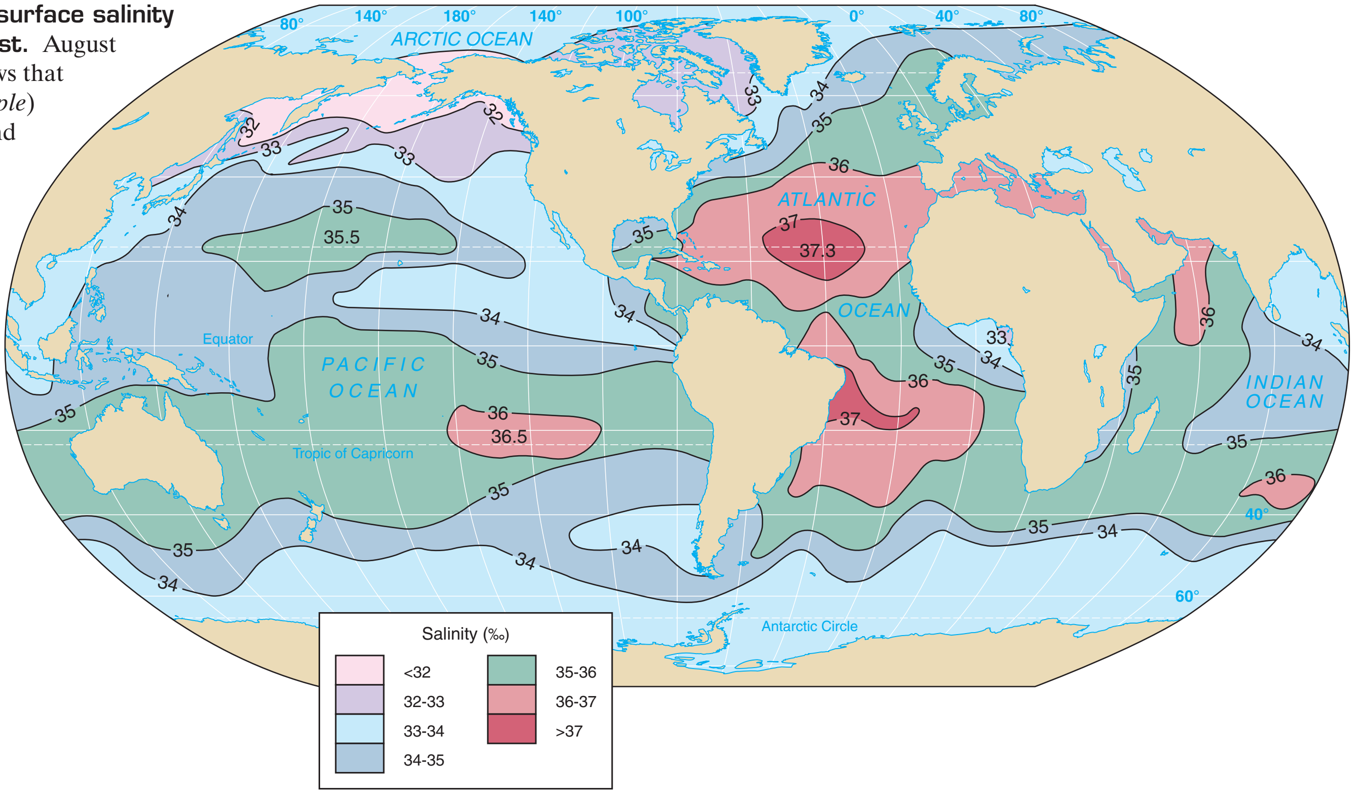
Figure 5.24 shows how seawater salinity varies with depth. The graph displays data for the open ocean far from land and shows one curve for high-latitude regions and one for low-latitude regions.

For low-latitude regions (such as in the tropics), the curve begins at the surface with relatively high salinity (as was discussed in the preceding section). Even at the equator, surface salinity is still relatively high. With increasing depth, the curve swings toward an intermediate salinity value.

For high-latitude regions (such as near Antarctica or in the Gulf of Alaska), the curve begins at the surface with relatively low salinity (again, see the discussion in the preceding section). With increasing depth, the curve also swings toward an intermediate salinity value that approaches the value of the low-latitude salinity curve at the same depth.

These two curves, which together resemble the outline of a wide Champagne glass, show that salinity varies widely at the surface, but very little in the deep ocean. Why is this so? It occurs because all the processes that affect seawater salinity (precipitation, runoff, melting icebergs, melting sea ice, sea ice forming, and evaporation) occur at the *surface* and thus have no effect on deep water below.

FIGURE 5.23 Average surface salinity of the oceans in August. August surface salinity map shows that the lowest salinities (*purple*) occur in high latitudes and the highest salinities (*red and pink*) occur near the tropics, while the equator has a slightly reduced salinity. Values in parts per thousand (‰).



Halocline

Both curves in Figure 5.24 show a rapid change in salinity between the depths of about 300 meters (980 feet) and 1000 meters (3300 feet). For the low-latitude curve, the change is a *decrease* in salinity. For the high-latitude curve, the change is an *increase* in salinity. In both cases, this layer of rapidly changing salinity with depth is called a **halocline** (*halo* = salt, *cline* = slope). Haloclines separate layers of different salinity in the ocean.

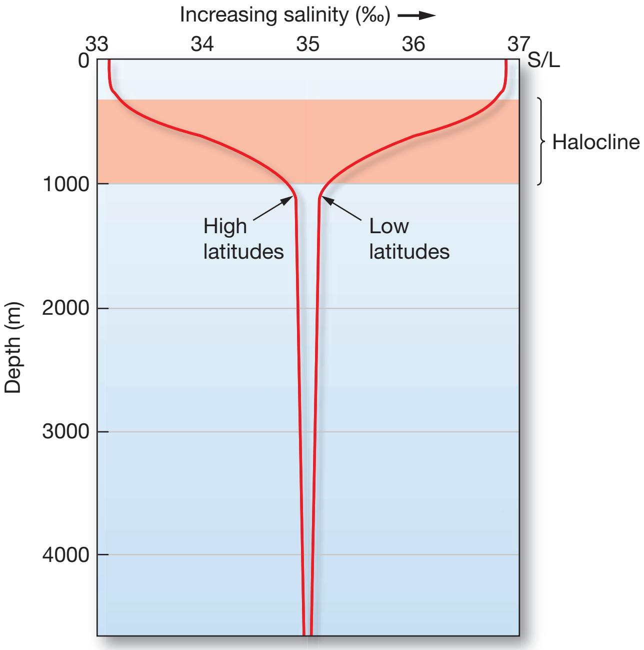


FIGURE 5.24 Salinity variation with depth. Vertical open-ocean profile showing high- and low-latitude salinity variation (horizontal scale in ‰) with depth (vertical scale in meters with sea level at the top). The layer of rapidly changing salinity with depth is the halocline.

5.7 How Does Seawater Density Vary With Depth?

The density of pure water is 1.000 gram per cubic centimeter (g/cm^3) at 4°C (39°F). This value serves as a standard against which the density of all other substances can be measured. Seawater contains various dissolved substances that increase its density. In the open ocean, seawater density averages between 1.022 and 1.030 g/cm^3 (depending on its salinity). Thus, the density of seawater is 2 to 3% greater than pure water. Unlike freshwater, seawater continues to increase in density until it freezes at a temperature of −1.9°C (28.6°F) (recall that below 4°C, the density of freshwater actually *decreases*; see Figure 5.12). At its freezing point, however, seawater behaves in a similar fashion to freshwater: Its density decreases dramatically, which is why sea ice floats, too.

Density is an important property of ocean water because density differences determine the vertical position of ocean water and cause water masses to float or sink, thereby creating deep-ocean currents. For example, if seawater with a density of 1.030 g/cm^3 were added to freshwater with a density of 1.000 g/cm^3, the denser seawater would sink below the freshwater, initiating a deep current.

Factors Affecting Seawater Density

The ocean, like Earth's interior, is layered according to density. Low-density water exists near the surface and higher density water occurs below. Except for some shallow inland seas with a high rate of evaporation that creates high salinity water, the highest-density water is found at the deepest ocean depths. Let's examine how temperature, salinity, and pressure influence seawater density by expressing the relationships using arrows (up arrow = increase, down arrow = decrease):

- As temperature increases (↑), seawater density decreases (↓)[11] (due to thermal expansion).
- As salinity increases (↑), seawater density increases (↑) (due to the addition of more dissolved material).
- As pressure increases (↑), seawater density increases (↑) (due to the compressive effects of pressure).

Of these three factors, only temperature and salinity influence the density of surface water. Pressure influences seawater density only when very high pressures are encountered, such as in deep-ocean trenches. Still, the density of seawater in the deep ocean is only about 5% greater than at the ocean surface, showing that despite tons of pressure per square centimeter, water is nearly incompressible. Unlike air, which can be compressed and put in a tank for scuba diving, the molecules in liquid water are already close together and cannot be compressed much more. Therefore, pressure has the least effect on influencing the density of surface water and can largely be ignored.

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[11] A relationship where one variable *decreases* as a result of another variable's *increase* is known as an inverse relationship, in which the two variables are *inversely proportional*.

KEY CONCEPT

Differences in ocean density cause the ocean to be layered. Seawater density increases with decreased temperature, increased salinity, and increased pressure.

Temperature, on the other hand, has the greatest influence on surface seawater density because the range of surface seawater temperature is greater than that of salinity. In fact, only in the extreme polar areas of the ocean, where temperatures are low and remain relatively constant, does salinity significantly affect density. Cold water that also has high salinity is some of the highest-density water in the world. The density of seawater—the result of its salinity and temperature—influences currents in the deep ocean because high-density water sinks below less-dense water.

Temperature and Density Variation with Depth

The four graphs in Figure 5.25 show how seawater temperature and density vary with depth in both low-latitude and high-latitude regions. Let's examine each graph individually.

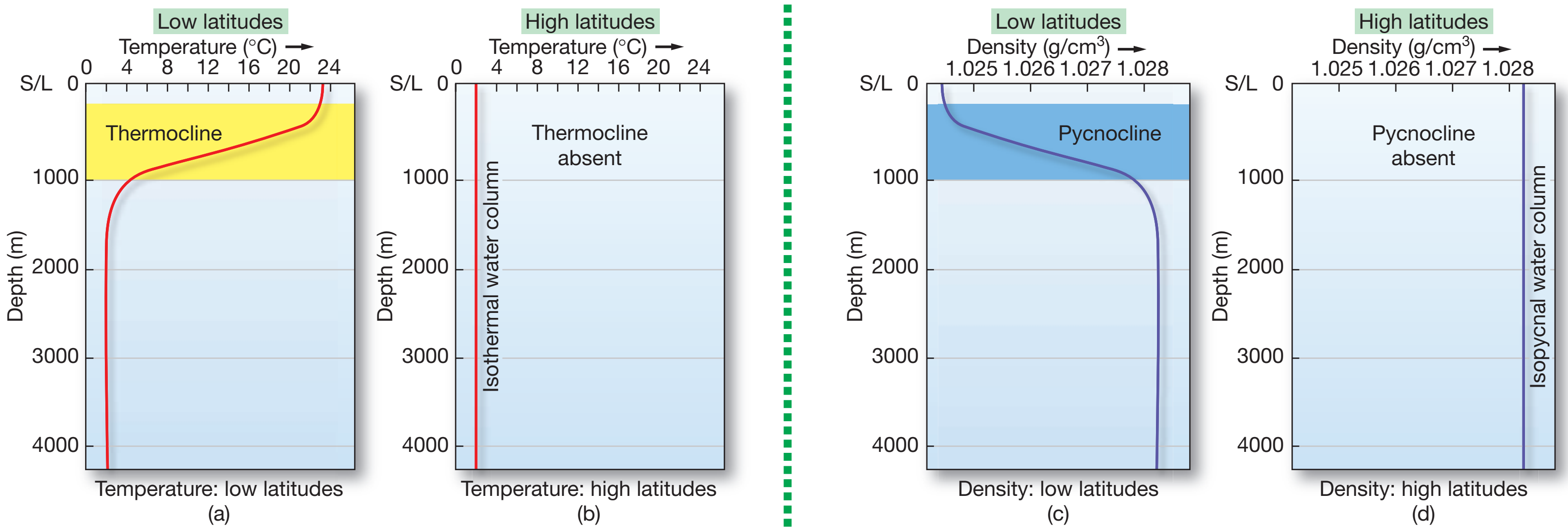
Figure 5.25a shows how temperature varies with depth in low-latitude regions, where surface waters are warmed by high Sun angles and constant length of days. However, the Sun's energy does not penetrate very far into the ocean. Surface water temperatures remain relatively constant until a depth of about 300 meters (980 feet) because of good surface mixing mechanisms such as surface currents, waves, and tides. Below 300 meters (980 feet), the temperature decreases rapidly until a depth of about 1000 meters (3300 feet). Below 1000 meters, the water's low temperature again remains constant down to the ocean floor.

Figure 5.25b shows how temperature varies with depth in high-latitude regions, where surface waters remain cool year-round and deep-water temperatures are about the same as the surface. The temperature curve for high-latitude regions, therefore, is a straight vertical line, which indicates uniform conditions at the surface and at depth.

The density curve for low-latitude regions in Figure 5.25c shows that density is relatively low at the surface. Density is low because surface water temperatures are high. (Remember that temperature has the greatest influence on density and temperature is inversely proportional to density.) Below the surface, density remains constant also until a depth of about 300 meters (980 feet) because of good surface mixing. Below 300 meters (980 feet), the density increases rapidly until a depth of about 1000 meters (3300 feet). Below 1000 meters, the water's low density again remains constant down to the ocean floor.

The density curve for high-latitude regions (Figure 5.25d) shows very little variation with depth. Density is relatively high at the surface because surface water temperatures are low. Density is high below the surface, too, because water

FIGURE 5.25 Temperature and density variations with depth. (a) Temperature variation with depth in low-latitude regions. The layer of rapidly changing temperature with depth is the thermocline. **(b)** Temperature variation with depth in high-latitude regions. Because the water column is isothermal, there is no thermocline. **(c)** Density variation with depth in low-latitude regions. The layer of rapidly changing density with depth is the pycnocline. **(d)** Density variation with depth in high-latitude regions. Because the water column is isopycnal, there is no pycnocline. Note the inverse relationship between temperature and density by comparing curves *a* and *c* (low latitudes) and curves *b* and *d* (high latitudes).



Temperature: low latitudes
(a)

Temperature: high latitudes
(b)

Density: low latitudes
(c)

Density: high latitudes
(d)

temperature is also low. The density curve for high-latitude regions, therefore, is a straight vertical line, which indicates uniform conditions at the surface and at depth. These conditions allow cold high-density water to form at the surface, sink, and initiate deep-ocean currents.

Temperature is the most important factor influencing seawater density, so the temperature graphs (Figure 5.25a and Figure 5.25b) strongly resemble the corresponding density graphs (Figure 5.25c and Figure 5.25d, respectively). The only difference is that they are *a mirror image of each other*, illustrating that temperature and density are inversely proportional to one another.

Thermocline and Pycnocline

Analogous to the halocline (the layer of rapidly changing salinity shown in Figure 5.24), the low-latitude temperature graph in Figure 5.25a displays a curving line that indicates a layer of rapidly changing temperature called a **thermocline** (*thermo* = heat, *cline* = slope). Similarly, the low-latitude density graph in Figure 5.25c displays a curving line that indicates a **pycnocline** (*pycno* = density, *cline* = slope), which is a layer of rapidly changing density. Note that the high-latitude graphs of temperature (Figure 5.25b) and density (Figure 5.25d) lack both a thermocline and a pycnocline, respectively, because these lines show a constant value with depth (they are straight lines and don't curve). Like a halocline, a thermocline and a pycnocline typically occur between about 300 meters (980 feet) and 1000 meters (3300 feet) below the surface. The temperature difference between water above and below the thermocline can be used to generate electricity (see Web Box 5.1).

When a pycnocline is established in an area, it presents an incredible barrier to mixing between low-density water above and high-density water below. A pycnocline has a high gravitational stability and thus physically isolates adjacent layers of water.[12] The pycnocline results from the combined effect of the thermocline and the halocline, because temperature and salinity influence density. The interrelation of these three layers determines the degree of separation between the upper-water and deep-water masses.

There are essentially three distinct water masses based on density. The **mixed surface layer** occurs above a strong permanent thermocline (and corresponding pycnocline). The water is uniform because it is well mixed by surface currents, waves, and tides. The thermocline and pycnocline occur in a relatively low-density layer called the **upper water**, which is well developed throughout the low and middle latitudes. Denser and colder **deep water** extends from below the thermocline/pycnocline to the deep-ocean floor.

At depths above the main thermocline, divers often experience lesser thermoclines (and corresponding pycnoclines) when descending into the ocean. Thermoclines can develop in swimming pools, ponds, and lakes, too. During the spring and fall, when nights are cool but days can be quite warm, the Sun heats the surface water of the pool yet the water below the surface can be quite cold. If the pool has not been mixed, a thermocline isolates the warm surface layer from the deeper cold water. The cold water below the thermocline can be quite a surprise for anyone who dives into the pool!

In high-latitude regions, the temperature of the surface water remains cold year round, so there is very little difference between the temperature at the surface and in deep water below. Thus, a thermocline and corresponding pycnocline rarely develop in high-latitude regions. Only during the short summer when the days are long does the Sun begin to heat surface waters. Even then, the water

KEY CONCEPT

A halocline is a layer of rapidly changing salinity, a thermocline is a layer of rapidly changing temperature, and a pycnocline is a layer of rapidly changing density.

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[12] This is similar to a temperature inversion in the atmosphere, which traps cold (high-density) air underneath warm (low-density) air.

does not heat up very much. Nearly all year, then, the water column in high latitudes is **isothermal** (*iso* = same, *thermo* = heat) and **isopycnal** (*iso* = same, *pycno* = density) allowing good vertical mixing between surface and deeper waters.

5.8 What Methods Are Used to Desalinate SeaWater?

Earth's expanding population uses freshwater in greater volumes each year. In the future, experts warn that more people worldwide will experience water shortage problems, even with increased water use efficiency. Because human need for water is growing yet its supply is dwindling, several countries have begun to use the ocean as a source of freshwater. **Desalination**, or salt removal from seawater, can provide freshwater for business, home, and agricultural use.

Although seawater is mostly just water molecules, its ability to form hydrogen bonds, easily dissolve so many substances, and resist changes in temperature and state makes seawater difficult to desalinate. This is why desalination is so energy-intensive and so expensive. The high cost of desalination, however, is only one issue. Recent studies, for example, indicate that desalination can have unwanted environmental effects through the production of highly concentrated salt brines and potential negative effects on marine life. Still, using the sea as a source of freshwater is attractive to many coastal communities that have few other sources.

Currently, there are more than 13,000 desalination plants worldwide, with the majority very small and located in arid regions of the Middle East, Caribbean, and Mediterranean. The United States produces only about 10% of the world's desalted water, primarily in Florida. To date, only a limited number of desalination plants have been built along the California coast, primarily because the cost of desalination is generally higher than the costs of other water supply alternatives available in California (such as water transfers and groundwater pumping) but also because the extensive permitting process is an impediment to building desalination facilities. However, as drought conditions occur and concern over water availability increases, desalination projects are being proposed at numerous locations in the state.

Because desalinated water requires a lot of energy and thus is expensive to produce, most desalination plants are small-scale operations. In fact, desalination plants provide less than 0.5% of human water needs. More than half of the world's desalination plants use *distillation* to purify water, while most of the remaining plants use *membrane processes*.

Distillation

The process of **distillation** (*distillare* = to trickle) is shown schematically in Figure 5.26. In distillation, saltwater is boiled and the resulting water vapor is passed through a cooling condenser where it condenses and is collected as freshwater. This simple procedure is very efficient at purifying seawater. For instance, distillation of 35‰ seawater produces freshwater with a salinity of only 0.03‰, which is about 10 times fresher than bottled water, so it needs to be mixed with less pure water to make it taste better. Distillation is expensive, however, because it requires large amounts of heat energy to boil the saltwater. Because of water's high latent heat of vaporization, it takes 540 calories to convert only 1 gram (0.035 ounce) of water at the boiling point to the vapor state.[13] Increased efficiency, such as using the waste heat from a power plant, is required to make distillation practical on a large scale.

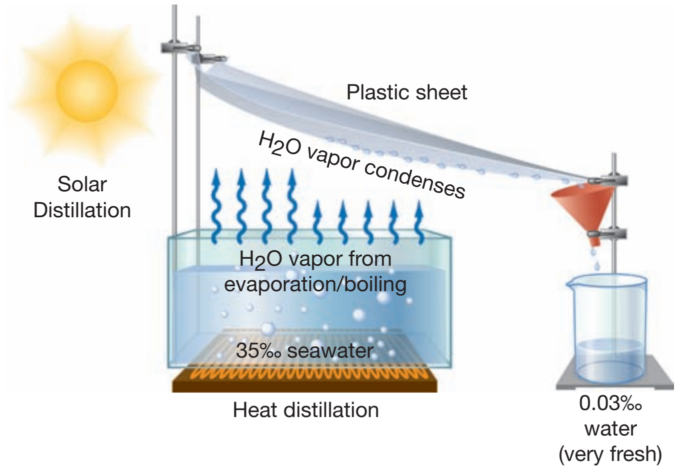

FIGURE 5.26 Distillation. The process of distillation requires boiling saltwater (*heat distillation*) or using the Sun's energy to evaporate seawater (*solar distillation*). In either case, the water vapor is captured and condensed in a process that produces very pure water.

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[13] Even at 100% efficiency, it still requires a whopping *540,000 calories* of heat energy to produce 1 liter (about 1 quart) of distilled water.

Solar distillation, which is also known as **solar humidification**, does not require supplemental heating and has been used successfully in small-scale agricultural experiments in arid regions such as Israel, West Africa, and Peru. Solar humidification is similar to distillation in that saltwater is evaporated in a covered container, but the water is heated by direct sunlight instead (Figure 5.26). Saltwater in the container evaporates, and the water vapor that condenses on the cover runs into collection trays. The major difficulty lies in effectively concentrating the energy of sunlight into a small area to speed evaporation.

Membrane Processes

Electrolysis can be used to desalinate seawater, too. In this method, two volumes of freshwater—one containing a positive electrode and the other a negative electrode—are placed on either side of a volume of seawater. The seawater is separated from each of the freshwater reservoirs by semipermeable membranes. These membranes are permeable to salt ions but not to water molecules. When an electrical current is applied, positive ions such as sodium ions are attracted to the negative electrode, and negative ions such as chloride ions are attracted to the positive electrode. In time, enough ions are removed through the membranes to convert the seawater to freshwater. The major drawback to electrolysis is that it requires large amounts of energy.

Reverse osmosis (*osmos* = to push) may have potential for large-scale desalination. In osmosis, water molecules naturally pass through a thin, semipermeable membrane from a freshwater solution to a saltwater solution. In reverse osmosis, water on the salty side is highly pressurized to drive water molecules—but not salt and other impurities—through the membrane to the freshwater side (Figure 5.27). A significant problem with reverse osmosis is that the membranes are flimsy, become clogged, and must be replaced frequently. Advanced composite materials may help eliminate these problems because they are sturdier, provide better filtration, and last up to 10 years.

Worldwide, at least 30 countries are operating reverse osmosis units. Saudi Arabia—where energy from oil is cheap but water is scarce—has the world's largest reverse osmosis plant, which produces 485 million liters (128 million gallons) of desalted water daily. The largest plant in the United States opened in 2008 in Tampa Bay, Florida, and produces up to 95 million liters (25 million gallons) of freshwater per day, which provides about 10% of the drinking water supply of the Tampa Bay region. Once permits are obtained, a new facility in Carlsbad, California, is designed to produce twice as much freshwater as the Tampa Bay plant. Reverse osmosis is also used in many household water purification units and aquariums.

Other Methods of Desalination

Seawater selectively excludes dissolved substances as it freezes—a process called **freeze separation**. As a result, the salinity of sea ice (once it is melted) is typically 70% lower than seawater. To make this an effective desalination technique, though, the water must be frozen and thawed multiple times, with the salts washed from the ice between each thawing. Like electrolysis, freeze separation requires large amounts of energy, so it may be impractical except on a small scale.

Yet another way to obtain freshwater is to melt naturally formed ice. Imaginative thinkers have proposed towing large icebergs to coastal waters off countries that need freshwater. Once there, the freshwater produced as the icebergs melt could be captured and pumped ashore. Studies have shown that towing large Antarctic icebergs to arid regions would be technologically feasible and, for certain Southern Hemisphere locations, economically feasible, too.

Other novel approaches to desalination include crystallization of dissolved components directly from seawater, solvent demineralization using chemical catalysts, and even making use of salt-eating bacteria!

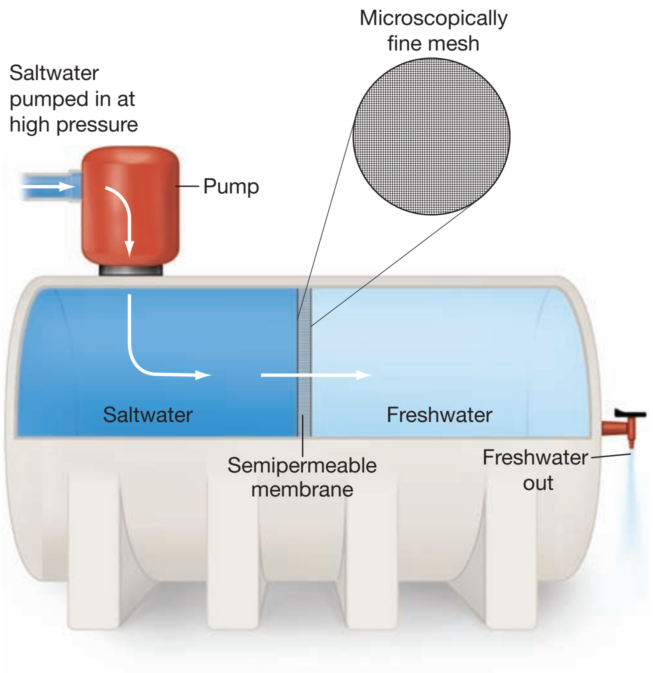

FIGURE 5.27 Reverse osmosis. The process of reverse osmosis involves applying pressure to salt water and forcing it through a semipermeable membrane, thus removing the salts and producing freshwater.

KEY CONCEPT

Although desalination of seawater is costly, desalination plants use the methods of distillation, solar humidification, electrolysis, freeze separation, and reverse osmosis to purify seawater for domestic use.

Chapter in Review

- *Water's remarkable properties help make life as we know it possible on Earth*. These properties include the arrangement of its atoms, how its molecules stick together, its ability to dissolve almost everything, and its heat storage capacity.
- *The water molecule is composed of one atom of oxygen and two atoms of hydrogen* (H_2O). The two hydrogen atoms, which are covalently bonded to the oxygen atom, are *attached to the same side of the oxygen atom* and produce a bend in the geometry of a water molecule. This geometry makes water molecules *polar*, which allows them to form hydrogen bonds with other water molecules or other substances and gives water its remarkable properties. Water, for example, is *"the universal solvent"* because it can hydrate charged particles (ions), thereby dissolving them.
- *Water is one of the few substances that exists naturally on Earth in all three states of matter* (solid, liquid, gas). Hydrogen bonding gives water *unusual thermal properties*, such as a high freezing point (0°C [32°F]) and boiling point (100°C [212°F]), a high heat capacity and high specific heat (1 calorie per gram), a high latent heat of melting (80 calories per gram), and a high latent heat of vaporization (540 calories per gram). Water's high heat capacity and latent heats have important implications in regulating global thermostatic effects.
- *The density of water increases as temperature decreases*, similar to most substances, and reaches a maximum density at 4°C (39°F). *Below 4°C, however, water density decreases with temperature*, due to the formation of bulky ice crystals. *As water freezes, it expands by about 9%* in volume, so *ice floats* on water.
- *Salinity is the amount of dissolved solids in ocean water*. It averages about 35 grams of dissolved solids per kilogram of ocean water (*35 parts per thousand* [‰]) but ranges from brackish to hypersaline. Six ions—chloride, sodium, sulfate, magnesium, calcium, and potassium—account for over 99% of the dissolved solids in ocean water. These ions always occur in a *constant proportion* in any seawater sample, so salinity can be determined by measuring the concentration of only one—typically, the chloride ion.
- *The physical properties of pure water and seawater are remarkably similar*, with a few notable exceptions. Compared to pure water, seawater has a higher pH, density, and boiling point (but a lower freezing point).
- *Dissolved components in seawater are added and removed by a variety of processes*. Precipitation, runoff, and the melting of icebergs and sea ice add freshwater to seawater and decrease its salinity. The formation of sea ice and evaporation remove freshwater from seawater and increase its salinity. The hydrologic cycle includes all the reservoirs of water on Earth, including the *oceans*, which contain *97% of Earth's water*. The residence time of various elements indicates how long they stay in the ocean and implies that *ocean salinity has remained constant through time*.
- *A natural buffering system* based on the chemical reaction of carbon dioxide in water *exists in the ocean*. This buffering system *regulates any changes in pH*, creating a stable ocean environment.
- *The salinity of surface water varies considerably due to surface processes*, with the maximum salinity found near the Tropics of Cancer and Capricorn and the minimum salinity found in high-latitude regions. Salinity also varies with depth down to about 1000 meters (3300 feet), but below that the salinity of deep water is very consistent. *A halocline is a layer of rapidly changing salinity*.
- *Seawater density increases as temperature decreases and salinity increases*, though temperature influences surface seawater density more strongly than salinity (the influence of pressure is negligible). Temperature and density vary considerably with depth in low-latitude regions, creating a *thermocline* (*layer of rapidly changing temperature*) and corresponding *pycnocline* (*layer of rapidly changing density*), both of which are generally absent in high latitudes.
- Although *desalination of seawater* is costly, it *provides freshwater* for business, home, and agricultural use. *Distillation, solar humidification, electrolysis, freeze separation*, and *reverse osmosis* are methods currently used to desalinate seawater.

Key Terms

Acid (p. 149)
Alkaline (p. 149)
Atom (p. 129)
Base (p. 149)
Boiling point (p. 133)
Brackish (p. 144)
Buffering (p. 150)
Calorie (p. 132)
Challenger, HMS (p. 142)
Chlorinity (p. 141)
Cohesion (p. 131)
Condensation point (p. 133)
Condense (p. 133)
Continental effect (p. 137)
Covalent bond (p. 130)
Deep water (p. 155)
Desalination (p. 156)
Dipolar (p. 130)
Distillation (p. 156)
Electrolysis (p. 157)
Electron (p. 130)
Electrostatic attraction (p. 131)
Evaporation (p. 135)
Freeze separation (p. 157)
Freezing point (p. 133)
Goiter (p. 141)
Halocline (p. 153)
Heat (p. 132)
Heat capacity (p. 134)
Hydrogen bond (p. 131)
Hydrologic cycle (p. 145)
Hypersaline (p. 144)
Ion (p. 130)
Ionic bond (p. 131)
Isopycnal (p. 156)
Isothermal (p. 156)
Kinetic energy (p. 132)
Latent heat of condensation (p. 135)
Latent heat of evaporation (p. 135)
Latent heat of freezing (p. 136)
Latent heat of melting (p. 135)
Latent heat of vaporization (p. 135)
Marine effect (p. 137)
Melting point (p. 133)
Mixed surface layer (p. 155)
Molecule (p. 130)
Neutral (p. 150)
Neutron (p. 130)
Nucleus (p. 129)
Parts per thousand (‰) (p. 139)
pH scale (p. 149)
Polarity (p. 130)
Potential energy (p. 132)
Precipitation (p. 136)
Principle of constant proportions (p. 140)
Proton (p. 130)
Pycnocline (p. 155)
Residence time (p. 147)
Reverse osmosis (p. 157)

Runoff (p. 144)
Salinity (p. 138)
Salinometer (p. 141)
Solar distillation (p. 157)
Solar humidification (p. 157)
Specific heat (p. 134)
Surface tension (p. 131)
Temperature (p. 133)
Thermal contraction (p. 137)
Thermocline (p. 155)
Thermostatic effect (p. 136)
Upper water (p. 155)
van der Waals force (p. 132)
Vapor (p. 133)

Review Questions

1. Sketch a model of an atom, showing the positions of the subatomic particles protons, neutrons, and electrons.

2. Describe what condition exists in water molecules to make them dipolar.

3. Sketch several water molecules, showing all covalent and hydrogen bonds. Be sure to indicate the polarity of each water molecule.

4. How does hydrogen bonding produce the surface tension phenomenon of water?

5. Discuss how the dipolar nature of the water molecule makes it such an effective solvent of ionic compounds.

6. Why are the freezing and boiling points of water higher than would be expected for a compound of its molecular makeup?

7. How does the specific heat capacity of water compare with that of other substances? Describe the effect this has on climate.

8. The heat energy added as latent heat of melting and latent heat of vaporization does not increase water temperature. Explain why this occurs and where the energy is used.

9. Why is the latent heat of vaporization so much greater than the latent heat of melting?

10. Describe how excess heat energy absorbed by Earth's low-latitude regions is transferred to heat-deficient higher latitudes through a process that uses water's latent heat of evaporation.

11. As water cools, two distinct changes take place in the behavior of molecules: Their slower movement tends to increase density, whereas the formation of bulky ice crystals decreases density. Describe how the relative rates of their occurrence cause pure water to have a temperature of maximum density at 4°C (39.2°F) and make ice less dense than liquid water.

12. What is your state sales tax, in parts per thousand?

13. What are goiters? How can they be avoided?

14. What condition of salinity makes it possible to determine the total salinity of ocean water by measuring the concentration of only one constituent, the chloride ion?

15. List some major achievements of the voyage of HMS *Challenger*.

16. What physical conditions create brackish water in the Baltic Sea and hypersaline water in the Red Sea?

17. Describe the ways in which dissolved components are added and removed from seawater.

18. List the components (reservoirs) of the hydrologic cycle that hold water on Earth and the percentage of Earth's water in each one. Describe the processes by which water moves among these reservoirs.

19. Explain the difference between an acid and an alkali (base) substance. How does the ocean's buffering system work?

20. Why is there such a close association between (a) the curve showing seawater density variation with ocean depth and (b) the curve showing seawater temperature variation with ocean depth?

Critical Thinking Exercises

1. Describe the differences between the three states of matter, using the arrangement of molecules in your explanation.

2. Explain why there is such a wide variation of surface salinity but such a narrow range of salinity at depth.

3. Compare and contrast the following seawater desalination methods: distillation, solar humidification, and reverse osmosis.

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