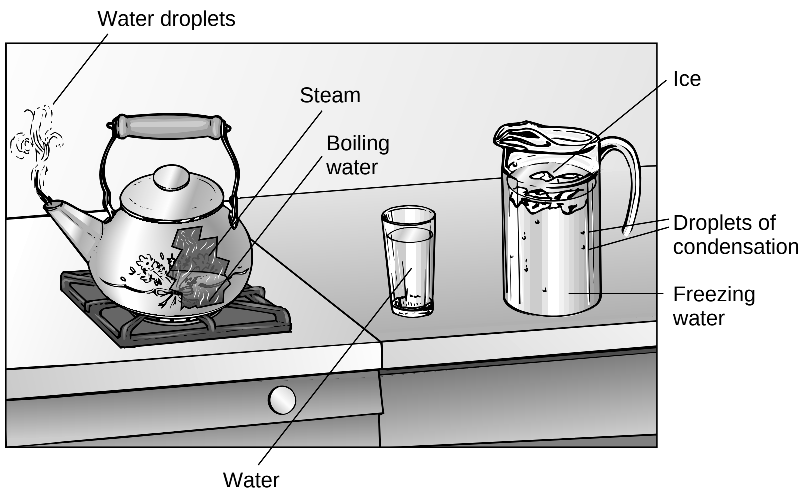

Section 15.1

Water, Steam, and Ice

Water is probably the single most important chemical in our daily lives. It plays such a central role in biology, climate, commerce, industry, and entertainment that it merits a whole section of its own. It also exhibits the three classic phases of matter, solid, liquid, and gas, so that much of what we learn about water is applicable to other materials as well. However, there are a few aspects of water that are almost unique in nature. Water really is a remarkable substance.

Questions to Think About: *Why does adding ice make a drink cold? Why is ice so slippery? How does snow disappear from the ground even when it's very cold outside? Why do icebergs float on water? How does perspiration cool you off? What is the difference between evaporating and boiling?*

Experiments to Do: *Experimenting with water is easy. Pick up an ice cube with wet fingers. What happens if the ice cube has just come out of the freezer? What if it has been melting on the table for a few minutes? Why is there a difference? Which ice cube feels the most slippery when you touch it with dry fingers? Put both ice cubes in water. Why do they float?*

Now heat tap water in a pot. Shortly before the water starts to boil, you will see mist begin to form above it. What is this mist and why does it form? Carefully feel the mist, but don't burn yourself! The mist feels damp because it contains water. How can water be leaving the pot before the water boils? Notice the small gas bubbles on the walls of the pot. It's not steam. Where did this gas come from?

Once the water boils, gaseous water or steam will appear in the mist. Don't touch this steam because it can burn you quickly. Why does steam release so much thermal energy when it touches your skin?

Solid, Liquid, and Gas: the Phases of Matter

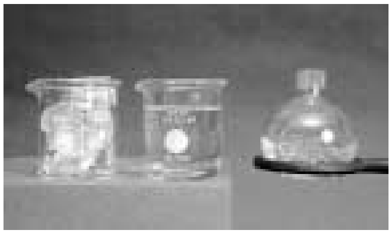

Fig. 15.1.1 - The three phases of water: solid (ice), liquid (water), and gas (steam).

Like most substances, water exists in three distinct forms or **phases**: solid ice, liquid water, and gaseous steam (Fig. 15.1.1). These phases differ in how easily their shapes and volumes can be changed. Solid ice is rigid and incompressible, preventing you from altering an ice cube’s shape or volume. Liquid water is an incompressible fluid, so that you can reshape a water drop but can’t change its volume. Gaseous steam is a compressible fluid, so that you can vary both the shape and volume of the steam in a teakettle.

These different characteristics reflect the different microscopic structures of steam, water, and ice. *Steam* is a gas, a collection of independent molecules that is also called *water vapor*. These water molecules bounce around their container, periodically colliding with one another or with the walls. The water molecules fill the container uniformly and easily accommodate any changes in its shape or size. Enlarging the container simply decreases the steam’s density and lowers its pressure.

To remain independent of one another, the water molecules in steam need a certain amount of thermal energy. Without this energy, they stick together to form water. *Water* is a liquid, a disorderly, fluid collection of molecules that is held together by chemical bonds. These bonds pack the molecules tightly together and give water its fixed volume. However, water still has enough thermal energy for its individual molecules to separate briefly and form new bonds with different partners. Its evolving microscopic structure makes water fluid. While its volume is fixed, its shape is not.

When the water molecules have even less thermal energy, they are unable to rearrange and cling together stiffly as ice. *Ice* is a solid, a rigid collection of chemically bound molecules. Like most solids, ice is crystalline, with orderly arrangements of molecules that extend over large distances. This order produces the beautiful crystalline facets of snowflakes and frost.

Just as an orderly arrangement of cannon balls takes up less volume than a disorderly one, a crystalline solid usually occupies less volume than its corresponding liquid. That’s why the solid phase of a typical substance sinks in the liquid phase of that same substance.

But there is an exception to this rule: water. Ice’s crystalline structure is unusually open and its density is surprisingly low. Almost unique in nature, solid ice is slightly less dense than liquid water so that ice floats on water. That’s why icebergs float on the open ocean and ice cubes float in your drink. In fact, water reaches its greatest density at about 5 °C (40 °F). Heated above that temperature, water behaves normally and expands. However water also expands as you cool it below that temperature, a very unusual effect.

CHECK YOUR UNDERSTANDING #1: Skating Anyone?
As a pond freezes, where does the ice begin to form?

Melting Ice and Freezing Water

When you heat ice, it remains solid until its rising temperature reaches 0 °C. At that point, the ice stops getting warmer and begins to melt. **Melting** is a **phase transition**, a transformation from the ordered solid phase to the disordered liquid phase. This transition occurs when heat breaks some of the chemical bonds between water molecules and permits the molecules to move about. The ice transforms into water, losing its rigid shape and crystalline structure.

0 °C is ice’s **melting temperature**, the temperature at which heat goes into

breaking bonds and converting ice into water, rather than making the ice hotter. The ice-water mixture remains at 0 °C until all of the ice has melted. When only water remains, heating it can again cause its temperature to rise.

The heat used to transform a certain mass of solid into liquid, without changing its temperature, is called the **latent heat of melting**. The bonds between water molecules are relatively strong, so that water has an enormous latent heat of melting: it takes about 333,000 J of heat to convert 1 kg of ice at 0 °C into 1 kg of water at 0 °C. That same amount of heat would raise the temperature of 1 kg of liquid water by about 80 °C so that it takes almost as much heat to melt an ice cube as it does to warm the resulting water all the way to boiling.

The latent heat of melting reappears when you cool the water back to its melting temperature and it begins to **solidify**. As you remove heat from water at 0 °C, the water freezes into ice rather than becoming colder. Because the water molecules release energy as they bind together to form ice crystals, the water releases heat as it freezes. The heat released when transforming a certain mass of liquid into solid, without changing its temperature, is again the latent heat of melting. You must add a certain amount of heat to ice to melt it and you must remove that same amount of heat from water to solidify it.

Ice's huge latent heat of melting is what keeps a mixture of water and ice at 0 °C. As long as both water and ice are present together in your glass, they are in the process of either melting or freezing. Any heat you add to the glass goes into melting more ice, not into raising its temperature. Any heat you remove from the glass comes from freezing more water, not from lowering its temperature. With ice floating in your drink, it will remain at 0 °C, even in the hottest or coldest weather (Fig. 15.1.2).

Fig. 15.1.2 - Ice and water can coexist only at 0 °C, ice's melting temperature.

CHECK YOUR UNDERSTANDING #2: Keeping Crops Warm

Fruit growers often spray their crop with water to protect it from freezing in unusually cold weather. How does liquid water keep the fruit from freezing?

Sublimation, Evaporation, and Condensation

The surface of solid ice or liquid water is a busy place. Water molecules are constantly leaving it as water vapor and returning to it as ice or water. When molecules leave the surface, the water is *evaporating* or the ice is *subliming*. When molecules return to the surface, the water vapor is *condensing*. This simple picture of water molecules taking off and landing on the surface of ice or water explains many familiar phenomena. But to complete the picture, we must follow the flow of energy in this system.

Ice and water both contain thermal energy, which is exchanged between neighboring molecules and keeps them in motion. While the average water molecule is unable to break free from the surface, molecules occasionally obtain enough thermal energy from their neighbors to break their bonds and leave as water vapor. In doing so, these molecules carry away more than their fair share of the water or ice's thermal energy and it becomes colder. Your body uses this effect to keep cool on hot summer days, when perspiration that evaporates from your skin draws heat from you and lowers your temperature.

If you add enough heat to water or ice, it will evaporate or sublime without getting colder. The heat needed to transform a certain mass of solid or liquid into gas, without changing its temperature, is called the **latent heat of vaporization**. Water's latent heat of vaporization is truly enormous because water molecules are surprisingly hard to separate. About 2,300,000 J of heat are needed to convert 1 kg of water at 100 °C into 1 kg of water vapor or steam at 100 °C. That same amount of heat would raise the temperature of 1 kg of water by more than

500 °C! Even more heat is required to convert ice directly into water vapor.

Although we are more familiar with the evaporation of water, we occasionally observe the sublimation of ice. For example, snow gradually disappears from the ground without ever melting and food that is left unprotected in a frostless freezer gradually dries out. While this “freezer burn” is a nuisance at home, sublimation from frozen food is used commercially to prepare freeze-dried food.

Evaporation and sublimation are most rapid when there isn’t much water vapor above the water or ice. When the air above the water or ice is full of water molecules, they collide frequently with the surface and experience its attractive forces. Many molecules stick and the water vapor condenses into water or ice.

The latent heat of vaporization reappears when water vapor condenses. The gaseous water molecules bind together as water or ice and release their chemical potential energy as heat. The heat released when transforming a certain mass of gas into liquid or solid, without changing its temperature, is again the latent heat of vaporization. You must add a certain amount of heat to water or ice to vaporize it and you must remove that same amount of heat from water vapor to liquefy or solidify it.

The huge amount of heat released by condensing water vapor is often used to cook food or warm radiators in older buildings. When you steam vegetables, you are allowing water vapor to condense on the vegetables and transfer heat to them. A double-boiler uses condensing steam to transfer heat from a burner to a cooking container in a controlled manner.

CHECK YOUR UNDERSTANDING #3: Tea Time
A kettle of water heats up rapidly on the stove but takes quite a while to boil away. Why does the water take so long to turn into steam?

Relative Humidity

Something is missing from our description of evaporation, sublimation, and condensation. Neglecting ice for the moment, we have seen that sometimes water molecules leave the surface of water and sometimes they return. What determines whether water molecules are leaving or returning?

The answer is that both processes usually occur simultaneously. The airport picture for the surface of water, with water molecules constantly taking off and landing, is a pretty accurate one. There is a continual exchange of water molecules between the gaseous and liquid phases. What matters most is whether there is a net transfer from one phase to the other. If more molecules leave the surface than return, the water is evaporating. If more molecules return to the surface than leave, the water vapor is condensing.

The simplest indicator of whether water molecules will evaporate or condense is relative humidity. **Relative humidity** measures the returning rate as a percentage of the leaving rate. When the relative humidity is 100%, the two rates are equal and there is neither evaporation nor condensation. If the relative humidity is less than 100%, the returning rate is less than the leaving rate and the water evaporates. If the relative humidity is more than 100%, the returning rate is more than the leaving rate and the water vapor condenses.

Relative humidity depends on the temperature and on the density of water vapor in the air. The higher the temperature, the more frequently water molecules leave the water’s surface as water vapor. Thus, increasing the temperature reduces the relative humidity. But the higher the density of water vapor, the more frequently water molecules collide with the surface and become liquid. Therefore, increasing the density of water vapor in the air increases the relative humidity.

Relative humidity plays an important role in countless experiences of everyday life. When the relative humidity is low, water evaporates quickly and the air feels dry. Perspiration cools you effectively. When the relative humidity is high (near 100%), water barely evaporates at all and the air feels damp. Perspiration clings to your skin and doesn't cool you much. The cooling of a wet surface can actually be used to measure relative humidity (Fig. 15.1.3).

When the relative humidity exceeds 100%, perhaps because of a sudden loss of temperature, water vapor begins to condense everywhere. Water droplets grow on surfaces as dew or form directly in the air as fog, mist, or clouds. When the humidity remains high, these droplets grow larger and eventually fall as rain. If the temperature is cold enough, water vapor can condense directly to ice, forming frost on surfaces or snowflakes or hail in the air.

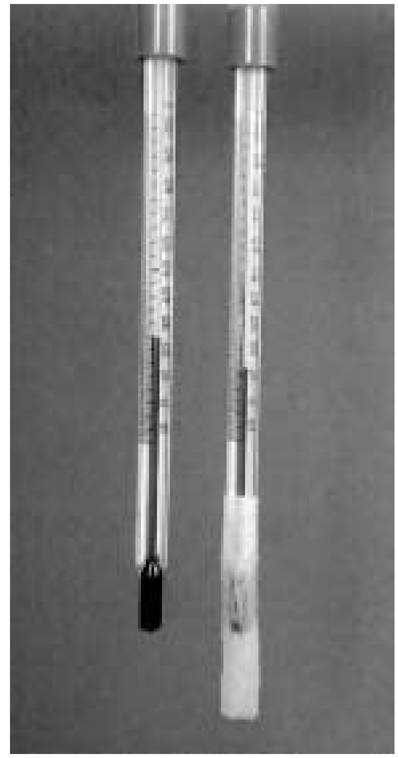

Fig. 15.1.3 - You can determine the air's relative humidity using two thermometers—one of which has a wet cloth wrapped around its bulb (right). Evaporation cools the wet bulb thermometer by an amount related to the air's relative humidity. The dryer the air, the colder the wet bulb thermometer becomes.

CHECK YOUR UNDERSTANDING #4: Seeing Your Breath
When you breathe out on a cold day, you often see mist appearing from your mouth. Explain.

Boiling Water

If you seal water inside a container and keep its temperature constant, it will evaporate until the relative humidity inside the container reaches 100%. At that point, water molecules will return to the liquid as often as they leave it. The air will be **saturated** with water, containing just the right density of water molecules to ensure this balanced exchange.

Water molecules also generate part of the air pressure in the container. Water's contribution to the overall air pressure depends on its density in the air. When the air is saturated with water, water contributes its **saturated vapor pressure** to the air's overall pressure.

Water's saturated vapor pressure depends only on its temperature. The hotter the container, the more water vapor it will contain and the higher the water's saturated vapor pressure. Near room temperature, water's saturated vapor pressure is much less than atmospheric pressure. But as water approaches 100 °C, its saturated vapor pressure approaches atmospheric pressure. Something special can happen right when water's saturated vapor pressure reaches atmospheric pressure. The water can begin to boil (Fig. 15.1.4).

Boiling occurs when bubbles of water vapor appear *inside* liquid water. A bubble forms when a number of water molecules simultaneously break free of their neighbors and create a tiny pocket of gaseous water inside the liquid water. The sudden appearance of this bubble allows evaporation to occur inside the water. More water molecules evaporate into the bubble and it grows in size and mass. Since gaseous water is much less dense than liquid water, the bubble floats upward to the water's surface.

The only gas present inside these bubbles is water vapor. If the water isn't hot enough, its saturated vapor pressure will be less than atmospheric pressure and any bubbles that begin to form will be crushed immediately by the surrounding water. Only when the saturated vapor pressure of water reaches atmospheric pressure can a steam bubble survive long enough to grow, float upward, and leave the surface of the water. When that occurs, the water is **boiling**. The temperature at which water's saturated vapor pressure reaches atmospheric pressure is water's **boiling temperature**.

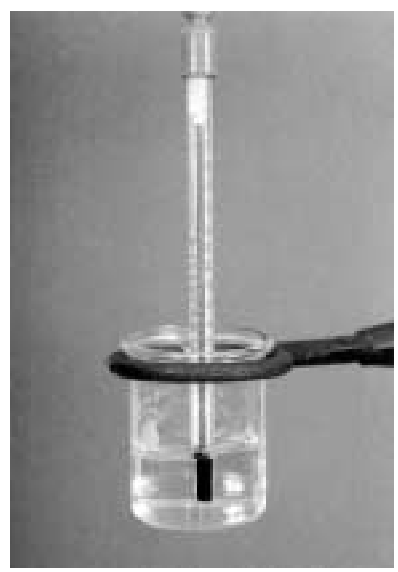

Fig. 15.1.4 - Water boils at 100 °C, when atmospheric pressure can no longer smash the bubbles of water vapor.

Boiling converts water to steam much more rapidly than evaporation, so it takes an enormous input of heat to keep water boiling. Once water has reached its boiling temperature, the conversion of liquid water to gaseous steam consumes all of the heat you add to the water. In an open pot, the water's temperature stops increasing and remains constant at its boiling temperature until all of the liquid water is gone. Only when the water is entirely gaseous does its temperature begin to increase.

The constant, well-defined temperature of boiling water allows you to cook vegetables or an egg at a particular rate. When you place an egg in boiling water, it cooks in 3 minutes because it's in contact with water at its boiling temperature.

But water's boiling temperature depends on the ambient pressure above the water. For an open pot or pan, that pressure is atmospheric pressure. However atmospheric pressure decreases with altitude and depends somewhat on the weather. While water boils at 100 °C near sea level, it boils at only 90 °C at an altitude of 3,000 m. This reduction in the water's boiling temperature with altitude explains why many recipes must be adapted for use at higher elevations. At 3,000 m, an egg cooks slowly in boiling water because it's surrounded by 90 °C water, not 100 °C water. The same problem slows the cooking of rice, beans, and many other foods at high altitudes.

In fact, at sufficiently low pressures, water boils even at room temperature. With a vacuum chamber, you can actually get a glass of ice water to boil. At the other extreme, high pressures can make water boil at very high temperatures. In a boiler that produces high-pressure steam for a power plant, water may boil at a temperature exceeding 300 °C.

One way to decrease cooking times is to use a pressure cooker. A pressure cooker seals in the steam so that the pressure inside can exceed atmospheric pressure. This increased pressure prevents boiling from occurring until the water's temperature is well above 100 °C. If you subject water to twice sea-level atmospheric pressure, it doesn't boil until it reaches 121 °C. An egg cooks very quickly at that temperature, as do vegetables and other foods.

One final note about boiling: it doesn't always happen. Just because water *can* boil, doesn't mean that it *will* boil. Boiling requires not only that steam bubble be stable but also that they form in the first place. Those first few gaseous water molecules, the nucleus of a bubble, rarely come together naturally in water near its boiling temperature. Since such spontaneous bubble formation is so unlikely, it usually requires assistance. Most of the seed bubbles required for boiling form or *nucleate* only with the help of defects on the container, contaminants in the liquid, or hotspots on the cooking surfaces. That's why the bubbles of boiling water, like those in soda or champagne, often stream upward from specific spots on their container—the nucleation sites.

When you heat water uniformly in a clean, glass container, it may not begin to boil properly at its boiling temperature. Although a few steam bubbles may form out of air bubbles caught in the water, the bubbling then ceases. That's because without nucleation sites, no true seed bubbles ever form. Instead, the water superheats—its temperature rises above the boiling temperature but without it actually boiling.

Superheated water, which forms easily and often in a microwave oven, can be extremely dangerous. The slightest trigger—a fork, a pinch of sugar or salt, or even a tap on the container—can initiate violent or even explosive boiling. The more the water's temperature exceeds the boiling temperature, the more energy it can release suddenly if it abruptly begins to boil. Be careful when you heat water in a microwave oven, particularly in a glass container. If it doesn't appear to be boiling properly despite being very hot, consider that it may be superheated. Your safest bet is to stay away from it until it has cooled down.

CHECK YOUR UNDERSTANDING #5: Getting Into Not-So-Hot Water

If you put a thermometer in boiling water at the top of a 3,000 m mountain, what temperature will the thermometer read?

Why Ice Is Slippery

Pressure also affects the melting of ice. When you squeeze ice, you lower its melting temperature slightly so that it melts more easily. This behavior, in which pressure can induced melting, is extremely unusual. Almost every other material in nature becomes more difficult to melt as you squeeze it. That's because most materials expand during melting and have to do work against the pressure surrounding them. Added pressure makes it harder for them to melt and you can even make them freeze by squeezing them. But ice shrinks during melting because ice is less dense than water. Rather than keeping ice from melting, pressure can actually liquefy ice, even below its normal melting temperature.

But while pressure-induced melting has long been used to explain ice's slipperiness (i.e. that stepping on ice causes its surface to melt so that you find yourself standing on a thin layer of liquid), that explanation appears to be incorrect. Instead, ice's slipperiness is now thought to be caused by frictional melting of ice's first few molecular layers. Such melting is particularly easy at the surface of ice, where the crystalline structure is incomplete and disordered. Because they lack a full complement of neighbors, the outermost water molecules are relatively mobile and already have a liquid-like character. When heated by sliding friction, this layer melts entirely and acts as a lubricant to make ice extremely slippery. Since the layer is so thin, very little heat is needed to melt it and a tiny bit of frictional heating is all it takes to get something sliding along the ice.

Temperature affects both ice's liquid-like surface and this frictional melting. At warmer temperatures, the liquid-like layer is thicker and melts more easily when heated by friction. That's why ice is most slippery when the temperature is close to freezing and the ice has a “wet” surface. When the temperature is extremely cold, ice has a “dry” surface and is barely slippery at all.

Ice's liquid-like surface also permits ice crystals to fuse together easily, so that snowflakes pack together into snowballs or glaciers. Here, too, the thickness of the liquid-like surface is important. It's hard to make snowballs when the weather is cold and the snow is dry. Snow “packs” best when its temperature is just below freezing.

These behaviors are very unusual. Most solids aren't slippery and don't stick together well. Their surfaces are solid and don't melt easily due to frictional heating. Instead, they experience strong frictional forces when they slide across one another. They also don't fuse together on contact. Imagine “ice skating” on a slate sidewalk or trying to make a ball by packing dry sand together. Not likely.

CHECK YOUR UNDERSTANDING #6: Grooming the Ice
The ice of a speed skating rink is kept very close to ice's melting temperature, so that the skaters can travel as fast as possible. Explain.

Dissolving Chemicals in Water

Many chemicals dissolve in water. Those that dissolve best contain atoms, molecules, or other particles that bind strongly to water molecules, more strongly than they bind to themselves. When you add a water-soluble chemical to water, water molecules surround its particles and carry them around individually. The chemical becomes separate particles and **dissolves** in the water.

To understand this process better, let's look at the structure of water itself. A water molecule consists of two hydrogen atoms and an oxygen atom. For reasons that we discussed in Chapter 11 and will continue below, the water mole-

❒ Concentrated sugar water boils at a much higher temperature than pure water. Boiling temperature is often used as a measure of sugar concentration when preparing candies. The higher the boiling temperature of the mixture, the less water there is in it. Eventually, the water evaporates completely and the temperature rises rapidly toward the melting and caramelizing temperatures of sugar itself.

cule has electrically charged ends. Its oxygen atom is negatively charged while its hydrogen atoms are positively charged. Since opposite electric charges attract, the oxygen atom of one water molecule tends to stick to a hydrogen atom of another water molecule. This attachment is called a **hydrogen bond**. Hydrogen bonds are quite strong and the large energy needed to break them explains water's enormous latent heats of melting and vaporization.

Chemicals that dissolve well in water are those that have electric charges of their own. Salt dissolves easily in water because it consists of **ions**, electrically charged atoms or molecules, that attract the water molecules. When you put salt in water, it breaks up into its constituent ions and the water molecules carry around these ions. Sugar dissolves well in water because its oxygen and hydrogen atoms are electrically charged, just as they are in water molecules. Sugar molecules form hydrogen bonds with water molecules and are carried around in the water.

There are also a few gases that dissolve well in water because they bind with water molecules. Carbon dioxide gas is quite soluble in water, where it produces carbonated water or soda. But the second law of thermodynamics allows even those gases that aren't attracted to water molecules to dissolve. They dissolve in small amounts because they increase the overall disorder and entropy of the liquid. This increase in entropy is most important in cold water, so that it dissolves more gas than hot water. When you heat cold water, its dissolved gases often reappear as bubbles on the walls of the pot.

The presence of dissolved chemicals in water lowers its melting temperature. This effect is a consequence of the second law of thermodynamics. Ice crystals are essentially pure water so the chemical-water mixture separates during freezing. Since this separation lowers the mixture's entropy, an additional source of order is needed to satisfy the second law. The extra order is provided by cooling the mixture below ice's normal melting temperature. That's why salt water freezes at a lower temperature than fresh water. And because a mixture of salt and water tends to be liquid at temperatures somewhat below water's normal melting temperature, you can melt ice on the sidewalk by sprinkling it with salt. Actually, any chemical that dissolves easily in water will melt ice. Even sugar will do the trick.

Dissolved chemicals also raise water's boiling temperature. The chemicals interfere with evaporation, changing the balance between evaporation and condensation and lowering water's saturated vapor pressure. The mixture won't boil until water's saturated vapor pressure reaches the ambient pressure, so you must heat the mixture above water's normal boiling temperature (see ❒). Only if you add a chemical that boils more easily than water will the mixture boil at or below its normal boiling temperature.

CHECK YOUR UNDERSTANDING #7: Safe Sidewalks
Why does salt melt ice while sand does not?

Water Molecules

Water is a simple molecule, consisting of one oxygen atom and two hydrogen atoms. It's frequently written in chemical shorthand as H_2O, where the H_2 refers to the two hydrogen atoms and the O refers to the oxygen atom.

Oxygen is a relatively small and simple atom. It has a nucleus containing 8 positively charged protons and at least 8 uncharged neutrons, surrounded by an orbiting contingent of 8 negatively charged electrons. For complicated reasons arising from quantum physics, the oxygen atom tends to borrow or share two additional electrons to complete a structure called an **electronic shell**. Completing electronic shells is a common theme in atomic and molecular systems because completed shells tend to minimize the available potential energy of a system and make it very stable. This goal of completing shells resembles brick patio construction, where missing or extra bricks make the patio less robust.

The tendency to complete its electronic shell is so strong that an oxygen atom will often share two of its outermost or **valence electrons** with neighboring atoms in order to have them share two of their electrons with it. This need to complete its shell explains why oxygen atoms typically pair up to form oxygen molecules, such as those that you are presently breathing. Each oxygen atom in an oxygen molecule has shared two of its electrons with the other atom, so that both atoms have completed their electronic shells (Fig. 15.1.5a). This sharing of electrons between atoms is common in nature and is called **covalent bonding**.

Hydrogen is the smallest and simplest atom of all. Its nucleus contains only one positively charge proton and that nucleus is orbited by a single electron. To complete a smaller electronic shell, the hydrogen atom tends to borrow or share one additional electron and will share its electron if necessary. Hydrogen atoms also tend to form pairs (Fig. 15.1.5*b*). Each atom in a pair shares its electron with its partner and they form a stable hydrogen molecule.

Oxygen and hydrogen atoms can also combine to form water molecules. The oxygen atom in a water molecule shares electrons with the two hydrogen atoms and the three atoms are held together by covalent bonds. Because of quantum physics, the two hydrogen atoms both reside on the same side of the larger oxygen atom (Fig. 15.1.5*c*). All three atoms complete their electronic shells and a very stable molecule is formed. In fact, the atoms are more tightly bound as water molecules than they were as molecular oxygen and hydrogen molecules. As a result, energy is released during the formation of water from hydrogen and oxygen, an observation made painfully obvious by the fires that consumed many hydrogen-filled airships, including the Hindenburg Zeppelin.

While a water molecule is very light, it behaves differently from other light atoms or molecules. Other molecules of similar mass, including neon, methane, ammonia, and carbon dioxide, are gases at room temperature. Why is water a liquid at room temperature?

The answer is that water molecules interact with one another through the hydrogen bonds mentioned earlier (Fig. 15.1.6). When the hydrogen atom shares its only electron with the oxygen atom, it moves that electron toward the oxygen atom and leaves the positively charged proton in its nucleus partly uncovered. Thus, while the hydrogen atom is well bound to the oxygen, its nucleus is attracted toward any negative charges that might pass by. The oxygen atom in a water molecule has eight electrons, including four that tend to remain far away from the molecule's hydrogen atoms. These negatively charged electrons are attracted to the uncovered protons of nearby water molecules.

Water molecules easily become attached to one another by these hydrogen bonds. Although a hydrogen bond isn't nearly as strong as a covalent bond, it's much stronger than the weak **van der Waals bonds** experienced by most other small molecules. Van der Waals bonds have their origins in quantum physics and can only keep small molecules together at very low temperatures. In contrast, hydrogen bonds are strong enough to keep water molecules together as a liquid at room temperature and above.

Another way to see why water tends to form hydrogen bonds is to look at how electric charge is distributed on the water molecule. The hydrogen side of the molecule is somewhat positively charged while the oxygen side is somewhat negatively charged. This charge imbalance occurs because negatively charged electrons shift away from the hydrogen atoms and toward the oxygen atom as the covalent bonds are formed. The water molecule is polarized, just like a bat-

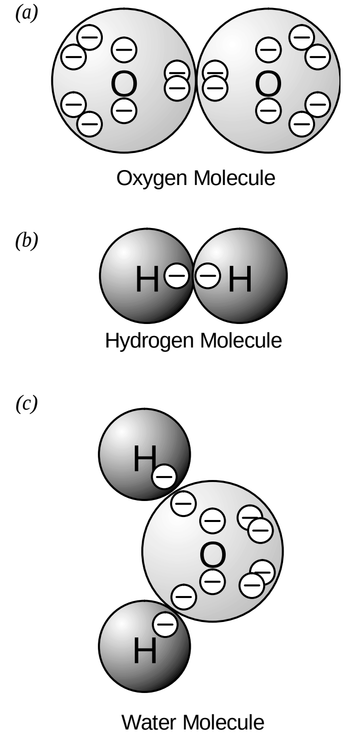

Fig. 15.1.5 - Oxygen, hydrogen, and water molecules are all bound by the sharing of outermost or valence electrons, a process called covalent bonding. Each oxygen atom shares two of its eight electrons with neighboring atoms while each hydrogen atom shares its one electron. In an oxygen atom, four additional valence electrons are located on the side of the atom farthest from the covalent bonds.

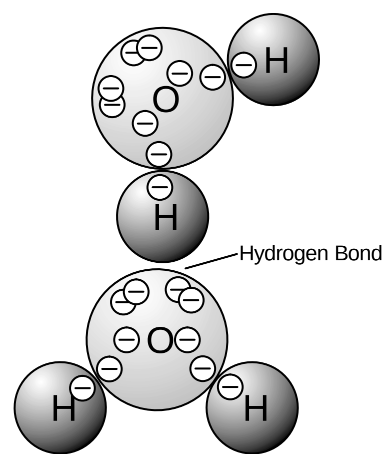

Fig. 15.1.6 - A hydrogen bond is formed when the positively charged hydrogen end of a water molecule is attracted to the negatively charged oxygen end of another water molecule.

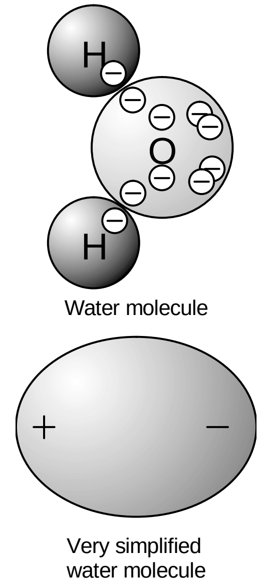

Fig. 15.1.7 - Water is a polar molecule. The hydrogen atom side of the molecule is positively charged while the oxygen atom side is negatively charged.

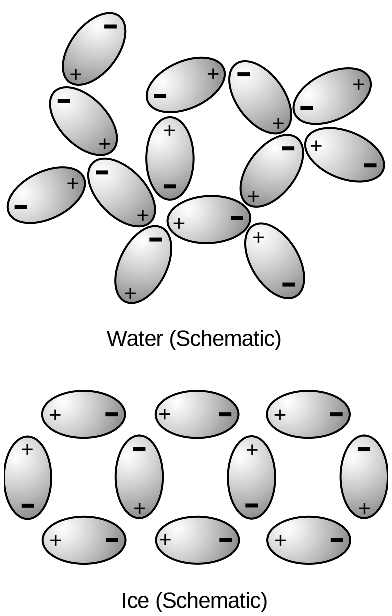

Fig. 15.1.8 - Liquid water contains random loops and chains of the polar water molecules, held together head to tail by hydrogen bonds. Ice is also bound by hydrogen bonds, but it's an orderly arrangement of loops. Both diagrams are schematic only, since it's difficult to visualize the true three-dimensional nature of the arrangements.

tery, and is called a **polar molecule**. Figure 15.1.7 includes a simplified drawing of a water molecule as a single object with positively and negatively charged ends.

Because they are polar, water molecules tend to stick together, positive end to negative end. This attraction between the oppositely charged ends of nearby water molecules is simply the hydrogen bond, described in a slightly different way. Water molecules tend to line up in interlocking chains and loops, with each water molecule hydrogen bonded to its neighbors.

If the temperature is near room temperature, the loops and chains are random in shape and size (Fig. 15.1.8*a*). Molecules constantly drift in and out of the loops and chains, creating a fluid material that we call "water." If the temperature is low enough, the chains and loops become fairly rigid and orderly (Fig. 15.1.8*a*). Few molecules move between the different chains and loops of the solid material we call "ice."

CHECK YOUR UNDERSTANDING #8: Some Things Don't Stick

A carbon dioxide molecule consists of 2 oxygen atoms and 1 carbon atom, bound together by covalent bonds. Unlike water, carbon dioxide is a gas at room temperature, rather than a liquid. Why?

More About Dissolving Materials in Water

Water's polar nature is what allows it to **dissolve** certain materials, such as table salt or sugar. When something dissolves in water, its atoms or molecules separate and become caught up in the loops and chains of water molecules. Not everything dissolves in water—materials that dissolve best are those that contain atoms or molecules that adhere well to water molecules. Salt and sugar both contain such atoms and molecules and both dissolve extremely well in water.

Table salt contains equal numbers of two types of atoms, sodium (Na) and chlorine (Cl). When sodium and chlorine are mixed together, one electron moves from each sodium atom to a nearby chlorine atom. This transfer of electrons is caused by the atoms' tendency to complete electronic shells. Removing one electron from a sodium atom leaves it with a completed electronic shell while adding one electron to a chlorine atom also leaves it with a completed electronic shell.

Because of this electron transfer, the sodium atoms become positively charged sodium **ions** (Na^+) and the chlorine atoms become negatively charged chlorine ions (Cl^-). Since oppositely charged particles attract, the sodium ions and chlorine ions are attracted toward one another. They bond together to form solid salt (Fig. 15.1.9). This type of bonding is called **ionic bonding**.

When you put table salt in water, the water molecules approach the salt crystal, remove the sodium and chlorine ions, and carry them away. Each ion is wrapped up in a **solvation shell** of water molecules. The water molecules in this shell are oriented so that they are attracted to the ion (Fig. 15.1.10). A positive sodium ion attracts the negative ends of the surrounding water molecules while a negative chlorine ion attracts the positive ends. Water does a wonderful job of dissolving most salts, as anyone who has tastes ocean water will discover—there are a great many salts other than sodium chloride and dozens of these salts can be found dissolved in ocean water.

Two special ions that also dissolve well in water are the positive hydrogen ion (H^+) and the negative hydroxyl ion (OH^-). These ions can be formed by pulling a hydrogen nucleus (a proton) away from a water molecule. The positively charged proton is the H^+ ion and the remaining negatively charged fragment is the OH^- ion. In pure water, about 1 in every 55 million water molecules has spontaneously broken up in this way. The resulting ions are carried around in water solvation shells. (In reality, a H^+ ion in water normally attaches itself to a water molecule to form a hydronium ion (H_30^+), but we'll make life simpler by continuing to think of it as a free H^+ ion.)

The density of hydrogen ions in the water is characterized by the water's **pH**. The more hydrogen ions the lower the pH, and the fewer hydrogen ions the higher the pH. Each step on the pH scale corresponds to a 10-fold change in the concentration of hydrogen ions—there are 10 times more hydrogen ions in water with pH=7 than in water with pH=8.

Pure water has a pH of 7 but added chemicals can change its pH. If water's pH is less than 7, it's **acidic**. If its pH is more than 7, the water is **basic**. Chemicals that increase the density of hydrogen ions in water are called **acids** while those that decrease the density of hydrogen ions are called **bases**.

A strong acid such as sulfuric or hydrochloric acid may reduce water's pH to 1 or 2. A strong base such as lye may increase water's pH to 11 or 12. Liquids with such extreme pH's are chemically very aggressive and you certainly don't want them on your skin or hair. A "pH balanced" soap or shampoo is one that has been chemically adjusted to make sure that its density of hydrogen ions is approximately that of pure water and its pH is about 7. Whether this pH is actually important to cleaning skin or hair isn't so clear, but it seems to be good for advertising.

Sugar also dissolves in water, but for a somewhat different reason than salt. Sugar is a carbohydrate and its molecules include carbon, oxygen, and hydrogen atoms. Most of the oxygen and hydrogen atoms are found right next to one another, where they form the same sorts of polar structures found in water molecules. As a result, sugar molecules can hydrogen bond with water molecules.

When you put sugar in water, the water molecules can carry away the individual sugar molecules in solvation shells—water molecules form hydrogen bonds with each sugar molecule and tangle it up in the loops and chains of liquid water. Because its molecules bind so well to water molecules, sugar dissolves easily in water. Water doesn't dissolve hydrocarbons such as oils because it can't hydrogen bond with them—they have no oxygen atoms.

Gases are also soluble in water. Most gases don't bind strongly to water molecules or to themselves and enter the water only because doing so increases the disorder and entropy of the system. A single region of water with dissolved gas is somewhat more disordered than two separate regions of pure water and pure gas. The gas molecules bounce around between the water molecules and attempt to bind with them as much as possible.

At higher temperatures, the disordering effect of dissolved gas becomes less important and the gas tends to come back out of solution. That's why warm soda goes flat very quickly. The carbon dioxide gas in soda dissolves much better in cold water than in hot water. Similarly, the first bubbles that you see as you heat water in a pan are dissolved air coming out of solution.

Pressure also increases the solubilities of gases. If you push hard enough on the gas, it will tend to work its way into the water. If you expose water to high-pressure carbon dioxide gas, the gas quickly goes into solution in the water and you make soda. If you then release the pressure, the gas comes back out of solution as bubbles and fizz. This same effect appears with nitrogen gas in scuba diving. At great depths and pressures, nitrogen dissolves extensively in your blood and tissues. If you come up too quickly, dissolved nitrogen comes out of solution as tiny bubbles and gives you decompression sickness.

Some gases dissolve in water and then rearrange into ions. These ions are immediately bound up in solvation shells and remain in the water for a long time. Carbon dioxide does this rearrangement to a modest extent, creating the

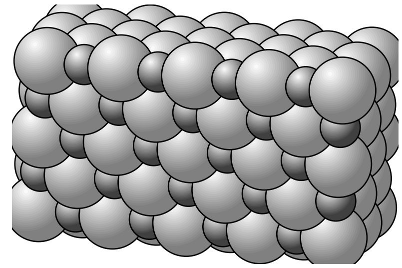

Fig. 15.1.9 - Table salt consists of sodium and chlorine atoms, arranged in an orderly fashion. Each sodium atom has donated an electron to a nearby chlorine atom. The attractive forces between the positively charged sodium ions (dark spheres) and the negatively charged chlorine ions (light spheres) hold the ionic crystal together.

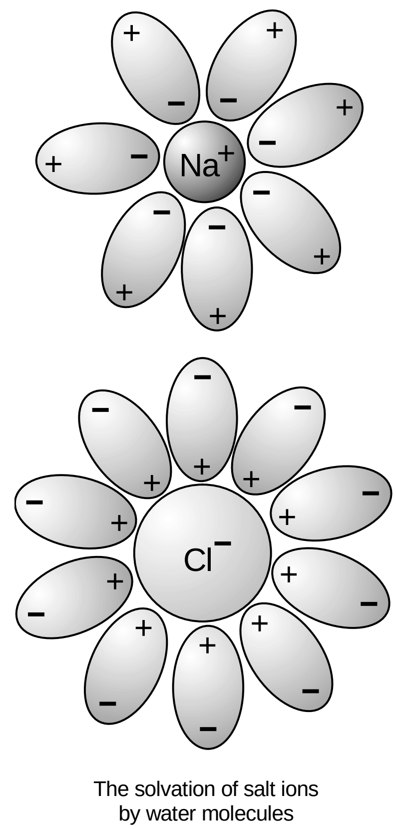

Fig. 15.1.10 - Table salt dissolves in water when the water molecules wrap each salt ion in a solvation shell. The negative ends of water molecules (the oxygen atoms) are aligned near a positive sodium ion and the positive ends (the hydrogen atoms) are aligned near a negative chlorine ion.

carbonic acid that gives soda its biting taste. Ammonia gas dissolves in water to form household ammonia (ammonium hydroxide), a weak base.

But there are also interesting environmental examples of gases dissolved in water. As raindrops pass through air, they dissolve the sulfur trioxide and nitrogen dioxide emitted by power plants and cars and quickly form sulfuric and nitric acids. These two strong acids are the main constituents of acid rain.

CHECK YOUR UNDERSTANDING #9: Taking a Mineral Bath
Why does Epsom salt dissolve in water?

Relative Humidity and Weather

In the atmosphere, changes in air temperature and relative humidity dominate the weather. Rain occurs when warm, moist air cools until its relative humidity exceeds 100%. Air can cool by contact with colder air or by moving upward and expanding. Thus it often rains when hot and cold air masses collide or when warm, moist air blows upward into the mountains and cools. Rain is often found at the edges of hot and cold weather fronts and on the windward slopes of mountains.

In contrast, air that blows downward out of the mountains becomes warmer and its relative humidity drops. Downslope winds, called *katabatic winds*, are very dry and warm. Examples of these katabatic winds include the chinook winds of the eastern Rocky Mountains and the Santa Ana winds of southern California. The chinook winds are famous for melting and evaporating snow at a furious pace, while the Santa Ana winds have fanned many disastrous fires in the chaparral between Los Angeles and San Diego.

A final way in which water vapor and humidity influence the weather is by storing energy. Water vapor contains an enormous amount of chemical potential energy, energy that is released when the gaseous water molecules bind together to form liquid water. This latent heat of vaporization is an essential source of energy that powers thunderstorms and hurricanes. When water vapor condenses inside one of these storms, it releases its latent heat of vaporization and warms the surrounding air. The air expands and floats upward. This heating due to water's condensation produces the ferocious updrafts that are present inside the storm.