B

Water as a Substance

Forces acting on water cause it to move through the hydrologic cycle, and the physical properties of water determine the qualitative and quantitative relations between those forces and the resulting motion. The physical properties of water, as well as its interactions with the environment, are in turn determined by its atomic and molecular structures. Thus, although the detailed study of these structures and properties is outside the traditional scope of hydrology, it is important for the student of hydrology to have some understanding of them.

Water is a very unusual substance with anomalous properties. This strangeness is the reason it is so common at the earth's surface—a topic that is entertainingly elaborated on by van Hylckama (1979). The abundance of water, and its existence in all three phases, makes our planet unique (see Figure B-1) and makes the science of hydrology vital to understanding and managing the environment and our relation to it.

B.1 STRUCTURE OF WATER

B.1.1 Molecular and Inter-Molecular Structure

The water molecule is formed by the combination of two hydrogen atoms (Group Ia, with one electron in the outer shell) with one oxygen atom (Group VIa, with six electrons in the outer shell); hence, it has the chemical formula H_2O. As is shown in Figure B-2a, the outer shell of oxygen can accommodate eight electrons, so it has two vacancies. The outer (and only) shell of hydrogen can hold two electrons, so it has one vacancy. The electron vacancies of two hydrogen atoms and one oxygen atom can be mutually filled by the sharing of outer-shell electrons, shown schematically in Figure B-2b. This sharing is known as a **covalent bond.**

The two most important features of the water molecule are (1) that its covalent bonds are very strong (i.e., much energy is needed to break them) and (2) that the molecular structure is asymmetric, with the hydrogen atoms attached on one "side" of the oxygen atom and an angle of about 105° between the two hydrogens (Figure B-3).

The asymmetry of the water molecule causes it to have a positively charged end (the "side" where the hydrogens are attached) and a negatively charged end (the "side" opposite the hydrogens), much like the poles of a magnet. Most of the unusual properties of water are ultimately the result of its being made up of these **polar molecules.** The polarity produces an attractive force between the positively charged end of one molecule and the negatively charged end of another, as is shown in Figure B-4. This force, called a **hydrogen bond,** is absent from most other liquids.

Although the hydrogen bond is only about one-twentieth the strength of the covalent bond (Stillinger 1980), it is far stronger than the intermolecular bonds that are present in liquids with symmetrical, nonpolar molecules. We get an idea of this strength when we compare the melting/freezing temperature and the boiling/condensation temperature of the hydrides of all the Group VIa elements: oxygen (O), sulfur (S), selenium (Se), and tellurium

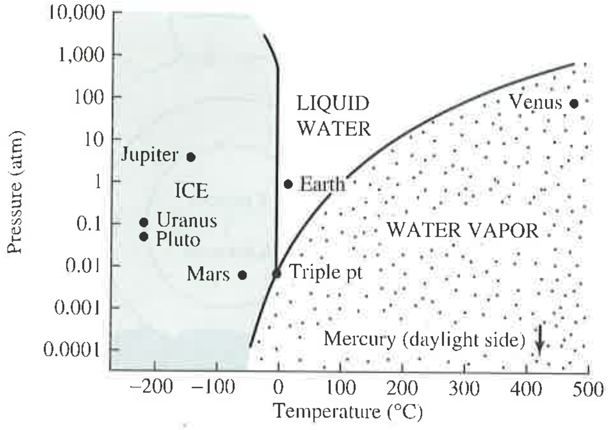


FIGURE B-1
Surface temperatures and pressures of the planets plotted on the phase diagram for water. From *Opportunities in the Hydrologic Sciences* © 1991 by the National Academy of Sciences. Reprinted with permission of the National Academy Press.

(Te). These elements are all characterized by an outer electron shell that can hold eight electrons but has two vacancies. Thus they all form covalent bonds with two hydrogens but, except for water, the resulting molecules are nearly symmetrical and therefore nonpolar.

In the absence of strong intermolecular forces that result from polar molecules, the melting/freezing and boiling/condensation temperatures of these compounds would be expected to rise as their atomic weights increased. As is shown in Figure B-5, these expectations are fulfilled, except—strikingly—in the case of H_2O. The reason for this departure from expectations is the hydrogen bonds, which attract one molecule to another and which can be loosened (as in melting) or broken (as in evaporation) only when the vibratory energy of the molecules is large—that is, when the temperature is high. Because of its high melting and boiling temperatures, water is one of the very few substances that exists in three physical states—solid, liquid, and gas—at earth-surface temperatures.

B.1.2 Freezing and Melting

At temperatures below 0 °C, the vibratory energy of water molecules is sufficiently low that the hydrogen bonds can lock the molecules into the regular three-dimensional crystal lattice of ice (Figure B-6). In this lattice, each molecule is hydrogen-bonded to four adjacent molecules. The angle between the hydrogen atoms in each molecule remains at 105°, but each molecule is oriented so that a puckered honeycomb of perfect hexagons is visible when the lattice is viewed from one direction. Thus ice is a hexagonal crystal, and snowflakes show infinite variation on a theme of six-fold symmetry.

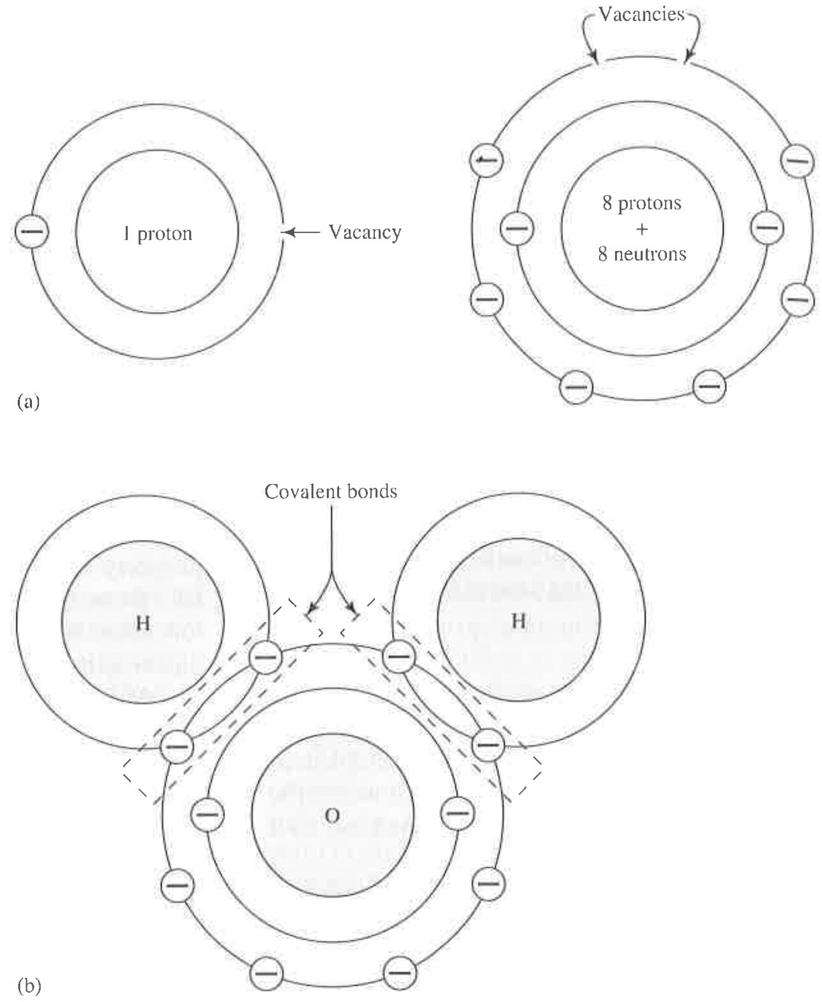
When ice is warmed to 0 °C, further additions of heat cause melting, in which about 15% of the hydrogen bonds break (Stillinger 1980). Because of the rupturing of some of the hydrogen bonds, the rigid ice lattice partially collapses, and a given number of molecules takes up less space in the liquid phase than in the solid. As a result, the density of ice is less than that of water (91.7% of the density of liquid water at 0 °C). Very few substances have a lower density in the solid state than in the liquid, and this property is of immense importance: If rivers and lake froze from the bottom up instead of from the top down, biological and hydrological conditions in higher latitudes would be markedly altered.

Although melting always occurs when ice at earth-surface pressure is warmed to 0 °C, freezing may not always take place when liquid water is cooled to 0 °C. If the liquid contains no impurities and is not in contact with preexisting ice, it is possible to supercool it to temperatures as low as −41 °C. This resistance to freezing occurs because the water molecules form various types of nonhexagonal hydrogen-bonded polyhedra (Figure B-4), which prevent the formation of the ice lattice. However, if ice particles (or common impurities like clay minerals, which have a crystal structure like that of ice) are present, they provide templates that act as growth nuclei to trigger the formation of ice at 0 °C. Significant supercooling is quite common in clouds, where effective growth nuclei may be lacking, and it is an important factor in the development of cloud particles into raindrops and snowflakes (Section D.5). In rivers, supercooling of 0.01 to 0.1 °C often occurs during the formation of ice in fast-flowing reaches (Michel 1971).

B.1.3 Evaporation and Condensation

At temperatures less than 100 °C, some molecules at the liquid–air interface with greater-than-average energy sever all hydrogen bonds with their neighbors and fly off from the bulk liquid. Thus evaporation from ice and liquid water can take place at all temperatures below the boiling point, and the loss of

FIGURE B-2
(a) Schematic diagram of a hydrogen atom (left) and an oxygen atom (right). (b) Schematic diagram of a water molecule showing covalent bonding. ⊖ symbols represent electrons. From *Fluvial Hydrology* by S. Lawrence Dingman. © 1984 by W.H. Freeman and Company. Reprinted with permission.



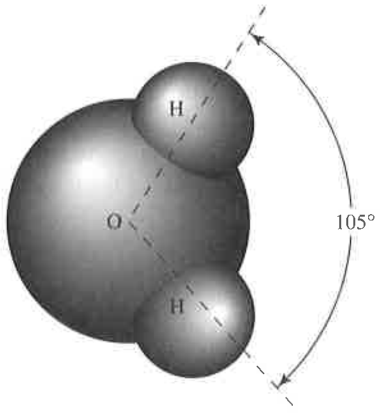


FIGURE B-3
Diagram of a water molecule, showing the angle between the hydrogen atoms. After Davis and Day (1961).

these high-energy molecules results in a lowering of the average energy, and hence the temperature, of the remaining solid or liquid. When liquid water is heated to 100 °C, further additions of energy cause the eventual breaking of all the remaining hydrogen bonds, and the liquid is transformed entirely into a gas consisting of relatively widely spaced, mostly non-bonded individual H_2O molecules.

B.1.4 Dissociation

An **ion** is an elemental or molecular species with a net positive or negative electrical charge. At any given instant, a fraction of the molecules of liquid water are **dissociated** into positively charged hydrogen ions, designated H^{+1}, and negatively charged hydroxide ions, designated OH^{-1}. In spite of their

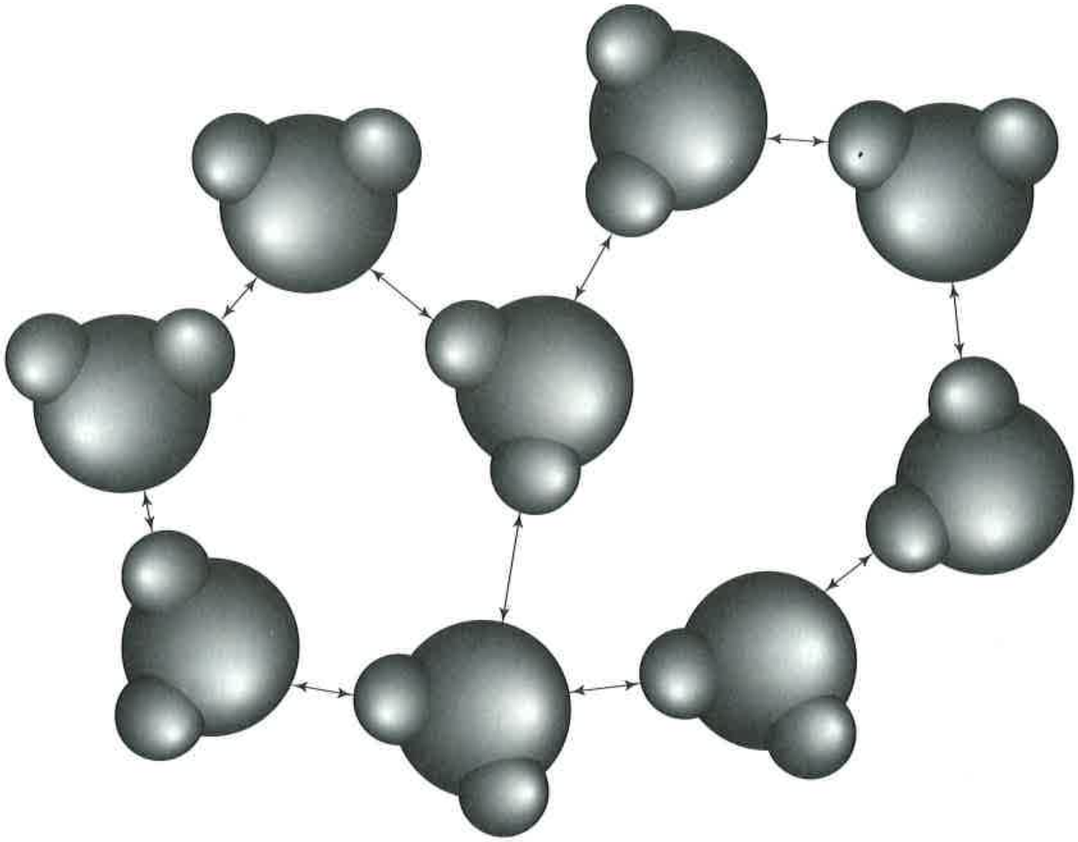


FIGURE B-4
Schematic diagram of water molecules in the liquid state. Arrows indicate hydrogen bonds between the oppositely charged ends of adjacent molecules. From *Fluvial Hydrology* by S. Lawrence Dingman. © 1984 by W.H. Freeman and Company. Reprinted with permission.

generally very low concentrations, these ions participate in many important chemical reactions.

Hydrogen ions are responsible for the acidity of water, and acidity is usually measured in terms of the quantity called **pH,** which is defined as

$$\mathrm{pH} \equiv -\log_{10}([\mathrm{H}^{+1}]), \qquad \textbf{(B-1)}$$

where $[\mathrm{H}^{+1}]$ designates the concentration of hydrogen ions in mg L^{-1}. The concentration of hydrogen ions in pure water at 25 °C is $10^{-7.00}$ mg L^{-1} ($\mathrm{pH} = 7.00$). As $[\mathrm{H}^{+1}]$ increases above this value (pH decreases below 7.00), water becomes more acidic; as $[\mathrm{H}^{+1}]$ decreases ($\mathrm{pH} > 7.00$), it becomes more basic.

Certain chemical reactions change the concentration of hydrogen ions, causing the water to become more or less acidic. The degree of acidity, in turn, determines the propensity of the water to dissolve many elements. The pH of cloud water in equilibrium with the carbon dioxide in the atmosphere is about 5.7; additional reactions make the natural pH of rainwater fall in the range from 4.5 to 5.6, depending on location (Turk 1983).

B.1.5 Isotopes

Isotopes of an element have the same number of protons and electrons, but differing numbers of neutrons; thus they have similar chemical behavior, but differ in atomic weight. Some isotopes are **radioactive,** and decay naturally to other atomic forms at a characteristic rate, while others are **stable.**

Table B-1 gives the properties and abundances of the isotopes of hydrogen and oxygen, from which

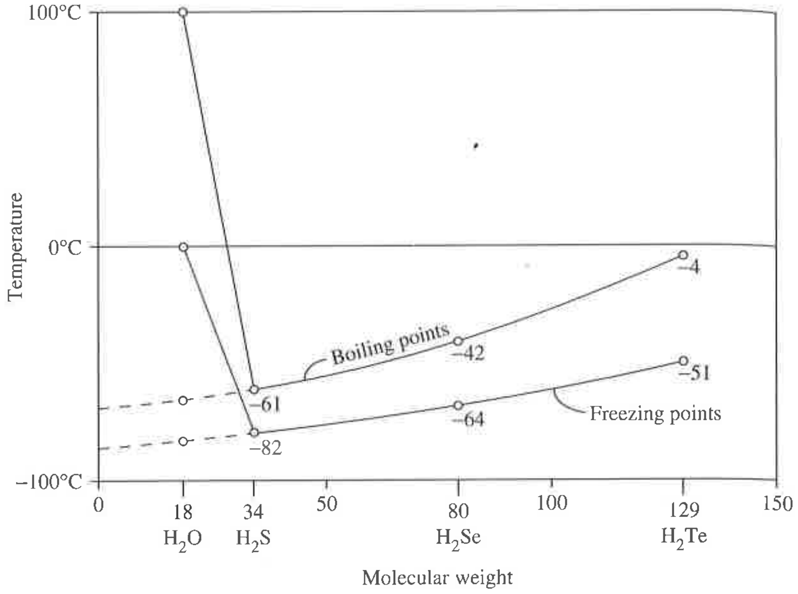


FIGURE B-5
Melting/freezing and boiling/condensation temperatures of Group VIa hydrides. In the absence of hydrogen bonds, water would have a melting/freezing point of −100 °C and a boiling/condensation point of −91 °C. After Davis and Day (1961).

it can be calculated that 99.73 % of all water consists of "normal" $^{1}H_2{}^{16}O$.[1]

The various isotopes are involved in differing proportions in phase changes and chemical and biological reactions, so they are **fractionated** as water moves through the hydrologic cycle. (See Fritz and Fontes 1980; Drever 1982). Thus the relative concentrations of these isotopes can be used in some hydrologic situations to identify the sources of water in glaciers, aquifers, or streams. (See, e.g., Perry and Montgomery 1982; also see Box 9-1.)

The isotope ^{3}H, called **tritium,** is radioactive; it decays with a half life of 12.5 yr to ^{3}He. It is produced in very small concentrations by natural processes and in larger concentrations by nuclear reactions. It has potential for use in dating relatively recent water in aquifers and glaciers. (See, e.g., Davis and Murphy 1987; also see Section 8.5.1.) Figure B-7 shows the annual rates of production of bomb tritium.

B.2 PROPERTIES OF WATER

In this section, we briefly describe the bulk properties of liquid water that influence its movement through the hydrologic cycle and its interactions with the terrestrial environment. More detailed discussions of these properties can be found in Dorsey (1940) and Davis and Day (1961). Properties of water in the solid and vapor form are discussed where relevant in the text.

The variation of water's properties with temperature is important in many hydrologic contexts. Thus, in the following discussion, the values of each property at 0 °C are given in the three unit systems, and their relative variations with temperature are

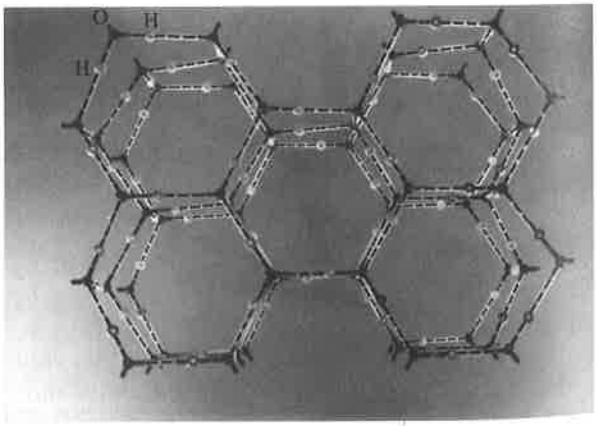


FIGURE B-6
A model of the crystal lattice of ice, showing its hexagonal structure. White circles are hydrogen atoms, dark circles are oxygen atoms, dashed lines are hydrogen bonds. Photo by author.

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[1] By convention, the atomic weight is written to the upper left of the element symbol.

TABLE B-1

Characteristics of Isotopes of Hydrogen and Oxygen

Isotope	Natural Abundance (%)	Stability
^{1}H	99.985	stable
^{2}H (Deuterium)	0.015	stable
^{3}H (Tritium)	trace	radioactive
^{16}O	99.76	stable
^{17}O	0.04	stable
^{18}O	0.20	stable

shown in Table B-2. Empirical equations for computing the values of the properties as functions of temperature are also given.

B.2.1 Density

Mass density, ρ, is the mass per unit volume [M L^{-3}] of a substance; **weight density,** γ, is the weight per unit volume [F L^{-3}]. These are related by Newton's Second Law (i.e., force equals mass time acceleration), so that

$$\gamma = \rho \cdot g, \tag{B-2}$$

where g is the acceleration due to gravity [L T^{-2}] (g = 9.81 m s^{-2} = 32.2 ft s^{-2}).

In the SI system of units, the kilogram is defined as the mass of 1 m^3 of pure water at its temperature of maximum density, 3.98 °C. For water at 0 °C,

$$\rho = 999.87\,\text{kg m}^{-3} = 0.99987\,\text{g cm}^{-3} = 1.9397\,\text{slug ft}^{-3}$$

and

$$\gamma = 9799\ \text{N m}^{-3} = 979.9\ \text{dyn cm}^{-3} = 62.46\ \text{lb ft}^{-3}.$$

The **specific gravity** of a substance is the ratio of its weight density to the weight density of pure water at 3.98 °C; thus it is dimensionless.

Because gravitational force and momentum are proportional to mass, and pressure depends on weight, either ρ or γ appears in most equations describing the motion of fluids.

The change in density of water with temperature is unusual (Table B-2) and environmentally significant. As was noted, liquid water at 0 °C is denser than ice. As liquid water is warmed from 0 °C its density initially *increases*, whereas most other substances become less dense as they warm. This anomalous increase continues until density reaches a maximum value of 1000 kg m^{-3} at 3.98 °C; beyond this, the density decreases with temperature, as with most other substances. These density variations can be approximated as

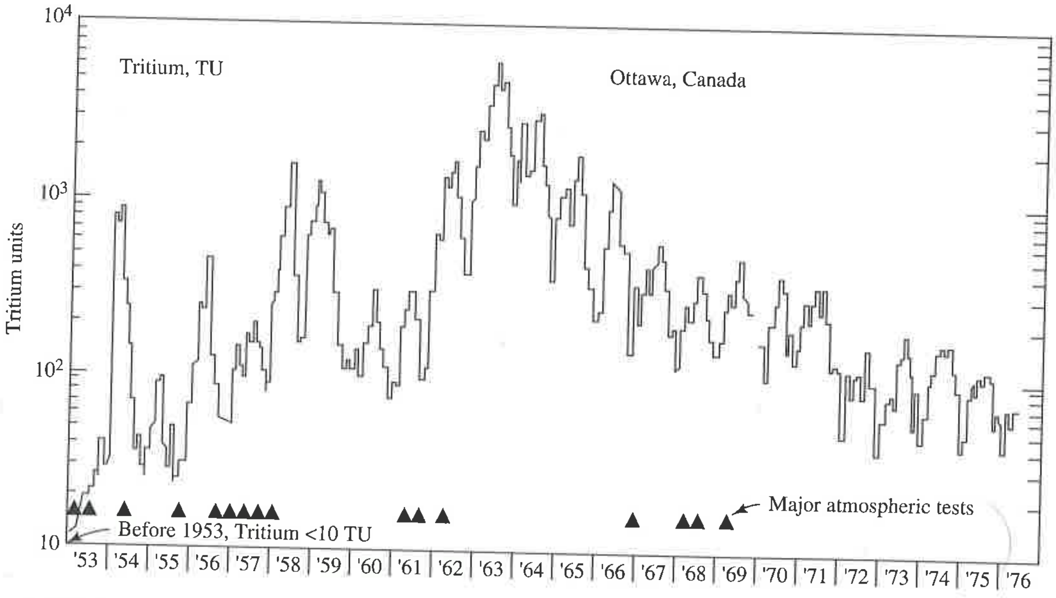

FIGURE B-7

Tritium concentrations in precipitation at Ottawa, Canada. This is the longest continuous record; it is representative of trends in the northern hemisphere. From R. Allan Freeze and John A. Cherry, *Groundwater*, © 1979, p. 137. Reprinted with permission of Prentice Hall, Englewood Cliffs, NJ.

TABLE B-2
Relative Values of Properties of Pure Liquid Water as Functions of Temperature

Temperature (°C)	ρ, γ	μ	ν	σ	c_p	λ_v
0	1.00000	1.0000	1.0000	1.0000	1.0000	1.0000
3.98	1.00013					
5	1.00012	0.8500	0.8500	0.9907	0.9963	0.9953
10	0.99986	0.7314	0.7315	0.9815	0.9940	0.9904
15	0.99926	0.6374	0.6379	0.9722	0.9924	0.9857
20	0.99836	0.5607	0.5616	0.9630	0.9915	0.9810
25	0.99720	0.4983	0.4997	0.9524	0.9910	0.9763
30	0.99580	0.4463	0.4482	0.9418	0.9907	0.9715

ρ = mass density, γ = weight density, μ = dynamic viscosity, ν = kinematic viscosity, σ = surface tension, λ_v = latent heat of vaporization, c_p = specific heat.

$$\rho = 1000 - 0.019549 \cdot |T - 3.98|^{1.68}, \qquad \textbf{(B-3)}$$

where T is temperature in °C and ρ is in kg m^{-3} (Heggen 1983). The variation of γ with temperature can be approximated via Equations (B-2) and (B-3).

In lakes where temperatures reach 3.98 °C, the density maximum controls the vertical distribution of temperature and causes an annual or semi-annual overturn of water that has a major influence on biological and physical processes. However, except for lakes, the variation of density with temperature is small enough that it can usually be neglected in hydrological calculations.[2]

The addition of dissolved or suspended solids to water increases its density in proportion to the density of the solids and their concentration. Again, the density effects of dissolved materials can be important in lakes, but they are not usually significant in other environments.[3] However, high concentrations of suspended matter can significantly increase the effective density of water in rivers.

Note that the kilogram and gram are commonly used as units of force as well as of mass: 1 kilogram (gram) of force is the weight of a mass of 1 kilogram (gram) at the earth's surface, where $g = 9.81$ m s^{-2} (981 cm s^{-2}). Thus, from Equation (B-2), 1 kilogram of force = 9.810 N; 1 gram of force = 981.0 dyn.

B.2.2 Surface Tension

Molecules in the surface of liquid water are subjected to a net inward force due to hydrogen bonding with the molecules below the surface (Figure B-8). **Surface tension,** σ, is equal to the magnitude of that force divided by the distance over which it acts; thus its dimensions are [F L^{-1}]. It can also be viewed as the work required to increase the surface area of a liquid by a unit amount ([F L]/[L^2] = [F L^{-1}]).

Surface tension and the closely related phenomenon of **capillarity** significantly influence fluid motion where a water surface is present and where the flow scale is less than a few millimeters—for example, in porous media that are partially saturated or ones in which there is an interface between water and an immiscible liquid (e.g., hydrocarbons).

As might be expected from the strong intermolecular forces, water has a surface tension higher than that of most other liquids; its value at 0 °C is

$$\sigma = 0.0756 \text{ N m}^{-1} = 75.6 \text{ dyn cm}^{-1} = 0.00518 \text{ lb ft}^{-1}.$$

Surface tension decreases rapidly as temperature increases (Table B-2), and this effect can be important when one is considering the movement of water in soils. (See Section 6.5.4.) The temperature effect can be approximated as

$$\sigma = 0.001 \cdot (20987 - 92.613 \cdot T)^{0.4348}, \qquad \textbf{(B-4)}$$

where T is in °C and σ is in N m^{-1} (Heggen 1983).

Dissolved substances can also increase or decrease surface tension, and certain organic compounds have a major effect on its value.

Capillary Rise

Consider the small (diameter a few millimeters or less) cylindrical tube immersed in a body of water with a free surface[4] shown in Figure B-9. If the ma-

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[2] An exception to this generalization can occur when one is considering deep regional ground-water flow systems on geologic time scales. In these systems, density variations related to geothermal heat flow can have significant effects on the flow.

[3] Density gradients due to the dissolution of minerals like salt can significantly affect deep ground-water flows on geological time scales.

[4] A "free surface" is a surface of liquid water at atmospheric pressure. In diagrams, such a surface is designated by the inverted triangular *hydrat* symbol, ∇.

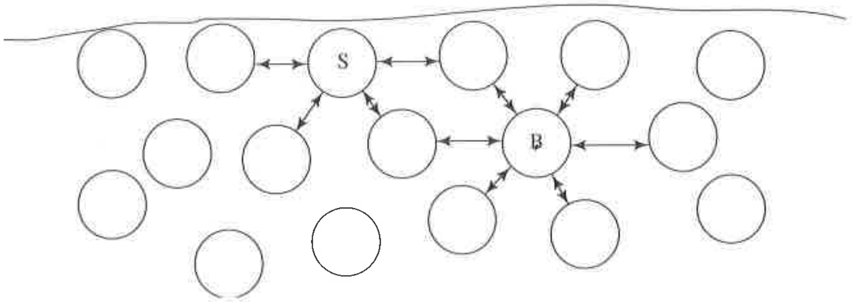

FIGURE B-8
Intermolecular forces acting on typical surface (S) and nonsurface (B) molecules. From *Fluvial Hydrology* by S. Lawrence Dingman. © 1984 by W.H. Freeman and Company. Reprinted with permission.

terial of the tube is such that the hydrogen bonds of the water are attracted to it, the molecules in contact with the tube are drawn upward. The degree of attraction between the water and the tube is reflected in the **contact angle** between the water surface, or **meniscus,** and the tube: the stronger the attraction, the smaller the angle. Because of the intermolecular hydrogen bonds, the entire mass of water within the tube will be also drawn upward until the adhesive force between the molecules of the tube and those of the water is balanced by the downward force due to the weight of the water suspended within the tube.

The height to which the water will rise in the tube can thus be calculated by equating the upward and downward forces. The upward force, F_u, equals the vertical component of the surface tension times the distance over which that force acts:

$$F_u = \sigma \cdot \cos(\theta_c) \cdot 2 \cdot \pi \cdot r, \qquad \textbf{(B-5)}$$

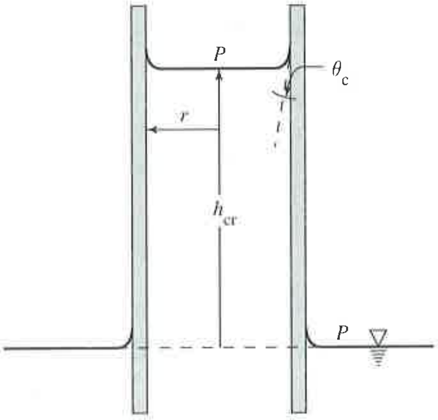

FIGURE B-9
Sketch of phenomenon of capillary rise in a circular tube of radius r, θ_c is the contact angle between the meniscus and the wall, and P is atmospheric pressure. After Dingman (1984).

where θ_c is the contact angle between the water surface and the tube and r is the radius of the tube. The downward force due to the weight of the column of water, F_d, is

$$F_d = \gamma \cdot \pi \cdot r^2 \cdot h_{cr}, \qquad \textbf{(B-6)}$$

where γ is the weight density of water and h_{cr} is the height of the column.

Equating F_u and F_d and solving for h_{cr} yields

$$h_{cr} = \frac{2 \cdot \sigma \cdot \cos(\theta_c)}{\gamma \cdot r}. \qquad \textbf{(B-7)}$$

Thus the height of capillary rise is inversely proportional to the radius of the tube and directly proportional to the surface tension and the cosine of the contact angle.

Table B-3 gives the contact angle for water in contact with air and selected solids. Note that the contact angle for most soil materials is close to 0°.

Under natural conditions, the interconnected pores between particles of granular geologic materials act as capillary tubes and cause a **tension-satu-**

TABLE B-3
Surface-Tension Contact Angles for Water–Air Interfaces Against Various Solids

Solid	Contact Angle, θ_c (°)	cos (θ_c)
Glass	0	1.0000
Most silicate minerals	0	1.0000
Ice	20	0.9397
Platinum	63	0.4540
Gold	68	0.3746
Talc	86	0.0698
Paraffin	105 to 110	−0.2588 to −0.3420
Shellac	107	−0.2924
Carnauba wax	107	−0.2924

Data from Dorsey (1940) and Jellinek (1972).

rated zone or **capillary fringe** above the free surface (Section 6.4.2). The height of this fringe can be calculated approximately via Equation (B-7), with r equal to the radius of the granular particles; it ranges from a few centimeters in gravel and sand to well over a meter in silt. For clay soils, however, additional electrical forces become significant and Equation (B-7) does not hold.

Pressure Relations

In unsaturated porous media of silt size and larger ($\geq$ 0.002 mm; see Section 6.1.1), pressure is determined by the radius of curvature of menisci, and the relation between the two can be developed from further consideration of the capillary-rise phenomenon.

When capillary rise has ceased in Figure B-9, the column of water is suspended from the meniscus, which is in turn attached to the walls by hydrogen bonds. Thus the water is under **tension,** which is defined as negative (i.e., less than atmospheric) pressure. The weight of the water suspended beneath the plane that is tangent to the lowest point of the meniscus is equal to F_d [Equation (B-6)]. The area of that plane is $\pi \cdot r^2$, so the pressure in the suspended water at that plane, P_m, is

$$P_m = -\frac{\gamma \cdot \pi \cdot r^2 \cdot h_{cr}}{\pi \cdot r^2} = -\gamma \cdot h_{cr}. \qquad \textbf{(B-8)}$$

Substituting Equation (B-7) into Equation (B-8) gives

$$P_m = -\frac{2 \cdot \sigma \cdot \cos(\theta_c)}{r}. \qquad \textbf{(B-9)}$$

Thus the pressure difference across a meniscus, like the height of capillary rise, is inversely proportional to the radius of a capillary opening. In unsaturated soils of silt size or larger, this radius is determined by the soil texture and by the amount of water present. (See Section 6.3.2.) The moisture-characteristics curve (Figure 6-7) is the empirical manifestation of the pressure–radius relation of Equation (B-9) in the range $0 > P_m > -0.9$ atmospheres ($= -930$ cm $= -91.2$ kPa). Tensions many times stronger than 1 atmosphere can exist in soils, but they are due to attractive forces between the water and the mineral grains and do not represent capillary pressure (Gray and Hassanizadeh 1991).

B.2.3 Boundary-Layer Flow, Viscosity, and Turbulence

Boundary-Layer and Potential Flows

When water flows over a solid boundary, hydrogen bonds cause the fluid molecules adjacent to the boundary to "stick" so that the velocity at the boundary is zero. This phenomenon is called the **no-slip condition.** This condition produces a frictional drag that is transmitted through the fluid for considerable distances normal to the boundary.

In virtually all open-channel flows of interest to hydrologists, the retarding effects of the boundary are present throughout the flow, and such flows are therefore called **boundary-layer flows.**

Flows within the pores of porous media are also boundary-layer flows, but mathematically they are not treated as such. As described in Section 8.1.1, Darcy's Law describes bulk flow through a representative "chunk" of the medium rather than through individual pores. Thus the effects of intrapore viscous resistance are represented by the bulk hydraulic conductivity, and the boundary of the overall flow (i.e., the aquifer boundary) is not a source of flow resistance. Thus ground-water flow is mathematically represented as a **potential flow.**

Viscosity and Laminar Flow

For flows that are very close to a boundary and are very slow, the water moves in parallel layers (Figure B-10a), and the flow is called **laminar.** The **dynamic viscosity,** μ, is the friction between the layers that transmits the boundary friction through the fluid and resists the forces tending to cause flow. The layers move successively faster as one moves away from the boundary, so that a velocity gradient exists, and μ is the proportionality between this gradient and the frictional force per unit area, called the **shear stress,** $\tau(y)$. Thus we have

$$\tau(y) = \mu \cdot \frac{d\,u(y)}{d\,y}, \qquad \textbf{(B-10)}$$

where $u(y)$ is the velocity at a distance y from the boundary.

Equation (B-10) shows that, for a given boundary shear stress, a higher viscosity causes a smaller velocity gradient—that is, the velocity increases more slowly as one moves away from the boundary.

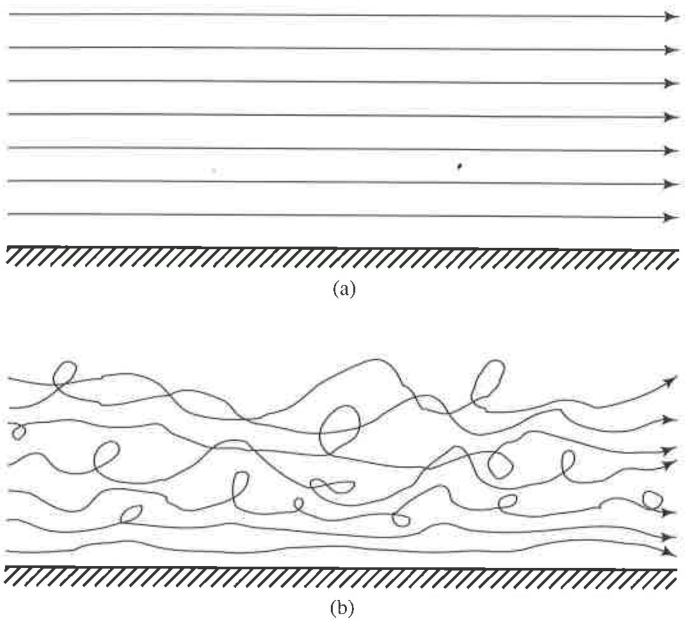

FIGURE B-10
Paths of individual water particles in (a) laminar flow, (b) turbulent flow (highly schematic). After Dingman (1984).

Note that, because $\tau(y)$ has the dimensions $[\mathrm{F\,L^{-2}}]$, μ has the dimensions $[\mathrm{F\,T\,L^{-2}}] = [\mathrm{M\,L^{-1}\,T^{-1}}]$.

The dynamic viscosity is due to inter-molecular attractions, and so it is also called the **molecular viscosity.** In spite of the strength of the hydrogen bonds, water's viscosity is relatively low because of the rapidity with which the hydrogen bonds spontaneously break and reform (about once every 10^{-12} s). Viscosity at 0 °C is

$$\mu = 0.001787 \mathrm{~N\,s\,m^{-2}} = 0.01787 \mathrm{~dyn\,s\,cm^{-2}}$$
$$= 3.735\ 10^{-5} \mathrm{~lb\,s\,ft^{-2}}.$$

In many contexts, the ratio μ/ρ arises; thus it is convenient to define the **kinematic viscosity,** $\nu\ [\mathrm{L^2\,T^{-1}}]$:

$$\nu \equiv \frac{\mu}{\rho}. \qquad \textbf{(B-11)}$$

Values of ν at 0 °C are

$$\nu = 1.787 \times 10^{-6} \mathrm{~m^2\,s^{-1}} = 1.787 \times 10^{-2} \mathrm{~cm^2\,s^{-1}}$$
$$= 1.926 \times 10^{-5} \mathrm{~ft^2\,s^{-1}}.$$

As is shown in Table B-2, viscosity decreases rapidly as temperature increases. The temperature effect can be approximated as

$$\mu = 2.0319 \times 10^{-4} + 1.5883 \times 10^{-3} \cdot \exp\left[-\left(\frac{T^{0.9}}{22}\right)\right], \qquad \textbf{(B-12)}$$

where T is in °C and μ is in $\mathrm{N\,s\,m^{-2}}$ (Heggen 1983).

Some dissolved constituents increase viscosity, others decrease it, but these effects are usually negligible at the concentrations found in nature.[5] However, moderate-to-high concentrations of suspended material can significantly increase the effective viscosity of the fluid.

Turbulence

As distance from a boundary and velocity increase, the inertia of the moving water increasingly overcomes viscous friction and the flow paths of individual water "particles" increasingly deviate from the parallel layers of laminar flow. At relatively modest distances and velocities, all semblance of parallel flow disappears and the water moves in highly irregular eddies. This is the phenomenon called **turbulence** (Figure B-10b).

Because the water in turbulent eddies moves in directions other than the main flow direction, turbulence consumes some of the energy that would otherwise drive the main flow. This energy loss can be thought of as an addition to the internal friction of the fluid and therefore as an addition to the ef-

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[5] As with density, thermal and solute effects on viscosity can have a significant influence on deep ground-water flow over geologic time scales.

W sphere. Unlike oxygen and nitrogen—the two most abundant components of the atmosphere—water can change from one state of matter to another (solid, liquid, or gas) at the temperatures and pressures experienced on Earth. Because of this unique property, water freely leaves the oceans as a gas and returns again as a liquid.

As you observe day-to-day weather changes, you might ask: Why is it generally more humid in the summer than in the winter? Why do clouds form on some occasions but not on others? Why do some clouds look thin and harmless whereas others form gray and ominous towers? Answers to these questions involve the role of water vapor in the atmosphere, the central theme of this chapter.

FOCUS ON CONCEPTS

To assist you in learning the important concepts in this chapter, focus on the following questions:

- What are the processes that cause water to change from one state of matter to another?
- What is latent heat?
- What is humidity? What is the most common method used to express humidity?
- What processes cause relative humidity to change?
- What four mechanisms cause air to rise?
- How is atmospheric stability determined?
- What are the necessary conditions for condensation?
- Which two criteria are used for cloud classification?
- What is fog? How do the various types of fog form?
- What two processes produce precipitation in a cloud?
- What are the different types of precipitation and how does each form?
- How is precipitation measured?

Water's Changes of State

GEODe EARTH SCIENCE

Earth's Dynamic Atmosphere

▸ Moisture and Cloud Formation

Water is the only substance that exists in the atmosphere as a solid, liquid, and gas (Figure 17.1). It is made of hydrogen and oxygen atoms that are bonded together to form water molecules (H_2O). In all three states of matter (even ice) these molecules are in constant motion—the higher the temperature, the more vigorous the movement. The chief difference among liquid water, ice, and water vapor is the arrangement of the water molecules.

Ice, Liquid Water, and Water Vapor

Ice is composed of water molecules that are held together by mutual molecular attractions. Here the molecules form a tight, orderly network, as shown in Figure 17.2. As a consequence, the water molecules in ice are not free to move relative to each other but rather vibrate about fixed sites. When ice is heated, the molecules oscillate more rapidly. When the rate of molecular movement increases sufficiently, the bonds between some of the water molecules are broken, resulting in melting.

In the liquid state, water molecules are still tightly packed but are moving fast enough that they are able to slide past one another. As a result, liquid water is fluid and will take the shape of its container.

As liquid water gains heat from its environment, some of the molecules will acquire enough energy to break the remaining molecular attractions and escape from the surface, becoming water vapor. Water-vapor molecules are widely spaced compared to liquid water and exhibit very energetic random motion. What distinguishes a gas from a liquid is its compressibility (and expandability). For example, you can easily put more and more air into a tire and increase its volume only slightly. However, don't try to put 10 gallons of gasoline into a five-gallon can.

To summarize, when water changes state, it does not turn into a different substance; only the distances and interactions among the water molecules change.

Latent Heat

Whenever water changes state, heat is exchanged between water and its surroundings. When water evaporates, heat is absorbed (Figure 17.2). Meteorologists often measure heat energy in calories. One **calorie** is the amount of heat required to raise the temperature of 1 gram of water 1° C (1.8° F). Thus, when 10 calories

FIGURE 17.1 Caught in a downpour. (Photo by ImageState/Alamy)

of heat are absorbed by 1 gram of water, the molecules vibrate faster and a 10° C (18° F) temperature rise occurs.

Under certain conditions, heat may be added to a substance without an accompanying rise in temperature. For example, when a glass of ice water is warmed, the temperature of the ice-water mixture remains a constant 0° C (32° F) until all the ice has melted. If adding heat does not raise the temperature, where does this

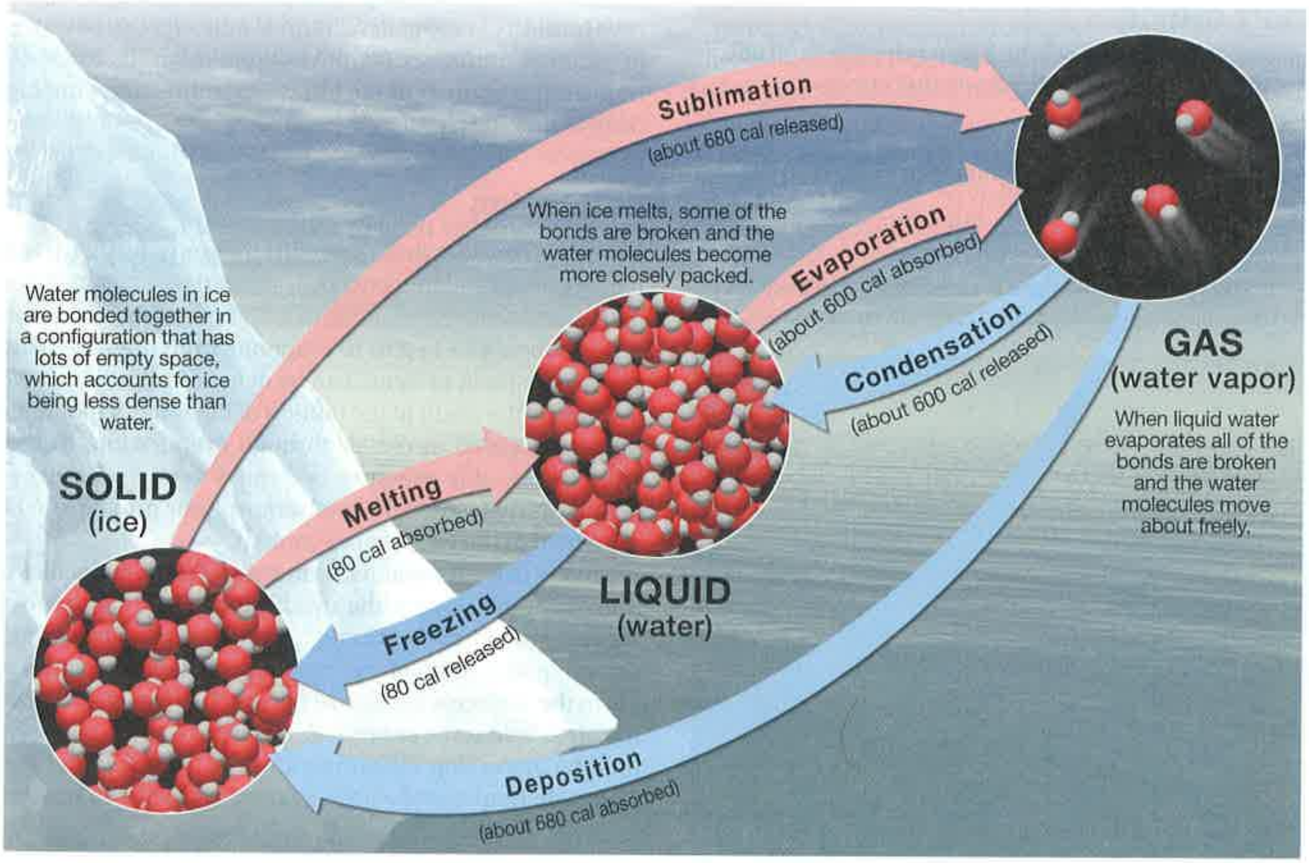
ecular attractions between the water molecules in the ice cubes.

Because the heat used to melt ice does not produce a temperature change, it is referred to as **latent heat**. (Latent means "hidden," like the latent fingerprints hidden at a crime scene.) This energy can be thought of as being stored in liquid water, and it is not released to its surroundings as heat until the liquid returns to the solid state.

It requires 80 calories to melt one gram of ice, an amount referred to as *latent heat of melting. Freezing*, the reverse process, releases these 80 calories per gram to the environment as *latent heat of fusion*.

Evaporation and Condensation We saw that heat is absorbed when ice is converted to liquid water. Heat is also absorbed during **evaporation**, the process of converting a liquid to a gas (vapor). The energy absorbed by water molecules during evaporation is used to give them the motion needed to escape the surface of the liquid and become a gas. This energy is referred to as the *latent heat of vaporization*. During the process of evaporation, it is the higher-temperature (faster-moving) molecules that escape the surface. As a result, the average molecular motion (temperature) of the remaining water is reduced—hence the common expression, "evaporation is a cooling process." You have undoubtedly experienced this cooling effect on stepping dripping wet from a swimming pool or bathtub. In this situation the energy used to evaporate water comes from your skin—hence, you feel cool.

FIGURE 17.2 Changes of state always involve an exchange of heat. The numbers shown here are the approximate number of calories either absorbed or released when 1 gram of water changes from one state of matter to another.



changes to the liquid state. During condensation, water-vapor molecules release energy (*latent heat of condensation*) in an amount equivalent to what was absorbed during evaporation. When condensation occurs in the atmosphere, it results in the formation of such phenomena as fog and clouds.

As you will see, latent heat plays an important role in many atmospheric processes. In particular, when water vapor condenses to form cloud droplets, latent heat of condensation is released, warming the surrounding air and giving it buoyancy. When the moisture content of air is high, this process can spur the growth of towering storm clouds.

Sublimation and Deposition You are probably least familiar with the last two processes illustrated in Figure 17.2—sublimation and deposition. **Sublimation** is the conversion of a solid directly to a gas without passing through the liquid state. Examples you may have observed include the gradual shrinking of unused ice cubes in the freezer and the rapid conversion of dry ice (frozen carbon dioxide) to wispy clouds that quickly disappear.

Deposition refers to the reverse process, the conversion of a vapor directly to a solid. This change occurs, for example, when water vapor is deposited as ice on solid objects such as grass or windows (**Figure 17.3**). These deposits are called *white frost* or *hoar frost* and are frequently referred to simply as *frost*. A household example of the process of deposition is the "frost" that accumulates in a freezer. As shown in Figure 17.2, deposition releases an amount of energy equal to the total amount released by condensation and freezing.

CONCEPT CHECK 17.1

① **Summarize the processes by which water changes from one state of matter to another. Indicate whether energy is absorbed or released.**
② **What is *latent heat*?**
③ **What is a common example of sublimation?**
④ **How does frost form?**

FIGURE 17.3 Frost on a window pane is an example of deposition. (Photo by Stockxpert/Jupiter Images Unlimited)

Students Sometimes Ask...

What is freezer burn?

"Freezer burn" is a common expression used to describe foods that have been left in a frost-free refrigerator for extended periods of time. Frost-free refrigerators circulate comparatively dry air through the freezer compartments. This causes ice on the freezer walls to sublimate (change from a solid to a gas) so it can be removed by the circulating air. Unfortunately, this process also removes moisture from those frozen foods that are not in airtight containers. Consequently, over a period of a few months these foods begin to dry out rather than actually being burned.

Humidity: Water Vapor in the Air

Earth's Dynamic Atmosphere
▸ Moisture and Cloud Formation

Water vapor constitutes only a small fraction of the atmosphere, varying from as little as one-tenth of 1 percent up to about 4 percent by volume. But the importance of water in the air is far greater than these small percentages would indicate. Indeed, scientists agree that *water vapor* is the most important gas in the atmosphere when it comes to understanding atmospheric processes.

Humidity is the general term for the amount of water vapor in air. Meteorologists employ several methods to express the water-vapor content of the air; we examine three: mixing ratio, relative humidity, and dew-point temperature.

Saturation

Before we consider these humidity measures further, it is important to understand the concept of **saturation**. Imagine a closed jar containing water overlain by dry air, both at the same temperature. As the water begins to evaporate from the water surface, a small increase in pressure can be detected in the air above. This increase is the result of the motion of the water-vapor molecules that were added to the air through evaporation. In the open atmosphere, this pressure is termed **vapor pressure** and is defined as that part of the total atmospheric pressure that can be attributed to the water-vapor content.

In the closed container, as more and more molecules escape from the water surface, the steadily increasing vapor pressure in the air above forces more and more of these molecules to return to the liquid. Eventually the number of vapor molecules returning to the surface will balance the number leaving. At that point, the air is said to be *saturated*. If we add heat to the container, thereby increasing the temperature of the water and air, more water will evaporate before a balance is reached. Consequently,