Chapter 8: Detailed Summary

How Water Behaves

This chapter applies the many concepts developed in earlier chapters to the behavior of one of our most important molecules—water. Solid water (ice) is discussed first, followed by liquid water, followed by gaseous water (water vapor).

The most remarkable feature of ice is its crystalline structure, in which water molecules align forming hexagonal open spaces. Because of this formation, ice occupies more volume than liquid water. Ice, therefore, is less dense, which is why it floats on water. Ice is slippery when you hold it in your hands because the warmth from your hands is melting the ice. Ice, however, is also slippery even at temperatures well below freezing because the hexagonal crystalline structure falls apart at the surface where water molecules have little to cling onto. Ice skaters skate on this thin film of water.

Freezing and melting occur at the same time within the interface between solid and liquid water. As water molecules join together to form ice they release energy. Conversely, energy is required for two joined water molecules to come apart. Add heat to a mixture of ice and water and you favor the separation of water molecules (melting). Take heat away and you favor the joining of water molecules (freezing). At the special temperature of 0°C the rate of freezing and melting are the same, which means that ice and liquid water can coexist indefinitely.

Adding a solute to a mixture of ice and water inhibits water molecules from joining the crystal structure because the solute particles decrease the number of liquid water molecules in contact with the ice. The rate of freezing goes down while the rate of melting remains the same. The overall effect is a melting of the ice. A lower temperature is needed to get to the point where melting and freezing rates are the same. The solute, therefore, has the effect of lowering the freezing/melting point.

Just as there is liquid water on the surface of ice below freezing temperatures, there are microscopic ice crystals found within water several degrees above melting. These microscopic ice crystals give the water more volume. As the water warms, these crystals fall apart such that the volume decreases (density increases). Concurrently, the water's volume increases because of thermal expansion. These two opposing trends combine to show a profile whereby water is most dense at 4°C. Because of this, lakes and ponds freeze from the top down.

The dipole–dipole attractions occurring among water molecules are a *cohesive* force, which are the intermolecular attractions occurring within a material. Water can also stick to other materials, such as glass. Such attractions between two different materials are *adhesive* forces. Water's cohesive forces give rise to a strong *surface tension*, which makes the surface of water behave like an elastic skin. Adhesive forces between water and other materials can give rise to capillary action in which water is seen to climb up the walls of the other material. Roots of trees are able to use the groundwater pulled up to them by capillary action within the soil.

Evaporation is the process whereby water molecules escape from the liquid to the gaseous phase. Only the faster-moving water molecules can do this. As the water loses these faster-moving molecules the water cools down. As these faster-moving molecules leave they must overcome the cohesive forces. In doing so they lose energy, which has a cooling effect on the surrounding air. No matter how you look at it, evaporation is a cooling process. The converse holds true such that *condensation* (gas to liquid) is a warming process.

The evaporation of water can take place beneath the surface of water when the water is heated to high temperatures. A bubble of water vapor forms and rises to the surface. This process is called *boiling,* and at 1 atmospheric pressure it occurs for water at 100°C. Higher atmospheric pressures require higher temperatures for boiling. Water boils at lower temperatures with lower atmospheric pressures. In a high vacuum water can be made to boil at its freezing temperature, 0°C.

When heat is applied to liquid water, much of the heat is consumed in breaking hydrogen bonds. Likewise, as water cools down, heat is released as hydrogen bonds form. The net result is that water is most sluggish when it comes to any increase or decrease in temperature. For every degree of temperature change, water holds or releases a lot of energy. Water's capacity to hold on to heat is its *specific heat*, which is relatively high compared to most other substances. Because of water's high specific heat, ocean temperatures do not vary much from summer to winter, which has a significant effect on our climate.

The overall relationship between water and energy is summarized in the final section of this chapter. One gram of ice at 0°C requires 335 joules to melt. This is water's *heat of melting*. To bring this 1 gram of liquid water from 0°C to 100°C requires 418 joules, which is obtained from water's specific heat of 4.18 joules per gram per °C. To transform 100°C liquid water to 100°C water vapor requires an astounding 2259 joules. This is water's *heat of vaporization*, which is large because it involves the complete separation of water molecules from one another.

Summary of Terms

Adhesive Force An attractive force between molecules of two different substances.

Capillary Action The rising of liquid into a small vertical space due to the interplay of cohesive and adhesive forces.

Cohesive Force An attractive force between molecules of the same substance.

Heat Of Condensation The energy released by a substance as it transforms from gas to liquid.

Heat Of Freezing The heat energy released by a substance as it transforms from liquid to solid.

Heat Of Melting The heat energy absorbed by a substance as it transforms from solid to liquid.

Heat Of Vaporization The heat energy absorbed by a substance as it transforms from liquid to gas.

Meniscus The curving of the surface of a liquid at the interface between the liquid surface and its container.

Specific Heat The quantity of heat required to change the temperature of 1 gram of a substance by 1 Celsius degree.

Surface Tension The elastic tendency found at the surface of a liquid.