Specific heat, heat of vaporization, and density of water

Specific heat capacity and heat of vaporization of water. Evaporative cooling. Why ice floats.

Introduction

Let's imagine that it's a hot day. You've just been out in the sun for a while, and you're sweating quite a bit as you sit down and grab a glass of cool ice water. You idly notice both the sweat beads on your arms and the chunks of ice floating at the top of your water glass. Thanks to your hard work studying the properties of water, you recognize both the sweat on your arms and the floating ice cubes in your glass as examples of water's amazing capacity for hydrogen bonding. How does that work? Water molecules are very good at forming hydrogen bonds, weak associations between the partially positive and partially negative ends of the molecules. Hydrogen bonding explains both the effectiveness of evaporative cooling (why sweating cools you off) and the low density of ice (why ice floats).

Here, we'll take a closer look at the role of hydrogen bonding in temperature changes, freezing, and vaporization of water.

Water: Solid, liquid, and gas

Water has unique chemical characteristics in all three states—solid, liquid, and gas—thanks to the ability of its molecules to hydrogen bond with one another. Since living things, from human beings to bacteria, have a high water content, understanding the unique chemical features of water in its three states is key to biology.

In liquid water, hydrogen bonds are constantly being formed and broken as the water molecules slide past each other. The breaking of these bonds is caused by the energy of motion (kinetic energy) of the water molecules due to the heat contained in the system.

When the heat is raised (for instance, as water is boiled), the higher kinetic energy of the water molecules causes the hydrogen bonds to break completely and allows water molecules to escape into the air as gas. We observe this gas as water vapor or steam.

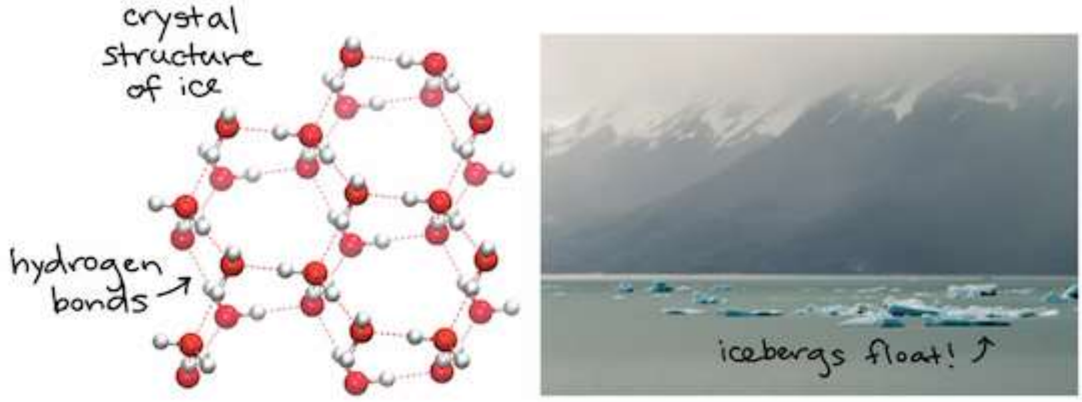
On the other hand, when the temperature drops and water freezes, water molecules form a crystal structure maintained by hydrogen bonding (as there is too little heat energy left to break the hydrogen bonds). This structure makes ice less dense than liquid water.

Density of ice and water

Water's lower density in its solid form is due to the way hydrogen bonds are oriented as it freezes. Specifically, in ice, the water molecules are pushed farther apart than they are in liquid water.

That means water expands when it freezes. You may have seen this for yourself if you've ever put a sealed glass container containing a mostly-watery food (soup, soda, etc.) into the freezer, only to have it crack or explode as the liquid water inside froze and expanded.

With most other liquids, solidification—which occurs when the temperature drops and kinetic (motion) energy of molecules is reduced—allows molecules to pack more tightly than in liquid form, giving the solid a greater density than the liquid. Water is an anomaly (that is, a weird standout) in its lower density as a solid.



(Left) Crystal structure of ice, with water molecules held in a regular 3D structure by hydrogen bonds. (Right) Image of icebergs floating on the surface of the ocean.

Image: modified from OpenStax Biology. Modifications of work by Jane Whitney (left), image created using Visual Molecular Dynamics (VMD) software (Humphrey, 1996), and by Carlos Ponte (right).

Because it is less dense, ice floats on the surface of liquid water, as we see for an iceberg or the ice cubes in a glass of iced tea. In lakes and ponds, a layer of ice forms on top of the liquid water, creating an insulating barrier that protects the animals and plant life in the pond below from freezing.

Why is it harmful for living things to freeze? We can understand this by thinking back to the case of a bottle of soda pop cracking in the freezer. When a cell freezes, its watery contents expand and its membrane (just like the soda bottle) is broken into pieces.

Heat capacity of water

It takes a lot of heat to increase the temperature of liquid water because some of the heat must be used to break hydrogen bonds between the molecules. In other words, water has a high **specific heat capacity**, which is defined as the amount of heat needed to raise the temperature of one gram of a substance by one degree Celsius. The amount of heat needed to raise the temperature of 1 g water by 1 °C is has its own name, the **calorie**.

This calorie is similar to that one, but not exactly the same. The “food calorie” you see on the back of a candy bar is is actually a **kilocalorie**, or a thousand small calories.

Because of its high heat capacity, water can minimize changes in temperature. For instance, the specific heat capacity of water is about five times greater than that of sand. The land cools faster than the sea once the sun goes down, and the slow-cooling water can release heat to nearby land during the night. Water is also used by warm-blooded animals to distribute heat through their bodies: it acts similarly to a car’s cooling system, moving heat from warm places to cool places, helping the body keep an even temperature.

Heat of vaporization of water

Just as it takes a lot of heat to increase the temperature of liquid water, it also takes an unusual amount of heat to vaporize a given amount of water, because hydrogen bonds must be broken in order for the molecules to fly off as gas. That is, water has a high **heat of vaporization**, the amount of energy needed to change one gram of a liquid substance to a gas at constant temperature.

Water’s heat of vaporization is around 540 cal/g at 100 °C, water's boiling point. Note that *some* molecules of water – ones that happen to have high kinetic energy – will escape from the surface of the water even at lower temperatures.

As water molecules evaporate, the surface they evaporate from gets cooler, a process called **evaporative cooling**. This is because the molecules with the highest kinetic energy are lost to evaporation. In humans and other organisms, the evaporation of sweat, which is 90% water, cools the body to maintain a steady temperature.