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Melting Point, Freezing Point, Boiling Point

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Melting Point and Freezing Point

Pure, crystalline solids have a characteristic **melting point**, the temperature at which the solid melts to become a liquid. The transition between the solid and the liquid is so sharp for small samples of a pure substance that melting points can be measured to 0.1ºC. The melting point of solid oxygen, for example, is -218.4ºC.

Liquids have a characteristic temperature at which they turn into solids, known as their **freezing point**. In theory, the melting point of a solid should be the same as the freezing point of the liquid. In practice, small differences between these quantities can be observed.

It is difficult, if not impossible, to heat a solid above its melting point because the heat that enters the solid at its melting point is used to convert the solid into a liquid. It is possible, however, to cool some liquids to temperatures below their freezing points without forming a solid. When this is done, the liquid is said to be *supercooled*.

An example of a supercooled liquid can be made by heating solid sodium acetate trihydrate ($NaCH_3CO_2$ 3 H_2O). When this solid melts, the sodium acetate dissolves in the water that was trapped in the crystal to form a solution. When the solution cools to room temperature, it should solidify. But it often doesn't. If a small crystal of sodium acetate trihydrate is added to the liquid, however, the contents of the flask solidify within seconds.

A liquid can become supercooled because the particles in a solid are packed in a regular structure that is characteristic of that particular substance. Some of these solids form very easily; others do not. Some need a particle of dust, or a seed crystal, to act as a site on which the crystal can grow. In order to form crystals of sodium acetate trihydrate, Na^+ ions, $CH_3CO_2^-$ ions, and water molecules must come together in the proper orientation. It is difficult for these particles to organize themselves, but a seed crystal can provide the framework on which the proper arrangement of ions and water molecules can grow.

Because it is difficult to heat solids to temperatures above their melting points, and because pure solids tend to melt over a very small temperature range, melting points are often used to help identify compounds. We can distinguish between the three sugars known as *glucose* (*MP* = 150ºC), *fructose* (*MP* = 103-105ºC), and *sucrose* (*MP* = 185-186ºC), for example, by determining the melting point of a small sample.

Measurements of the melting point of a solid can also provide information about the purity of the substance. Pure, crystalline solids melt over a very narrow range of temperatures, whereas mixtures melt over a broad temperature range. Mixtures also tend to melt at temperatures below the melting points of the pure solids.

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Boiling Point

When a liquid is heated, it eventually reaches a temperature at which the vapor pressure is large enough that bubbles form inside the body of the liquid. This temperature is called the **boiling point**. Once the liquid starts to boil, the temperature remains constant until all of the liquid has been converted to a gas.

The normal boiling point of water is 100ºC. But if you try to cook an egg in boiling water while camping in the Rocky Mountains at an elevation of 10,000 feet, you will find that it takes longer for the egg to cook because water boils at only 90ºC at this elevation.

In theory, you shouldn't be able to heat a liquid to temperatures above its normal boiling point. Before microwave ovens became popular, however, pressure cookers were used to decrease the amount of time it took to cook food. In a typical pressure cooker, water can remain a liquid at temperatures as high as 120ºC, and food cooks in as little as one-third the normal time.

To explain why water boils at 90ºC in the mountains and 120ºC in a pressure cooker, even though the normal boiling point of water is 100ºC, we have to understand why a liquid boils. By definition, a liquid boils when the vapor pressure of the gas escaping from the liquid is equal to the pressure exerted on the liquid by its surroundings, as shown in the figure below.

graph

Liquids boil when their vapor pressure is equal to the pressure exerted on the liquid by its surroundings.

The normal boiling point of water is 100ºC because this is the temperature at which the vapor pressure of water is 760 mmHg, or 1 atm.

Under normal conditions, when the pressure of the atmosphere is approximately 760 mmHg, water boils at 100ºC. At 10,000 feet above sea level, the pressure of the atmosphere is only 526 mmHg. At these elevations, water boils when its vapor pressure is 526 mmHg, which occurs at a temperature of 90ºC.

Pressure cookers are equipped with a valve that lets gas escape when the pressure inside the pot exceeds some fixed value. This valve is often set at 15 psi, which means that the water vapor inside the pot must reach a pressure of 2 atm before it can escape. Because water doesn't reach a vapor pressure of 2 atm until the temperature is 120ºC, it boils in this container at 120ºC.

Liquids often boil in an uneven fashion, or *bump*. They tend to bump when there aren't any scratches on the walls of the container where bubbles can form. Bumping is easily prevented by adding a few boiling chips to the liquid, which provide a rough surface upon which bubbles can form. When boiling chips are used, essentially all of the bubbles that rise through the solution form on the surface of these chips.

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