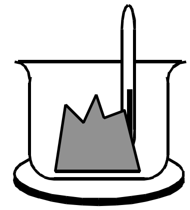
# PhyzGuide: Phase Changes

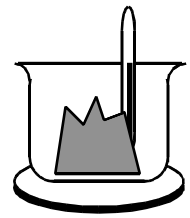
## MELTING AND FREEZING

When heat is added to an object, the object generally undergoes an increase in temperature. A hot dog goes from 10 °C to 100 °C as it absorbs heat from the boiling water surrounding it. (The rate of temperature increase is determined by the specific heat of the substance.)



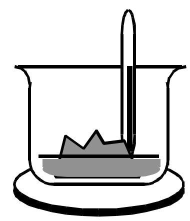
**Heat is added to ice at −10°C, and its temperature increases**

There are special temperatures at which substances no longer undergo increases in temperature, and instead undergo changes of state (or phase). When heat is added to ice at 0 °C, the ice's temperature does not increase. Instead, the ice melts and becomes water at 0 °C. There will be no increase in temperature until all* the ice has melted. There has been a change of phase from the solid state to the liquid state. Zero degrees Celsius is designated as the **melting point**—denoted MP or $T_m$—of ice (solid water). The process works equally well in reverse: removing heat from water at 0 °C does not result in a decrease in the temperature of the water. It results in a change of phase: from liquid at 0 °C to solid at 0 °C—the water becomes ice. Only after all the water has changed to ice can there be a reduction in temperature. Zero degrees Celsius is also designated as the **freezing point** of water.



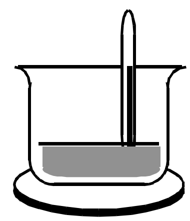
**Heat is added to ice at 0°C, and its temperature does not increase!**

For a given substance, the terms "melting point" and "freezing point" refer to the same temperature. The terms simply distinguish the direction of the phase change: if the substance is going from solid to liquid, the term "melting point" is used; if the substance is going from liquid to solid, the term "freezing point" is used.



**The temperature stays constant as the ice becomes water**

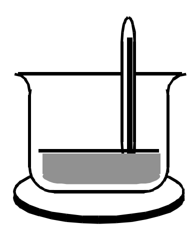
This "change of state" process is most peculiar to those with active curiosities (i.e. "physics types"). Just how much heat has to be added (or removed) while the change of state occurs? It varies for different substances. Physicists call this quantity of energy **heat of fusion.** It is denoted $L_f$ and is measured in units of J/kg. The heat of fusion of water is 335,000 J/kg. That means it takes 335,000J of heat to change 1 kg of ice at 0 °C to 1 kg of water at 0 °C. If it takes 335,000 J to completely melt *1 kg* of ice, how much energy do you suppose is required to melt *2 kg* of ice? That's right: 2 × 335,000 = 670,000. The amount of energy $Q$ that must be added to a mass $m$ of a substance to melt it is equal to the amount of energy that must be removed to freeze it, and is given by the following relation.



**The temperature stays constant until all* the ice becomes water**

$$Q = mL_f$$

Q is the total energy to be added (if melting) or removed (if freezing)
m is the mass of the sample undergoing the phase change
$L_f$ is the heat of fusion of the given substance



**Further heat added to the liquid results in an increase in temperature**

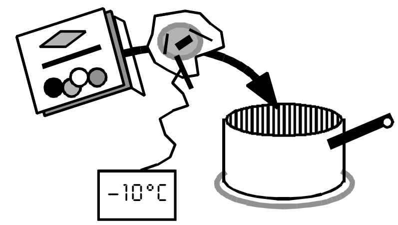
* "nearly all" would be more accurate; "microscopic slush" persists slightly beyond 4°C

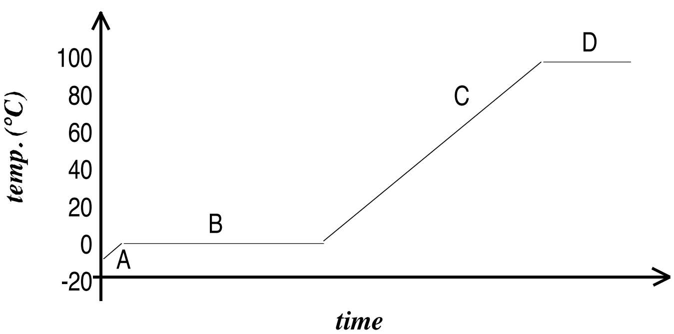
## DINNER AT DEAN'S

*An illustrative example to further confuse these strange new terms:*

One night Mr. B. prepared to fix one of his famous "home-cooked" meals. He poured a kilogram or so of water into a pot and set it on his thermal activator/molecule agitator (i.e., his stove) and began adding energy to it. Soon enough, the water reached its boiling point. Mr. B. promptly tossed his Swanson's™ brand "Boil-N-Bag" of Chicken à la King into the bubbling water. Of course, he'd attached a temperature sensor to this boil-n-bag so he could observe the physics of this amazing culinary process. As any good phyzteacher would do, he quickly graphed the results and added comments of his own to aid in your understanding.




**Boil-n-bag with temperature probe**




A. First, the SOLID Chicken à la King undergoes an increase in temperature.

B. As the SOLID melts to LIQUID, the temperature remains constant.

C. The LIQUID Chicken à la King undergoes an increase in temperature until it reaches the temperature of the boiling water at D.

## VAPORIZING AND CONDENSING

Why is it that when you heat water on the stove (or in the lab) it doesn't get any hotter than 100 °C? The heat you added when the water was at 20 °C increased the water's temperature, but the heat you add at 100 °C seems to do nothing. Why doesn't the temperature continue to rise? Why does the temperature stop rising at 100 °C and not 90 °C or 127.3 °C?

Another phase change occurs for water at 100 °C. Just as solid water (ice) changes to liquid at 0 °C, liquid water changes to vapor (steam) at 100 °C. We call this process—the change from liquid to vapor—**vaporization**. The temperature at which this occurs is called the **boiling point** (BP or $T_b$). For water, that temperature is 100 °C. If steam is allowed to cool, it will begin to **condense** (change from vapor to liquid) at 100 °C. During vaporization and condensation, temperature remains constant. The quantity of energy required for a phase change of one kilogram of a substance at its boiling point is called **heat of vaporization** $L_v$. The amount of energy $Q$ that must be added to a mass $m$ of a substance to vaporize it is equal to the amount of energy that must be removed to condense it, and is given by the following relation.

$$Q = mL_v$$

Q is the total energy to be added (if vaporizing) or removed (if condensing)
m is the mass of the sample undergoing the phase change
$L_V$ is the heat of vaporization of the given substance