<u>THERMODYNAMICS - Heat and Heat Transfer</u>: Heat (***Q***) is a form of <u>Energy</u> that is transferred between an object and another object or its surrounding environment due to a ***difference*** *in Temperature*. Heat is <u>always</u> transferred **from higher T to lower T**.

Heat can be transferred through 3 primary mechanisms:

1. <u>Conduction</u>: Requires physical contact. Primary means of transferring heat in **solids**. Example: frying an egg in a skillet.

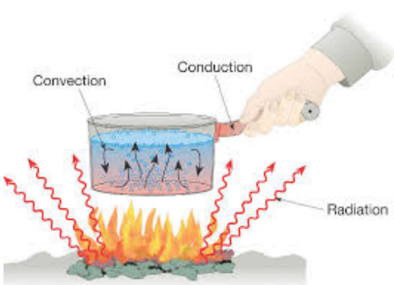
   > When you touch something, the sensation of "hot" or "cold" indicates the direction of heat flow; <u>Hot</u> means heat is transferred <u>to</u> you, *Cold* means heat is transferred *away from* you. This is due to the vibrating molecules (*recall: temperature measures molecular KE ⇒ motion*) colliding and transferring some KE during the collision.

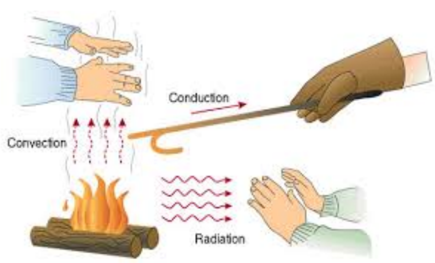
2. <u>Convection</u>: Heat transfer using "currents". Primary means of transferring heat in **liquids**. Example: boiling water in a pot.

   > When you heat a fluid (*i.e., water*), the hotter material will generally become less dense than the colder material, and have a tendency to rise, pulling colder material downwards. As the hot stuff rises it cools; as the cold stuff falls, it warms up, and the cycle repeats. This forms **Convection Currents**, and these are very important for studying weather.

3. <u>Radiation</u>: Heat transfer by IR radiation. How heat is transferred through space; also works well in **gases**. Example: the heat from the sun.

   > A hot object will give off IR radiation; the hotter it is, the more intense the radiation (*e.g., a bonfire vs. a match*). We can feel the transfer of heat as the warmth we feel near a fire.





These diagrams illustrates how Conduction, Convection and Radiation occur near an open flame.

<u>Heat Transfer - Calculation</u>:

To determine the amount of heat energy that is transferred, in relation to the change in temperature, we use the relationship

$$\boxed{\mathbf{Q = mc\Delta T}}$$

where **Q** is the amount of Heat [J], **m** is the mass of the sample [kg], **c** is a material property known as the *Specific Heat Capacity* [J/kg*K], and Δ**T** is the change in temperature.

Note that, when* ΔT** *is positive, Q is positive; when* Δ**T** *is negative, Q is negative.*

**Recall: Conservation of Energy implies that energy can neither be created nor destroyed; any energy that is absorbed must come from somewhere, and energy that is emitted must go somewhere. Often we will compare two substances, equating the amount of energy that one loses with the amount of energy that the other gains.

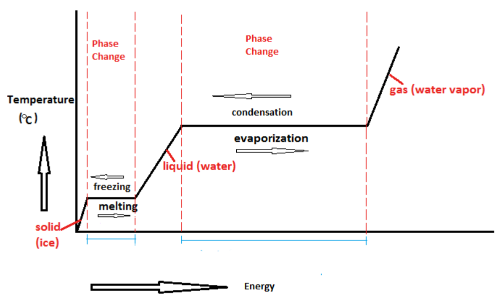
<u>Examples</u>:

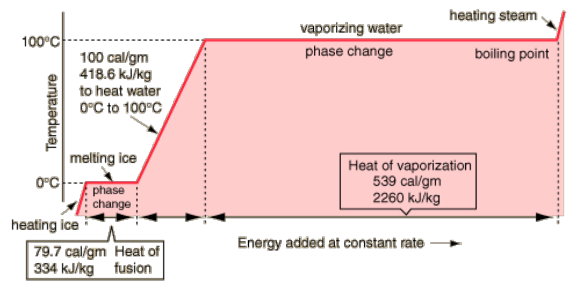
1. Water (c = 4,186 J/kg*K) at room temperature (20 °C) is heated until it reaches 40 °C. If the sample of water has a mass of 0.25 kg, how much heat energy is added to the water?

2. A cup of iced coffee, initially at a temperature of 5 °C, is left on a table, where it gradually warms to room temperature, 20 °C. If the mass of the cup of iced coffee is 0.10 kg, how much heat did the iced coffee lose? Assume iced coffee has the same specific heat as water.

3. In a blacksmith's shop, a very hot piece of iron is dunked into a bucket of water to rapidly cool it. Assume no energy escapes to the surrounding environment. The specific heat for iron is c_{Fe} = 450 J/kg*K, and for water is c_w = 4,186 J/kg*K. Assume the initial temperature of iron is 870 °C, and for the water it is 40 °C. Further, assume the mass of the piece of iron is m_{Fe} = 0.5 kg, and the mass of water in the bucket is 3 kg. Determine the final temperature of the iron-and-water. Express this temperature in °C.

<u>Heat Transfer - Phase Change</u>: When a substance reaches a certain threshold temperature, the continued transfer of heat will have no effect on the temperature, but instead will be used to alter the state of the substance by changing its phase. The primary processes we will investigate are Boiling/ Condensation, and Freezing/Melting.





- **<u>Freezing/Melting</u>:**

- When a substance reaches its freezing point (*e.g., 0 °C for water*), continually adding energy will convert that substance from a solid to a liquid, with no observable change in temperature.
  > When a *Liquid* is cooled to the "Freezing" point, removing heat will cause it to condense into a solid (*e.g., ice cubes*)

  > The phase from liquid ↔ solid is generally referred to as ***Fusion*** and the amount of energy required is determined by the mass of the substance, as well as its ***Latent Heat of Fusion*** $\mathbf{L_F}$ (for water, 3.34 E5 J/kg)

  – This implies that, to <u>completely convert</u> 1 kg of ice, at 0 °C, to water requires the addition of 334,000 J of energy!

  > In equation form, we have:

$$\boxed{\mathbf{Q = mL_F}}$$

Examples:

1. A small amount of water (0.05 kg) is to be frozen into an ice cube. If its temperature is 0 °C, how much energy would need to be *removed* from the water?

2. A 10 kg ice sculpture is at 0 °C. How much energy must it absorb for it to completely melt to liquid water?

3. *A 0.05 kg ice chunk is placed into a glass of iced tea. Assume the ice is at an initial temperature of 0 °C. If the ice totally melts, how much energy did it absorb? Assuming that the iced tea had an initial temperature of 15 °C and that its mass is 0.50 kg, what is the final temperature of the iced tea? (*Hint: first determine the amount of energy needed to melt the ice; then, this energy is* <u>*lost*</u> *by the iced tea; use this to determine the final temperature.*)

- **<u>Boiling/Condensation</u>:**

Boiling

Condensation

- When a substance reaches its boiling point (*e.g., 100 °C for water*), continually adding energy will convert that substance from a liquid to a gas, with no observable change in temperature.
  - > When a *Gas* is cooled to the "Boiling" point, removing heat will cause it to condense into a liquid (*e.g., the "sweat" on a cold glass on a hot day*)
  - > The phase from liquid ↔ gas is generally referred to as ***Vaporization***, and the amount of energy required is determined by the mass of the substance, as well as its ***Latent Heat of Vaporization,*** L_V (for water, 2.256 E6 J/kg)
    - – This implies that, to <u>completely convert</u> 1 kg of water, at 100 °C, to steam requires the addition of 2,256,000 J of energy! (*This is why you can get quite severely burned by steam; it has lots of heat energy!*)
  - > In equation form,

$$\boxed{Q = mL_V}$$

Examples:

1. A 0.1 kg sample of water, at its boiling point, is converted to steam. What is the amount of energy absorbed by the water?

2. A very small amount of water (0.002 kg) condenses on the side of a bottle of water. How much energy does the water vapor lose as it condenses?

3. In order to cook pasta, a 1.20 kg pot of water is brought from a temperature of 40 °C to its boiling point of 100 °C. How much energy does this require? After the pasta cooks, it is determined that 0.06 kg of water has been completely boiled into steam. How much energy does this require? Which takes more energy: raising the temperature of the water to the boiling point, or boiling the water itself? What does that tell you about the water as you begin to cook pasta?