Thermal Energy

Ck12 Science

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Printed: April 9, 2014

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CHAPTER 1 Thermal Energy

CHAPTER OUTLINE

We often think of temperature and heat as the same thing. While they both involve molecular motion, they are not in fact the same. Thermal energy, or heat, depends on more than just the temperature. In the image above of a river running across the surface of a glacier, the liquid water and the ice have the same temperature, and yet they are in different states. In this chapter, we will uncover the realities of thermal energy, heat, temperature, and changing states of matter.

1.1 Heat, Temperature, and Thermal Energy Transfer

- Define heat.
- Define temperature.
- Describe thermal energy transfer.
- Define Celsius and Kelvin temperature scales.
- Convert Celsius temperatures to Kelvin and vice versa.

The temperature of basalt lava at Kilauea (Hawaii) reaches 1,160 degrees Celsius (2,120 degrees Fahrenheit). A crude estimation of temperature can be determined by looking at the color of the rock: orange-to-yellow colors are emitted when rocks (or metals) are hotter than about 900 degrees Celsius; dark-to-bright cherry red is characteristic as material cools to 630 degrees Celsius; faint red glow persists down to about 480 degrees Celsius. For comparison, a pizza oven is commonly operated at temperatures ranging from 260 to 315 degrees Celsius.

Heat, Temperature, and Thermal Energy Transfer

The first theory about how a hot object differs from a cold object was formed in the 18th century. The suggested explanation was that when an object was heated, an invisible fluid called "caloric" was added to the object. Hot objects contained more caloric than cold objects. The caloric theory could explain some observations about heated objects (such as that the fact that objects expanded as they were heated) but could not explain others (such as why your hands got warm when you rub them together).

In the mid-19th century, scientists devised a new theory to explain heat. The new theory was based on the assumption that matter is made up of tiny particles that are always in motion. In a hot object, the particles move faster and therefore have greater kinetic energy. The theory is called the kinetic-molecular theory and is the accepted theory of heat. Just as a baseball has a certain amount of kinetic energy due to its mass and velocity, each molecule has a certain amount of kinetic energy due to its mass and velocity. Adding up the kinetic energy of all the molecules in an object yields the **thermal energy** of the object.

When a hot object and a cold object touch each other, the molecules of the objects collide along the surface where they touch. When higher kinetic energy molecules collide with lower kinetic energy molecules, kinetic energy is passed from the molecules with more kinetic energy to those with less kinetic energy. In this way, heat always flows from hot to cold and heat will continue to flow until the two objects have the same temperature. The movement of heat from one object to another by molecular collision is called **conduction**.

Heat is the energy that flows as a result of a difference in temperature. We use the symbol Q for heat. Heat, like all forms of energy, is measured in joules.

The **temperature** of an object is a measurement of the average kinetic energy of all the molecules of the object. You should note the difference between heat and temperature. Heat is the *sum* of all the kinetic energies of all the molecules of an object, while temperature is the *average* kinetic energy of the molecules of an object. If an object was composed of exactly three molecules and the kinetic energies of the three molecules are 50 J, 70 J, and 90 J, the heat would be 210 J and the temperature would be 70 J.

The terms *hot* and cold refer to temperature. A hot object has greater average kinetic energy but may not have greater total kinetic energy. Suppose you were to compare a milliliter of water near the boiling point with a bathtub full of water at room temperature. The bathtub contains a billion times as many water molecules, and therefore has a higher total kinetic energy and more heat. Nonetheless, we would consider the bathtub colder because its average kinetic energy, or temperature, is lower.

Temperature Scales: Celsius and Kelvin

A **thermometer** is a device used to measure temperature. It is placed in contact with an object and allowed to reach thermal equilibrium with the object (they will have the same temperature). The operation of a thermometer is based on some property, such as volume, that varies with temperature. The most common thermometers contain liquid mercury, or some other liquid, inside a sealed glass tube. The liquid expands and contracts faster than the glass tube. Therefore, when the temperature of the thermometer increases, the liquid volume expands faster than the glass volume, allowing the liquid to rise in the tube. The positions of the liquid in the tube can then be calibrated for accurate temperature readings. Other properties that change with temperature can also be used to make thermometers; liquid crystal colors and electrical conductivity change with temperature, and are also relatively common thermometers.

The most commonly used temperature scale in the United States is the Farenheit scale. However, this scale is rarely used throughout the world; the metric temperature scale is Celsius. This scale, based on the properties of water, was devised by the Swedish physicist, Anders Celsius (1704 –1744). The freezing point of water is 0°C and the boiling point of water was assigned to be 100°C. The kinetic energies between these two points was divided evenly into 100 “degrees Celsius”.

The Kelvin or “Absolute” temperature scale is the scale often used by chemists and physicists. It is based on the temperature at which all molecular motion ceases; this temperature is called absolute zero and is 0 K. This temperature corresponds to -273.15°C. Since absolute zero is the coldest possible temperature, there are no negative values on the Kelvin temperature scale. Conveniently, the Kelvin and Celsius scales have the same definition of a degree, which makes it very easy to convert from one scale to the other. The relationship between Celsius and Kelvin temperature scales is given by:

$$K = {}^\circ C + 273.15$$

On the Kelvin scale, water freezes at 273 K and boils at 373 K.

Example Problem: Convert 25°C to Kelvin.

Solution: $K = {}^\circ C + 273 = 25^\circ C + 273 = 298\ K$

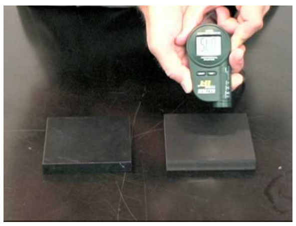
Summary

- The thermal energy, or heat, of an object is obtained by adding up the kinetic energy of all the molecules within it.
- Temperature is the average kinetic energy of the molecules.
- Absolute zero is the temperature where molecular motion stops and is the lowest possible temperature.
- Zero on the Celsius scale is the freezing point of water and 100°C is the boiling point of water.
- The relationship between Celsius and Kelvin temperature scales is given by K = °C + 273.15.

Practice

The following video demonstrates how some materials conduct heat better than others. Use this resource to answer the questions that follow.

http://www.youtube.com/watch?v=FmG1Sc0AS3s



MEDIA

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1. Which material was a better conductor of heat?
2. Explain why metals feel cold even when thay are at room temperature.

Heat and temperature practice problems and questions:

http://www.education.com/study-help/article/telllperature-heat_answer/

Review

1. Convert 4.22 K to °C.
2. Convert 37°C to K.
3. If you had beeswax attached to one end of a metal skewer and you placed the other end of the skewer in a flame, what would happen after a few minutes?
4. Which contains more heat, a coffee cup of boiling water or a bathtub of room temperature water?

- **thermal energy:** The total energy of a substance particles due to their translational movement or vibrations.
- **heat:** energy transferred from one body to another by thermal interactions.
- **temperature:** A measurement of the average kinetic energy of the molecules in an object or system and can be measured with a thermometer or a calorimeter.
- **conduction:** The transfer of thermal energy by the movement of particles that are in contact with each other.
- **absolute zero:** The lowest possible temperature, at which point the atoms of a substance transmit no thermal energy - they are completely at rest. It is 0 degrees on the Kelvin scale, which translates to -273.15 degrees Celsius (or -459.67 degrees Fahrenheit).

1.2 Specific Heat

- Define specific heat.
- Calculate heat transfer.

This image is of the Beehive Geyser in Yellowstone National Park. Underground water is heated by the earth's molten core and, when sufficient pressure is built up, the water shoots out of the ground in an amazing display.

Specific Heat

When heat flows into an object, its thermal energy increases and so does its temperature. The amount of temperature increase depends on three things: 1) how much heat was added, 2) the size of the object, and 3) the material of which the object is made. When you add the same amount of heat to the same mass of different substances, the amount of temperature increase is different. Each substance has a **specific heat,** which is the amount of heat necessary to raise one mass unit of that substance by one temperature unit.

In the SI system, specific heat is measured in J/kg•K. (Occasionally, you may also see specific heat expressed sometimes in J/g•K). The specific heat of aluminum is 903 J/kg•K. Therefore, it requires 903 J to raise 1.00 kg of aluminum by 1.00 K.

TABLE 1.1: Specific Heat of Some Common Substances

Material	Specific Heat (J/kg•K)
Aluminum	903
Brass	376
Carbon	710
Copper	385
Glass	664

TABLE 1.1: (continued)

Ice	2060
Lead	130
Methanol	2450
Water Vapor	2020
Water (liquid)	4180
Zinc	388

The amount of heat gained or lost by an object when its temperature changes can be calculated by the formula

$Q = mc\Delta t$,

where Q is the heat gained or lost, m is the mass of the object, c is its specific heat, and Δt is the change in temperature. You should note that the size of a Celsius degree and a Kelvin degree are exactly the same, and therefore Δt is the same whether measured in Celsius or Kelvin.

Example Problem: A 0.500 kg block of zinc is heated from 295 K to 350. K. How much heat was absorbed by the zinc?

Solution: $Q = mc\Delta t = (0.500\ kg)(388\ J/kg \cdot K)(350.\ K - 295\ K) = 10,600\ J$

Example Problem: 845 J of heat are added to a 0.200 kg block of aluminum at a temperature of 312.00 K. How high will the temperature of the aluminum rise?

Solution: $(t_2 - t_1) = \frac{Q}{mc} = \frac{845\ J}{(0.200\ kg)(903\ J/kg \cdot K)} = 4.68\ K$

$t_2 = t_1 + 4.68\ K = 312.00\ K + 4.68\ K = 316.68\ K$

Summary

- When heat flows into an object, its thermal energy increases and so does its temperature.
- The amount of temperature increase depends on three things: 1) how much heat was added, 2) the size of the object, and 3) the material of which the object is made.
- Each substance has a specific heat, which is the amount of heat necessary to raise one mass unit of that substance by one temperature unit.
- The amount of heat gained or lost by an object when its temperature changes can be calculated by the formula $Q = mc\Delta t$.

Practice

Interactive step by step demonstration of how to calculate the specific heat of a material.

http://www.chem.uiuc.edu/webfunchem/specificheat/newSample.htm

Practice problems in specific heat.

http://www.kwanga.net/chemnotes/specific-heat-practice.pdf

Review

1. How much heat is absorbed by 60.0 g of copper when it is heated from 20.0°C to 80.0°C?
2. A 40.0 kg block of lead is heated from -25°C to 200.°C. How much heat is absorbed by the lead block?
3. The cooling system of an automobile motor contains 20.0 kg of water. What is the Δt of the water if the engine operates until 836,000 J of heat have been added to the water?

- **specific heat:** The heat required to raise a unit mass of a substance by one unit temperature interval under specified conditions, such as constant pressure: usually measured in joules per kelvin per kilogram.

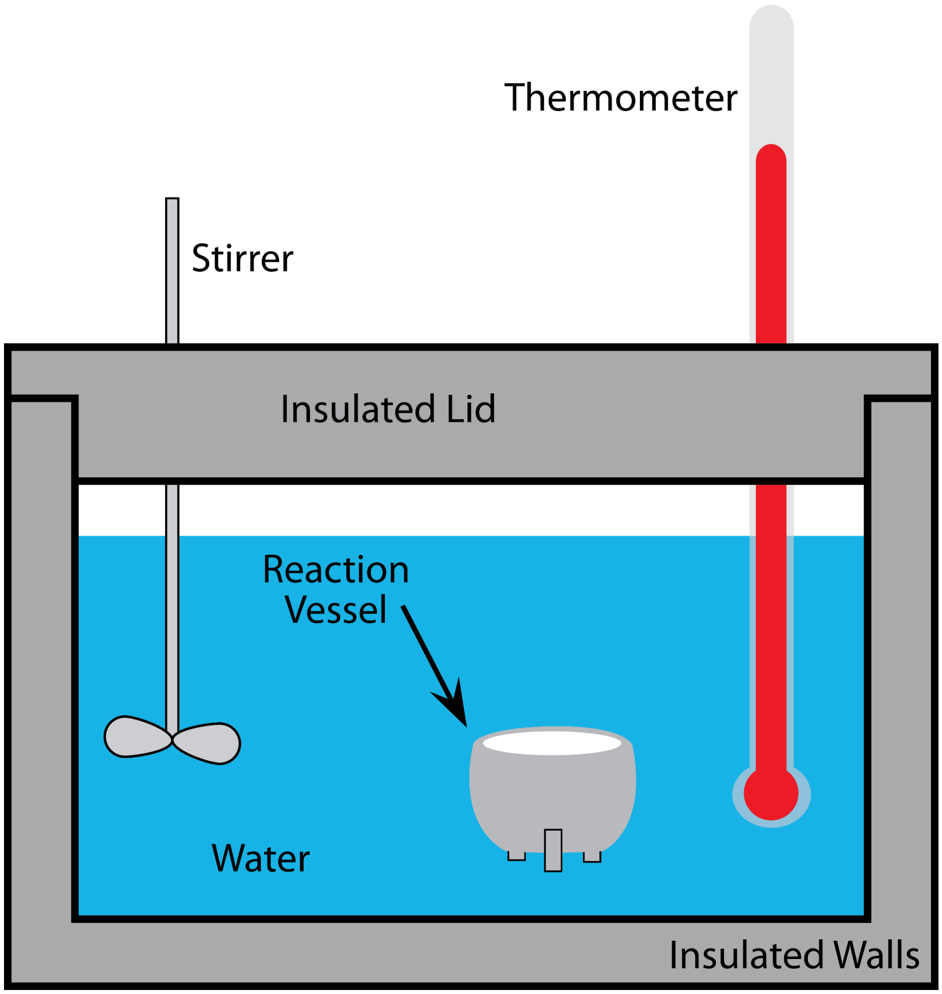
1.3 Calorimetry

- Understand the function of a calorimeter.
- Apply the law of conservation of energy to calorimeter calculations of temperature changes due to heat transfer.

Though not particularly beautiful machines, calorimeters are incredibly useful ones. They are used to determine the calories (food energy) in food, as well as the average heat yield from burning various grades of coal and oil. The price of coal is often dependent on the heat yield from samples burned in a calorimeter.

Calorimetry

A **calorimeter** is a device used to measure changes in thermal energy or heat transfer. More specifically, it measures **calories.** A calorie is the amount of energy required to raise one gram of water by one degree Celsius. As such, the calorimeter measures the temperature change of a known amount of water. If a reaction is carried out in the reaction vessel, or if a measured mass of heated substance is placed in the water of the calorimeter, the change in the water temperature allows us to calculate the change in thermal energy.



The function of the calorimeter depends on the conservation of energy in a closed, isolated system. Calorimeters are carefully insulated so that heat transfer in or out is negligible. Consider the following example.

Example Problem: A 0.500 kg sample of water in a calorimeter is at 15.0ºC. A 0.0400 kg block of zinc at 115.0ºC is placed in the water. The specific heat of zinc is 388 J/kg•ºC. Find the final temperature of the system.

Solution: The heat lost by the block of zinc will equal the heat gain by the water in the calorimeter. In order to set heat gain mathematically equal to heat loss, either one of the terms must be made negative or the temperature change must be reversed. You should also note that the final temperature of the water and the block of zinc will be the same when equilibrium is reached.

$m_W c_W (t_2 - t_1)_W = m_{Zn} c_{Zn} (t_1 - t_2)_{Zn}$

$(0.500 \text{ kg})(4180 \text{ J/kg}^\circ\text{C})(x - 15.0^\circ\text{C}) = (0.0400 \text{ kg})(388 \text{ J/kg}^\circ\text{C})(115.0^\circ\text{C} - x)$

$2090\,x - 31350 = 1785 - 15.52\,x$

$2105.52\,x = 33135$

$x = 15.7^\circ\text{C}$

Example Problem: A 100. g block of aluminum at 100.0ºC is placed in 100. g of water at 10.0ºC. The final temperature of the mixture is 25.0ºC. What is the specific heat of the aluminum as determined by the experiment?

Solution:

$m_W c_W (t_2 - t_1)_W = m_{Al} c_{Al} (t_1 - t_2)_{Al}$

$(0.100\ kg)(4180\ J/kg^\circ\text{C})(25.0^\circ\text{C} - 10.0^\circ\text{C}) = (0.100 \text{ kg})(x)(100.0^\circ\text{C} - 25.0^\circ\text{C})$

$6270 = 7.50\,x$

$x = 836 \text{ J/kg}^\circ\text{C}$

Summary

- A calorimeter is a device used to measure changes in thermal energy or heat transfer.
- If a reaction is carried out in the reaction vessel or if a measured mass of heated substance is placed in the water of the calorimeter, the change in the water temperature allows us to calculate the change in thermal energy.

Practice

The following video covers the calorimetry equation. Use this resource to answer the questions that follow.

http://youtu.be/L9spPoot3fU

1. What is the number 4.18 J/g°C in the video?
2. In the equation, q = mcΔt, what does c represent?
3. What does it mean if the temperature in the calorimeter goes down?

Solved calorimetry problems:

http://calorimetry-physics-problems.blogspot.com/2010/10/specific-heat-problems.html

Review

1. A 300.0 g sample of water at 80.0ºC is mixed with 300.0 g of water at 10.0ºC. Assuming no heat loss to the surroundings, what is the final temperature of the mixture?
2. A 400.0 g sample of methanol at 16.0ºC is mixed with 400.0 g of water at 85.0ºC. Assuming no heat loss to the surroundings, what is the final temperature of the mixture? The specific heat of methanol is 2450 J/kg•ºC.
3. A 100.0 g brass block at 100.0ºC is placed in 200.0 g of water at 20.0ºC. The specific heat of brass is 376 J/kg•ºC. Assuming no heat loss to the surroundings, what is the final temperature of the mixture?

- **calorimeter:** A device used to measure the heat flow of a chemical reaction or physical change.
- **calorimetry:** A way to measure the energy change of a reaction or the energy contained in matter.

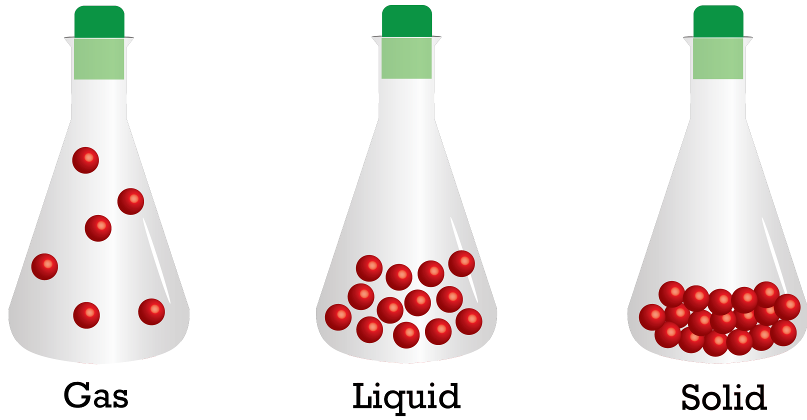
1.4 Change of State

- Define heat of fusion.
- Define heat of vaporization.
- Calculate heat transfers necessary for changes of state.

Before the internal combustion engine was invented, steam engines were the power source for ships, locomotives, tractors, lumber saws, and most industrial machines. Coal or wood was burned to boil water into steam, which ran the engine.

Change of State

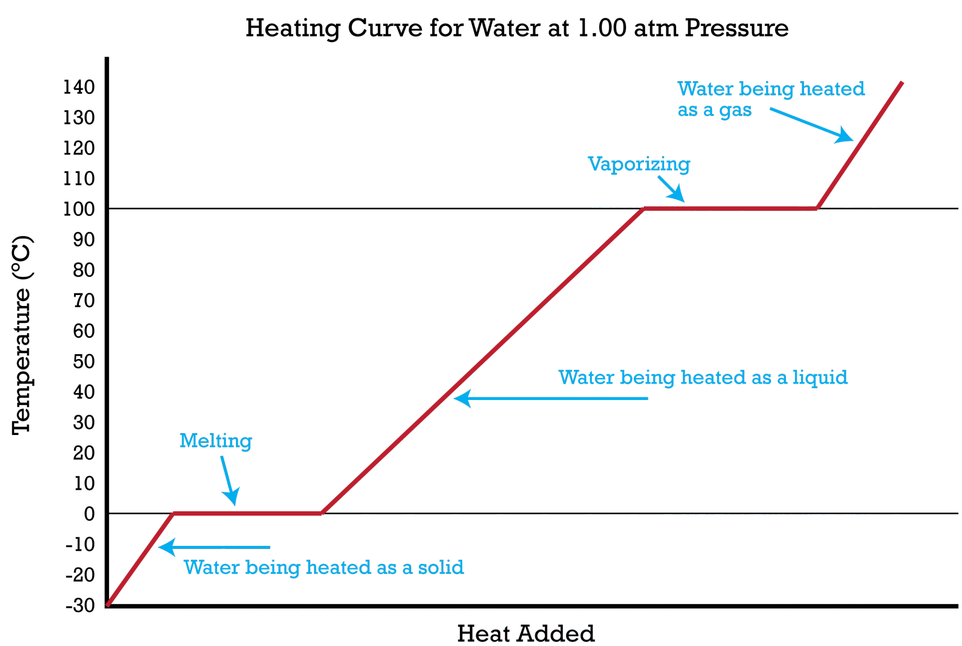
Most substances may exist in any of the three common states of matter. In the gaseous state, the molecular motion has completely overcome any attraction between the particles and the particles are totally separate from each other. There are large spaces between the particles and they move large distances between collisions. In the liquid state, the molecular motion and the molecular attractions are more balanced. While the particles stay more or less in contact with each other, they are still free to move and can slide past one another easily. In the solid state, the attractive forces dominate. The particles are pulled together into a tightly packed pattern which does not allow the particles to pass each other. The molecular motion in this form is essentially reduced to vibration in place. Increasing the temperature of a substance means increasing the molecular motion (kinetic energy) of the molecules in the substance. The phase in which a substance exists is the result of a competition between attractive forces and molecular motion.



For most substances, when the temperature of the solid is raised high enough, the substance changes to a liquid, and when the temperature of the liquid is raised high enough, the substance changes to a gas. We typically visualize a solid as tiny particles in constant motion held together by attractive forces. As we add heat to the solid, the motion, or the kinetic energy, of the particles increases. At some temperature, the motion of the particles becomes great enough to overcome the attractive forces. The thermal energy that was added to the solid up to this point was absorbed by the solid as kinetic energy, increasing the speed of the molecules. The lowest temperature at which the particles are able to exist in the liquid form is called the **melting point**.

In order for the molecules to actually separate from each other, more energy must be added. This energy, called **heat of fusion** or heat of melting, is absorbed by the particles as potential energy as the solid changes to a liquid. Recognize that, once the temperature of a solid has been raised to the melting point, it is still necessary for the solid to absorb additional thermal energy in the form of potential energy as the molecules separate.

The **boiling point** of a liquid is the temperature at which the particles have sufficient molecular motion to exist in the form of a gas. Once again, however, in order for the particles to separate to the gaseous form, they must absorb a sufficient amount of potential energy. The amount of potential energy necessary for a **phase change** to gaseous form is called the **heat of vaporization**. Consider the heating curve shown below.



The heating curve shown is for water but other substances have similarly shaped heating curves. Suppose you begin with solid water (ice) at -30°C and add heat at a constant rate. The heat you add in the beginning will be absorbed as kinetic energy and the temperature of the solid will increase. When you reach a temperature of 0°C (the melting point for water), the heat you add is no longer absorbed as kinetic energy. Instead, the added heat is absorbed as potential energy and the particles separate from each other. During the flat part of the curve labeled "melting", heat is being added constantly but the temperature does not increase. At the left edge of this flat line, the water is solid; by the time enough heat has been added to get to the right edge, the water is liquid, but maintains the same temperature. Once all the water is in the liquid form, the added heat will once again be absorbed as kinetic energy and the temperature will increase again. During the time labeled "water being heated as a liquid", all the added heat is absorbed as kinetic energy.

When a temperature of 100°C (the boiling point of water) is reached, the added heat is once again absorbed as potential energy and the molecules separate from liquid form into gaseous form. When all the substance has been converted into gas, the temperature will again begin to rise.

TABLE 1.2: Heats of Fusion and Vaporization of Some Common Substances

Substance	Heat of Fusion, H_f (J/kg)	Heat of Vaporization, H_v (J/kg)
Copper	2.05×10^5	5.07×10^6
Gold	6.30×10^4	1.64×10^6
Iron	2.66×10^5	6.29×10^6
Methanol	1.09×10^5	8.78×10^5
Water	3.34×10^5	2.26×10^6

When the temperature of a substance is changing, we can use the specific heat to determine the amount of heat that is being gained or lost. When a substance is changing phase, we can use the heat of fusion or heat of vaporization to determine the amount of heat being gained or lost. When a substance freezes from liquid to solid, the amount of heat given off is exactly the same as the amount of heat absorbed when the substance melts from solid to liquid. The equations for heat gained or lost are given here:

The heat gained or lost during a temperature change: $Q = mc\Delta t$.

The heat gained or lost during a phase change of solid to liquid: $Q = mH_f$.

The heat gained or lost during a phase change of liquid to gas: $Q = mH_v$.

Example Problem: 5000. Joules of heat is added to ice at 273 K. All the heat goes into changing solid ice into liquid water. How much ice is melted?

Solution: $m = \frac{Q}{H_f} = \frac{5000\ J}{3.34 \times 10^5\ J/kg} = 0.0150\ kg$

Example Problem: Beginning with 1.00 kg of ice at -20.0°C, heat is added until the substance becomes water vapor at 130.0°C. How much heat was added? The specific heat of ice is 2108 $J/kg°C$, the specific heat of liquid water is 4187 $J/kg°C$, and the specific heat of water vapor is 1996 $J/kg°C$.

Solution: 5 steps.

1. Calculate the heat required to raise the sample from -20.0°C to 0°C.
2. Calculate the heat required to melt the sample.
3. Calculate the heat required to raise the sample from 0°C to 100°C.
4. Calculate the heat required to vaporize the sample.
5. Calculate the heat required to raise the sample from 100°C to 130°C.

The solution is the sum of these steps.

1.$Q_{HS} = mc_{\text{ice}}\Delta t = (1.00\ kg)(2108\ J/kg.°C)(20.0°C) = 42160\ J$

2.$Q_{\text{Melt}} = mH_f = (1.00\ kg)(334000\ J/kg) = 334000\ J$

3.$Q_{HL} = mc_{\text{water}}\Delta t = (1.00\ kg)(4187\ J/kg.°C)(100.0°C) = 418700\ J$

4.$Q_{\text{Vap}} = mH_v = (1.00\ kg)(2260000\ J/kg) = 2260000\ J$

5.$Q_{\text{HV}} = mc_{\text{vapor}}\Delta t = (1.00\ kg)(1996\ J/kg.^\circ C)(30.0^\circ C) = 59880\ J$

Total Heat $= 3.11 \times 10^6\ J$

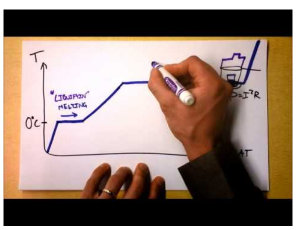
Summary

- Most substances may exist in any of the three common states of matter, solid, liquid, or gas.
- The phase in which a substance exists is the result of a competition between attractive forces and molecular motion.
- The potential energy absorbed by a solid as it changes to a liquid is called the heat of fusion or the heat of melting.
- The amount of potential energy necessary for a phase change to gaseous form is called the heat of vaporization.
- The heat gained or lost during a temperature change is given by, $Q = mc\Delta t$.
- The heat gained or lost during a phase change of solid to liquid is given by, $Q = mH_f$.
- The heat gained or lost during a phase change of liquid to gas is given by, $Q = mH_v$.

Practice

The following video explains heat of fusion and vaporization. Use this video to answer the questions that follow.

http://www.youtube.com/watch?v=6lAxBTLgYfU



MEDIA

Click image to the left for more content.

1. For water, which takes more energy, melting or evaporating?
2. When are there two phases present at the same time in the pot?

Practice problems involving phase changes:

http://www.slideshare.net/angel4all1/specific-heat-and-phase-change-ditto

Review

1. A 200. g sample of water at 60.0°C is heated to water vapor at 140.0°C. How much heat was absorbed?
2. A 175 g lump of molten lead at its melting point (327°C) is placed into 55.0 g of water at 20.0°C. The specific heat of lead is 130 $J/kg\cdot^\circ C$ and the $H_f$of lead is 20,400 J/kg.
   (a) When the lead has become solid but is still at the melting point, what is the temperature of the water?
   (b) When the lead and the water have reached equilibrium, what is the temperature of the mixture?

- **phase change:** When a substance changes from one state, or phase, of matter to another we say that it has undergone a change of state, or we say that it has undergone a change of phase.
- **heat of fusion:** The change in potential energy for the conversion of 1 mole or 1 gram of a solid to a liquid, at constant pressure and temperature and is usually denoted as ΔH_f.

- **heat of vaporization:** The change in potential energy for the conversion of 1 mole or 1 gram of a liquid to a gas, at constant pressure and temperature and is usually denoted as ΔH_v.

The calculation of heat change is not as simple as measuring the temperature of an object. Different objects have different specific heats, meaning they change temperature at different rates. Temperature and heat, though closely related, are not the same. While temperature is a measurement of the average kinetic energy of the molecules, the heat is the total kinetic energy. Calculations of changes in heat over phase changes depend not only on the specific heat, but also on the heat of fusion and the heat of vaporization.

1.5 References

1. Courtesy of Roger J. Braithwaite, The University of Manchester, UK/NASA. http://www.giss.nasa.gov/research/briefs/gornitz_09/ . Public Domain
2. Courtesy of the National Parks Service. http://commons.wikimedia.org/wiki/File:Volcano_q.jpg . Public Domain
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