

SCIENCE 1203 WORK AND ENERGY

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SPECIFIC HEAT

Temperature and heat are distinct from each other. Temperature is an indicator of an object's average molecular energy, without regard to mass. Heat is a quantitative measure of total molecular energy of a known mass.

In most instances, addition of heat to an object is accompanied by a rise in temperature. During changes of phase, however, heat is added to or removed from a substance unaccompanied by a temperature change.

Temperature. If you place one hand in cold water and then in temperate water, the temperate water will seem warm. If you transfer your hand from hot water to temperate water, the temperate water will seem cool. Such a subjective measure of temperature is not acceptable for scientific work; therefore, a thermometer is needed.

What is temperature? Temperature is a measure of the average motion of molecules measured in degrees Fahrenheit ($^{\circ}\text{F}$), Celsius ($^{\circ}\text{C}$), or Kelvin ($^{\circ}\text{K}$).

Water covers over two-thirds the face of the earth; therefore, it is a convenient substance by which to define temperatures. Pure water freezes at 0°C (32°F) and boils at 100°C (212°F). One hundred eighty Fahrenheit degrees and 100 Celsius degrees fall between freezing and boiling points of water (pure water at sea level).

Since temperature is the measure of the motion of molecules, the temperature at which all motion ceases is -273°C , absolute zero. The Kelvin degree is equivalent in size to the Celsius degree, but the Kelvin scale starts at absolute zero. The freezing point of water is 273°K and the boiling point is 373°K . To change from Celsius to Kelvin simply add 273° .

The amount of energy that will change 1 gram of water 1 Celsius degree is defined as 1 **calorie** (cal). The kilocalorie (kcal or Cal) is the amount of heat needed to change 1 kilogram 1 Celsius degree. In the English system the unit of heat is a *Btu* (British thermal unit), defined as the amount of heat required to change the temperature of 1 pound of water 1 Fahrenheit degree. For water, heat energy can be calculated by

$$\text{heat} = m \Delta t,$$

where m is the mass and Δt is its change in temperature on the Celsius scale.

When 1 calorie of heat is added to 1 gram of water (below its boiling point), its temperature changes by 1 Celsius degree. However, 1 calorie of heat produces a change of 5 Celsius degrees in 1 gram of aluminum, 10 Celsius degrees in 1 gram of copper, and 20 Celsius degrees in 1 gram of silver. Said another way, a 1°C change in aluminum, copper, or silver requires less heat than does a 1°C change in an equal mass of water. Specifically, $\frac{1}{5}$ calorie added to 1 gram of aluminum, $\frac{1}{10}$ calorie added to 1 gram of copper, or $\frac{1}{20}$ calorie added to 1 gram of silver will all produce a 1°C change. The amount of heat added (or removed) from one gram of matter to change its temperature by 1°C is the **specific heat** of the substance. Specific heat is given as kilocalories per kilogram-degree Celsius, $\frac{\text{kcal}}{\text{kg} \cdot ^{\circ}\text{C}}$, which is equivalent to $\frac{\text{cal}}{\text{g} \cdot ^{\circ}\text{C}}$.

$$\frac{\text{kcal}}{\text{kg} \cdot ^{\circ}\text{C}} = \frac{1000 \text{ cal}}{1000 \text{ g} \cdot ^{\circ}\text{C}} = \frac{\text{cal}}{\text{g} \cdot ^{\circ}\text{C}}$$

Substance	Specific Heat, $\frac{\text{kcal}}{\text{kg} \cdot ^{\circ}\text{C}}$	Substance	Specific Heat, $\frac{\text{kcal}}{\text{kg} \cdot ^{\circ}\text{C}}$
alcohol (ethyl)	0.58	marble	0.21
aluminum	0.22	mercury	0.03
copper	0.09	silver	0.05
glass	0.20	steam	0.48
gold	0.03	sulfuric acid	0.27
granite	0.19	turpentine	0.42
ice	0.50	water	1.00
iron	0.11	wood	0.42
lead	0.03	zinc	0.09

Specific Heats

More accurate values for the specific heats of aluminum, copper, and silver are shown in the table. By definition, water has a specific heat of $1 \frac{\text{kcal}}{\text{kg} \cdot ^{\circ}\text{C}}$. Therefore the heat energy for any material can be measured by the formula

$$\text{heat} = mC \Delta t$$

where C is the specific heat of the substance.

Phases. Some substances, such as water, metals, and atmospheric gases exist as solid, liquid, and gas. Their phase, or state, is temperature dependent. In the **solid phase** molecules are restricted in motion to vibrations about an equilibrium position (**vibrational motion**). The substance has a definite shape and volume and is at its lowest internal energy state. As heat is added, the temperature increases until molecular bonds begin to break. The molecules now have vibrational motion, **rotational motion**, and **translational motion**. In this **liquid phase** molecules are free to move around, but are restricted to the volume of the liquid.

In the **gaseous phase** molecules are in more rapid motion and are no longer confined to a particular volume, but occupy all space available. It is the highest state of matter.

Usually the study of matter stops here. However, most of the matter in the universe is not solid nor liquid nor gaseous; but it is in a **plasma phase**. If excessive heat energy is added to gas, one or more electrons attain enough energy to break away from the atom. The atom is then positively charged and is said to be *ionized*. Most of the matter composing the sun is plasma. Both a fluorescent light and the gas at the tip of a welding torch are composed of plasma.



Complete these activities.

3.1 Calculate the equivalent temperatures.

- a. 60°C = _____ $^{\circ}\text{K}$
- b. 300°K = _____ $^{\circ}\text{C}$
- c. 373°K = _____ $^{\circ}\text{F}$
- d. 200°C = _____ $^{\circ}\text{K}$

3.2 Calculate the energy needed to heat 4 kg of water from 25°C to 45°C .

3.3 Why does a large container of boiling water contain more heat than a kilogram of lead at 300°C ? _____



Match these items.

- | | | |
|-----|-------------------|-----------|
| 3.4 | _____ ice | a. solid |
| 3.5 | _____ steam | b. gas |
| 3.6 | _____ air | c. plasma |
| 3.7 | _____ stars | d. liquid |
| 3.8 | _____ molten iron | |



Answer this question.

3.9 What must be done to get copper from the solid phase to the liquid phase?



Complete these calculations.

3.10 Calculate the heat energy needed to change the temperature of 2 kg of copper from 10°C to 110°C.

3.11 Calculate the change in temperature of 20 kg of water if 30 kg of aluminum is dropped in the water and the aluminum changes temperature by 20°C (hint, use the principle of conservation of energy).

3.12 If 10 calories of energy are added to 2 grams of ice at -30°C, calculate the final temperature of the ice. (Notice that ice has a specific heat different from that of water.)

LATENT HEAT

Between a substance's phase boundaries, a given heat increase (or decrease) produces a given increase (or decrease) in temperature. Across phase boundaries, however, no such temperature change occurs. All heat that is added during a phase change goes toward altering the bonds that hold molecule to molecule. An investigation of water yields the following results.

Latent heat of fusion. If 1 gram of water is at 10°C, removing 1 calorie changes the

temperature by 1 Celsius degree to 9°C. Removal of 9 more calories drops the temperature to 0°C. At this point, however, removal of any more heat does not change the temperature at all; but now ice begins to form. In fact, 80 calories of heat must be removed for each gram of water (80 kilocalories from 1 kilogram) in order for the water to freeze at 0°C! The $\frac{80}{\text{gm}} \text{ cal}$ (or $\frac{80}{\text{kg}} \text{ kcal}$) is called the **latent** (or hidden) **heat of fusion** (L_f). The term *latent* is used because no change in temperature results, and the heat seems to be "hidden"!

Solid Substance	Melting Point ($^{\circ}\text{C}$)	Latent Heat of Fusion $\frac{\text{kcal}}{\text{kg}}$
ice	0	80
oxygen	-219	3.3
aluminum	660	93
copper	1,083	51
iron	1,539	65
tungsten	3,400	44
lead	330	6
mercury	-39	2.8
zinc	420	24
alcohol (ethyl)	-114	25

Latent Heats of Fusion

If a substance is in the solid phase (such as ice) the amount of heat needed to change it to a

liquid is found through the equation:

heat = $(m)(C)(\Delta t)$ the heat needed to raise it to the melting point
 + $(m)(L_f)$ the heat needed to melt it
 + $(m)(C)(\Delta t)$ the heat needed to raise it to its final temperature in the liquid phase

Example. How much heat is needed to raise 2 g of ice ($C_{\text{solid}} = 0.5$) at -15°C to water ($C_{\text{liquid}} = 1$) at 35°C ?

$$\begin{aligned} \text{heat} &= (2 \text{ g})(0.5)(15^{\circ}\text{C}) + (2 \text{ g})(80 \frac{\text{cal}}{\text{g}}) + \\ &\quad (2 \text{ g})(1)(35^{\circ}\text{C}) \\ &= 15 \text{ cal} + 160 \text{ cal} + 70 \text{ cal} \\ \text{heat} &= 245 \text{ cal} \end{aligned}$$

The first term accounts for the change in temperature to raise the ice from -15°C to the melting point (0°C). The second term accounts for the phase change. The third term accounts for the change in temperature of the melted ice (0°C) to its final temperature of 35°C .



Complete these activities.

- 3.13 How much energy is needed to heat and melt 3 kg of copper initially at 83°C ?
- 3.14 Calculate the total heat content of 10 kg of ice at -23°C . (Hint: absolute zero is -273°C .)
- 3.15 How much heat is contained in 100 kg of water at 60°C ?