CHAPTER 10

Heat

PHYSICS IN ACTION

Whether you pop corn by putting the kernels in a pan of hot oil or in a microwave oven, the hard kernels will absorb energy until, at a high temperature, they rupture. At this point, superheated water suddenly turns to steam and rushes outward, and the kernels burst open to form the fluffy, edible puffs of starch. But what actually happens when water turns into steam, and what do we mean when we talk about heat and temperature?

In this chapter you will study what distinguishes temperature and heat and how different substances behave when energy is added to or removed from them, causing a change in their temperature or phase.

- *Why do the kernels require steam, not just superheated water, to produce popcorn?*
- *What role does oil play in the preparation of popcorn?*

CONCEPT REVIEW

Work (Section 5-1)

Energy (Section 5-2)

Conservation of energy (Section 5-3)

10-1
Temperature and thermal equilibrium

10-1 SECTION OBJECTIVES

- **Relate temperature to the kinetic energy of atoms and molecules.**
- **Describe the changes in the temperatures of two objects reaching thermal equilibrium.**
- **Identify the various temperature scales, and be able to convert from one scale to another.**

DEFINING TEMPERATURE

When you hold a glass of lemonade with ice, like that shown in **Figure 10-1,** you feel a sharp sensation in your hand that we describe as "cold." Likewise, you experience a "hot" feeling when you touch a cup of hot chocolate. We often associate temperature with how hot or cold an object feels when we touch it. Our sense of touch serves as a qualitative indicator of temperature. However, this sensation of hot or cold also depends on the temperature of the skin and therefore is misleading. The same object may feel warm or cool, depending on the properties of the object and on the conditions of your body.

Determining an object's temperature with precision requires a standard definition of temperature and a procedure for making measurements that establish how "hot" or "cold" objects are.

Figure 10-1
Objects at low temperatures feel cold to the touch, while objects at high temperatures feel hot. However, the sensation of hot and cold can be misleading.

Adding or removing energy usually changes temperature

Consider what happens when you use an electric range to cook food. By turning the dial that controls the electric current delivered to the heating element, you can adjust the element's temperature. As the current is increased, the temperature of the element increases. Similarly, as the current is reduced, the temperature of the element decreases. In general, energy must be either added to or removed from a substance to change its temperature.

Quick Lab

Sensing Temperature

MATERIALS LIST

- ✔ 3 identical basins
- ✔ hot and cold tap water
- ✔ ice

SAFETY CAUTION

Use only hot tap water. The temperature of the hot water must not exceed 50°C (122°F).

Fill one basin with hot tap water. Fill another with cold tap water, and add ice until about one-third of the mixture is ice. Fill the third basin with an equal mixture of hot and cold tap water.

Place your left hand in the hot water and your right hand in the cold water for 15 s. Then place both hands in the basin of lukewarm water for 15 s. Describe whether the water feels hot or cold to either of your hands.

Temperature is proportional to the kinetic energy of atoms and molecules

In Section 9-2 you learned that temperature is proportional to the average kinetic energy of particles in a substance. A substance's temperature increases as a direct result of added energy being distributed among the particles of the substance, as shown in **Figure 10-2.**

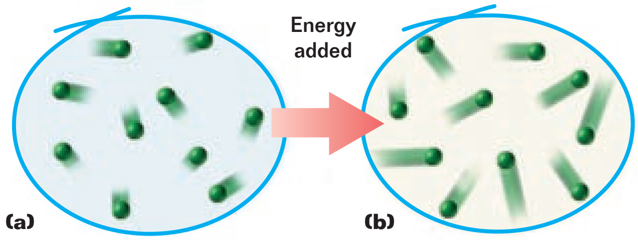


Figure 10-2
The low average kinetic energy of the particles **(a),** and thus the temperature of the gas, increases when energy is added to the gas **(b).**

For a monatomic gas, temperature can be understood in terms of the translational kinetic energy of the atoms in the gas. For other kinds of substances, molecules can rotate or vibrate, so rotational kinetic energy or vibrational kinetic and potential energies also exist (see **Table 10-1**).

The energies associated with atomic motion are referred to as **internal energy,** which is proportional to the substance's temperature. For an ideal gas, the internal energy depends only on the temperature of the gas. For gases with two or more atoms per molecule, as well as for liquids and solids, other properties besides temperature contribute to the internal energy. The symbol U stands for internal energy, and ΔU stands for a change in internal energy.

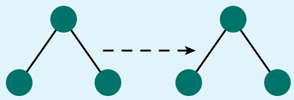
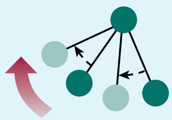
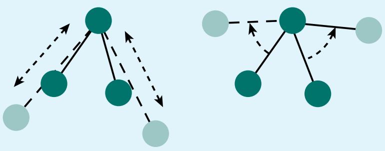
internal energy

the energy of a substance due to the random motions of its component particles and equal to the total energy of those particles

Temperature is meaningful only when it is stable

Imagine a can of warm fruit juice immersed in a large beaker of cold water. After about 15 minutes, the can of fruit juice will be cooler and the water surrounding it will be slightly warmer. Eventually, both the can of fruit juice and

Table 10-1 Examples of different forms of energy

Form of energy	Macroscopic examples	Microscopic examples	Energy type
Translational	airplane in flight, roller coaster at bottom of rise	CO_2 molecule in linear motion	kinetic energy 
Rotational	spinning top	CO_2 molecule spinning about its center of mass	kinetic energy 
Vibrational	plucked guitar string	bending and stretching of bonds between atoms in a CO_2 molecule	kinetic and potential energy 

thermal equilibrium

the state in which two bodies in physical contact with each other have identical temperatures

the water will be at the same temperature. That temperature will not change as long as conditions remain unchanged in the beaker. Another way of expressing this is to say that the water and can of juice are in **thermal equilibrium** with each other.

Thermal equilibrium is the basis for measuring temperature with thermometers. By placing a thermometer in contact with an object and waiting until the column of liquid in the thermometer stops rising or falling, you can find the temperature of the object. This is because the thermometer is at the same temperature as, or is in thermal equilibrium with, the object. Just as in the case of the can of fruit juice in the cold water, the temperature of any two objects at thermal equilibrium always lies between their initial temperatures.

Matter expands as its temperature increases

You have learned that increasing the temperature of a gas may cause the volume of the gas to increase. This occurs not only for gases, but also for liquids and solids. In general, if the temperature of a substance increases, so does its volume. This phenomenon is known as *thermal expansion.*

You may have noticed that the concrete roadway segments of a bridge are separated by gaps several centimeters wide. This is necessary because concrete expands with increasing temperature. Without these gaps, the force from the thermal expansion would cause the segments to push against each other, and they would eventually buckle and break apart.

Different substances undergo different amounts of expansion for a given temperature change. The thermal expansion characteristics of a material are indicated by a quantity called the *coefficient of volume expansion.* Gases have the largest values for this coefficient. Liquids have much smaller values.

In general, the volume of a liquid tends to increase with increasing temperature. However, the volume of water increases with decreasing temperature in the range between 0°C and 4°C. This explains why ice floats in liquid water. It also explains why a pond freezes from the top down instead of from the bottom up. If this did not happen, fish would likely not survive in freezing temperatures.

Solids typically have the smallest coefficient of volume expansion values. For this reason, liquids in solid containers expand more than the container. This property allows some liquids to be used to measure changes in temperature.

MEASURING TEMPERATURE

In order for a device to be used as a thermometer, it must make use of a change in some physical property that corresponds to changing temperature, such as the volume of a gas or liquid, or the pressure of a gas at constant volume. The most common thermometers use a glass tube containing a thin column of mer-

Conceptual Challenge

1. Hot chocolate

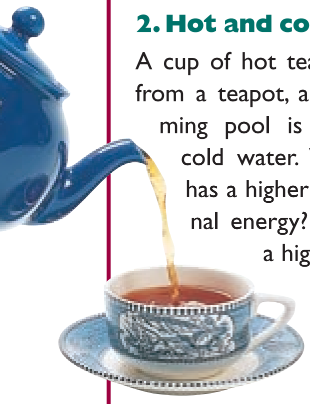
If two cups of hot chocolate, one at 50°C and the other at 60°C, are poured together in a large container, will the final temperature of the double batch be

a. less than 50°C?
b. between 50°C and 60°C?
c. greater than 60°C?

Explain your answer.

2. Hot and cold liquids

A cup of hot tea is poured from a teapot, and a swimming pool is filled with cold water. Which one has a higher total internal energy? Which has a higher average kinetic energy? Explain.



cury, colored alcohol, or colored mineral spirits. When the thermometer is heated, the volume of the liquid expands. Because the cross-sectional area of the tube remains nearly constant during temperature changes, the change in length of the liquid column is proportional to the temperature change (see **Figure 10-3**).

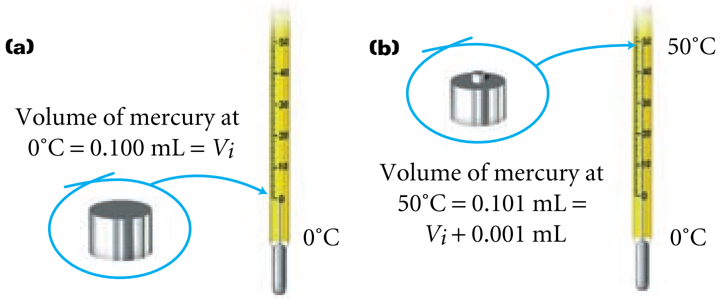


Figure 10-3
The change in the mercury's volume from a temperature of 0°C **(a)** to a temperature of 50°C **(b)** is small, but because the mercury is limited to expansion in only one direction, the linear change is large.

Calibrating thermometers requires fixed temperatures

A thermometer must be more than an unmarked, thin glass tube of liquid. For a thermometer to measure temperature in a variety of situations, the length of the liquid column at different temperatures must be known. One reference point is etched on the tube and refers to when the thermometer is in thermal equilibrium with a mixture of water and ice at one atmosphere of pressure. This temperature is called the *ice point* of water and is defined as zero degrees Celsius, or 0°C. A second reference mark is made at the point when the thermometer is in thermal equilibrium with a mixture of steam and water at one atmosphere of pressure. This temperature is called the *steam point* of water and is defined as 100°C.

A temperature scale can be made by dividing the distance between the reference marks into equally spaced units, called *degrees*. The scale assumes the expansion of the mercury is linear.

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Temperature units depend on the scale used

The temperature scales most widely used today are the Fahrenheit, Celsius, and Kelvin (or absolute) scales. The Fahrenheit scale is commonly used in the United States. The Celsius scale is used in countries that have adopted the metric system and by the scientific community worldwide.

Celsius and Fahrenheit temperature measurements can be converted to each other using this equation.

CELSIUS-FAHRENHEIT TEMPERATURE CONVERSION

$$T_F = \frac{9}{5}T_C + 32.0$$

$$\textbf{Fahrenheit temperature} = \left(\frac{9}{5} \times \textbf{Celsius temperature}\right) + \textbf{32.0}$$

Did you know?

When a thermometer reaches thermal equilibrium with an object, the object's temperature changes slightly. In most cases the object is so massive compared with the thermometer that the object's temperature change is insignificant.

The number 32.0 in the equation indicates the difference between the ice point value in each scale. The point at which water freezes is 0.0 degrees in the Celsius scale and 32.0 degrees in the Fahrenheit scale.

Temperature values in the Celsius and Fahrenheit scales can have positive, negative, or zero values. But because the kinetic energy of the atoms in a substance is positive, the absolute temperature that is proportional to that energy should be positive also. A temperature scale with only positive values is suggested

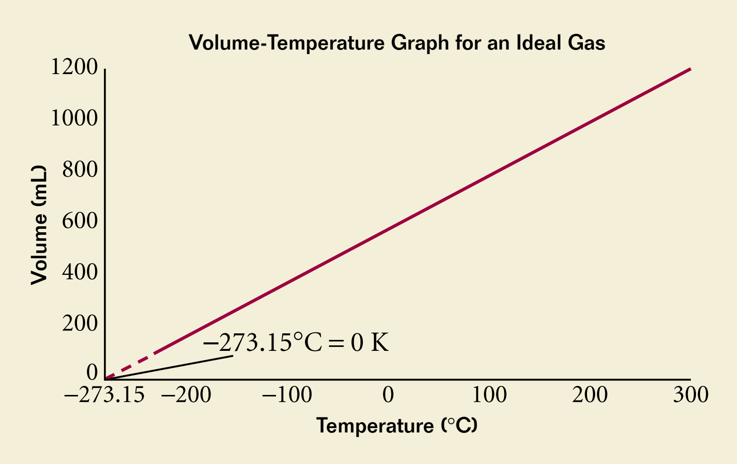

Figure 10-4
If an ideal gas could be compressed to zero volume, its temperature would be −273.15°C, or 0 K.

in the graph of volume versus temperature for an ideal gas, shown in **Figure 10-4.** As the temperature of the gas decreases, so does its volume. If it were possible to compress the matter in a gas to zero volume, the gas temperature would equal −273.15°C. This temperature is designated in the Kelvin scale as 0.00 K, where K is the symbol for the temperature unit called the kelvin. Temperatures in this scale are indicated by the symbol *T*.

A temperature difference of one degree is the same on the Celsius and Kelvin scales. The two scales differ only in the choice of zero point. Thus, the ice point (0.00°C) equals 273.15 K, and the steam point (100.00°C) equals 373.15 K (see **Table 10-2**). The Celsius temperature can therefore be converted to the Kelvin temperature by adding 273.15.

CELSIUS-KELVIN TEMPERATURE CONVERSION

$$T = T_C + 273.15$$

Kelvin temperature = Celsius temperature + 273.15

Kelvin temperatures for various physical processes can range from around 1 000 000 000 K (10^9 K), which is the temperature of the interiors of the most massive stars, to less than 1 K, which is slightly cooler than the boiling point of liquid helium. The temperature 0 K is often referred to as *absolute zero.* Absolute zero has never been reached, although laboratory experiments have reached temperatures of 0.000 001 K.

Table 10-2 Temperature scales and their uses

Scale	Ice point	Steam point	Applications
Fahrenheit	32°F	212°F	meteorology, medicine, and non-scientific uses (U.S.)
Celsius	0°C	100°C	meteorology, medicine, and non-scientific uses (outside U.S.); other sciences (international)
Kelvin (absolute)	273.15 K	373.15 K	physical chemistry, gas laws, astrophysics, thermodynamics, low-temperature physics

SAMPLE PROBLEM 10A

Temperature conversion

PROBLEM

What are the equivalent Celsius and Kelvin temperatures of 50.0°F?

SOLUTION

Given: $T_F = 50.0°F$

Unknown: $T_C = ?$ $T = ?$

Use the Celsius-Fahrenheit equation from page 361.

$$T_F = \frac{9}{5}T_C + 32.0$$

$$T_C = \frac{5}{9}(T_F - 32.0)$$

$$T_C = \frac{5}{9}(50.0 - 32.0)°C = 10.0°C$$

Use the Celsius-Kelvin equation from page 362.

$$T = T_C + 273.15$$

$$T = (10.0 + 273.2)K = 283.2\ K$$

$$\boxed{\begin{array}{l} T_C = 10.0°C \\ T = 283.2\ K \end{array}}$$

PRACTICE 10A

Temperature conversion

1. The lowest outdoor temperature ever recorded on Earth is −128.6°F, recorded at Vostok Station, Antarctica, in 1983. What is this temperature on the Celsius and Kelvin scales?

2. The temperatures of one northeastern state range from 105°F in the summer to −25°F in winter. Express this temperature range in degrees Celsius and in kelvins.

3. The normal human body temperature is 98.6°F. A person with a fever may record 102°F. Express these temperatures in degrees Celsius.

4. A pan of water is heated from 23°C to 78°C. What is the change in its temperature on the Kelvin and Fahrenheit scales?

5. Liquid nitrogen is used to cool substances to very low temperatures. Express the boiling point of liquid nitrogen (77.34 K at 1 atm of pressure) in degrees Fahrenheit and in degrees Celsius.

Section Review

1. Two gases that are in physical contact with each other consist of particles of identical mass. In what order should the images shown in **Figure 10-5** be placed to correctly describe the changing distribution of kinetic energy among the gas particles? Which group of particles has the highest temperature at any time? Explain.

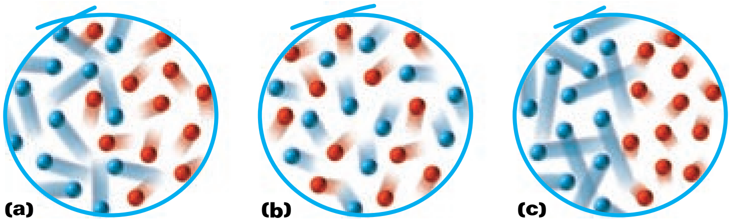

Figure 10-5

2. A hot copper pan is dropped into a tub of water. If the water's temperature rises, what happens to the temperature of the pan? How will you know when the water and copper pan reach thermal equilibrium?

3. Oxygen condenses into a liquid at approximately 90.2 K. To what temperature does this correspond on both the Celsius and Fahrenheit temperature scales?

4. The boiling point of sulfur is 444.6°C. Sulfur's melting point is 586.1°F lower than its boiling point.

a. Determine the melting point of sulfur in degrees Celsius.
b. Find the melting and boiling points in degrees Fahrenheit.
c. Find the melting and boiling points in kelvins.

5. **Physics in Action** Which of the following is true for the water molecules inside popcorn kernels during popping?

a. Their temperature increases.
b. They are destroyed.
c. Their kinetic energy increases.
d. Their mass changes.

6. **Physics in Action** Referring to **Figure 10-6,** determine which of the following pairs represent objects that are in thermal equilibrium with each other.

a. the hot plate and the glass pot
b. the hot oil and the popcorn kernels
c. the outside air and the hot plate

Figure 10-6

10-2
Defining heat

HEAT AND ENERGY

Thermal physics often appears mysterious at the macroscopic level. Hot objects become cool without any obvious cause. To understand thermal processes, it is helpful to shift attention to the behavior of atoms and molecules. Mechanics can be used to explain much of what is happening at the molecular, or microscopic, level. This in turn accounts for what you observe at the macroscopic level. Throughout this chapter, the focus will shift between these two viewpoints.

Recall the can of warm fruit juice immersed in the beaker of cold water (shown in **Figure 10-7**). The temperature of the can and the juice in it is lowered, and the water's temperature is slightly increased, until at thermal equilibrium both final temperatures are the same. Energy is transferred from the can of juice to the water because the two objects are at different temperatures. This energy that is transferred is defined as **heat.**

The word *heat* is sometimes used to refer to the *process* by which energy is transferred between objects because of a difference in their temperatures. This textbook will use *heat* to refer only to the energy itself.

10-2 SECTION OBJECTIVES

- **Explain heat as the energy transferred between substances that are at different temperatures.**
- **Relate heat and temperature change on the macroscopic level to particle motion on the microscopic level.**
- **Apply the principle of energy conservation to calculate changes in potential, kinetic, and internal energy.**

heat

the energy transferred between objects because of a difference in their temperatures

Energy is transferred between substances as heat

From a macroscopic viewpoint, energy transferred as heat always moves from an object at higher temperature to an object at lower temperature. This is similar to the mechanical behavior of objects moving from a higher gravitational potential energy to a lower gravitational potential energy. Just as a pencil will drop from your desk to the floor but will not jump from the floor to your

Figure 10-7
Energy is transferred as heat from objects with higher temperatures (the fruit juice and can) to those with lower temperatures (the cold water).

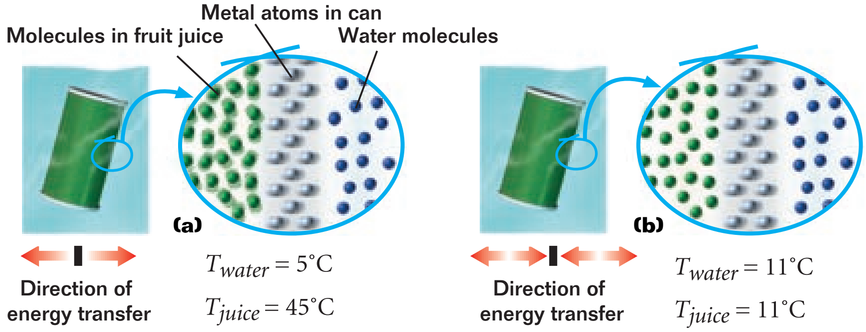

Figure 10-8
Energy is transferred as heat from the higher-energy particles to lower-energy particles **(a).** The net energy transferred is zero when thermal equilibrium is reached **(b).**

desk, so energy will travel spontaneously from an object at higher temperature to one at lower temperature and not the other way around.

The direction in which energy travels as heat can be explained at the atomic level. At first, the molecules in the fruit juice have higher average kinetic energies than do the water molecules that surround the can, as shown in **Figure 10-8.** This energy is transferred from the juice to the can by the molecules in the juice colliding with the metal atoms of the can. The atoms vibrate more because of their increased energy, and this energy is transferred to the surrounding water molecules.

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As the energy of the water molecules gradually increases, the energy of the molecules in the fruit juice and the atoms of the can decreases until all of the particles have, on the average, equal kinetic energies. However, it is also possible for some of the energy to be transferred through collisions from the lower-energy water molecules to the higher-energy metal atoms and fruit-juice particles. Therefore, energy can move in both directions. Because the average kinetic energy of particles is higher in the body at higher temperature, more energy is transferred out of it as heat than is transferred into it. The net result is that energy is transferred as heat in only one direction.

The transfer of energy as heat alters an object's temperature

Thermal equilibrium may be understood in terms of energy exchange between two objects at equal temperature. When the can of fruit juice and the surrounding water are at the same temperature, as depicted in **Figure 10-9,** the quantity of energy transferred from the can of fruit juice to the water is the same as the energy transferred from the water to the can of juice. The net energy transferred between the two objects is zero.

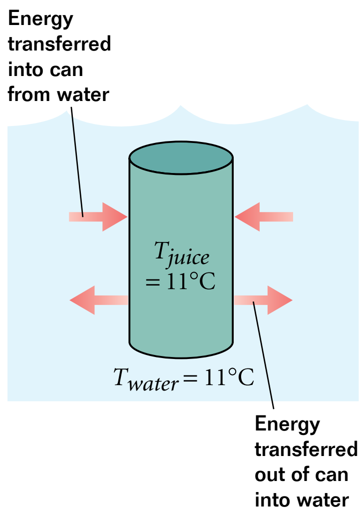


Figure 10-9
At thermal equilibrium, the net energy exchanged between two objects equals zero.

This reveals the difference between temperature and heat. The atoms of all objects are in continuous motion, so all objects have some internal energy. Because temperature is a measure of that energy, all objects have some temperature. Heat, on the other hand, is the energy transferred from one object to another because of the temperature difference between them. When there is no temperature difference between a substance and its surroundings, no net energy is transferred as heat.

Energy transfer depends on the difference of the temperatures of the two objects. The greater the temperature difference between two objects, the greater the amount of energy that is transferred between them as heat.

For example, in winter, energy is transferred as heat from a car's surface at 30°C to a cold raindrop at 5°C. In the summer, energy is transferred as heat from a car's surface at 45°C to a warm raindrop at 20°C. In each case, the amount of energy transferred is the same, because the substances and the temperature difference (25°C) are the same (see **Figure 10-10**).

The concepts of heat and temperature help to explain why hands held in separate bowls containing hot and cold water subsequently sense the temperature of lukewarm water differently. The nerves in the outer skin of your hand detect energy passing through the skin from objects with temperatures different than your body temperature. If one hand is at thermal equilibrium with cold water, more energy is transferred from the outer layers of your hand than can be replaced by the blood, which has a temperature of about 37.0°C (98.6°F). When the hand is immediately placed in water that is at a higher temperature, energy is transferred from the water to the cooler hand. The energy transferred into the skin causes the water to feel warm. Likewise, the hand that has been in hot water temporarily gains energy from the water. The loss of this energy to the lukewarm water makes that water feel cool.

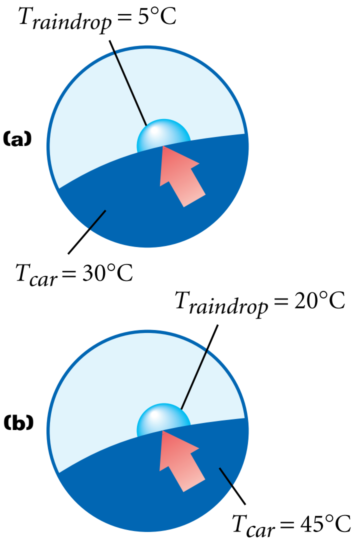


Figure 10-10
The energy transferred as heat from the car's surface to the raindrop is the same for low temperatures **(a)** as for high temperatures **(b),** provided the temperature differences are the same.

Heat has the units of energy

Before scientists arrived at the current model for heat, several different units for measuring heat had already been developed. These units are still widely used in many applications and therefore are listed in **Table 10-3.** Because heat, like work, is energy in transit, all heat units can be converted to joules, the SI unit for energy.

Just as other forms of energy have a symbol that identifies them (*PE* for potential energy, *KE* for kinetic energy, *U* for internal energy, *W* for work), heat is indicated by the symbol *Q*.

Table 10-3 Thermal units and their values in joules

Heat unit	Equivalent value	Uses
joule (J)	equal to 1 kg • $\left(\frac{m^2}{s^2}\right)$	SI unit of energy
calorie (cal)	4.186 J	non-SI unit of heat; found especially in older works of physics and chemistry
kilocalorie (kcal)	4.186×10^3 J	non-SI unit of heat
Calorie, or dietary Calorie	4.186×10^3 J = 1 kcal	food and nutritional science
British thermal unit (Btu)	1.055×10^3 J	English unit of heat; used in engineering, air-conditioning, and refrigeration
therm	1.055×10^8 J	equal to 100 000 Btu; used to measure natural-gas usage

HEAT AND WORK

Hammer a nail into a block of wood. After several minutes, pry the nail loose from the block and touch the side of the nail. It feels warm to the touch, indicating that energy is being transferred from the nail to your hand. Work is done in pulling the nail out of the wood. The nail encounters friction with the wood, and most of the energy required to overcome this friction is transformed into internal energy. The increase in the internal energy of the nail raises the nail's temperature, and the temperature difference between the nail and your hand results in the transfer of energy to your hand as heat.

Friction is just one way of increasing a substance's internal energy. In the case of solids, internal energy can be increased by deforming their structure. Common examples of this are when a rubber band is stretched or a piece of metal is bent.

Total energy is conserved

When the concept of mechanical energy was introduced in Chapter 5, you discovered that whenever friction between two objects exists, not all of the work done in overcoming friction appears as mechanical energy. Similarly, not all of the kinetic energy in inelastic collisions remains as kinetic energy. Some of this energy is absorbed by the objects as internal energy. This is why, in the case of the nail pulled from the wood, the nail (and if you could touch it, the wood inside the hole) feels warm. If changes in internal energy are taken into account along with changes in mechanical energy, the total energy is a universally conserved property.

CONSERVATION OF ENERGY

$$\Delta PE + \Delta KE + \Delta U = 0$$

the change in potential energy + the change in kinetic energy + the change in internal energy = 0

Quick Lab

Work and Heat

MATERIALS LIST

- ✔ 1 large rubber band about 7–10 mm wide

SAFETY CAUTION

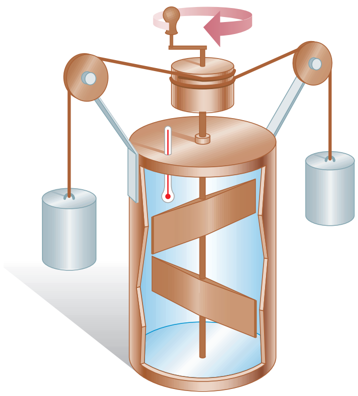
To avoid breaking the rubber band, do not stretch it more than a few inches. Do not point a stretched rubber band at another person.

Hold the rubber band between your thumbs. Touch the middle section of the rubber band to your lip and note how it feels. Rapidly stretch the rubber band and keep it stretched. Touch the middle section of the rubber band to your lip again. Notice whether the rubber band's temperature has changed. (You may have to repeat this procedure several times before you can clearly distinguish the temperature difference.)

Conservation of energy

PROBLEM

An arrangement similar to the one used to demonstrate energy conservation is shown at right. A vessel contains water. Paddles that are propelled by falling masses turn in the water. This agitation warms the water and increases its internal energy. The temperature of the water is then measured, giving an indication of the water's internal-energy increase. If a total mass of 11.5 kg falls 1.3 m and all of the mechanical energy is converted to internal energy, by how much will the internal energy of the water increase? (Assume no energy is transferred as heat out of the vessel to the surroundings or from the surroundings to the vessel's interior.)



SOLUTION

1. DEFINE **Given:** $m = 11.5$ kg $\quad h = 1.3$ m $\quad g = 9.81$ m/s^2

Unknown: $\Delta PE = ?$ $\quad \Delta KE = ?$ $\quad \Delta U = ?$

2. PLAN **Choose an equation(s) or situation:** The equation for conservation of energy can be expressed as the initial total energy equal to the final total energy. Because there is no kinetic energy in the apparatus when the mass is released or when it comes to rest, both KE_i and KE_f equal zero. Because all of the potential energy is assumed to be converted to internal energy, PE_i can be set equal to mgh if PE_f is set equal to zero.

$$\Delta PE + \Delta KE + \Delta U = 0$$

$$PE_i + KE_i + U_i = PE_f + KE_f + U_f$$

$$PE_i = mgh$$

$$PE_f = 0$$

$$KE_i = 0$$

$$KE_f = 0$$

$$mgh + 0 + U_i = 0 + 0 + U_f$$

$$\Delta U = U_f - U_i = mgh$$

3. CALCULATE **Substitute values into the equation(s) and solve:**

$$\Delta U = (11.5\ \text{kg})(9.81\ \text{m/s}^2)(1.3\ \text{m})$$
$$= 1.5 \times 10^2\ \text{J}$$

$$\boxed{\Delta U = 1.5 \times 10^2\ \text{J}}$$

CALCULATOR SOLUTION

Because the minimum number of significant figures in the data is two, the calculator answer, 146.6595 J, should be rounded to two digits.

4. EVALUATE The answer can be estimated using rounded values for m and g. If $m \approx 10$ kg and $g \approx 10$ m/s^2, then $\Delta U \approx 130$ J, which is close to the actual value calculated.

PRACTICE 10B

Conservation of energy

1. In the arrangement described in Sample Problem 10B, how much would the water's internal energy increase if the mass fell 6.69 m?
2. A worker drives a 0.500 kg spike into a rail tie with a 2.50 kg sledgehammer. The hammer hits the spike with a speed of 65.0 m/s. If one-third of the hammer's kinetic energy is converted to the internal energy of the hammer and spike, how much does the total internal energy increase?
3. A 3.0×10^{-3} kg copper penny drops a distance of 50.0 m to the ground. If 65 percent of the initial potential energy goes into increasing the internal energy of the penny, determine the magnitude of that increase.
4. A 2.5 kg block of ice at a temperature of 0.0°C and an initial speed of 5.7 m/s slides across a level floor. If 3.3×10^5 J are required to melt 1.0 kg of ice, how much ice melts, assuming that the initial kinetic energy of the ice block is entirely converted to the ice's internal energy?
5. The amount of internal energy needed to raise the temperature of 0.25 kg of water by 0.2°C is 209.3 J. How fast must a 0.25 kg baseball travel in order for its kinetic energy to equal this internal energy?

Section Review

1. A bottle of water at room temperature is placed in a freezer for a short time. An identical bottle of water that has been lying in the sunlight is placed in a refrigerator for the same amount of time. What must you know to determine which situation involves more energy transfer?
2. Use the microscopic interpretations of temperature and heat to explain how you can blow on your hands to warm them and also blow on a bowl of hot soup to cool it.
3. If a bottle of water is shaken vigorously, will the internal energy of the water change? Why or why not?
4. Water at the top of Niagara Falls has a temperature of 10.0°C. Assume that all of the potential energy goes into increasing the internal energy of the water and that it takes 4186 J/kg to increase the water's temperature by 1°C. If 505 kg of water falls a distance of 50.0 m, what will the temperature of the water be at the bottom of the falls?

10-3
Changes in temperature and phase

SPECIFIC HEAT CAPACITY

You have probably noticed on a hot day that the air around a swimming pool (like the one shown in **Figure 10-11**) is hot but the pool water is cool. This may seem odd, because both the air and water receive energy from sunlight. The water may be cooler than the air, in part because of evaporation, which is a cooling process. However, there is another property of all substances that causes their temperatures to vary by different amounts when equal amounts of energy are added to or removed from them.

This property can be explained in terms of the motion of atoms and molecules in a substance, which in turn affects how much the substance's temperature changes for a given amount of energy that is added or removed. Each substance has a unique value for the energy required to change the temperature of 1 kg of that substance by 1°C. This value, known as the **specific heat capacity** (or sometimes just specific heat) of the substance, relates mass, temperature change, and energy transferred as heat.

The specific heat capacity is related to energy transferred, mass, and temperature change by the following equation:

SPECIFIC HEAT CAPACITY

$$c_p = \frac{Q}{m\Delta T}$$

$$\text{specific heat capacity} = \frac{\text{energy transferred as heat}}{\text{mass} \times \text{change in temperature}}$$

The subscript p indicates that the specific heat capacity is measured at constant pressure. Maintaining constant pressure is an important detail when determining certain thermal properties of gases, which are much more affected by changes in pressure than are solids or liquids. Note that a temperature change of 1°C is equal in magnitude to a temperature change of 1 K, so that ΔT gives the temperature change in either scale.

The equation for specific heat capacity applies to both substances that absorb energy from their surroundings and those that transfer energy to their surroundings. When the temperature increases, ΔT and Q are taken to be positive, which corresponds to energy transferred into the substance. Likewise, when the temperature decreases, ΔT and Q are negative and energy is

10-3 SECTION OBJECTIVES

- Perform calculations with specific heat capacity.
- Perform calculations involving latent heat.
- Interpret the various sections of a heating curve.

Figure 10-11
The air around the pool and the water in the pool receive energy from sunlight. However, the increase in temperature is greater for the air than for the water.

specific heat capacity

the quantity of energy needed to raise the temperature of 1 kg of a substance by 1°C at constant pressure

Table 10-4 Specific heat capacities

Substance	c_p (J/kg•°C)	Substance	c_p (J/kg•°C)
aluminum	8.99×10^2	lead	1.28×10^2
copper	3.87×10^2	mercury	1.38×10^2
glass	8.37×10^2	silver	2.34×10^2
gold	1.29×10^2	steam	2.01×10^3
ice	2.09×10^3	water	4.186×10^3
iron	4.48×10^2		

transferred from the substance. **Table 10-4** lists specific heat capacities that have been determined for several substances.

Determining specific heat capacity

calorimetry

an experimental procedure used to measure the energy transferred from one substance to another as heat

To measure the specific heat capacity of a substance, it is necessary to measure mass, temperature change, and energy transferred as heat. Mass and temperature change are directly measurable, but the direct measurement of heat is difficult. However, the specific heat capacity of water (4.186 kJ/kg•°C) is well known, so the energy transferred as heat between an object of unknown specific heat capacity and a known quantity of water can be measured.

If a hot substance is placed in an insulated container of cool water, energy conservation requires that the energy the substance gives up must equal the energy absorbed by the water. Although some energy is transferred to the surrounding container, this effect is small and will be ignored in this discussion. Energy conservation can be used to calculate the specific heat capacity, $c_{p,x}$, of the substance (indicated by the subscript x). For simplicity, a subscript w will always stand for "water" in problems involving specific heat capacities.

energy absorbed by water = energy released by the substance

$$Q_w = Q_x$$

$$c_{p,w} m_w \Delta T_w = c_{p,x} m_x \Delta T_x$$

The energy gained by a substance is usually expressed as a positive quantity, and energy released usually has a negative value. The minus sign of the latter quantity can be eliminated if ΔT_x and ΔT_{water} are written as the larger temperature value minus the smaller one. Therefore, ΔT should always be written as a positive quantity for this equation.

This approach to determining a substance's specific heat capacity is called **calorimetry,** and devices that are used for making this measurement are called *calorimeters.* A calorimeter also contains a thermometer for measuring the final temperature when the substances are at thermal equilibrium and a stirrer to ensure the uniform mixture of energy throughout the water (see **Figure 10-12**).

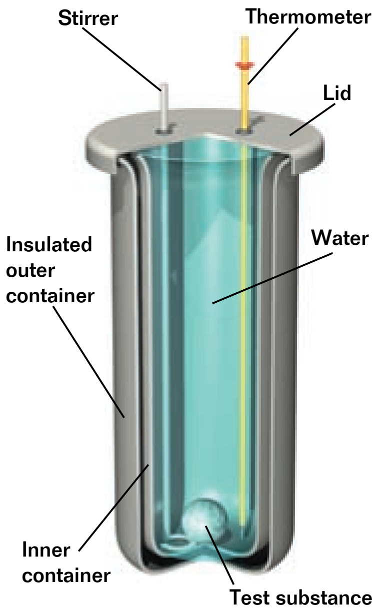


Figure 10-12
A simple calorimeter allows the specific heat capacity of a substance to be determined.

Calorimetry

PROBLEM

A 0.050 kg metal bolt is heated to an unknown initial temperature. It is then dropped into a beaker containing 0.15 kg of water with an initial temperature of 21.0°C. The bolt and the water then reach a final temperature of 25.0°C. If the metal has a specific heat capacity of 899 J/kg•°C, find the initial temperature of the metal.

SOLUTION

1. DEFINE **Given:**

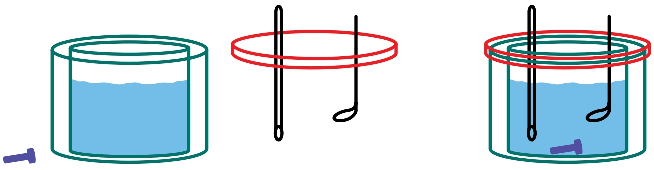
$m_{metal} = m_m = 0.050$ kg $c_{p,m} = 899$ J/kg•°C

$m_{water} = m_w = 0.15$ kg $c_{p,w} = 4186$ J/kg•°C

$T_{water} = T_w = 21.0°C$ $T_{final} = T_f = 25.0°C$

Unknown: $T_{metal} = T_m = ?$

Diagram:



2. PLAN **Choose an equation(s) or situation:** Equate the energy removed from the bolt to the energy absorbed by the water.

energy removed from metal = energy absorbed by water

$$c_{p,m} m_m \Delta T_m = c_{p,w} m_w \Delta T_w$$

Rearrange the equation to isolate the unknown:

$$\Delta T_m = \frac{m_w c_{p,w} \Delta T_w}{m_m c_{p,m}}$$

3. CALCULATE **Substitute values into the equation(s) and solve:**

Note that ΔT_w has been made positive.

$$\Delta T_w = T_f - T_w = 25.0°C - 21.0°C = 4.0°C$$

$$\Delta T_m = \frac{(0.15 \text{ kg})\left(\frac{4186 \text{ J}}{\text{kg}\bullet°\text{C}}\right)(4.0°\text{C})}{(0.050 \text{ kg})\left(\frac{899 \text{ J}}{\text{kg}\bullet°\text{C}}\right)}$$

continued on next page

$$\Delta T_m = 56°C$$

$$T_m = T_f + \Delta T_m$$

$$T_m = 25°C + 56°C = 81°C$$

$$\boxed{T_m = 81°C}$$

PRACTICE 10C

Calorimetry

1. What is the final temperature when a 3.0 kg gold bar at 99°C is dropped into 0.22 kg of water at 25°C?
2. A 0.225 kg sample of tin initially at 97.5°C is dropped into 0.115 kg of water initially at 10.0°C. If the specific heat capacity of tin is 230 J/kg•°C, what is the final equilibrium temperature of the tin-water mixture?
3. What is the final temperature when 0.032 kg of milk at 11°C is added to 0.16 kg of coffee at 91°C? Assume the specific heat capacities of the two liquids are the same as water, and disregard any energy transfer to the liquids' surroundings.
4. A cup is made of an experimental material that can hold hot liquids without significantly increasing its own temperature. The 0.75 kg cup has an initial temperature of 36.5°C when it is submerged in 1.25 kg of water with an initial temperature of 20.0°C. What is the cup's specific heat capacity if the final temperature is 24.4°C?
5. Brass is an alloy made from copper and zinc. A 0.59 kg brass sample at 98.0°C is dropped into 2.80 kg of water at 5.0°C. If the equilibrium temperature is 6.8°C, what is the specific heat capacity of brass?
6. The air temperature above coastal areas is profoundly influenced by the large specific heat capacity of water. How large of a volume of air can be cooled by 1.0°C if energy is transferred as heat from the air to the water, thus increasing the temperature of 1.0 kg of water by 1.0°C? The specific heat capacity of air is approximately 1000.0 J/kg•°C, and the density of air is approximately 1.29 kg/m^3.
7. A hot, just-minted copper coin is placed in 101 g of water to cool. The water temperature changes by 8.39°C and the temperature of the coin changes by 68.0°C. What is the mass of the coin? Disregard any energy transfer to the water's surroundings.

Heating and Cooling from the Ground Up

As the earliest cave dwellers knew, a good way to stay warm in the winter and cool in the summer is to go underground. Now scientists and engineers are using the same premise—and using existing technology in a new, more efficient way—to heat and cool aboveground homes for a fraction of the cost of conventional systems.

"At any given occasion, the earth temperature is the seasonal average temperature," said Gunnar Walmet, of the New York State Energy Research and Development Authority (NYSERDA). "In New York state, that's typically about 50°F all year long."

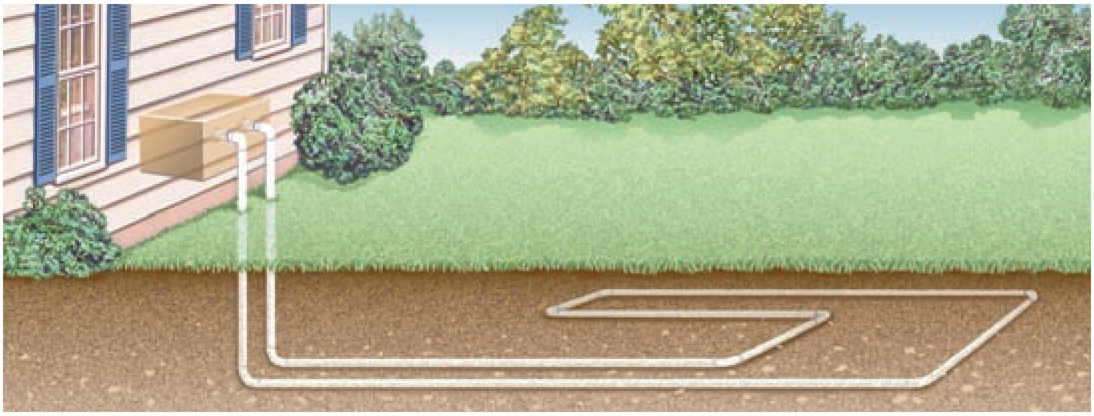
Although the average specific heat capacity of earth has a smaller value than the specific heat capacity of air, the earth has a greater density. That means there are more kilograms of earth than there are of air near a house and that a 1°C change in temperature involves transferring more energy to or from the ground than to or from the air. Thus, in the wintertime, the ground will probably have a higher temperature than the air above it, while in the summer, the ground will likely have a lower temperature than the air.

An earth-coupled heat pump enables homeowners to tap the earth's belowground temperature to heat their homes in the winter or cool them during the summer. The system includes a network of plastic pipes placed in trenches or inserted in holes drilled 2 to 3 m (6 to 10 ft) beneath the ground's surface. To heat a home, a fluid circulates through the pipe, absorbs energy from the surrounding earth, and transfers this energy to a heat pump inside the house.

The heat pump uses a compressor, tubing, and refrigerant to transfer the energy from the liquid to the air inside the house. A blower-and-duct system distributes the warm air through the home. According to NYSERDA, the system can deliver up to four times as much energy into the house as the electrical energy needed to drive it.

Like other heat pumps, the system is reversible. In the summer, it can transfer energy from the air in the house to the system of pipes belowground.

There are currently tens of thousands of earth-coupled heat pumps installed throughout the United States. Although the system can function anywhere on Earth's surface, it is most appropriate in severe climates, where dramatic temperature swings may not be ideal for air-based systems.



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LATENT HEAT

If you place an ice cube with a temperature of −25°C in a pan and then place the pan on a hot range element or burner, the temperature of the ice will increase until the ice begins to melt at 0°C. By knowing the mass and specific heat capacity of ice, you can calculate how much energy is being added to the ice from the element. However, this procedure only works as long as the ice remains ice and its temperature continues to rise as energy is transferred to it.

The graph in **Figure 10-13** and data in **Table 10-5** show how the temperature of 10.0 g of ice changes as energy is added. You can see that as the ice is heated there is a steady increase in temperature from −25°C to 0°C (segment **A** of the graph).

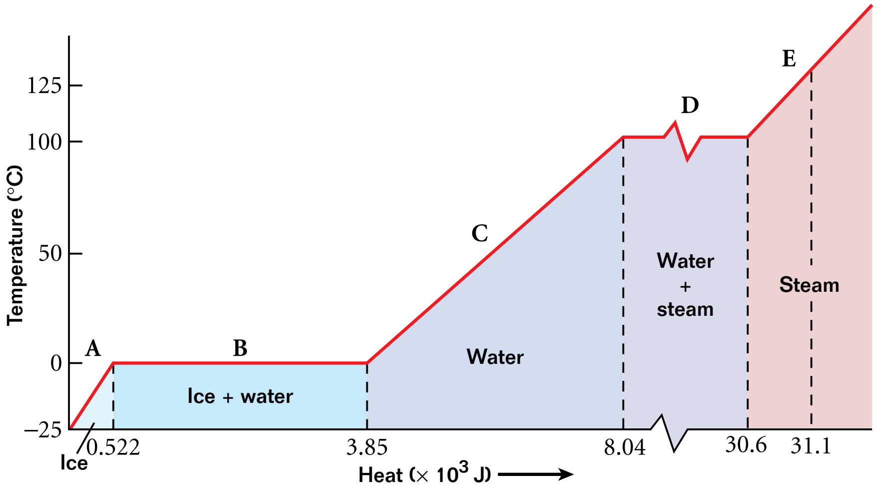


Figure 10-13
This idealized graph shows the temperature change of 10.0 g of ice as it is heated from −25°C in the ice phase to steam above 125°C at atmospheric pressure.

The situation at 0°C is very different. Despite the fact that energy is continuously being added, there is no change in temperature. Instead, the nature of the ice changes. The ice begins to melt and change into water at 0°C (segment **B**). The ice-and-water mixture remains at this temperature until all of the ice melts. From 0°C to 100°C, the water's temperature steadily increases (segment **C**). At 100°C, however, the temperature stops rising and the water turns into

Table 10-5 Changes occurring during the heating of 10.0 g of ice

Segment of graph	Type of change	Amount of energy transferred as heat	Temperature range of segment
A	temperature of ice increases	522 J	−25°C to 0°C
B	ice melts; becomes water	3.33×10^3 J	0°C
C	temperature of water increases	4.19×10^3 J	0°C to 100°C
D	water boils; becomes steam	2.26×10^4 J	100°C
E	temperature of steam increases	502 J	100°C to 125°C

steam (segment **D**). Once the water has completely vaporized, the temperature of the steam increases (segment **E**). Steam whose temperature is greater than the boiling point of water is referred to as *superheated.*

When substances melt, freeze, boil, condense, or sublime (change from a solid to vapor or from vapor to a solid), the energy added or removed changes the internal energy of the substance without changing its temperature. These changes in matter are called **phase changes.**

phase change

the physical change of a substance from one state (solid, liquid, or gas) to another at constant temperature and pressure

The existence of phase changes requires that the definition of heat be expanded. *Heat is the energy that is exchanged between two objects at different temperatures or between two objects at the same temperature when one of them is undergoing a phase change.*

Phase changes involve potential energy between particles

To understand the behavior of a substance undergoing a phase change, you will need to recall how energy is transferred in collisions. Potential energy is the energy an object has because of its position relative to another object. Examples of potential energy are a pencil about to fall from your desk or a rubber band held in a tightly stretched position. Potential energy is present among a collection of particles in a solid or a liquid in the form of attractive bonds. These bonds result from the charges within atoms and molecules. Potential energy is associated with the electric forces between these charges.

The equilibrium separation between atoms or molecules corresponds to a position at which there is a minimum potential energy. The potential energy increases with increasing atomic separation from the equilibrium position. This resembles the elastic potential energy of a spring, as discussed in Chapter 5. For this reason, a collection of individual atoms or molecules and the bonds between them are often modeled as masses at the ends of springs.

If the particles are far enough apart, the bonds between them can break. The work needed to increase potential energy and break a bond is provided by collisions with energetic atoms or molecules, as shown in **Figure 10-14.** Just as bonds can be broken, new bonds can be formed if atoms or molecules are brought close together. This involves the collection of particles going from a high potential energy (large average separation) to a lower potential energy (small average separation). This decrease in potential energy involves a release of energy in the form of increasing kinetic energy of nearby particles.

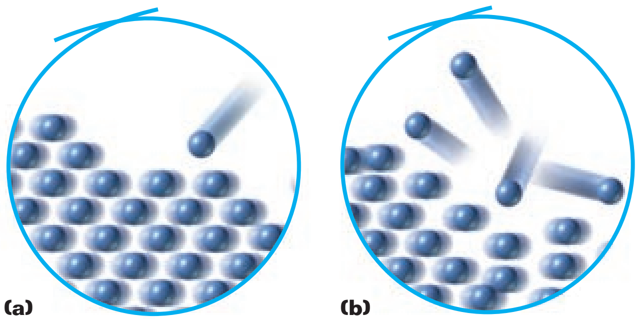


Figure 10-14
Energy added to a substance **(a)** can increase the vibrational kinetic energy of its particles or break the bonds between those particles **(b).**

Energy required to melt a substance goes into rearranging the molecules

Phase changes result from a change in the potential energy between particles of a substance. When energy is added to or removed from a substance undergoing a phase change, the particles of the substance rearrange themselves to make up for their change of energy. This occurs without a change in the average kinetic energy of the particles.

For instance, if ice is melting, the absorbed energy is sufficient to break the weak bonds that hold the water molecules together as a well-ordered crystal. New but different bonds form between the liquid water molecules that have separated from the crystal, so some of the absorbed energy is released again. The difference between the potential energies of the broken bonds and the newly formed bonds is equal to the net energy added to the ice, as shown in **Figure 10-15.** As a result, the energy used to rearrange the molecules is not available to increase the molecules' kinetic energy, and therefore no increase in the temperature of the ice-and-water mixture occurs.

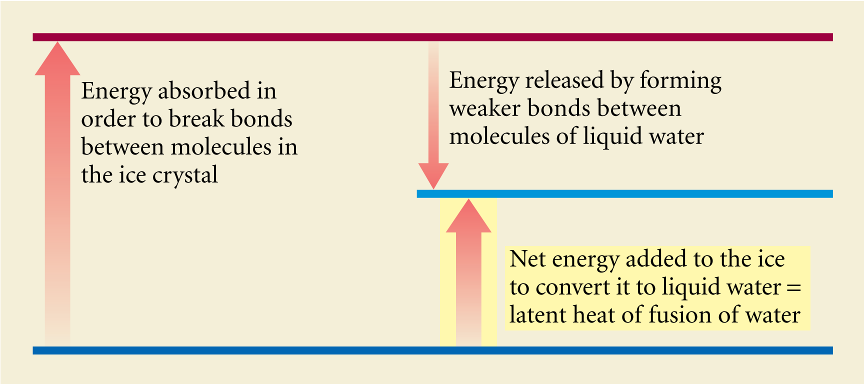


Figure 10-15
The heat of fusion is the difference between the energy needed to break bonds in a solid and the energy released when bonds form in a liquid.

For any substance, the energy added to the substance during melting equals the difference between the total potential energies for particles in the solid and the liquid phases. This energy per unit mass is called the **heat of fusion.**

heat of fusion

the energy per unit mass transferred in order to change a substance from solid to liquid or from liquid to solid at constant temperature and pressure

Energy required to vaporize a substance mostly goes into separating the molecules

As in a solid, the molecules are close together in a liquid, and in liquid water they are even closer together than in ice. The forces between liquid water molecules are stronger than those that exist between the more widely separated water molecules in steam. Therefore, at 100°C, all the energy absorbed by water goes into overcoming the attractive forces between the liquid water molecules. None is available to increase the kinetic energy of the molecules.

The energy added to a substance during vaporization equals the difference in the potential energy of attraction between the particles of a liquid and the potential energy of attraction between the gas particles (see **Figure 10-16**). This energy per unit mass is called the **heat of vaporization.**

heat of vaporization

the energy per unit mass transferred in order to change a substance from liquid to vapor or from vapor to liquid at constant temperature and pressure

Because they have few close neighbors, the particles in the gas phase gain very little energy from weak bonding. Therefore, more energy is required to vaporize a given mass of substance than to melt it. As a result, the heat of vaporization is much greater than the heat of fusion. Both the heat of fusion and the heat of vaporization are classified as **latent heat.**

latent heat

the energy per unit mass that is transferred during a phase change of a substance

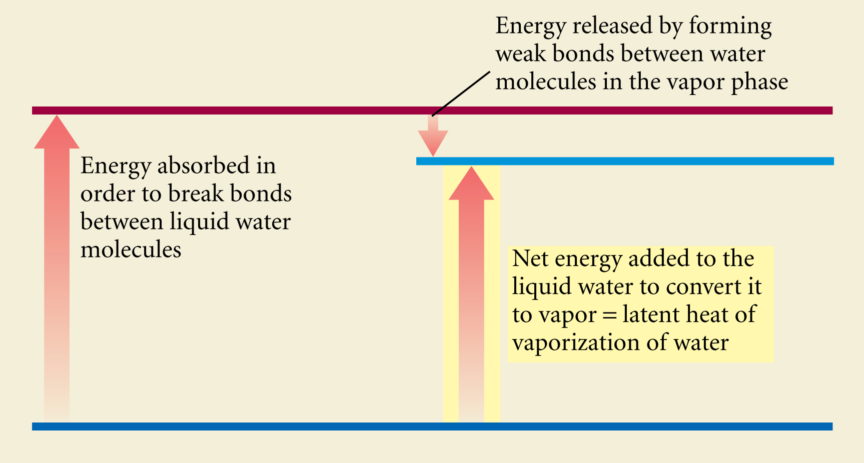


Figure 10-16
The heat of vaporization is mostly the energy required to separate molecules from the liquid phase.

LATENT HEAT

$$Q = mL$$

Energy transferred as heat during a phase change = mass × latent heat

For calculations involving melting or freezing, the latent heat of fusion is noted by the symbol L_f. Similarly, for calculations involving vaporizing or condensing, the symbol L_v is used for latent heat of vaporization. **Table 10-6** lists latent heats for a few substances.

Table 10-6 Latent heats of fusion and vaporization at standard pressure

Substance	Melting point (°C)	L_f (J/kg)	Boiling point (°C)	L_v (J/kg)
nitrogen	−209.97	2.55×10^4	−195.81	2.01×10^5
oxygen	−218.79	1.38×10^4	−182.97	2.13×10^5
ethyl alcohol	−114	1.04×10^5	78	8.54×10^5
water	0.00	3.33×10^5	100.00	2.26×10^6
lead	327.3	2.45×10^4	1745	8.70×10^5
aluminum	660.4	3.97×10^5	2467	1.14×10^7

SAMPLE PROBLEM 10D

Heat of phase change

PROBLEM

How much energy is removed when 10.0 g of water is cooled from steam at 133.0°C to liquid at 53.0°C?

SOLUTION

1. DEFINE

Given: $T_{steam} = T_s = 133.0\ °C \qquad T_{water} = T_w = 53.0°C$

$c_{p,steam} = c_{p,s} = 2.01 \times 10^3\ J/kg{\bullet}°C$

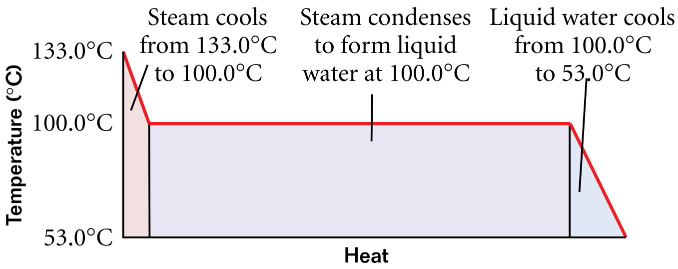
$c_{p,water} = c_{p,w} = 4.186 \times 10^3\ J/kg{\bullet}°C$

$L_v = 2.26 \times 10^6\ J/kg$

$m = 10.0\ g = 10.0 \times 10^{-3}\ kg$

Unknown: $Q_{total} = ?$

Diagram:



2. PLAN

Choose an equation(s) or situation: Heat is calculated by using $Q = mc_p\Delta T$ when no phase changes occur. When steam changes to liquid water, a phase change occurs and the equation for the heat of vaporization, $Q = mL_v$, must be used. Be sure that ΔT is positive for each step.

To cool the steam to 100.0°C: $Q_1 = mc_{p,s}\Delta T$

To change steam to water at 100.0°C: $Q_2 = mL_v$

To cool the water to 53.0°C: $Q_3 = mc_{p,w}\Delta T$

3. CALCULATE

Substitute values into the equation(s) and solve: Find ΔT for steam cooling and water cooling. Calculate Q for both cooling steps and the phase change.

For the cooling steam:

$$\Delta T_s = 133.0°C - 100.0°C = 33.0°C$$

$$Q_1 = mc_{p,s}\Delta T = (10.0 \times 10^{-3}\ kg)\left(2.01 \times 10^3 \frac{J}{kg{\bullet}°C}\right)(33.0°C)$$

$$= 663\ J$$

For the steam condensing to water:

$$Q_2 = mL_v = (10.0 \times 10^{-3}\ kg)\left(2.26 \times 10^6 \frac{J}{kg}\right)$$

$$= 2.26 \times 10^4\ J$$

For the cooling water:

$$\Delta T_w = 100.0°C - 53.0°C = 47.0°C$$

$$Q_3 = mc_{p,w}\Delta T = (10.0 \times 10^{-3}\ \text{kg})\left(4.186 \times 10^3 \frac{\text{J}}{\text{kg}\bullet°\text{C}}\right)(47.0°C)$$

$$= 1.97 \times 10^3\ \text{J}$$

$$Q_{total} = Q_1 + Q_2 + Q_3 =$$
$$663\ \text{J} + (2.26 \times 10^4\ \text{J})$$
$$+ (1.97 \times 10^3\ \text{J}) = 2.52 \times 10^4\ \text{J}$$

$$\boxed{Q_{total} = 2.52 \times 10^4\ \text{J removed}}$$

CALCULATOR SOLUTION

Because of the significant figure rule for addition, the calculator answer, 25233, should be rounded to 2.52×10^4.

4. EVALUATE Most of the energy is added to or removed from a substance during phase changes. In this example, about 90 percent of the energy removed from the steam is accounted for by the heat of vaporization.

PRACTICE 10D

Heat of phase change

1. How much energy is required to change a 42 g ice cube from ice at –11°C to steam at 111°C? (Hint: Refer to **Tables 10-4** and **10-6.**)

2. Liquid nitrogen, which has a boiling point of 77 K, is commonly used to cool substances to low temperatures. How much energy must be removed from 1.0 kg of gaseous nitrogen at 77 K for it to completely liquefy?

3. How much energy is needed to melt 0.225 kg of lead so that it can be used to make a lead sinker for fishing? The sample has an initial temperature of 27.3°C and is poured in the mold immediately after it has melted.

4. How much energy is needed to melt exactly 1000 aluminum cans, each with a mass of 14.0 g, for recycling? Assume an initial temperature of 26.4°C.

5. A 0.011 kg cube of ice at 0.0°C is added to 0.450 kg of soup at 80.0°C. Assuming that the soup has the same specific heat capacity as water, find the final temperature of the soup after the ice has melted. (Hint: There is a temperature change after the ice melts.)

6. At a foundry, 25 kg of molten aluminum with a temperature of 660.4°C is poured into a mold. If this is carried out in a room containing 130 kg of air at 25°C, what is the temperature of the air after the aluminum is completely solidified? Assume that the specific heat capacity of air is 1.0×10^3 J/kg•°C.

Section Review

1. A jeweler working with a heated 47 g gold ring must lower the ring's temperature to make it safe to handle. If the ring is initially at 99°C, what mass of water at 25°C is needed to lower the ring's temperature to 38°C?

2. Using the concepts of latent heat and internal energy, explain why it is difficult to build a fire with damp wood.

3. Why does steam at 100°C cause more severe burns than does liquid water at 100°C?

4. From the heating curve for a 15 g sample, as shown in **Figure 10-17,** estimate the following properties of the substance.
   - **a.** the specific heat capacity of the liquid
   - **b.** the latent heat of fusion
   - **c.** the specific heat capacity of the solid
   - **d.** the specific heat capacity of the vapor
   - **e.** the latent heat of vaporization

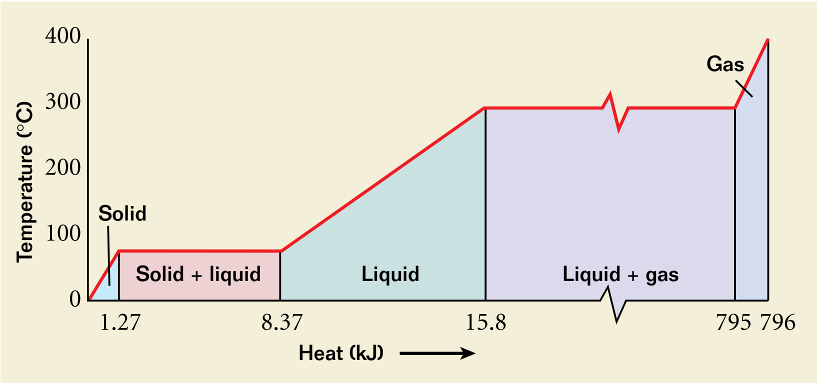


Figure 10-17

5. **Physics in Action** How much energy must be added to a bowl of 125 popcorn kernels in order for them to reach a popping temperature of 175°C? Assume that their initial temperature is 21°C, that the specific heat capacity of popcorn is 1650 J/kg•°C, and that each kernel has a mass of 0.105 g.

6. **Physics in Action** Because of the pressure inside a popcorn kernel, water does not vaporize at 100°C. Instead, it stays liquid until its temperature is about 175°C, at which point the kernel ruptures and the superheated water turns into steam. How much energy is needed to pop 95.0 g of corn if 14 percent of a kernel's mass consists of water? Assume that the latent heat of vaporization for water at 175°C is 0.90 times its value at 100°C and that the kernels have an initial temperature of 175°C.

10-4
Controlling heat

THERMAL CONDUCTION

When you first place an iron skillet on a range burner or element, the metal handle feels comfortable to the touch. But after a few minutes the handle becomes too hot to touch without a cooking mitt. During that time, energy is transferred as heat from the high-temperature burner to the skillet. This type of energy transfer is called **thermal conduction.**

Thermal conduction can be understood by the behavior of atoms in a metal. Before the skillet is placed on the heating element, the skillet's iron atoms have an energy proportional to the temperature of the room. As the skillet is heated, the atoms nearest the heating element vibrate with greater energy. These vibrating atoms jostle their less energetic neighbors and transfer some of their energy in the process. Gradually, iron atoms farther away from the element gain more energy.

The rate of thermal conduction depends on the properties of the substance being heated. A metal ice tray and a package of frozen food removed from the freezer are at the same temperature. However, the metal tray feels colder because metal conducts energy more easily and more rapidly than cardboard at the place where it comes into contact with your hand. In contrast, a piece of ceramic conducts energy very slowly, as may be seen in **Figure 10-18.** The end of the ceramic piece that is embedded in ice is barely affected by the energy of the flame surrounding the other end. Substances that rapidly transfer energy as heat are called *thermal conductors,* while those that slowly transfer energy as heat are called *thermal insulators.*

In general, metals are good thermal conductors. Materials such as asbestos, cork, ceramic, cardboard, and fiberglass are poor thermal conductors (and therefore good thermal insulators). Gases also are poor thermal conductors. The gas particles are so far apart with respect to their size that collisions between them are rare, and their kinetic energy is transferred slowly.

Although cooking oil is not any better as a thermal conductor than most nonmetals, it is useful for transferring energy uniformly around the surface of the food being cooked. When popping popcorn, for instance, coating the kernels with oil improves the energy transfer to each kernel, so that a higher percentage of them pop.

Figure 10-18
Ceramics are poor thermal conductors, as indicated in this photograph.

10-4 SECTION OBJECTIVES

- **Explain how energy is transferred as heat through the process of thermal conduction.**
- **Recognize how energy transfer can be controlled with clothing.**

thermal conduction

the process by which energy is transferred as heat through a material between two points at different temperatures internetconnect

SCiLINKS NSTA

TOPIC: Conduction and convection
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Convection and radiation also transfer energy

There are two other mechanisms for transferring energy between places or objects at different temperatures. *Convection* involves the displacement of cold matter by hot matter, such as when hot air over a flame rises upward. This mechanism does not involve heat alone. Instead, it uses the combined effects of pressure differences, conduction, and buoyancy. In the case of air over a flame, the air is heated through particle collisions (conduction), causing it to expand and its density to decrease. The warm air is then displaced by denser, colder air from above.

The other principal energy transfer mechanism is *electromagnetic radiation,* which includes visible light. Unlike convection, energy in this form does not involve the transfer of matter. Instead, objects reduce their internal energy by radiating electromagnetic radiation with particular wavelengths. This energy transfer often takes place from high temperature to low temperature, so that radiation is frequently associated with heat. The human body emits energy in the infrared portion of the electromagnetic spectrum.

CONCEPT PREVIEW

Electromagnetic radiation will be discussed in more detail in Chapter 14.

CLOTHING AND CLIMATE

To remain healthy, the human body must maintain a temperature close to 37.0°C (98.6°F). This becomes increasingly difficult as the surrounding air becomes hotter or colder than body temperature.

Without proper insulation, the body's temperature will drop in its attempt to reach thermal equilibrium with very cold surroundings. If this situation is not corrected in time, the body will enter a state of hypothermia, which lowers pulse, blood pressure, and respiration. Once the body temperature reaches 32.2°C (90.0°F), the person can lose consciousness. When the body temperature reaches 25.6°C (78.0°F), hypothermia is almost always fatal.

Insulating materials retain energy for cold climates

To prevent hypothermia, the transfer of energy from the human body to the surrounding air must be hindered. This is done by surrounding the body with heat-insulating material. An extremely effective and common thermal insulator is air. Like most gases, air is a very poor thermal conductor, so even a thin layer of air near the skin provides a barrier to energy transfer.

The Inupiat Eskimo people of northern Alaska have designed clothing to protect them from the severe Arctic climate, where average air temperatures range from 10°C (50°F) to –37°C (–35°F). The Inupiat clothing is made from animal skins that make use of air's insulating properties. Until recently, the traditional parka *(atigi)* was made from caribou skins. Two separate parkas are worn in layers, with the fur lining the inside of the inner parka and the outside of the outer parka. Insulation is provided by air trapped between the short inner hairs and within the long, hollow hairs of the fur. Today, inner parkas are made from sheepskin (see **Figure 10-19**).

Figure 10-19
The Inupiat parka, called an *atigi,* consists today of a canvas shell over sheepskin. The wool provides layers of insulating air between the wearer and the cold.

Evaporation aids energy transfer in hot climates

At the other extreme, the Bedouins of the Arabian Desert have developed clothing that permits them to survive another of the harshest environments on Earth. Bedouin garments cover most of the body, thus protecting the wearer from direct sunlight and preventing excessive loss of body water from evaporation. These clothes are also designed to cool the wearer. The Bedouins must keep their body temperatures from becoming too high in desert temperatures, which often are in excess of 38°C (100°F). Heat exhaustion or heatstroke will result if the body's temperature becomes too high.

Although there are a number of differences among the types of clothing worn by different tribes and by men and women within tribes, a few basic garments are common to all Bedouins. One of these is the loose-fitting, elongated linen shirt called a *dish-dash* or *dish-dasha,* depending on whether it is worn by men or women, respectively. This shirt is worn close to the body, usually over an undergarment.

The loose fit and flared cut of the *dish-dash* permits air to flow over the wearer's skin. This causes any perspiration that has collected on the skin's surface to evaporate. During evaporation, water molecules enter the vapor phase. Because of the high specific heat capacity and latent heat of vaporization for water, evaporation removes a good deal of energy from the skin and air, thus causing the skin to cool.

Another common article of clothing is the *kefiyah,* a headcloth worn by Bedouin men, as shown in **Figure 10-20.** A similar garment made of two separate cloths, which are called a *mandil* and a *hatta*, is worn by Bedouin women. Firmly wrapped around the head of the wearer, the cloth absorbs perspiration and cools the wearer during evaporation. This same garment is also useful during cold periods in the desert. Wound snugly around the head, the garment traps air within its folds, thus providing an insulating layer to keep the head warm.

Figure 10-20
The Bedouin headcloth, called a *kefiyah,* employs evaporation to remove energy from the air close to the head, thus cooling the wearer.

Section Review

1. Why do fluffy down comforters feel warmer than thin cloth blankets?
2. Explain how conduction causes water on the surface of a bridge to freeze sooner than water on the road surface on either side of the bridge.
3. On a camping trip, your friend tells you that fluffing up a down sleeping bag before you go to bed will keep you warmer than sleeping in the same bag when it is still crushed from being in its storage sack. Explain why this happens. (Hint: A large amount of air is present in an uncrushed sleeping bag.)

CHAPTER 10
Summary

KEY TERMS

calorimetry (p. 372)

heat (p. 365)

heat of fusion (p. 378)

heat of vaporization (p. 378)

internal energy (p. 359)

latent heat (p. 379)

phase change (p. 377)

specific heat capacity (p. 371)

thermal conduction (p. 383)

thermal equilibrium (p. 360)

KEY IDEAS

Section 10-1 Temperature and thermal equilibrium
- Temperature can be changed by transferring energy to or from a substance.
- Thermal equilibrium is the condition in which the temperature of two objects in physical contact with each other is the same.

Section 10-2 Defining heat
- Heat is energy that is transferred from objects at higher temperatures to objects at lower temperatures.
- Energy is conserved when mechanical energy and internal energy are taken into account.

Section 10-3 Changes in temperature and phase
- Specific heat capacity, which is a measure of the energy needed to change a substance's temperature, is described by the following formula:

$$\boxed{c_p = \frac{Q}{m\Delta T}}$$

- Latent heat, the energy required to change the phase of a substance, is described by the following formula:

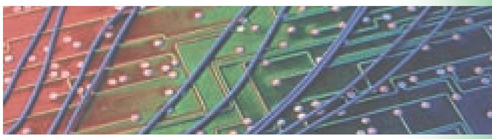
$$\boxed{L = \frac{Q}{m}}$$

Section 10-4 Controlling heat
- Energy is transferred by thermal conduction through particle collisions.

Variable symbols

Quantities		Units	
T	temperature (Kelvin)	K	kelvins
T_C	temperature (Celsius)	°C	degrees Celsius
T_F	temperature (Fahrenheit)	°F	degrees Fahrenheit
ΔU	change in internal energy	J	joules
Q	heat	J	joules
c_p	specific heat capacity at constant pressure	$\frac{\text{J}}{\text{kg} \bullet {}^\circ\text{C}}$	
L	latent heat	$\frac{\text{J}}{\text{kg}}$	

CHAPTER 10
Review and Assess



TEMPERATURE AND THERMAL EQUILIBRIUM

Review questions

1. What is the relationship between temperature and internal energy?

2. What property of two objects determines if the two are in a state of thermal equilibrium?

3. What are some physical properties that could be used in developing a temperature scale?

4. What property must a substance have in order to be used for calibrating a thermometer?

Conceptual questions

5. Which object in each of the following pairs has greater total internal energy, assuming that both objects in each pair are in thermal equilibrium? Explain your reasoning in each case.

- **a.** a metal knife in thermal equilibrium with a hot griddle
- **b.** a 1 kg block of ice at −25°C or seven 12 g ice cubes at −25°C

6. Assume that each pair of objects in item 5 has the same internal energy instead of the same temperature. Which item in each pair will have the higher temperature?

7. Why are the steam and ice points of water better fixed points for a thermometer than the temperature of a human body?

8. How does the temperature of a tub of hot water as measured by a thermometer differ from the water's temperature before the measurement is made? What property of a thermometer is necessary for the difference between these two temperatures to be minimized?

Practice problems

9. The highest recorded temperature on Earth was 136°F, at Azizia, Libya, in 1922. Express this temperature in degrees Celsius and in kelvins.
(See Sample Problem 10A.)

10. The melting point of gold is 1947°F. Express this temperature in degrees Celsius and in kelvins.
(See Sample Problem 10A.)

DEFINING HEAT

Review questions

11. Which drawing in **Figure 10-21** correctly shows the direction in which the net energy is transferred by heat between an ice cube and the freezer walls when the temperature of both is −10°C? Explain your answer.

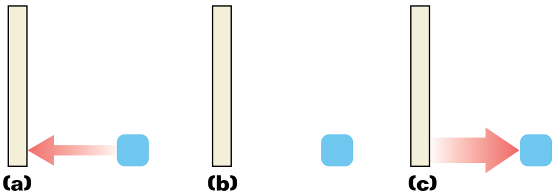


Figure 10-21

12. A glass of water has a temperature of 8°C. In which situation will more energy be transferred, when the air's temperature is 25°C or 35°C?

13. How much energy is transferred between a piece of toast and an oven when both are at a temperature of 55°C? Explain.

Conceptual questions

14. If water in a sealed, insulated container is stirred, is its temperature likely to increase slightly, decrease slightly, or stay the same? Explain your answer.

15. Given your answer to item 14, why does stirring a hot cup of coffee cool it down?

16. Given any two bodies, the one with the higher temperature contains more heat. What is wrong with this statement?

17. Use the kinetic theory of atoms and molecules to explain why energy that is transferred as heat always goes from objects at higher temperatures to those at lower temperatures.

18. In which of the two situations described is more energy transferred? Explain your answer.
    a. a cup of hot chocolate with a temperature of 40°C inside a freezer at −20°C
    b. the same cup of hot chocolate at 90°C in a room at 25°C

Practice problems

19. A force of 315 N is applied horizontally to a wooden crate in order to displace it 35.0 m across a level floor at a constant velocity. As a result of this work the crate's internal energy is increased by an amount equal to 14 percent of the crate's initial internal energy. Calculate the initial internal energy of the crate. (See Sample Problem 10B.)

20. A 0.75 kg spike is hammered into a railroad tie. The initial speed of the spike is equal to 3.0 m/s.
    a. If the tie and spike together absorb 85 percent of the spike's initial kinetic energy as internal energy, calculate the increase in internal energy of the tie and spike.
    b. What happens to the remaining energy?

    (See Sample Problem 10B.)

CHANGES IN TEMPERATURE AND PHASE

Review questions

21. What data are required in order to determine the specific heat capacity of an unknown substance by means of calorimetry?

22. What principle permits calorimetry to be used to determine the specific heat capacity of a substance? Explain.

23. Why does the temperature of melting ice not change even though energy is being transferred as heat to the ice?

Conceptual questions

24. Why does the evaporation of water cool the air near the water's surface?

25. Ethyl alcohol has about one-half the specific heat capacity of water. If equal masses of alcohol and water in separate beakers at the same temperature are supplied with the same amount of energy, which will have the higher final temperature?

26. Until refrigerators were invented, many people stored fruits and vegetables in underground cellars. Why was this more effective than keeping them in the open air?

27. During the winter, the people mentioned in item 26 would often place an open barrel of water in the cellar alongside their produce. Explain why this was done and why it would be effective.

28. During a cold spell, Florida orange growers often spray a mist of water over their trees during the night. What does this accomplish?

29. From the heating curve for a 23 g sample (see **Figure 10-22**), estimate the following properties of the substance.
    a. the specific heat capacity of the liquid
    b. the latent heat of fusion
    c. the specific heat capacity of the solid
    d. the specific heat capacity of the vapor
    e. the latent heat of vaporization

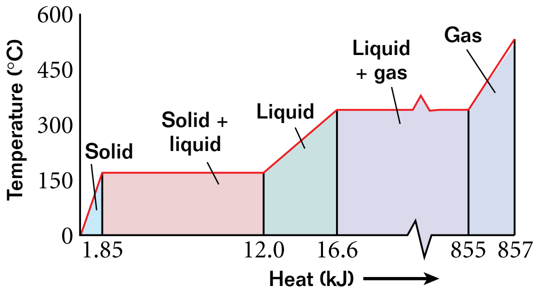


Figure 10-22

Practice problems

30. A 25.5 g silver ring ($c_p = 234$ J/kg•°C) is heated to a temperature of 84.0°C and then placed in a calorimeter containing 5.00×10^{-2} kg of water at 24.0°C. The calorimeter is not perfectly insulated, however, and 0.140 kJ of energy is transferred to the surroundings before a final temperature is reached. What is the final temperature?
(See Sample Problem 10C.)

31. When a driver brakes an automobile, friction between the brake disks and the brake pads converts part of the car's translational kinetic energy to internal energy. If a 1500 kg automobile traveling at 32 m/s comes to a halt after its brakes are applied, how much can the temperature rise in each of the four 3.5 kg steel brake disks? Assume the disks are made of iron ($c_p = 448$ J/kg•°C) and that all of the kinetic energy is distributed in equal parts to the internal energy of the brakes.
(See Sample Problem 10C.)

32. A plastic-foam container used as a picnic cooler contains a block of ice at 0°C. If 225 g of ice melts, how much heat passes through the walls of the container?
(See Sample Problem 10D.)

33. The largest of the Great Lakes, Lake Superior, contains about 1.20×10^{16} kg of water. If the lake had a temperature of 12.0°C, how much energy would have to be removed to freeze the whole lake at 0°C?
(See Sample Problem 10D.)

THERMAL CONDUCTION AND INSULATION

Review questions

34. How does a metal rod conduct energy from one end, which has been placed in a fire, to the other end, which is at room temperature?

35. How does air within winter clothing keep you warm on cold winter days?

Conceptual questions

36. A metal spoon is placed in one of two identical cups of hot coffee. Which cup will be cooler after a few minutes?

37. A tile floor may feel uncomfortably cold to your bare feet, but a carpeted floor in an adjoining room at the same temperature feels warm. Why?

38. Why is it recommended that several items of clothing be worn in layers on cold days?

39. Why does a fan make you feel cooler on a hot day?

40. A paper cup is filled with water and then placed over an open flame, as shown in **Figure 10-23.** Explain why the cup does not catch fire and burn.

Figure 10-23

MIXED REVIEW

41. Absolute zero on the Rankine temperature scale is $T_R = 0$°R, and the scale's unit is the same size as the Fahrenheit degree.

- **a.** Write a formula that relates the Rankine scale to the Fahrenheit scale.
- **b.** Write a formula that relates the Rankine scale to the Kelvin scale.

42. A 3.0 kg rock is initially at rest at the top of a cliff. Assuming the rock falls into the sea at the foot of the cliff and that its kinetic energy is transferred entirely to the water, how high is the cliff if the temperature of 1.0 kg of water is raised 0.10°C?

43. Convert the following temperatures to degrees Fahrenheit and kelvins.

- **a.** the boiling point of liquid hydrogen (–252.87°C)
- **b.** the temperature of a room at 20.5°C

44. The freezing and boiling points of water on the imaginary "Too Hot" temperature scale are selected to be exactly 50 and 200 degrees TH.

- **a.** Derive an equation relating the Too Hot scale to the Celsius scale. (Hint: Make a graph of one temperature scale versus the other, and solve for the equation of the line.)
- **b.** Calculate absolute zero in degrees TH.

45. Show that the temperature –40° is unique in that it has the same numerical value on the Celsius and Fahrenheit scales.

46. A hot-water heater is operated by solar power. If the solar collector has an area of 6.0 m^2 and the power delivered by sunlight is 550 W/m^2, how long will it take to increase the temperature of 1.0 m^3 of water from 21°C to 61°C?

47. A student drops two metallic objects into a 120 g steel container holding 150 g of water at 25°C. One object is a 253 g cube of copper that is initially at 85°C, and the other is a chunk of aluminum that is initially at 5°C. To the surprise of the student, the water reaches a final temperature of 25°C, its initial temperature. What is the mass of the aluminum chunk?

Technology & Learning

Graphing calculators

Refer to Appendix B for instructions on downloading programs for your calculator. The program "Chap10" allows you to analyze a graph of temperature versus energy absorbed for a sample with a known mass and specific heat capacity.

Specific heat capacity, as you learned earlier in this chapter, is described by the following equation:

$$c_p = \frac{Q}{m\Delta t}$$

The program "Chap10" stored on your graphing calculator makes use of the equation for specific heat capacity. Once the "Chap10" program is executed, your calculator will ask for the initial temperature, mass, and specific heat capacity of the sample. The graphing calculator will use the following equation to create a graph of temperature (Y_1) versus the energy absorbed (X).

$$Y_1 = T + (X/(MC))$$

a. The graphing calculator equation is the same as the specific heat capacity equation shown above. Specify what each variable in the graphing calculator equation represents.

Execute "Chap10" on the PRGM menu and press ENTER to begin the program. Enter the values for the mass, specific heat capacity, and initial temperature (shown below), pressing ENTER after each one.

The calculator will provide a graph of the temperature versus the energy absorbed. (If the graph is not visible, press WINDOW and change the settings for the graph window so that Xmin is the lowest energy value required and Xmax is the highest value required, then press ENTER.)

Press TRACE, and use the arrow keys to trace along the curve. The *x* value corresponds to the absorbed energy in joules, and the *y* value corresponds to the temperature in degrees Celsius.

Determine the temperature of a 0.050 kg piece of aluminum foil (specific heat capacity equals 899 J/kg•°C) originally at 25°C that absorbs the following amounts of energy by heat:

b. 75 J
c. 225 J
d. 475 J
e. 825 J
f. If the initial temperature were 10°C instead of 25°C, how would the graph be different?

Press 2nd QUIT to stop graphing. Press ENTER to input a new value or CLEAR to end the program.

48. At what Fahrenheit temperature are the Kelvin and Fahrenheit temperatures numerically equal?

(See Sample Problem 10A.)

49. A 250 g aluminum cup holds and is in thermal equilibrium with 850 g of water at 83°C. The combination of cup and water is cooled uniformly so that the temperature decreases by 1.5°C per minute. At what rate is energy being removed?

50. A jar of tea is placed in sunlight until it reaches an equilibrium temperature of 32°C. In an attempt to cool the liquid, which has a mass of 180 g, 112 g of ice at 0°C is added. At the time at which the temperature of the tea is 15°C, determine the mass of the remaining ice in the jar. Assume the specific heat capacity of the tea to be that of pure liquid water.

Alternative Assessment

Performance assessment

1. According to legend, Archimedes determined whether the king's crown was pure gold by comparing its water displacement with the displacement of a piece of pure gold of equal mass. But this procedure is difficult to apply to very small objects. Design a method for determining whether a ring is pure gold using the concept of specific heat capacity. Present your plan to the class, and ask others to suggest improvements to your design. Discuss each suggestion's advantages and disadvantages.

2. The host of a cooking show on television claims that you can greatly reduce the baking time for potatoes by inserting a nail through each potato. Explain whether this advice has a scientific basis. Would this approach be more efficient than wrapping the potatoes in aluminum foil? List all arguments and discuss their strengths and weaknesses.

3. The graph of decreasing temperature versus time of a hot object is called its cooling curve. Design and perform an experiment to determine the cooling curve of water in containers of various materials and shapes. Draw cooling curves for each one. Which trends represent good insulation? Use your findings and graphs to design a lunch box that keeps food warm or cold.

Portfolio projects

4. Research the life and work of James Prescott Joule, who is best known for his apparatus demonstrating the equivalence of work and heat and the conservation of energy. Many scientists of the day initially did not accept Joule's conclusions. Research the reasoning behind their objections. Prepare a presentation for a class discussion either supporting the objections of Joule's critics or defending Joule's conclusion before England's Royal Academy of Sciences.

5. Get information on solar water heaters available where you live. How does each type work? Compare prices and operating expenses for solar water heaters versus gas water heaters. What are some of the other advantages and limitations of solar water heaters? Prepare an informative brochure for homeowners interested in this technology.

6. Research how scientists measure the temperature of the following: the sun, a flame, a volcano, outer space, liquid hydrogen, mice, and insects. Find out what instruments are used in each case and how they are calibrated to known temperatures. Using what you learn, prepare a chart or other presentation on the tools used to measure temperature and the limitations on their ranges.

CHAPTER 10
Laboratory Exercise

SPECIFIC HEAT CAPACITY

In this experiment, you will use calorimetry to identify various metals. In each trial, you will heat a sample of metal by placing it above a bath of water and bringing the water to a boil. When the sample is heated, you will place it in a calorimeter containing cold water. The water in the calorimeter will be warmed by the metal as the metal cools. According to the principle of energy conservation, the total amount of energy transferred out of the metal sample as it cools equals the energy transferred into the water and calorimeter as they are warmed. In this lab, you will use your measurements to determine the specific heat capacity and identity of each metal.

OBJECTIVES

- Measure temperature.
- Apply the specific heat capacity equation for calorimetry to calculate the specific heat capacity of a metal.
- Identify unknown metals by comparing their specific heat capacities with accepted values for specific heat capacities.

MATERIALS LIST

- ✔ 2 beakers
- ✔ samples of various metals
- ✔ hot plate
- ✔ metal calorimeter and stirring rod
- ✔ ice
- ✔ balance
- ✔ metal heating vessel with metal heating dipper
- ✔ small plastic dish

PROCEDURE

CBL AND SENSORS

- ✔ CBL
- ✔ graphing calculator with link cable
- ✔ 2 temperature probes

THERMOMETER

- ✔ hand-held magnifying lens
- ✔ 2 thermometers

SAFETY

- **When using a burner or hot plate, always wear goggles and an apron to protect your eyes and clothing. Tie back long hair, secure loose clothing, and remove loose jewelry. If your clothing catches on fire, walk to the emergency lab shower and use the shower to put out the fire.**
- **Never leave a hot plate unattended while it is turned on.**
- **If a thermometer breaks, notify the teacher immediately.**
- **Do not heat glassware that is broken, chipped, or cracked. Use tongs or a mitt to handle heated glassware and other equipment because it does not always look hot when it is hot. Allow all equipment to cool before storing it.**
- **Never put broken glass or ceramics in a regular waste container. Use a dustpan, brush, and heavy gloves to carefully pick up broken pieces and dispose of them in a container specifically provided for this purpose.**

PREPARATION

1. Determine whether you will be using the CBL and sensors procedure or the thermometers. Read the entire lab for the appropriate procedure, and plan what steps you will take. Plan efficiently. Make sure you know which steps can be performed while you are waiting for the water to heat.

2. Prepare a data table with four columns and eight rows in your lab notebook. In the first row, label the second through fourth columns *Trial 1, Trial 2,* and *Trial 3.* In the first column, label the second through eighth

rows *Sample number, Mass of metal, Mass of calorimeter cup and stirrer, Mass of water, Initial temperature of metal, Initial temperature of water and calorimeter,* and *Final temperature of metal, water, and calorimeter.*

3. In Appendix E, look up the specific heat capacity of the material the calorimeter is made of and record the information in the top left corner of your data table.

Thermometer procedure begins on page 395.

PROCEDURE

CBL AND SENSORS

Finding the specific heat capacity of a metal

4. Choose a location where you can set up the experiment away from the edge of the table and away from other groups. Make sure the hot plate is in the "off" position before you plug it in.

5. Fill a metal heating vessel with 200 mL of water and place it on the hot plate. Turn on the hot plate and adjust the heating controls to heat the water.

6. Set up the temperature probe, CBL, and graphing calculator as shown in **Figure 10-24.** Connect the CBL to the graphing calculator with the unit-to-unit link cable using the I/O ports located on each unit. Connect the first temperature probe to the CH1 port. Connect the second temperature probe to the CH2 port. Turn on the CBL and the graphing calculator. Start the program PHYSICS on the graphing calculator.

   a. Select option *SET UP PROBES* from the MAIN MENU. Enter 2 for the number of probes. Select the temperature probe from the list. Enter 1 for the channel number. Select the temperature probe from the list again, and enter 2 for the channel number.

   b. Select the *COLLECT DATA* option from the MAIN MENU. Select the *TRIGGER* option from the DATA COLLECTION menu.

7. Obtain about 100 g of the metal sample. First find the mass of the small plastic dish. Place the metal shot in the dish and determine the mass of the shot. Record the number and the mass of the sample in your data table. Place one temperature probe in the metal heating dipper, and carefully pour the sample into the metal heating dipper. Make sure the temperature probe is surrounded by the metal sample.

Figure 10-24
Step 5: Start heating the water before you set up the CBL and temperature probes. Never leave a hot plate unattended when it is turned on.
Step 7: Be very careful when pouring the metal sample in the dipper around the temperature probe.
Step 17: Begin taking temperature readings a few seconds before adding the sample to the calorimeter.
Step 19: Record the *highest* temperature reached by the water, sample, and calorimeter combination, not the final temperature.

8. Place the dipper with metal contents into the top of the heating vessel, as shown in **Figure 10-24.** Make sure the temperature probe leads do not touch the hot plate or any heated surface.

9. While the sample is heating, find the mass of the empty inner cup of the calorimeter and the stirring rod. Record the mass in your data table. Do not leave the hot plate unattended.

10. For the water in the calorimeter, you will need about 100 g of water that is a little colder than room temperature. Put the water in a beaker. Use the second CBL temperature probe to find room temperature. Look at the temperatures on the CBL and read the temperature reading for the second probe. Place the second sensor in the water to check the water's temperature. (Do not use water colder than 5°C below room temperature. You may need to use ice to get the initial temperature low enough, but make sure all the ice has melted before pouring the water into the calorimeter.)

11. Place the calorimeter and stirrer on the balance and carefully add 100 g of the water. Record the mass of the water in your data table. Replace the cup in its insulating shell, and cover.

12. When the CBL displays a constant temperature for several readings, press TRIGGER on the CBL to collect the temperature reading of the metal sample. Record the temperature reading in your data table as the metal's initial temperature. Select *CONTINUE* from the TRIGGER menu on the graphing calculator.

13. Carefully remove the first temperature probe from the dipper. Set the probe aside to cool.

14. Use the stirring rod to stir the water in the calorimeter. Place the second temperature probe in the calorimeter. When the CBL displays a constant temperature for several readings for the calorimeter water, press TRIGGER on the CBL to collect the temperature readings. Record the initial temperature of the water and calorimeter. Select *STOP* from the TRIGGER menu on the graphing calculator. Leave the probe in the calorimeter.

15. Select *COLLECT DATA* from the MAIN MENU on the graphing calculator. Select the *TIME GRAPH* option from the DATA COLLECTION menu. Enter 2.0 for the time between samples. Enter 99 for the number of samples. Check the values you entered, and then press ENTER. Press ENTER to continue. If you made a mistake entering the time values, select *MODIFY SETUP*, reenter the values, and continue.

16. From the TIME GRAPH menu, select *LIVE DISPLAY.* Enter 0 for *Ymin*, enter 100 for *Ymax*, and enter 5 for *Yscl*.

17. Press ENTER on the graphing calculator to begin collecting the temperature readings for the water in the calorimeter.

18. Quickly transfer the metal sample to the calorimeter of cold water and replace the cover. Use a mitt when handling the metal heating dipper. Use the stirring rod to gently agitate the sample and to stir the water in the calorimeter. If you are not doing any more trials, make sure the hot plate is turned off. Otherwise, make sure there is plenty of water in the heating vessel, and do not leave the hot plate unattended.

19. When the CBL displays DONE, use the arrow keys to trace the graph. Time in seconds is graphed on the *x*-axis, and the temperature readings are graphed on the *y*-axis. Record the highest temperature reading from the CBL in your data table.

20. Press ENTER on the graphing calculator. On the REPEAT? menu, select *NO.* If you are going to perform another trial, select the *COLLECT DATA* option from the MAIN MENU. Select *TRIGGER* from the DATA COLLECTION menu.

21. If time permits, make additional trials with other samples. Record data for all trials in your data table.

Analysis and Interpretation begins on page 396.

PROCEDURE

THERMOMETER

Finding the specific heat capacity of a metal

4. Choose a location where you can set up the experiment away from the edge of the table and from other groups. Make sure the hot plate is in the "off" position before you plug it in.

5. Fill a metal heating vessel with 200 mL of water and place it on the hot plate, as shown in **Figure 10-25.** Turn on the hot plate and adjust the heating control to heat the water.

6. Measure out about 100 g of the metal sample. Record the number of the metal sample in your data table. Hold the thermometer in the metal heating dipper, and very carefully pour the sample into the metal heating dipper. Make sure the bulb of the thermometer is surrounded by the metal. Place the dipper with metal contents into the heating vessel. Hold the thermometer while the sample is heating.

7. While the sample is heating, determine the mass of the stirring rod and empty inner cup of the calorimeter. Record the mass in your data table. Do not leave the hot plate unattended.

8. Use the second thermometer to measure room temperature. For the water in the calorimeter, you will need about 100 g of water that is a little colder than room temperature. Put the water in a beaker. Place the thermometer in the water to check the temperature of the water. (Do not use water colder than 5°C below room temperature. You may need to use ice to get the initial temperature low enough, but make sure all the ice has melted before pouring the water into the calorimeter.)

9. Place the calorimeter and stirrer on the balance, and carefully add 100 g of the water. Record the mass of the water in your data table. Replace the cup in its insulating shell, and cover.

10. Use the thermometer to measure the temperature of the sample when the water is boiling and the sample reaches a constant temperature. Record this temperature as the initial temperature of the metal sample. (Note: When making temperature readings, take care not to touch the hot plate and the water.) Use the hand-held magnifying lens to estimate to the nearest 0.5°C. Make sure that the

Figure 10-25
Step 5: Start heating the water before you begin the rest of the lab. Never leave a hot plate unattended when it is turned on.
Step 6: Be very careful when pouring the metal sample in the dipper around the thermometer. Make sure the thermometer bulb is surrounded by the metal sample.
Step 12: Begin taking temperature readings a few seconds before adding the sample to the calorimeter.
Step 15: Record the *highest* temperature reached by the water, sample, and calorimeter combination.

thermometer bulb is completely surrounded by the metal sample, and keep your line of sight at a right angle to the stem of the thermometer. Reading the thermometer at an angle will cause considerable errors in your measurements. Carefully remove the thermometer and set it aside in a secure place.

11. Use the stirring rod to gently stir the water in the calorimeter. **Do not use the thermometer to stir the water.**

12. Place the second thermometer in the covered calorimeter. Measure the temperature of the water in the calorimeter to the nearest 0.1°C. Record this temperature in your data table as the initial temperature of the water and calorimeter.

13. Quickly transfer the sample to the cold water in the calorimeter and replace the cover. Use a mitt when handling the metal heating dipper. If you are not doing any more trials, make sure the hot plate is turned off. Otherwise, make sure there is plenty of water in the heating vessel, and do not leave the hot plate unattended.

14. Use the stirring rod to gently agitate the sample and stir the water in the calorimeter. **Do not use the thermometer to stir the water.**

15. Take readings every 5.0 s until five consecutive readings are the same. Record the highest reading in your data table.

16. If time permits, make additional trials with other metals. Record the data for all trials in your data table.

17. Clean up your work area. Put equipment away safely so that it is ready to be used again.

ANALYSIS AND INTERPRETATION

Calculations and data analysis

1. **Organizing data** For each trial, calculate the temperature change of the water and calorimeter.

2. **Analyzing data** Use your data for each trial.
   a. Calculate the energy transferred to the calorimeter cup and stirring rod as heat, using the value for the specific heat capacity you found in step 3.
   b. Calculate the energy transferred to the water as heat.

3. **Applying ideas** Calculate the total energy transferred as heat into the water and the calorimeter.

4. **Analyzing results** For each trial, find the temperature change of the sample and calculate the specific heat capacity of the sample.

Conclusions

5. **Evaluating data** Use the accepted values for the specific heat capacities of various metals in **Table 10-4** on page 372, to determine what metal each sample is made of.

6. Evaluating results Calculate the absolute and relative errors of the experimental values. Check with your teacher to see if you have correctly identified the metals.

a. Use the following equation to compute the absolute error:

$$\text{absolute error} = |\text{experimental} - \text{accepted}|$$

b. Use the following equation to compute the relative error:

$$\text{relative error} = \frac{(\text{experimental} - \text{accepted})}{\text{accepted}}$$

7. Evaluating methods Explain why the energy transferred as heat into the calorimeter and the water is equal to the energy transferred as heat from the metal sample.

8. Evaluating methods Explain why it is important to calculate the temperature change using the highest temperature as the final temperature, rather than the last temperature recorded.

9. Evaluating methods Why should the water be a few degrees colder than room temperature when the initial temperature is taken?

10. Applying conclusions How would your results be affected if the initial temperature of the water in the calorimeter were 50°C instead of slightly cooler than room temperature?

11. Relating ideas How is the temperature change of the calorimeter and the water within the calorimeter affected by the specific heat capacity of the metal? Did a metal with a high specific heat capacity raise the temperature of the water and the calorimeter more or less than a metal with a low specific heat capacity?

12. Building models An environmentally conscious engineering team wants to design tea kettles out of a metal that will allow the water to reach its boiling point using the least possible amount of energy from a range or other heating source. Using the values for specific heat capacity in **Table 10-4** on page 372, choose a material that would work well, considering only the implications of transfer of energy as heat. Explain how the specific heat capacity of water will affect the operation of the tea kettle.

Extensions

13. Evaluating methods What is the purpose of the outer shell of the calorimeter and the insulating ring in this experiment?

14. Designing experiments If there is time and your teacher approves, design an experiment to measure the specific heat capacity of the calorimeter. Compare this measured value with the accepted value from **Table 10-4** on page 372. Are they the same? If not, how would using the experimental value affect your results in this lab?

Climatic Warming

Scientists typically devise solutions to problems and then test the solution to determine if it indeed solves the problem. But sometimes the problem is only suggested by the evidence, and there are no chances to test the solutions. A current example of such a problem is climatic warming.

Data recorded from various locations around the world over the past century indicate that the average atmospheric temperature is 0.5°C higher now than it was 100 years ago. Although this sounds like a small amount, such an increase can have pronounced effects. Increased temperatures may eventually cause the ice in polar regions to melt, causing ocean levels to increase, which in turn may flood some coastal areas.

Small changes in temperature can also affect living organisms. Most trees can tolerate only about a 1°C increase in average temperature. If a tree does not reproduce often or easily enough to “migrate” through successive generations to a cooler location, it can become extinct in that region. Any organisms dependent on that type of tree also will suffer.

But such disasters depend on whether global temperatures continue to increase. Historical studies indicate that some short-term fluctuations in climate are natural, like the “little ice age” of the seventeenth century. If the current warming trend is part of a natural cycle, the dire predictions may be overstated or wrong.

Even if the warming is continuous, climatic systems are very complex and involve many unexpected factors. For example, if polar ice melts, a sudden increase in humidity may result in snow in polar areas. This could counter the melting, thus causing ocean levels to remain stable.

Greenhouse Gases

Most of the current attention and concern about climatic warming has been focused on the increase in the amount of “greenhouse gases,” primarily carbon dioxide and methane, in the atmosphere. Molecules of these gases absorb energy that is radiated from Earth’s surface, causing their temperature to rise. These molecules then release energy as heat, causing the atmosphere to be warmer than it would be without these gases.

While carbon dioxide and methane are natural components of the air, their levels have increased rapidly during the last hundred years. This has been determined by analyzing air trapped in the ice layers of Greenland. Deeper sections of the ice contain air from earlier times. During the last ice age, there were about 185 ppm of carbon dioxide, CO_2, in the air, but the concentration from 130 years ago was slightly below 300 ppm. Today, the levels are 350 ppm, an increase that can be accounted for by the increase in combustion reactions, primarily from coal and petroleum burning, and by the decrease in CO_2-consuming trees through deforestation.

But does the well-documented increase in greenhouse gas concentrations enable detailed predictions? Atmospheric physicists have greatly improved their models in recent years, and they are able to correctly predict past ice ages and account for the energy-absorbing qualities of oceans. But such models remain oversimplified, partly because of a lack of detailed long-term data. In addition, the impact of many variables, such as fluctuations in solar energy output and volcanic processes, are poorly understood and cannot be factored into predictions. To take all factors into account would require more-complex models and more-sophisticated supercomputers than are currently available. As a result, many question whether meaningful decisions and planning can occur.

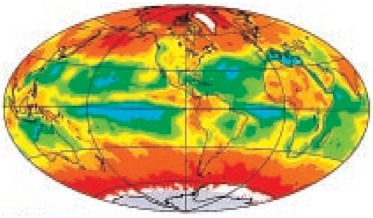
Risk of Action and Inaction

The evidence for climatic warming is suggestive but not conclusive. What should be done? Basically, there are two choices: either do something or do nothing.

The risks of doing nothing are that the situation may worsen. But it is also possible that waiting for better evidence will allow for a greater consensus among the world's nations about how to solve the problem efficiently. Convincing the world's population that action taken now will have the desired benefit decades from now will not be easy.

Acting now also involves risks. Gas and coal could be rationed or taxed to limit consumption. The development of existing energy-efficient technologies, such as low-power electric lights and more efficient motors and engines, could cut use of coal and gasoline in half. However, the economic effects could be as severe as those resulting from climatic warming.

But none of these options can guarantee results. Even if the trend toward climatic warming stops, it will be hard to prove whether this was due to human reduction in greenhouse gases, to natural cyclic patterns, or to other causes.



This map shows the reflecting properties of the Earth's surface. Regions colored in blue or green absorb much of the energy striking them.

Researching the Issue

1. Carbon dioxide levels in the atmosphere have varied during Earth's history. Research the roles of volcanoes, plants, and limestone formation, and determine whether these processes have any bearing on the current increase in CO_2 concentrations. Can you think of any practical means of using these processes to reduce CO_2 concentrations? What would be the advantages and disadvantages?

2. Find out what technological developments have been suggested for slowing climatic warming. Can they be easily implemented? What are the drawbacks of these methods?

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