Chapter

12

Thermal Energy

What You'll Learn

- You will learn how temperature relates to the potential and kinetic energies of atoms and molecules.
- You will distinguish heat from work.
- You will calculate heat transfer and the absorption of thermal energy.

Why It's Important

Thermal energy is vital for living creatures, chemical reactions, and the working of engines.

Solar Energy A strategy used to produce electric power from sunlight concentrates the light with many mirrors onto one collector that becomes very hot. The energy collected at a high temperature is then used to drive an engine, which turns an electric generator.

Think About This ▶

What forms of energy does light from the Sun take in the process of converting solar energy into useful work through an engine?

Physics Online

physicspp.com

CORBIS

LAUNCH Lab

What happens when you provide thermal energy by holding a glass of water?

Question

What happens to the temperature of water when you hold a glass of water in your hand?

Procedure

1. You will need to use a 250-mL beaker and 150 mL of water.
2. Fill the beaker with the 150 mL of water.
3. Record the initial temperature of the water by holding a thermometer in the water in the beaker. Note that the bulb end of the thermometer must not touch the bottom or sides of the beaker, nor should it touch a table or your hands.
4. Remove the thermometer and hold the beaker of water for 2 min by cupping it with both hands, as shown in the figure.
5. Have your lab partner record the final temperature of the water by placing the thermometer in the beaker. Be sure that the bulb end of the thermometer is not touching the bottom or sides of the beaker.

Analysis

Calculate the change in temperature of the water. If you had more water in the beaker, would it affect the change in temperature?

Critical Thinking Explain what caused the water temperature to change.

12.1 Temperature and Thermal Energy

The study of heat transformations into other forms of energy, called thermodynamics, began with the eighteenth-century engineers who built the first steam engines. These steam engines were used to power trains, factories, and water pumps for coal mines, and thus they contributed greatly to the Industrial Revolution in Europe and in the United States. In learning to design more efficient engines, the engineers developed new concepts about how heat is related to useful work. Although the study of thermodynamics began in the eighteenth century, it was not until around 1900 that the concepts of thermodynamics were linked to the motions of atoms and molecules in solids, liquids, and gases.

Today, the concepts of thermodynamics are widely used in various applications that involve heat and temperature. Engineers use the laws of thermodynamics to continually develop higher performance refrigerators, automobile engines, aircraft engines, and numerous other machines.

▶ **Objectives**

- **Describe** thermal energy and compare it to potential and kinetic energies.
- **Distinguish** temperature from thermal energy.
- **Define** specific heat and **calculate** heat transfer.

▶ **Vocabulary**

conduction
thermal equilibrium
heat
convection
radiation
specific heat

Horizons Companies

Thermal Energy

You already have studied how objects collide and trade kinetic energies. For example, the many molecules present in a gas have linear and rotational kinetic energies. The molecules also may have potential energy in their vibrations and bending. The gas molecules collide with each other and with the walls of their container, transferring energy among each other in the process. There are numerous molecules moving freely in a gas, resulting in many collisions. Therefore, it is convenient to discuss the total energy of the molecules and the average energy per molecule. The total energy of the molecules is called thermal energy, and the average energy per molecule is related to the temperature of the gas.

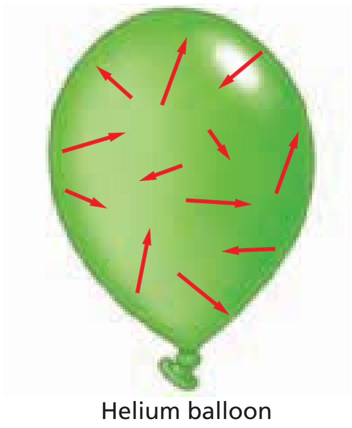


Figure 12-1 Helium atoms in a balloon collide with the rubber wall and cause the balloon to expand.

CLICK HERE

CLICK HERE

CLICK HERE

Hot objects What makes an object hot? When you fill up a balloon with helium, the rubber in the balloon is stretched by the repeated pounding from helium atoms. Each of the billions of helium atoms in the balloon collides with the rubber wall, bounces back, and hits the other side of the balloon, as shown in **Figure 12-1.** If you put a balloon in sunlight, you might notice that the balloon gets slightly larger. The energy from the Sun makes each of the gas atoms move faster and bounce off the rubber walls of the balloon more often. Each atomic collision with the balloon wall puts a greater force on the balloon and stretches the rubber. Thus, the balloon expands.

On the other hand, if you refrigerate a balloon, you will find that it shrinks slightly. Lowering the temperature slows the movement of the helium atoms. Hence, their collisions do not transfer enough momentum to stretch the balloon quite as much. Even though the balloon contains the same number of atoms, the balloon shrinks.

Solids The atoms in solids also have kinetic energy, but they are unable to move freely as gas atoms do. One way to illustrate the molecular structure of a solid is to picture a number of atoms that are connected to each other by springs. Because of the springs, the atoms bounce back and forth, with some bouncing more than others. Each atom has some kinetic energy and some potential energy from the springs that are attached to it. If a solid has *N* number of atoms, then the total thermal energy in the solid is equal to the average kinetic and potential energy per atom times *N*.

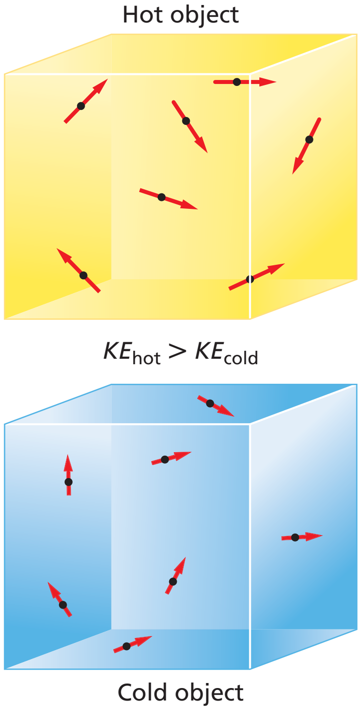


Figure 12-2 Particles in a hot object have greater kinetic and potential energies than particles in a cold object do.

Thermal Energy and Temperature

According to the previous discussion of gases and solids, a hot object has more thermal energy than a similar cold object, as shown in **Figure 12-2.** This means that, as a whole, the particles in a hot object have greater thermal energy than the particles in a cold object. This does not mean that all the particles in an object have exactly the same amount of energy; they have a wide range of energies. However, the average energy of the particles in a hot object is higher than the average energy of the particles in a cold object. To understand this, consider the heights of students in a twelfth-grade class. Although the students' heights vary, you can calculate the average height of the students in the class. This average is likely to be greater than the average height of students in a ninth-grade class, even though some ninth-grade students may be taller than some twelfth-grade students.

Temperature Temperature depends only on the average kinetic energy of the particles in the object. Because temperature depends on average kinetic energy, it does not depend on the number of atoms in an object. To understand this, consider two blocks of steel. The first block has a mass of 1 kg, and the second block has a mass of 2 kg. If the 1-kg block is at the same temperature as the 2-kg block, the average kinetic energy of the particles in each block is the same. However, the 2-kg block has twice the mass of the 1-kg block. Hence, the 2-kg block has twice the amount of particles as the 1-kg block. Thus, the total amount of kinetic energy of the particles in the 2-kg block is twice that of the 1-kg mass. Total kinetic energy is divided by the total number of particles in an object to calculate its average kinetic energy. Therefore, the thermal energy in an object is proportional to the number of particles in it. Temperature, however, is not dependent on the number of particles in an object.

Equilibrium and Thermometry

How do you measure your body temperature? For example, if you suspect that you have a fever, you might place a thermometer in your mouth and wait for a few minutes before checking the thermometer for your temperature reading. The microscopic process involved in measuring temperature involves collisions and energy transfers between the thermometer and your body. Your body is hot compared to the thermometer, which means that the particles in your body have greater thermal energy and are moving faster than the particles in the thermometer. When the cold glass tube of the thermometer touches your skin, which is warmer than the glass, the faster-moving particles in your skin collide with the slower-moving particles in the glass. Energy is then transferred from your skin to the glass particles by the process of **conduction,** which is the transfer of kinetic energy when particles collide. The thermal energy of the particles that make up the thermometer increases, while at the same time, the thermal energy of the particles in your skin decreases.

Thermal equilibrium As the particles in the glass gain more energy, they begin to give some of their energy back to the particles in your body. At some point, the rate of transfer of energy between the glass and your body becomes equal, and your body and the thermometer are then at the same temperature. At this point, your body and the thermometer are said to have reached **thermal equilibrium,** the state in which the rate of energy flow between two objects is equal and the objects are at the same temperature, as shown in **Figure 12-3.**

The operation of a thermometer depends on some property, such as volume, which changes with temperature. Many household thermometers contain colored alcohol that expands when heated and rises in a narrow tube. The hotter the thermometer, the more the alcohol expands and the higher it rises in the tube. In liquid-crystal thermometers, such as the one shown in **Figure 12-4,** a set of different kinds of liquid crystals is used. Each crystal's molecules rearrange at a specific temperature, which causes the color of the crystal to change and indicates the temperature by color. Medical thermometers and the thermometers that monitor automobile engines use very small, temperature-sensitive electronic circuits to take rapid measurements.

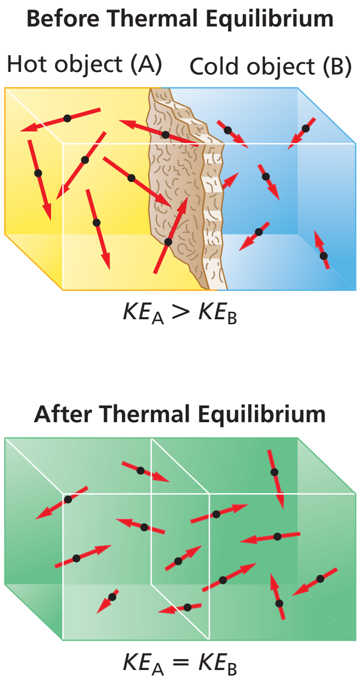


Figure 12-3 Thermal energy is transferred from a hot object to a cold object. When thermal equilibrium is reached, the transfer of energy between objects is equal.

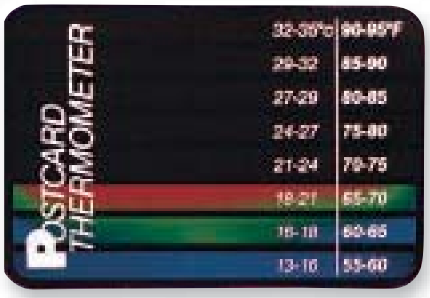


Figure 12-4 Thermometers use a change in physical properties to measure temperature. A liquid-crystal thermometer changes color with a temperature change.

Tom Pantages

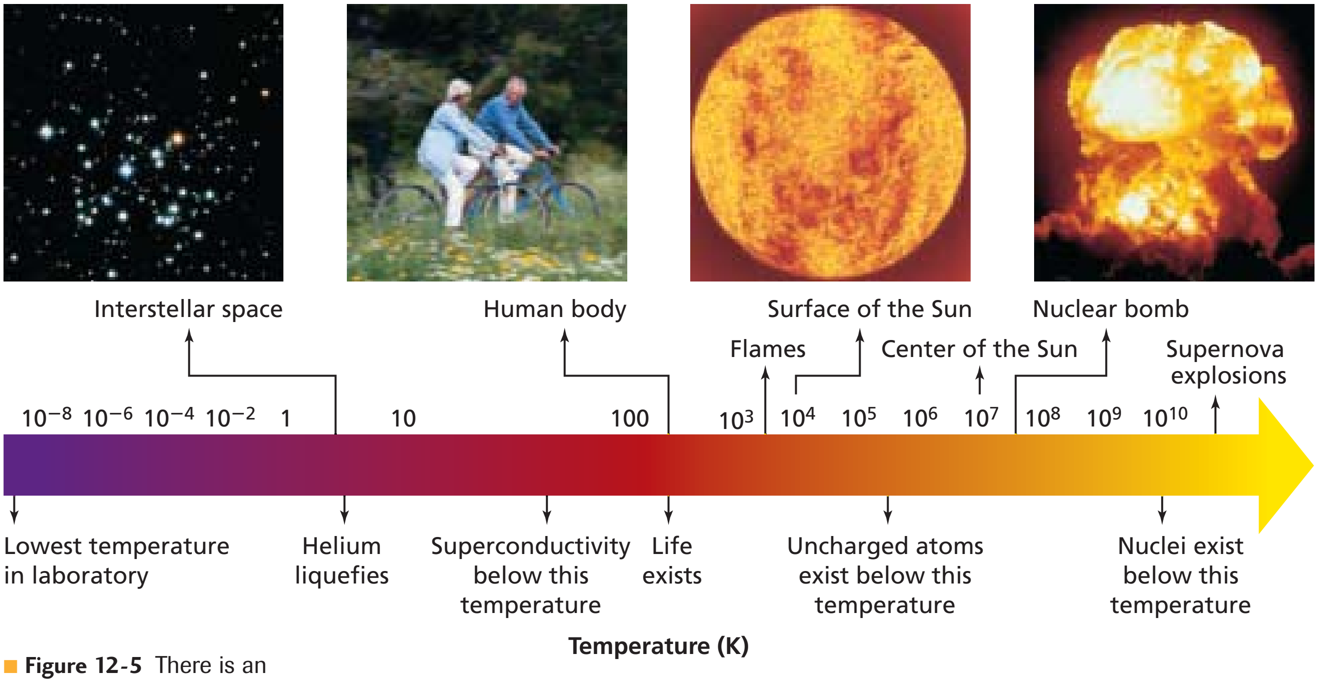


Figure 12-5 There is an extremely wide range of temperatures throughout the universe. Note that the scale has been expanded in areas of particular interest.

Temperature Scales: Celsius and Kelvin

Over the years, scientists developed temperature scales so that they could compare their measurements with those of other scientists. A scale based on the properties of water was devised in 1741 by Swedish astronomer and physicist Anders Celsius. On this scale, now called the Celsius scale, the freezing point of pure water is defined to be 0°C. The boiling point of pure water at sea level is defined to be 100°C.

Temperature limits The wide range of temperatures present in the universe is shown in **Figure 12-5.** Temperatures do not appear to have an upper limit. The interior of the Sun is at least 1.5×10^{7}°C. Temperatures do, however, have a lower limit. Generally, materials contract as they cool. If an ideal gas, such as the helium in a balloon is cooled, it contracts in such a way that it occupies a volume that is only the size of the helium atoms at −273.15°C. At this temperature, all the thermal energy that can be removed has been removed from the gas. It is impossible to reduce the temperature any further. Therefore, there can be no temperature lower than −273.15°C, which is called absolute zero.

The Celsius scale is useful for day-to-day measurements of temperature. It is not conducive for working on science and engineering problems, however, because it has negative temperatures. Negative temperatures suggest a molecule could have negative kinetic energy, which is not possible because kinetic energy is always positive. The solution to this issue is to use a temperature scale based on absolute zero.

The zero point of the Kelvin scale is defined to be absolute zero. On the Kelvin scale, the freezing point of water (0°C) is about 273 K and the boiling point of water is about 373 K. Each interval on this scale, called a kelvin, is equal to 1°C. Thus, $T_C + 273 = T_K$. **Figure 12-6** shows representative temperatures on the three most-common scales: Fahrenheit, Celsius, and Kelvin.

Very cold temperatures are reached by liquefying gases. Helium liquefies at 4.2 K, or −269°C. Even colder temperatures can be reached by making use of special properties of solids, helium isotopes, and atoms and lasers.

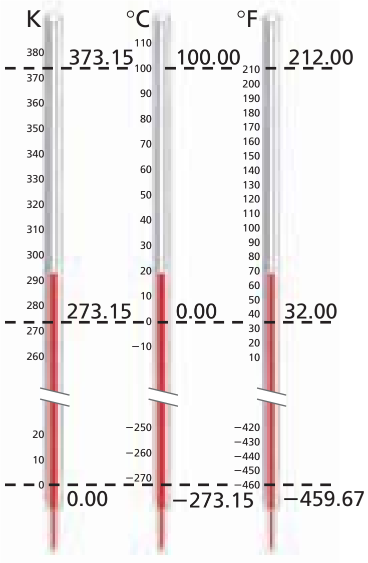


Figure 12-6 The three most-common temperature scales are Kelvin, Celsius, and Fahrenheit.

(cl)Getty Images, (r)FPG/Getty Images, (others)CORBIS

PRACTICE Problems **Additional Problems, Appendix B**

1. Convert the following Kelvin temperatures to Celsius temperatures.
   - **a.** 115 K
   - **b.** 172 K
   - **c.** 125 K
   - **d.** 402 K
   - **e.** 425 K
   - **f.** 212 K
2. Find the Celsius and Kelvin temperatures for the following.
   - **a.** room temperature
   - **b.** a typical refrigerator
   - **c.** a hot summer day in North Carolina
   - **d.** a winter night in Minnesota

Heat and the Flow of Thermal Energy

When two objects come in contact with each other, they transfer energy. This energy that is transferred between the objects is called **heat.** Heat is described as the energy that always flows from the hotter object to the cooler object. Left to itself heat never flows from a colder object to a hotter object. The symbol *Q* is used to represent an amount of heat, which, like other forms of energy, is measured in joules. If *Q* has a negative value, heat has left the object; if *Q* has a positive value, heat has been absorbed by the object.

Conduction If you place one end of a metal rod in a flame, the hot gas particles in the flame conduct heat to the rod. The other end of the rod also becomes warm within a short period of time. Heat is conducted because the particles in the rod are in direct contact with each other.

Convection Thermal energy transfer can occur even if the particles in an object are not in direct contact with each other. Have you ever looked into a pot of water just about to boil? The water at the bottom of the pot is heated by conduction and rises to the top, while the colder water at the top sinks to the bottom. Heat flows between the rising hot water and the descending cold water. This motion of fluid in a liquid or gas caused by temperature differences is called **convection.** Atmospheric turbulence is caused by convection of gases in the atmosphere. Thunderstorms are excellent examples of large-scale atmospheric convection. Ocean currents that cause changes in weather patterns also result from convection.

Radiation The third method of thermal transfer, unlike the first two, does not depend on the presence of matter. The Sun warms Earth from over 150 million km away via **radiation,** which is the transfer of energy by electromagnetic waves. These waves carry the energy from the hot Sun through the vacuum of space to the much cooler Earth.

Meteorology Connection

APPLYING PHYSICS

► **Steam Heating** In a steam heating system of a building, water is turned into steam in a boiler located in a maintenance area or the basement. The steam then flows through insulated pipes to each room in the building. In the radiator, the steam is condensed as liquid water and then flows back through pipes to the boiler to be revaporized. The hot steam physically carries the heat from the boiler, and that energy is released when the steam condenses in the radiator. Some disadvantages of steam heating are that it requires expensive boilers and pipes must contain steam under pressure. ◄

Specific Heat

Some objects are easier to heat than others. On a bright summer day, the Sun warms the sand on a beach and the ocean water. However, the sand on the beach gets quite hot, while the ocean water stays relatively cool. When heat flows into an object, its thermal energy and temperature increase. The amount of the increase in temperature depends on the size of the object and on the material from which the object is made.

Table 12-1

Specific Heat of Common Substances

Material	Specific Heat (J/kg·K)	Material	Specific Heat (J/kg·K)
Aluminum	897	Lead	130
Brass	376	Methanol	2450
Carbon	710	Silver	235
Copper	385	Steam	2020
Glass	840	Water	4180
Ice	2060	Zinc	388
Iron	450		

The **specific heat** of a material is the amount of energy that must be added to the material to raise the temperature of a unit mass by one temperature unit. In SI units, specific heat, represented by *C*, is measured in J/kg·K. **Table 12-1** provides values of specific heat for some common substances. For example, 897 J must be added to 1 kg of aluminum to raise its temperature by 1 K. The specific heat of aluminum is therefore 897 J/kg·K.

The heat gained or lost by an object as its temperature changes depends on the mass, the change in temperature, and the specific heat of the substance. By using the following equation, you can calculate the amount of heat, *Q*, that must be transferred to change the temperature of an object.

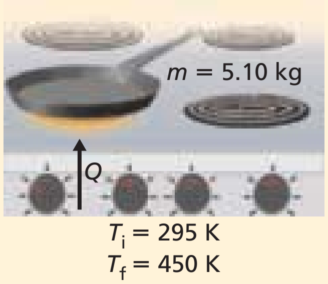
Heat Transfer $Q = mC\Delta T = mC(T_f - T_i)$

Heat transfer is equal to the mass of an object times the specific heat of the object times the difference between the final and initial temperatures.

Liquid water has a high specific heat compared to the other substance in Table 12-1. When the temperature of 10.0 kg of water is increased by 5.0 K, the heat absorbed is $Q = (10.0\ \text{kg})(4180\ \text{J/kg·K})(5.0\ \text{K}) = 2.1\times10^5$ J. Remember that the temperature interval for kelvins is the same as that for Celsius degrees. For this reason, you can calculate ΔT in kelvins or in degrees Celsius.

▶ EXAMPLE Problem 1

Heat Transfer A 5.10-kg cast-iron skillet is heated on the stove from 295 K to 450 K. How much heat had to be transferred to the iron?



1 Analyze and Sketch the Problem

- Sketch the flow of heat into the skillet from the stove top.

Known:		Unknown:
$m = 5.10$ kg	$C = 450$ J/kg·K	$Q = ?$
$T_i = 295$ K	$T_f = 450$ K	

2 Solve for the Unknown

$$Q = mC(T_f - T_i)$$
$$= (5.10\ \text{kg})(450\ \text{J/kg·K})(450\ \text{K} - 295\ \text{K})$$
$$= 3.6\times10^5\ \text{J}$$

Substitute m = 5.10 kg, C = 450 J/kg·K, T_f = 450 K, T_i = 295 K

3 Evaluate the Answer

- **Are the units correct?** Heat is measured in J.
- **Does the sign make sense?** Temperature increased, so *Q* is positive.

Math Handbook
Order of Operations
page 843

PRACTICE **Problems**

- Additional Problems, Appendix B
- Solutions to Selected Problems, Appendix C

3. When you turn on the hot water to wash dishes, the water pipes have to heat up. How much heat is absorbed by a copper water pipe with a mass of 2.3 kg when its temperature is raised from 20.0°C to 80.0°C?

4. The cooling system of a car engine contains 20.0 L of water (1 L of water has a mass of 1 kg).

  a. What is the change in the temperature of the water if the engine operates until 836.0 kJ of heat is added?

  b. Suppose that it is winter, and the car's cooling system is filled with methanol. The density of methanol is 0.80 g/cm^3. What would be the increase in temperature of the methanol if it absorbed 836.0 kJ of heat?

  c. Which is the better coolant, water or methanol? Explain.

5. Electric power companies sell electricity by the kWh, where 1 kWh $= 3.6\times10^6$ J. Suppose that it costs $0.15 per kWh to run an electric water heater in your neighborhood. How much does it cost to heat 75 kg of water from 15°C to 43°C to fill a bathtub?

Calorimetry: Measuring Specific Heat

A simple calorimeter, such as the one shown in **Figure 12-7,** is a device used to measure changes in thermal energy. A calorimeter is carefully insulated so that heat transfer to the external world is kept to a minimum. A measured mass of a substance that has been heated to a high temperature is placed in the calorimeter. The calorimeter also contains a known mass of cold water at a measured temperature. The heat released by the substance is transferred to the cooler water. The change in thermal energy of the substance is calculated using the resulting increase in the water temperature. More sophisticated types of calorimeters are used to measure chemical reactions and the energy content of various foods.

The operation of a calorimeter depends on the conservation of energy in an isolated, closed system. Energy can neither enter nor leave this system. As a result, if the energy of one part of the system increases, the energy of another part of the system must decrease by the same amount. Consider a system composed of two blocks of metal, block A and block B, shown in **Figure 12-8a** on the next page. The total energy of the system is constant, as represented by the following equation.

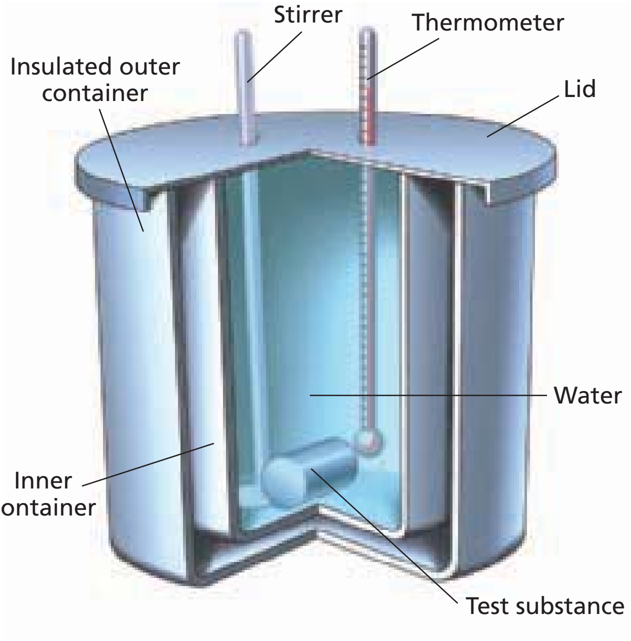


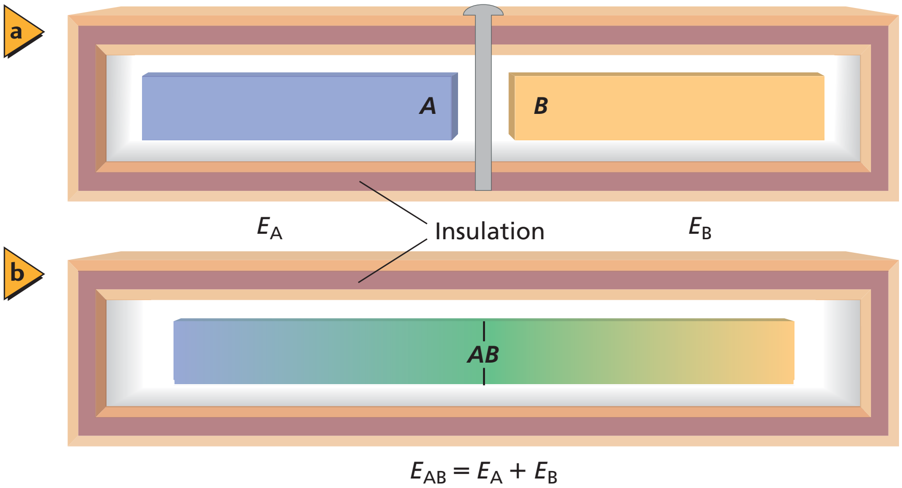
Figure 12-7 A calorimeter provides an isolated, closed system in which to measure energy transfer.

CLICK HERE

Conservation of Energy $E_A + E_B = \text{constant}$

In an isolated, closed system, the thermal energy of object A plus the thermal energy of object B is constant.

Figure 12-8 A system is composed of two model blocks at different temperatures that initially are separated **(a).** When the blocks are brought together, heat flows from the hotter block to the colder block **(b).** Total energy remains constant.



Suppose that the two blocks initially are separated but can be placed in contact with each other. If the thermal energy of block A changes by an amount ΔE_A, then the change in thermal energy of block B, ΔE_B, must be related by the equation, $\Delta E_A + \Delta E_B = 0$. Thus, $\Delta E_A = -\Delta E_B$. The change in energy of one block is positive, while the change in energy of the other block is negative. For the block whose thermal energy change is positive, the temperature of the block rises. For the block whose thermal energy change is negative, the temperature falls.

Assume that the initial temperatures of the two blocks are different. When the blocks are brought together, heat flows from the hotter block to the colder block, as shown in **Figure 12-8b.** The heat flow continues until the blocks are in thermal equilibrium, which is when the blocks have the same temperature.

In an isolated, closed system, the change in thermal energy is equal to the heat transferred because no work is done. Therefore, the change in energy for each block can be expressed by the following equation:

$$\Delta E = Q = mC\Delta T$$

The increase in thermal energy of block A is equal to the decrease in thermal energy of block B. Thus, the following relationship is true:

$$m_A C_A \Delta T_A + m_B C_B \Delta T_B = 0$$

The change in temperature is the difference between the final and initial temperatures; that is, $\Delta T = T_f - T_i$. If the temperature of a block increases, $T_f > T_i$, and ΔT is positive. If the temperature of the block decreases, $T_f < T_i$, and ΔT is negative. The final temperatures of the two blocks are equal. The following is the equation for the transfer of energy:

$$m_A C_A (T_f - T_A) + m_B C_B (T_f - T_B) = 0$$

To solve for T_f, expand the equation.

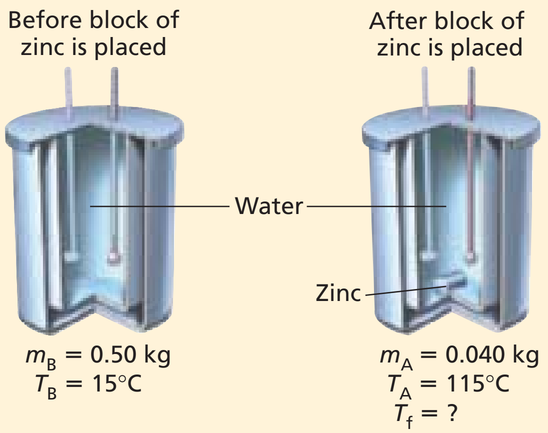
$$m_A C_A T_f - m_A C_A T_A + m_B C_B T_f - m_B C_B T_B = 0$$

$$T_f(m_A C_A + m_B C_B) = m_A C_A T_A + m_B C_B T_B$$

$$T_f = \frac{m_A C_A T_A + m_B C_B T_B}{m_A C_A + m_B C_B}$$

EXAMPLE Problem 2

Transferring Heat in a Calorimeter A calorimeter contains 0.50 kg of water at 15°C. A 0.040-kg block of zinc at 115°C is placed in the water. What is the final temperature of the system?



1 Analyze and Sketch the Problem

- Let zinc be sample A and water be sample B.
- Sketch the transfer of heat from the hotter zinc to the cooler water.

Known:	**Unknown:**
$m_A = 0.040$ kg	$T_f = ?$
$C_A = 388$ J/kg·°C	
$T_A = 115$°C	
$m_B = 0.50$ kg	
$C_B = 4180$ J/kg·°C	
$T_B = 15.0$°C	

Math Handbook
Operations with Significant Digits
pages 835–836

2 Solve for the Unknown

Determine the final temperature using the following equation.

$$T_f = \frac{m_A C_A T_A + m_B C_B T_B}{m_A C_A + m_B C_B}$$

$$= \frac{(0.040\text{ kg})(388\text{ J/kg·°C})(115\text{°C}) + (0.50\text{ kg})(4180\text{ J/kg·°C})(15.0\text{°C})}{(0.040\text{ kg})(388\text{ J/kg·°C}) + (0.50\text{ kg})(4180\text{ J/kg·°C})}$$

Substitute m_A = 0.040 kg, C_A = 388 J/kg·°C, T_A = 115°C, m_B = 0.50 kg, C_B = 4180 J/kg·°C, T_B = 15°C

$$= 16\text{°C}$$

3 Evaluate the Answer

- **Are the units correct?** Temperature is measured in Celsius.
- **Is the magnitude realistic?** The answer is between the initial temperatures of the two samples, as is expected when using a calorimeter.

PRACTICE Problems

• Additional Problems, Appendix B
• Solutions to Selected Problems, Appendix C

6. A 2.00×10^2-g sample of water at 80.0°C is mixed with 2.00×10^2 g of water at 10.0°C. Assume that there is no heat loss to the surroundings. What is the final temperature of the mixture?

7. A 4.00×10^2-g sample of methanol at 16.0°C is mixed with 4.00×10^2 g of water at 85.0°C. Assume that there is no heat loss to the surroundings. What is the final temperature of the mixture?

8. Three metal fishing weights, each with a mass of 1.00×10^2 g and at a temperature of 100.0°C, are placed in 1.00×10^2 g of water at 35.0°C. The final temperature of the mixture is 45.0°C. What is the specific heat of the metal in the weights?

9. A 1.00×10^2-g aluminum block at 100.0°C is placed in 1.00×10^2 g of water at 10.0°C. The final temperature of the mixture is 25.0°C. What is the specific heat of the aluminum?

Figure 12-9 A lizard regulates its body temperature by hiding under a rock when the atmosphere is hot **(a)** and sunbathing when the atmosphere gets cold **(b).**

Biology Connection

Animals can be divided into two groups based on their body temperatures. Most are cold-blooded animals whose body temperatures depend on the environment. The others are warm-blooded animals whose body temperatures are controlled internally. That is, a warm-blooded animal's body temperature remains stable regardless of the temperature of the environment. In contrast, when the temperature of the environment is high, the body temperature of a cold-blooded animal also becomes high. A cold-blooded animal, such as the lizard shown in **Figure 12-9,** regulates this heat flow by hiding under a rock or crevice, thereby reducing its body temperature. Humans are warm-blooded and have a body temperature of about 37°C. To regulate its body temperature, a warm-blooded animal increases or decreases the level of its metabolic processes. Thus, a warm-blooded animal may hibernate in winter and reduce its body temperature to approach the freezing point of water.

12.1 Section Review

10. Temperature Make the following conversions.
- **a.** 5°C to kelvins
- **b.** 34 K to degrees Celsius
- **c.** 212°C to kelvins
- **d.** 316 K to degrees Celsius

11. Conversions Convert the following Celsius temperatures to Kelvin temperatures.
- **a.** 28°C
- **b.** 154°C
- **c.** 568°C
- **d.** −55°C
- **e.** −184°C

12. Thermal Energy Could the thermal energy of a bowl of hot water equal that of a bowl of cold water? Explain your answer.

13. Heat Flow On a dinner plate, a baked potato always stays hot longer than any other food. Why?

14. Heat The hard tile floor of a bathroom always feels cold to bare feet even though the rest of the room is warm. Is the floor colder than the rest of the room?

15. Specific Heat If you take a plastic spoon out of a cup of hot cocoa and put it in your mouth, you are not likely to burn your tongue. However, you could very easily burn your tongue if you put the hot cocoa in your mouth. Why?

16. Heat Chefs often use cooking pans made of thick aluminum. Why is thick aluminum better than thin aluminum for cooking?

17. Heat and Food It takes much longer to bake a whole potato than to cook french fries. Why?

18. Critical Thinking As water heats in a pot on a stove, the water might produce some mist above its surface right before the water begins to roll. What is happening, and where is the coolest part of the water in the pot?

Physics Online physicspp.com/self_check_quiz

(l)John Cancalosi/Peter Arnold, Inc., (r)Jenny Hager/The Image Works

12.2 Changes of State and the Laws of Thermodynamics

Eighteenth-century steam-engine builders used heat to turn liquid water into steam. The steam pushed a piston to turn the engine, and then the steam was cooled and condensed into a liquid again. Adding heat to the liquid water changed not only its temperature, but also its structure. You will learn that changing state means changing form as well as changing the way in which atoms store thermal energy.

▶ Objectives

- **Define** heats of fusion and vaporization.
- **State** the first and second laws of thermodynamics.
- **Distinguish** between heat and work.
- **Define** entropy.

▶ Vocabulary

heat of fusion
heat of vaporization
first law of thermodynamics
heat engine
entropy
second law of thermodynamics

Changes of State

The three most common states of matter are solids, liquids, and gases. As the temperature of a solid is raised, it usually changes to a liquid. At even higher temperatures, it becomes a gas. How can these changes be explained? Consider a material in a solid state. When the thermal energy of the solid is increased, the motion of the particles also increases, as does the temperature.

Figure 12-10 diagrams the changes of state as thermal energy is added to 1.0 g of water starting at 243 K (ice) and continuing until it reaches 473 K (steam). Between points A and B, the ice is warmed to 273 K. At some point, the added thermal energy causes the particles to move rapidly enough that their motion overcomes the forces holding the particles together in a fixed location. The particles are still touching each other, but they have more freedom of movement. Eventually, the particles become free enough to slide past each other.

Melting point At this point, the substance has changed from a solid to a liquid. The temperature at which this change occurs is the melting point of the substance. When a substance is melting, all of the added thermal energy goes to overcome the forces holding the particles together in the solid state. None of the added thermal energy increases the kinetic energy of the particles. This can be observed between points B and C in Figure 12-10, where the added thermal energy melts the ice at a constant 273 K. Because the kinetic energy of the particles does not increase, the temperature does not increase between points B and C.

Boiling point Once a solid is completely melted, there are no more forces holding the particles in the solid state. Adding more thermal energy again increases the motion of the particles, and the temperature of the liquid rises. In the diagram, this process occurs between points C and D. As the temperature increases further, some particles in the liquid acquire enough energy to break free from the other particles. At a specific temperature, known as the boiling point, further addition of energy causes the substance to undergo another change of state. All the added thermal energy converts the substance from the liquid state to the gaseous state.

■ **Figure 12-10** A plot of temperature versus heat added when 1.0 g of ice is converted to steam. Note that the scale is broken between points D and E.

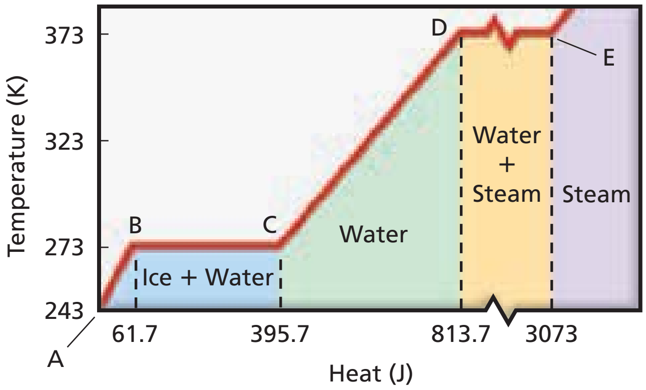


Table 12-2

Heats of Fusion and Vaporization of Common Substances

Material	Heat of Fusion H_f (J/kg)	Heat of Vaporization H_v (J/kg)
Copper	2.05×10^5	5.07×10^6
Mercury	1.15×10^4	2.72×10^5
Gold	6.30×10^4	1.64×10^6
Methanol	1.09×10^5	8.78×10^5
Iron	2.66×10^5	6.29×10^6
Silver	1.04×10^5	2.36×10^6
Lead	2.04×10^4	8.64×10^5
Water (ice)	3.34×10^5	2.26×10^6

As in melting, the temperature does not rise while a liquid boils. In Figure 12-10, this transition is represented between points D and E. After the material is entirely converted to gas, any added thermal energy again increases the motion of the particles, and the temperature rises. Above point E, steam is heated to temperatures greater than 373 K.

Heat of fusion The amount of energy needed to melt 1 kg of a substance is called the **heat of fusion** of that substance. For example, the heat of fusion of ice is 3.34×10^5 J/kg. If 1 kg of ice at its melting point, 273 K, absorbs 3.34×10^5 J, the ice becomes 1 kg of water at the same temperature, 273 K. The added energy causes a change in state but not in temperature. The horizontal distance in Figure 12-10 from point B to point C represents the heat of fusion.

Heat of vaporization At normal atmospheric pressure, water boils at 373 K. The thermal energy needed to vaporize 1 kg of a liquid is called the **heat of vaporization.** For water, the heat of vaporization is 2.26×10^6 J/kg. The distance from point D to point E in Figure 12-10 represents the heat of vaporization. Every material has a characteristic heat of vaporization.

Between points A and B, there is a definite slope to the line as the temperature is raised. This slope represents the specific heat of the ice. The slope between points C and D represents the specific heat of water, and the slope above point E represents the specific heat of steam. Note that the slope for water is less than those of both ice and steam. This is because water has a greater specific heat than does ice or steam. The heat, *Q*, required to melt a solid of mass *m* is given by the following equation.

Heat Required to Melt a Solid $Q = mH_f$

The heat required to melt a solid is equal to the mass of the solid times the heat of fusion of the solid.

Similarly, the heat, *Q*, required to vaporize a mass, *m*, of liquid is given by the following equation.

Heat Required to Vaporize a Liquid $Q = mH_v$

The heat required to vaporize a liquid is equal to the mass of the liquid times the heat of vaporization of the liquid.

When a liquid freezes, an amount of heat, $Q = -mH_f$, must be removed from the liquid to turn it into a solid. The negative sign indicates that the heat is transferred from the sample to the external world. In the same way, when a vapor condenses to a liquid, an amount of heat, $Q = -mH_v$, must be removed from the vapor. The values of some heats of fusion, H_f, and heats of vaporization, H_v, are shown in **Table 12-2.**

MINI LAB

Melting

1. Label two foam cups *A* and *B*.
2. Measure and pour 75 mL of room-temperature water into each cup. Wipe up any spilled liquid.
3. Add an ice cube to cup A, and add ice water to cup B until the water levels are equal.
4. Measure the temperature of the water in each cup at 1-min intervals until the ice has melted.
5. Record the temperatures in a data table and plot a graph.

Analyze and Conclude

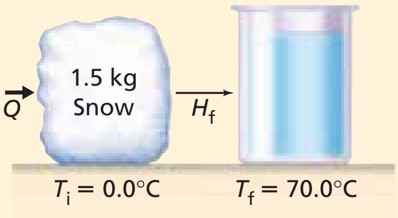
6. Do the samples reach the same final temperature? Why?

EXAMPLE Problem 3

Heat Suppose that you are camping in the mountains. You need to melt 1.50 kg of snow at 0.0°C and heat it to 70.0°C to make hot cocoa. How much heat will be needed?

1 Analyze and Sketch the Problem

- Sketch the relationship between heat and water in its solid and liquid states.
- Sketch the transfer of heat as the temperature of the water increases.



Known:

$m = 1.50$ kg

$T_i = 0.0°C$

$C = 4180$ J/kg·°C

$H_f = 3.34\times10^5$ J/kg

$T_f = 70.0°C$

Unknown:

$Q_{\text{melt ice}} = ?$

$Q_{\text{heat liquid}} = ?$

$Q_{\text{total}} = ?$

Physics Online

Personal Tutor For an online tutorial on heat, visit physicspp.com.

2 Solve for the Unknown

Calculate the heat needed to melt ice.

$$\begin{aligned}Q_{\text{melt ice}} &= mH_f\\ &= (1.50\text{ kg})(3.34\times10^5\text{ J/kg}) \\ &= 5.01\times10^5\text{ J}\\ &= 5.01\times10^2\text{ kJ}\end{aligned}$$

Substitute $m = 1.50$ kg, $H_f = 3.34\times10^5$ J/kg

Calculate the temperature change.

$$\begin{aligned}\Delta T &= T_f - T_i\\ &= 70.0°C - 0.0°C\\ &= 70.0°C\end{aligned}$$

Substitute $T_f = 70.0°C$, $T_i = 0.0°C$

Calculate the heat needed to raise the water temperature.

$$\begin{aligned}Q_{\text{heat liquid}} &= mC\Delta T\\ &= (1.50\text{ kg})(4180\text{ J/kg·°C})(70.0°C)\\ &= 4.39\times10^5\text{ J}\\ &= 4.39\times10^2\text{ kJ}\end{aligned}$$

Substitute $m = 1.50$ kg, $C = 4180$ J/kg·°C, $\Delta T = 70.0°C$

Calculate the total amount of heat needed.

$$\begin{aligned}Q_{\text{total}} &= Q_{\text{melt ice}} + Q_{\text{heat liquid}}\\ &= 5.01\times10^2\text{ kJ} + 4.39\times10^2\text{ kJ}\\ &= 9.40\times10^2\text{ kJ}\end{aligned}$$

Substitute $Q_{\text{melt ice}} = 5.01\times10^2$ kJ, $Q_{\text{heat liquid}} = 4.39\times10^2$ kJ

3 Evaluate the Answer

- **Are the units correct?** Energy units are in joules.
- **Does the sign make sense?** Q is positive when heat is absorbed.
- **Is the magnitude realistic?** The amount of heat needed to melt the ice is greater than the amount of heat needed to increase the water temperature by 70.0°C. It takes more energy to overcome the forces holding the particles in the solid state than to raise the temperature of water.

PRACTICE Problems

- Additional Problems, Appendix B
- Solutions to Selected Problems, Appendix C

19. How much heat is absorbed by 1.00×10^2 g of ice at −20.0°C to become water at 0.0°C?

20. A 2.00×10^2-g sample of water at 60.0°C is heated to steam at 140.0°C. How much heat is absorbed?

21. How much heat is needed to change 3.00×10^2 g of ice at −30.0°C to steam at 130.0°C?

The First Law of Thermodynamics

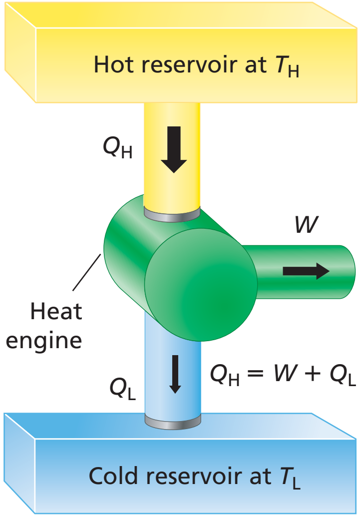
Before thermal energy was linked to the motion of atoms, the study of heat and temperature was considered to be a separate science. The first law developed for this science was a statement about what thermal energy is and where it can go. As you know, you can heat a nail by holding it over a flame or by pounding it with a hammer. That is, you can increase the nail's thermal energy by adding heat or by doing work on it. We do not normally think that the nail does work on the hammer. However, the work done by the nail on the hammer is equal to the negative of the work done by the hammer on the nail. The **first law of thermodynamics** states that the change in thermal energy, ΔU, of an object is equal to the heat, Q, that is added to the object minus the work, W, done by the object. Note that ΔU, Q, and W are all measured in joules, the unit of energy.

The First Law of Thermodynamics $\Delta U = Q - W$

The change in thermal energy of an object is equal to the heat added to the object minus the work done by the object.

Thermodynamics also involves the study of the changes in thermal properties of matter. The first law of thermodynamics is merely a restatement of the law of conservation of energy, which states that energy is neither created nor destroyed, but can be changed into other forms.

Another example of changing the amount of thermal energy in a system is a hand pump used to inflate a bicycle tire. As a person pumps, the air and the hand pump become warm. The mechanical energy in the moving piston is converted into thermal energy of the gas. Similarly, other forms of energy, such as light, sound, and electric energy, can be changed into thermal energy. For example, a toaster converts electric energy into heat when it toasts bread, and the Sun warms Earth with light from a distance of over 150 million km away.



■ **Figure 12-11** A heat engine transforms heat at high temperature into mechanical energy and low-temperature waste heat.

Heat engines The warmth that you experience when you rub your hands together is a result of the conversion of mechanical energy into thermal energy. The conversion of mechanical energy into thermal energy occurs easily. However, the reverse process, the conversion of thermal energy into mechanical energy, is more difficult. A device that is able to continuously convert thermal energy to mechanical energy is called a **heat engine.**

A heat engine requires a high-temperature source from which thermal energy can be removed; a low-temperature receptacle, called a sink, into which thermal energy can be delivered; and a way to convert the thermal energy into work. A diagram of a heat engine is shown in **Figure 12-11.** An automobile internal-combustion engine, such as the one shown in **Figure 12-12,** is one example of a heat engine. In the engine, a mixture of air and gasoline vapor is ignited and produces a high-temperature flame. Input heat, Q_H, flows from the flame to the air in the cylinder. The hot air expands and pushes on a piston, thereby changing thermal energy into mechanical energy. To obtain continuous mechanical energy, the engine must be returned to its starting condition. The heated air is expelled and replaced by new air, and the piston is returned to the top of the cylinder.

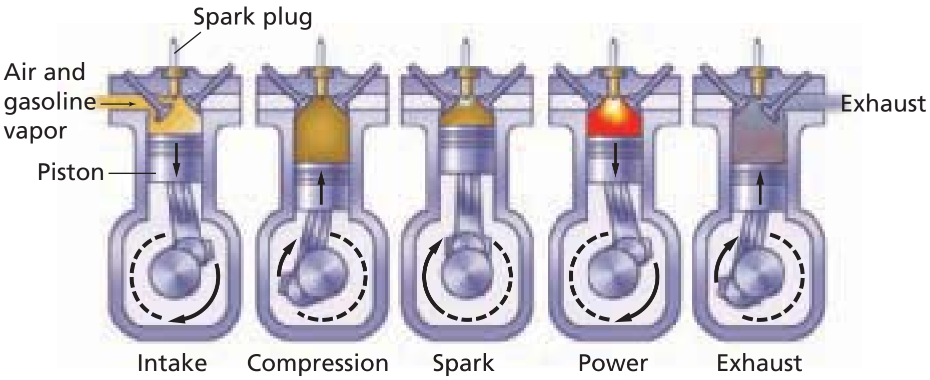


Figure 12-12 The heat produced by burning gasoline causes the gases that are produced to expand and to exert force and do work on the piston.

Concepts In Motion

Interactive Figure To see an animation on a heat engine, visit physicspp.com.

CLICK HERE

The entire cycle is repeated many times each minute. The thermal energy from the burning of gasoline is converted into mechanical energy, which eventually results in the movement of the car.

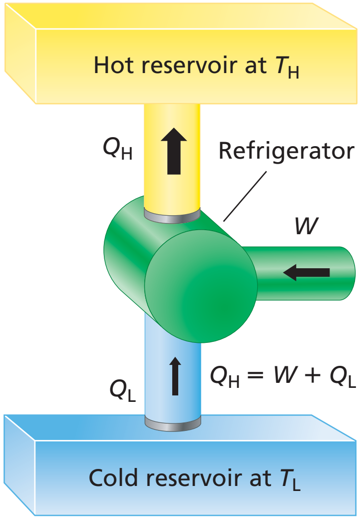
Not all of the thermal energy from the high-temperature flame in an automobile engine is converted into mechanical energy. When the automobile engine is functioning, the exhaust gases and the engine parts become hot. As the exhaust comes in contact with outside air and transfers heat to it, the temperature of the outside air is raised. In addition, heat from the engine is transferred to a radiator. Outside air passes through the radiator and the air temperature is raised.

All of this energy, Q_L, transferred out of the automobile engine is called waste heat, that is, heat that has not been converted into work. When the engine is working continuously, the internal energy of the engine does not change, or $\Delta U = 0 = Q - W$. The net heat going into the engine is $Q = Q_H - Q_L$. Thus, the work done by the engine is $W = Q_H - Q_L$. In an automobile engine, the thermal energy in the flame produces the mechanical energy and the waste heat that is expelled. All heat engines generate waste heat, and therefore no engine can ever convert all of the energy into useful motion or work.

Efficiency Engineers and car salespeople often talk about the fuel efficiency of automobile engines. They are referring to the amount of the input heat, Q_H, that is turned into useful work, W. The actual efficiency of an engine is given by the ratio W/Q_H. The efficiency could equal 100 percent only if all of the input heat were turned into work by the engine. Because there is always waste heat, even the most efficient engines fall short of 100-percent efficiency.

In solar collectors, heat is collected at high temperatures and used to drive engines. The Sun's energy is transmitted as electromagnetic waves and increases the internal energy of the solar collectors. This energy is then transmitted as heat to the engine, where it is turned into useful work and waste heat.

Refrigerators Heat flows spontaneously from a warm object to a cold object. However, it is possible to remove thermal energy from a colder object and add it to a warmer object if work is done. A refrigerator is a common example of a device that accomplishes this transfer with the use of mechanical work. Electric energy runs a motor that does work on a gas and compresses it.



■ **Figure 12-13** A refrigerator absorbs heat, Q_L, from the cold reservoir and gives off heat, Q_H, to the hot reservoir. Work, W, is done on the refrigerator.

The gas draws heat from the interior of the refrigerator, passes from the compressor through the condenser coils on the outside of the refrigerator, and cools into a liquid. Thermal energy is transferred into the air in the room. The liquid reenters the interior, vaporizes, and absorbs thermal energy from its surroundings. The gas returns to the compressor and the process is repeated. The overall change in the thermal energy of the gas is zero. Thus, according to the first law of thermodynamics, the sum of the heat removed from the refrigerator's contents and the work done by the motor is equal to the heat expelled, as shown in **Figure 12-13.**

Heat pumps A heat pump is a refrigerator that can be run in two directions. In the summer, the pump removes heat from a house and thus cools the house. In the winter, heat is removed from the cold outside air and transferred into the warmer house. In both cases, mechanical energy is required to transfer heat from a cold object to a warmer one.

PRACTICE **Problems**

• Additional Problems, Appendix B
• Solutions to Selected Problems, Appendix C

22. A gas balloon absorbs 75 J of heat. The balloon expands but stays at the same temperature. How much work did the balloon do in expanding?

23. A drill bores a small hole in a 0.40-kg block of aluminum and heats the aluminum by 5.0°C. How much work did the drill do in boring the hole?

24. How many times would you have to drop a 0.50-kg bag of lead shot from a height of 1.5 m to heat the shot by 1.0°C?

25. When you stir a cup of tea, you do about 0.050 J of work each time you circle the spoon in the cup. How many times would you have to stir the spoon to heat a 0.15-kg cup of tea by 2.0°C?

26. How can the first law of thermodynamics be used to explain how to reduce the temperature of an object?

The Second Law of Thermodynamics

Many processes that are consistent with the first law of thermodynamics have never been observed to occur spontaneously. Three such processes are presented in **Figure 12-14.** For example, the first law of thermodynamics does not prohibit heat flowing from a cold object to a hot object. However, when hot objects have been placed in contact with cold objects, the hot objects have never been observed to become hotter. Similarly, the cold objects have never been observed to become colder.

Entropy If heat engines completely converted thermal energy into mechanical energy with no waste heat, then the first law of thermodynamics would be obeyed. However, waste heat is always generated, and randomly distributed particles of a gas are not observed to spontaneously arrange themselves in specific ordered patterns. In the nineteenth century, French engineer Sadi Carnot studied the ability of engines to convert thermal energy into mechanical energy. He developed a logical proof that even an ideal engine would generate some waste heat. Carnot's result is best described in terms of a quantity called **entropy,** which is a measure of the disorder in a system.

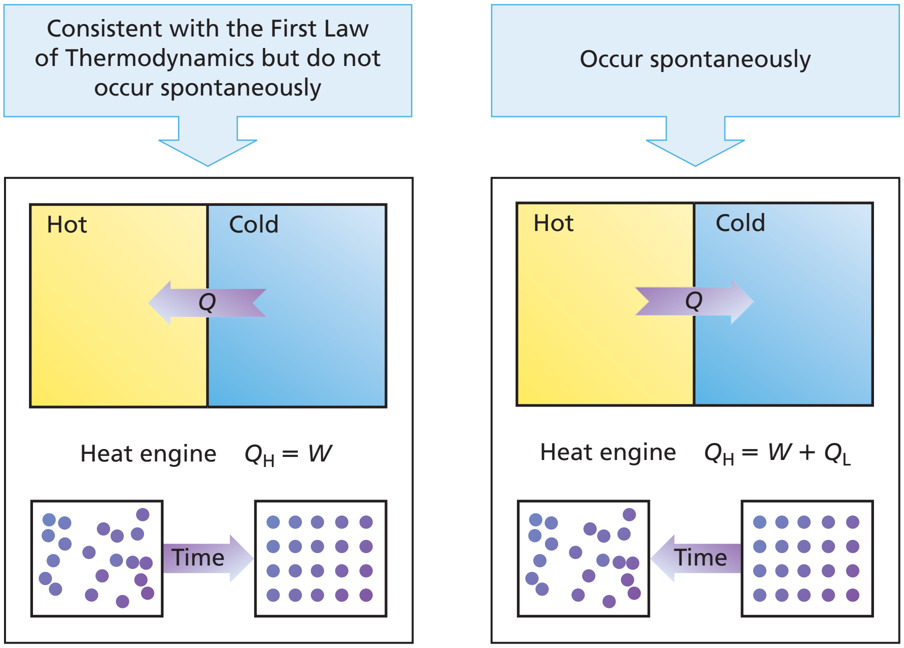


Figure 12-14 Many processes that do not violate the first law of thermodynamics do not occur spontaneously. The spontaneous processes obey both the first and second law of thermodynamics.

When a baseball is dropped and falls due to gravity, it possesses potential and kinetic energies that can be recovered to do work. However, when the baseball falls through the air, it collides with many air molecules that absorb some of its energy. This causes air molecules to move in random directions and at random speeds. The energy absorbed from the baseball causes more disorder among the molecules. The greater the range of speeds exhibited by the molecules, the greater the disorder, which in turn increases the entropy. It is highly unlikely that the molecules that have been dispersed in all directions will come back together, give their energies back to the baseball, and cause it to rise.

Entropy, like thermal energy, is contained in an object. If heat is added to an object, entropy is increased. If heat is removed from an object, entropy is decreased. If an object does work with no change in temperature, the entropy does not change, as long as friction is ignored. The change in entropy, ΔS, is expressed by the following equation, in which entropy has units of J/K and the temperature is measured in kelvins.

Change in Entropy $\Delta S = \frac{Q}{T}$

The change in entropy of an object is equal to the heat added to the object divided by the temperature of the object in kelvins.

CHALLENGE PROBLEM

Entropy has some interesting properties. Compare the following situations. Explain how and why these changes in entropy are different.

1. Heating 1.0 kg of water from 273 K to 274 K.
2. Heating 1.0 kg of water from 353 K to 354 K.
3. Completely melting 1.0 kg of ice at 273 K.
4. Heating 1.0 kg of lead from 273 K to 274 K.

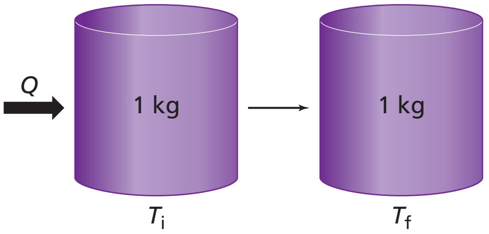


Figure 12-15 The spontaneous mixing of the food coloring and water is an example of the second law of thermodynamics.

The **second law of thermodynamics** states that natural processes go in a direction that maintains or increases the total entropy of the universe. That is, all things will become more and more disordered unless some action is taken to keep them ordered. The increase in entropy and the second law of thermodynamics can be thought of as statements of the probability of events happening. **Figure 12-15** illustrates an increase in entropy as food-coloring molecules, originally separate from the clear water, are thoroughly mixed with the water molecules over time. **Figure 12-16** shows an example of the second law of thermodynamics that might be familiar to many teenagers.

The second law of thermodynamics predicts that heat flows spontaneously only from a hot object to a cold object. Consider a hot iron bar and a cold cup of water. On the average, the particles in the iron will be moving very fast, whereas the particles in the water will be moving slowly. When the bar is plunged into the water and thermal equilibrium is eventually reached, the average kinetic energy of the particles in the iron and the water will be the same. More particles now have an increased random motion than was true for the initial state. This final state is less ordered than the initial state. The fast particles are no longer confined solely to the iron, and the slower particles are no longer confined only to the water; all speeds are evenly distributed. The entropy of the final state is greater than that of the initial state.

Violations of the second law We take for granted many daily events that occur spontaneously, or naturally, in one direction. We would be shocked, however, if the reverse of the same events occurred spontaneously. You are not surprised when a metal spoon, heated at one end, soon becomes uniformly hot. Consider your reaction, however, if a spoon lying on a table suddenly, on its own, became red hot at one end and icy cold at the other. If you dive into a swimming pool, you take for granted that you push the water molecules away as you enter the water. However, you would be amazed if you were swimming in the pool and all the water molecules spontaneously threw you up onto the diving board. Neither of these imagined reverse processes would violate the first law of thermodynamics. They are simply examples of the countless events that do not occur because their processes would violate the second law of thermodynamics.

Figure 12-16 If no work is done on a system, entropy spontaneously reaches a maximum.

(t)Doug Martin, (others)Richard Hutchings/CORBIS

The second law of thermodynamics and the increase in entropy also give new meaning to what has been commonly called the *energy crisis.* The energy crisis refers to the continued use of limited resources of fossil fuels, such as natural gas and petroleum. When you use a resource, such as natural gas to heat your home, you do not use up the energy in the gas. As the gas ignites, the internal chemical energy contained in the molecules of the gas is converted into thermal energy of the flame. The thermal energy of the flame is then transferred to thermal energy in the air of your home. Even if this warm air leaks to the outside, the energy is not lost. Energy has not been used up. The entropy, however, has increased.

The chemical structure of natural gas is very ordered. As you have learned, when a substance becomes warmer, the average kinetic energy of the particles in the substance increases. In contrast, the random motion of warmed air is very disordered. While it is mathematically possible for the original chemical order to be reestablished, the probability of this occurring is essentially zero. For this reason, entropy often is used as a measure of the unavailability of useful energy. The energy in the warmed air in a home is not as available to do mechanical work or to transfer heat to other objects as the original gas molecules were. The lack of usable energy is actually a surplus of entropy.

12.2 Section Review

27. Heat of Vaporization Old-fashioned heating systems sent steam into radiators in each room of a house. In the radiators, the steam condensed back to water. Analyze this process and explain how it heated a room.

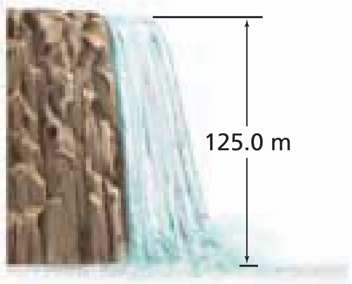
28. Heat of Vaporization How much heat is needed to change 50.0 g of water at 80.0°C to steam at 110.0°C?

29. Heat of Vaporization The specific heat of mercury is 140 J/kg·°C. Its heat of vaporization is 3.06×10^5 J/kg. How much energy is needed to heat 1.0 kg of mercury metal from 10.0°C to its boiling point and vaporize it completely? The boiling point of mercury is 357°C.

30. Mechanical Energy and Thermal Energy James Joule carefully measured the difference in temperature of water at the top and bottom of a waterfall. Why did he expect a difference?

31. Mechanical Energy and Thermal Energy A man uses a 320-kg hammer moving at 5.0 m/s to smash a 3.0-kg block of lead against a 450-kg rock. When he measured the temperature he found that it had increased by 5.0°C. Explain how this happened.

32. Mechanical Energy and Thermal Energy Water flows over a fall that is 125.0 m high, as shown in **Figure 12-17.** If the potential energy of the water is all converted to thermal energy, calculate the temperature difference between the water at the top and the bottom of the fall.



■ **Figure 12-17**

33. Entropy Evaluate why heating a home with natural gas results in an increased amount of disorder.

34. Critical Thinking A new deck of cards has all the suits (clubs, diamonds, hearts, and spades) in order, and the cards are ordered by number within the suits. If you shuffle the cards many times, are you likely to return the cards to their original order? Explain. Of what physical law is this an example?

PHYSICS LAB

Heating and Cooling

When a beaker of water is set on a hot plate and the hot plate is turned on, heat is transferred. It first is transferred to the beaker and then to the water at the bottom of the beaker by conduction. The water then transfers heat from the bottom to the top by moving hot water to the top through convection. Once the heat source is removed or shut off, the water radiates thermal energy until it reaches room temperature. How quickly the water heats up is a function of the amount of heat added, the mass of the water, and the specific heat of water.

Alternate CBL instructions can be found on the Web site.
physicspp.com

QUESTION

How does the constant supply of thermal energy affect the temperature of water?

Objectives

- **Measure,** in SI, temperature and mass.
- **Make and use graphs** to help describe the change in temperature of water as it heats up and cools down.
- **Explain** any similarities and differences in these two changes.

Safety Precautions

- **Be careful when using a hot plate. It can burn the skin.**

Materials

hot plate (or Bunsen burner)
250-mL ovenproof glass beaker
50–200 g of water
two thermometers (non-mercury)
stopwatch (or timer)

Procedure

1. Set the hot plate to the highest setting, or as recommended by your teacher. Allow a few minutes for the plate to heat up.
2. Measure the mass of the empty beaker.
3. Pour 150 mL of water into the beaker and measure the combined mass of the water and the beaker.
4. Calculate and record the mass of the water in the beaker.
5. Create a data and observations table.
6. Record the initial temperature of the water and the air in the classroom. Note that the bulb end of the thermometers must not touch the bottom or sides of the beaker, nor should it touch a table or your hands.
7. Place the beaker on the hot plate and record the temperature every minute for 5 min.
8. Carefully remove the beaker from the hot plate and record the temperature every minute for the next 10 min.
9. At the end of 10 min, record the temperature of the air.
10. Turn off the hot plate.
11. When finished, allow the equipment to cool and dispose of the water as instructed by your teacher.

Horizons Companies

Data Table

Mass of water	
Initial air temperature	
Final air temperature	
Change in air temperature	

Time (min)	Temperature (°C)	Heating or Cooling

Analyze

1. Calculate the change in air temperature to determine if air temperature may be an extraneous variable.
2. Make a scatter-plot graph of temperature (vertical axis) versus time (horizontal axis). Use a computer or a calculator to construct the graph, if possible.
3. **Calculate** What was the change in water temperature as the water heated up?
4. **Calculate** What was the drop in water temperature when the heat source was removed?
5. Calculate the average slope for the temperature increase by dividing change in temperature by the amount of time the water was heating up.
6. Calculate the average slope for the temperature decrease by dividing change in temperature by the amount of time the heat source was removed.

Conclude and Apply

1. **Summarize** What was the change in water temperature when a heat source was applied?
2. **Summarize** What was the change in water temperature once the heat source was removed?
3. What would happen to the water temperature after the next 10 min? Would it continue cooling down forever?
4. Did the water appear to heat up or cool down quicker? Why do you think this is so? *Hint: Examine the slopes you calculated.*
5. **Hypothesize** Where did the thermal energy in the water go once the water began to cool down? Support your hypothesis.

Going Further

1. Does placing your thermometer at the top of the water in your beaker result in different readings than if it is placed at the bottom of the beaker? Explain.
2. Hypothesize what the temperature changes might look like if you had the following amounts of water in the beaker: 50 mL, 250 mL.
3. Suppose you insulated the beaker you were using. How would the beaker's ability to heat up and cool down be affected?

Real-World Physics

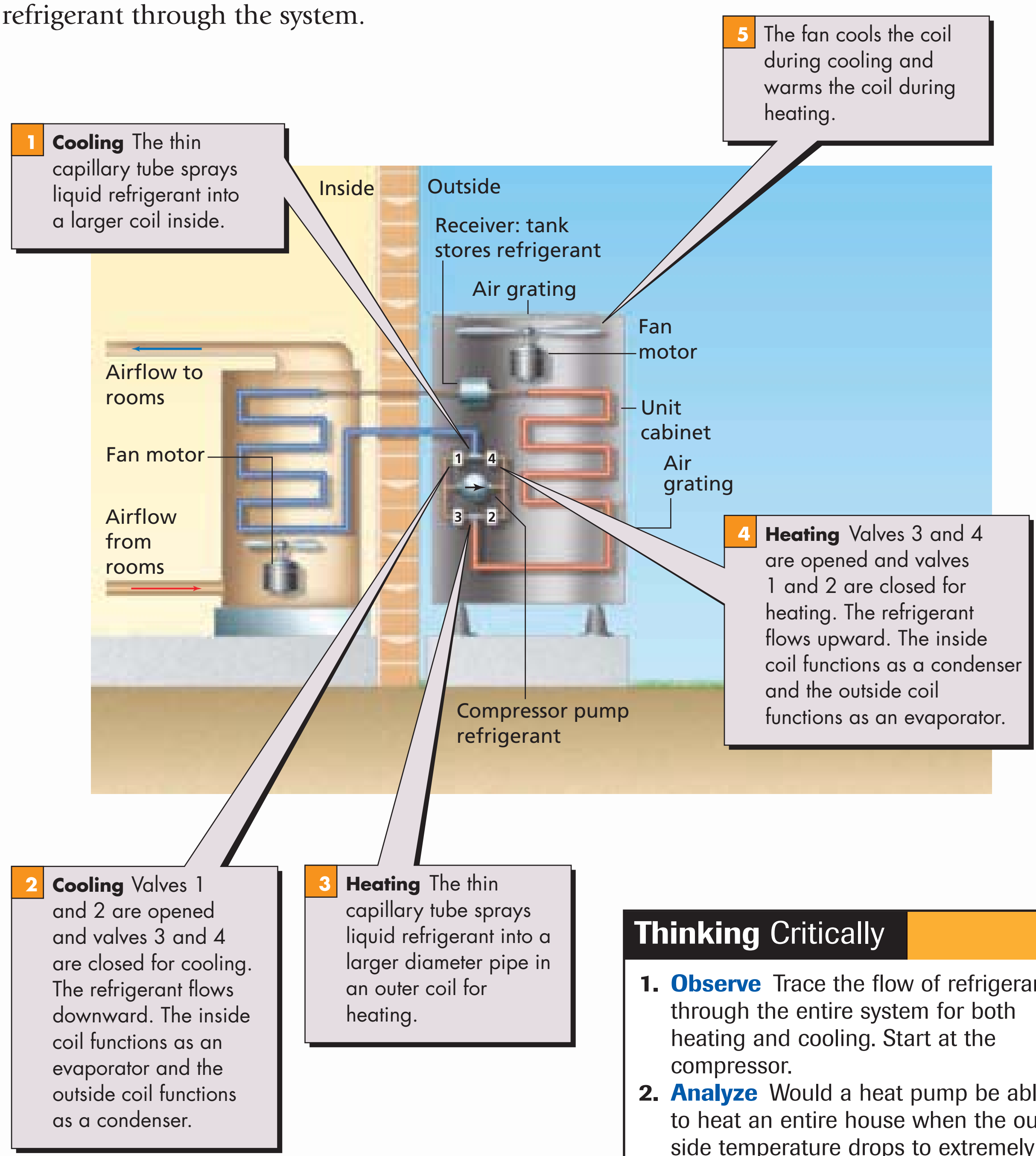
1. Suppose you were to use vegetable oil in the beaker instead of water. Hypothesize what the temperature changes might look like if you were to follow the same steps and perform the experiment.
2. If you were to take soup at room temperature and cook it in a microwave oven for 3 min, would the soup return to room temperature in 3 min? Explain your answer.

Physics Online

To find out more about thermal energy, visit the Web site: **physicspp.com**

HOW it Works The Heat Pump

Heat pumps, also called reversible air conditioners, were invented in the 1940s. They are used to heat and cool homes and hotel rooms. Heat pumps change from heaters to air conditioners by reversing the flow of refrigerant through the system.



1 **Cooling** The thin capillary tube sprays liquid refrigerant into a larger coil inside.

2 **Cooling** Valves 1 and 2 are opened and valves 3 and 4 are closed for cooling. The refrigerant flows downward. The inside coil functions as an evaporator and the outside coil functions as a condenser.

3 **Heating** The thin capillary tube sprays liquid refrigerant into a larger diameter pipe in an outer coil for heating.

4 **Heating** Valves 3 and 4 are opened and valves 1 and 2 are closed for heating. The refrigerant flows upward. The inside coil functions as a condenser and the outside coil functions as an evaporator.

5 The fan cools the coil during cooling and warms the coil during heating.

Thinking Critically

1. **Observe** Trace the flow of refrigerant through the entire system for both heating and cooling. Start at the compressor.
2. **Analyze** Would a heat pump be able to heat an entire house when the outside temperature drops to extremely cold levels?

Chapter 12

Study Guide

12.1 Temperature and Thermal Energy

Vocabulary

- conduction (p. 315)
- thermal equilibrium (p. 315)
- heat (p. 317)
- convection (p. 317)
- radiation (p. 317)
- specific heat (p. 318)

Key Concepts

- The temperature of a gas is proportional to the average kinetic energy of its particles.
- Thermal energy is a measure of the internal motion of an object's particles.
- A thermometer reaches thermal equilibrium with the object that it comes in contact with, and then a temperature-dependent property of the thermometer indicates the temperature.
- The Celsius and Kelvin temperature scales are used in scientific work. The magnitude of 1 K is equal to the magnitude of 1°C.
- At absolute zero, no more thermal energy can be removed from a substance.
- Heat is energy transferred because of a difference in temperature.

$$Q = mC\Delta T = mC(T_f - T_i)$$

- Specific heat is the quantity of heat required to raise the temperature of 1 kg of a substance by 1 K.
- In a closed, isolated system, heat may flow and change the thermal energy of parts of the system, but the total energy of the system is constant.

$$E_A + E_B = \text{constant}$$

12.2 Changes of State and the Laws of Thermodynamics

Vocabulary

- heat of fusion (p. 324)
- heat of vaporization (p. 324)
- first law of thermodynamics (p. 326)
- heat engine (p. 326)
- entropy (p. 328)
- second law of thermodynamics (p. 330)

Key Concepts

- The heat of fusion is the quantity of heat needed to change 1 kg of a substance from its solid to liquid state at its melting point.

$$Q = mH_f$$

- The heat of vaporization is the quantity of heat needed to change 1 kg of a substance from its liquid to gaseous state at its boiling point.

$$Q = mH_v$$

- Heat transferred during a change of state does not change the temperature of a substance.
- The change in energy of an object is the sum of the heat added to it minus the work done by the object.

$$\Delta U = Q - W$$

- A heat engine continuously converts thermal energy to mechanical energy.
- A heat pump and a refrigerator use mechanical energy to transfer heat from a region of lower temperature to one of higher temperature.
- Entropy is a measure of the disorder of a system.
- The change in entropy of an object is defined to be the heat added to the object divided by the temperature of the object.

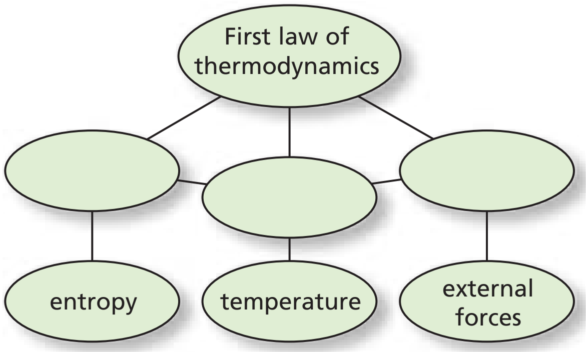
$$\Delta S = \frac{Q}{T}$$

Physics Online physicspp.com/vocabulary_puzzlemaker

Chapter 12 Assessment

Concept Mapping

35. Complete the following concept map using the following terms: *heat, work, internal energy.*



Mastering Concepts

36. Explain the differences among the mechanical energy of a ball, its thermal energy, and its temperature. (12.1)

37. Can temperature be assigned to a vacuum? Explain. (12.1)

38. Do all of the molecules or atoms in a liquid have the same speed? (12.1)

39. Is your body a good judge of temperature? On a cold winter day, a metal doorknob feels much colder to your hand than a wooden door does. Explain why this is true. (12.1)

40. When heat flows from a warmer object in contact with a colder object, do the two have the same temperature changes? (12.1)

41. Can you add thermal energy to an object without increasing its temperature? Explain. (12.2)

42. When wax freezes, does it absorb or release energy? (12.2)

43. Explain why water in a canteen that is surrounded by dry air stays cooler if it has a canvas cover that is kept wet. (12.2)

44. Which process occurs at the coils of a running air conditioner inside a house, vaporization or condensation? Explain. (12.2)

Applying Concepts

45. **Cooking** Sally is cooking pasta in a pot of boiling water. Will the pasta cook faster if the water is boiling vigorously or if it is boiling gently?

46. Which liquid would an ice cube cool faster, water or methanol? Explain.

47. Equal masses of aluminum and lead are heated to the same temperature. The pieces of metal are placed on a block of ice. Which metal melts more ice? Explain.

48. Why do easily vaporized liquids, such as acetone and methanol, feel cool to the skin?

49. Explain why fruit growers spray their trees with water when frost is expected to protect the fruit from freezing.

50. Two blocks of lead have the same temperature. Block A has twice the mass of block B. They are dropped into identical cups of water of equal temperatures. Will the two cups of water have equal temperatures after equilibrium is achieved? Explain.

51. **Windows** Often, architects design most of the windows of a house on the north side. How does putting windows on the south side affect the heating and cooling of the house?

Mastering Problems

12.1 Temperature and Thermal Energy

52. How much heat is needed to raise the temperature of 50.0 g of water from 4.5°C to 83.0°C?

53. A 5.00×10^{2}-g block of metal absorbs 5016 J of heat when its temperature changes from 20.0°C to 30.0°C. Calculate the specific heat of the metal.

54. **Coffee Cup** A 4.00×10^{2}-g glass coffee cup is 20.0°C at room temperature. It is then plunged into hot dishwater at a temperature of 80.0°C, as shown in **Figure 12-18.** If the temperature of the cup reaches that of the dishwater, how much heat does the cup absorb? Assume that the mass of the dishwater is large enough so that its temperature does not change appreciably.

Figure 12-18

55. A 1.00×10^{2}-g mass of tungsten at 100.0°C is placed in 2.00×10^{2} g of water at 20.0°C. The mixture reaches equilibrium at 21.2°C. Calculate the specific heat of tungsten.

56. A 6.0×10^2-g sample of water at 90.0°C is mixed with 4.00×10^2 g of water at 22.0°C. Assume that there is no heat loss to the surroundings. What is the final temperature of the mixture?

57. A 10.0-kg piece of zinc at 71.0°C is placed in a container of water, as shown in **Figure 12-19.** The water has a mass of 20.0 kg and a temperature of 10.0°C before the zinc is added. What is the final temperature of the water and the zinc?

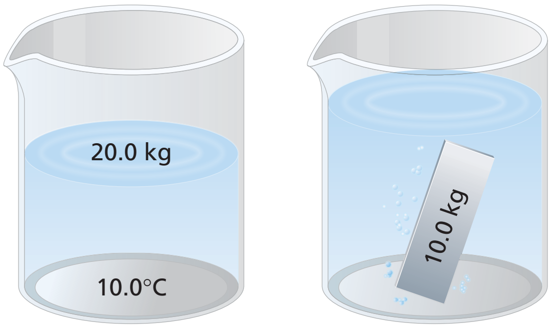


Figure 12-19

58. The kinetic energy of a compact car moving at 100 km/h is 2.9×10^5 J. To get a feeling for the amount of energy needed to heat water, what volume of water (in liters) would 2.9×10^5 J of energy warm from room temperature (20.0°C) to boiling (100.0°C)?

59. Water Heater A 3.0×10^2-W electric immersion heater is used to heat a cup of water, as shown in **Figure 12-20.** The cup is made of glass, and its mass is 3.00×10^2 g. It contains 250 g of water at 15°C. How much time is needed to bring the water to the boiling point? Assume that the temperature of the cup is the same as the temperature of the water at all times and that no heat is lost to the air.

Figure 12-20

60. Car Engine A 2.50×10^2-kg cast-iron car engine contains water as a coolant. Suppose that the engine's temperature is 35.0°C when it is shut off, and the air temperature is 10.0°C. The heat given off by the engine and water in it as they cool to air temperature is 4.40×10^6 J. What mass of water is used to cool the engine?

12.2 Changes of State and the Laws of Thermodynamics

61. Years ago, a block of ice with a mass of about 20.0 kg was used daily in a home icebox. The temperature of the ice was 0.0°C when it was delivered. As it melted, how much heat did the block of ice absorb?

62. A 40.0-g sample of chloroform is condensed from a vapor at 61.6°C to a liquid at 61.6°C. It liberates 9870 J of heat. What is the heat of vaporization of chloroform?

63. A 750-kg car moving at 23 m/s brakes to a stop. The brakes contain about 15 kg of iron, which absorbs the energy. What is the increase in temperature of the brakes?

64. How much heat is added to 10.0 g of ice at −20.0°C to convert it to steam at 120.0°C?

65. A 4.2-g lead bullet moving at 275 m/s strikes a steel plate and comes to a stop. If all its kinetic energy is converted to thermal energy and none leaves the bullet, what is its temperature change?

66. Soft Drink A soft drink from Australia is labeled "Low-Joule Cola." The label says "100 mL yields 1.7 kJ." The can contains 375 mL of cola. Chandra drinks the cola and then wants to offset this input of food energy by climbing stairs. How high would Chandra have to climb if she has a mass of 65.0 kg?

Mixed Review

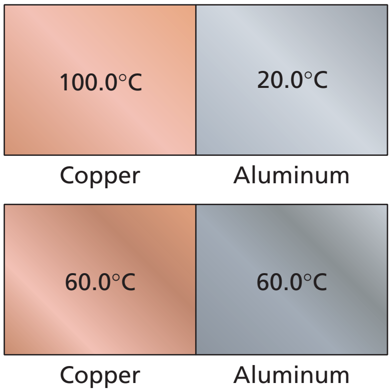
67. What is the efficiency of an engine that produces 2200 J/s while burning enough gasoline to produce 5300 J/s? How much waste heat does the engine produce per second?

68. Stamping Press A metal stamping machine in a factory does 2100 J of work each time it stamps out a piece of metal. Each stamped piece is then dipped in a 32.0-kg vat of water for cooling. By how many degrees does the vat heat up each time a piece of stamped metal is dipped into it?

69. A 1500-kg automobile comes to a stop from 25 m/s. All of the energy of the automobile is deposited in the brakes. Assuming that the brakes are about 45 kg of aluminum, what would be the change in temperature of the brakes?

70. **Iced Tea** To make iced tea, you start by brewing the tea with hot water. Then you add ice. If you start with 1.0 L of 90°C tea, what is the minimum amount of ice needed to cool it to 0°C? Would it be better to let the tea cool to room temperature before adding the ice?

71. A block of copper at 100.0°C comes in contact with a block of aluminum at 20.0°C, as shown in **Figure 12-21.** The final temperature of the blocks is 60.0°C. What are the relative masses of the blocks?



■ **Figure 12-21**

72. A 0.35-kg block of copper sliding on the floor hits an identical block moving at the same speed from the opposite direction. The two blocks come to a stop together after the collision. Their temperatures increase by 0.20°C as a result of the collision. What was their speed before the collision?

73. A 2.2-kg block of ice slides across a rough floor. Its initial velocity is 2.5 m/s and its final velocity is 0.50 m/s. How much of the ice block melted as a result of the work done by friction?

Thinking Critically

74. **Analyze and Conclude** A certain heat engine removes 50.0 J of thermal energy from a hot reservoir at temperature $T_H = 545$ K and expels 40.0 J of heat to a colder reservoir at temperature $T_L = 325$ K. In the process, it also transfers entropy from one reservoir to the other.

a. How does the operation of the engine change the total entropy of the reservoirs?

b. What would be the total entropy change in the reservoirs if $T_L = 205$ K?

75. **Analyze and Conclude** During a game, the metabolism of basketball players often increases by as much as 30.0 W. How much perspiration must a player vaporize per hour to dissipate this extra thermal energy?

76. **Analyze and Conclude** Chemists use calorimeters to measure the heat produced by chemical reactions. For instance, a chemist dissolves 1.0×10^{22} molecules of a powdered substance into a calorimeter containing 0.50 kg of water. The molecules break up and release their binding energy to the water. The water temperature increases by 2.3°C. What is the binding energy per molecule for this substance?

77. **Apply Concepts** All of the energy on Earth comes from the Sun. The surface temperature of the Sun is approximately 10^4 K. What would be the effect on our world if the Sun's surface temperature were 10^3 K?

Writing in Physics

78. Our understanding of the relationship between heat and energy was influenced by a soldier named Benjamin Thompson, Count Rumford, and a brewer named James Prescott Joule. Both relied on experimental results to develop their ideas. Investigate what experiments they did and evaluate whether or not it is fair that the unit of energy is called the joule and not the thompson.

79. Water has an unusually large specific heat and large heats of fusion and vaporization. Our weather and ecosystems depend upon water in all three states. How would our world be different if water's thermodynamic properties were like other materials, such as methanol?

Cumulative Review

80. A rope is wound around a drum with a radius of 0.250 m and a moment of inertia of 2.25 kg·m^2. The rope is connected to a 4.00-kg block. (Chapter 8)

a. Find the linear acceleration of the block.

b. Find the angular acceleration of the drum.

c. Find the tension, F_T, in the rope.

d. Find the angular velocity of the drum after the block has fallen 5.00 m.

81. A weight lifter raises a 180-kg barbell to a height of 1.95 m. How much work is done by the weight lifter in lifting the barbell? (Chapter 10)

82. In a Greek myth, the man Sisyphus is condemned by the gods to forever roll an enormous rock up a hill. Each time he reaches the top, the rock rolls back down to the bottom. If the rock has a mass of 215 kg, the hill is 33 m in height, and Sisyphus can produce an average power of 0.2 kW, how many times in 1 h can he roll the rock up the hill? (Chapter 11)

For more problems, go to Additional Problems, Appendix B.

Standardized Test Practice

Multiple Choice

1. Which of the following temperature conversions is incorrect?

Ⓐ −273°C = 0 K
Ⓑ 273°C = 546 K
Ⓒ 298 K = 571°C
Ⓓ 88 K = −185°C

2. What are the units of entropy?

Ⓐ J/K
Ⓑ K/J
Ⓒ J
Ⓓ kJ

3. Which of the following statements about thermal equilibrium is false?

Ⓐ When two objects are at equilibrium, heat radiation between the objects continues to occur.
Ⓑ Thermal equilibrium is used to create energy in a heat engine.
Ⓒ The principle of thermal equilibrium is used for calorimetry calculations.
Ⓓ When two objects are not at equilibrium, heat will flow from the hotter object to the cooler object.

4. How much heat is required to heat 87 g of methanol ice at 14 K to vapor at 340 K? (melting point = −97.6°C, boiling point = 64.6°C)

Ⓐ 17 kJ
Ⓑ 69 kJ
Ⓒ 1.4×10^2 kJ
Ⓓ 1.5×10^2 kJ

5. Which statement is true about energy, entropy, and changes of state?

Ⓐ Freezing ice increases in energy as it gains molecular order as a solid.
Ⓑ The higher the specific heat capacity of a substance, the higher its melting point will be.
Ⓒ States of matter with increased kinetic energy have higher entropy.
Ⓓ Energy and entropy cannot increase at the same time.

6. How much heat is needed to warm 363 mL of water in a baby bottle from 24°C to 38°C?

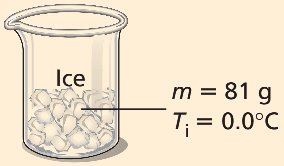
Ⓐ 21 kJ
Ⓑ 36 kJ
Ⓒ 121 kJ
Ⓓ 820 kJ

7. Why is there always some waste heat in a heat engine?

Ⓐ Heat cannot flow from a cold object to a hot object.
Ⓑ Friction slows the engine down.
Ⓒ The entropy increases at each stage.
Ⓓ The heat pump uses energy.

8. How much heat is absorbed from the surroundings when 81 g of 0.0°C ice in a beaker melts and warms to 10°C?

Ⓐ 0.34 kJ
Ⓑ 27 kJ
Ⓒ 30 kJ
Ⓓ 190 kJ



9. You do 0.050 J of work on the coffee in your cup each time you stir it. What would be the increase in entropy in 125 mL of coffee at 65°C when you stir it 85 times?

Ⓐ 0.013 J/K
Ⓑ 0.050 J
Ⓒ 0.095 J/K
Ⓓ 4.2 J

Extended Answer

10. What is the difference in heat required to melt 454 g of ice at 0.00°C, and to turn 454 g of water at 100.0°C into steam? Is the amount of this difference greater or less than the amount of energy required to heat the 454 g of water from 0.00°C to 100.0°C?

✓ Test-Taking TIP

Your Mistakes Can Teach You

The mistakes you make before the test are helpful because they show you areas in which you need more work. When calculating the heat needed to melt and warm a substance, remember to calculate the heat needed for melting as well as the heat needed for raising the temperature of the substance.