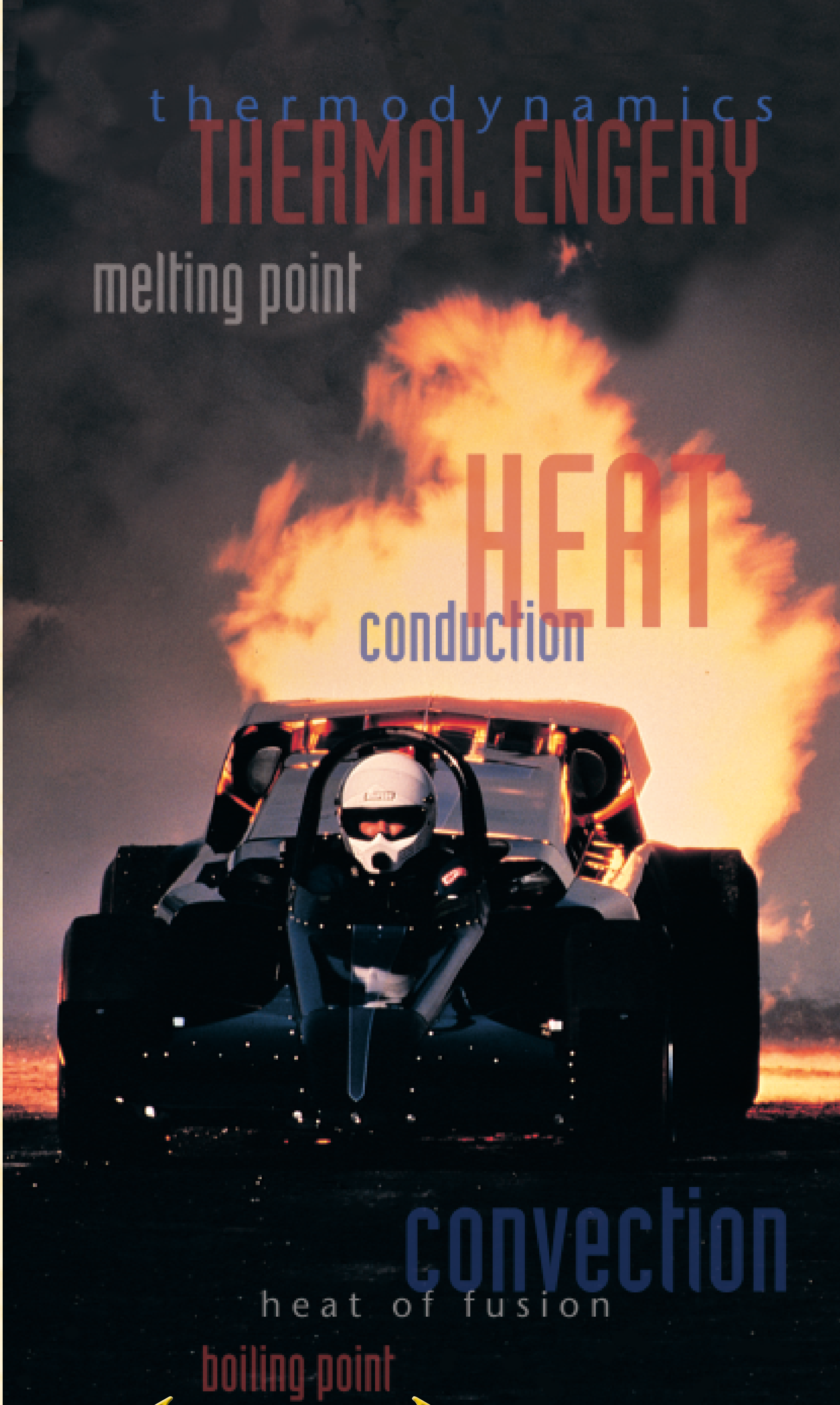
Perpetual Motion?

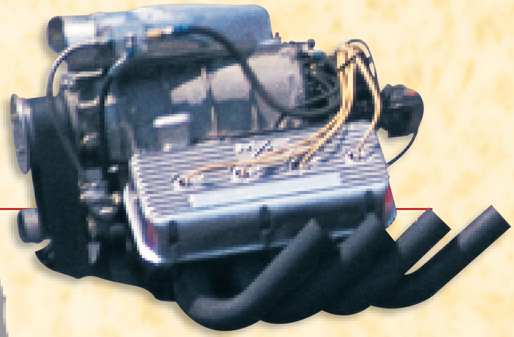
The engine converts the chemical energy stored in the fuel and oxygen into kinetic energy. Could we ever invent an engine that converts all the energy into useful energy of motion?

Look at the text on page 290 for the answer.



CHAPTER

12 Thermal Energy



In Chapter 11, you learned that one of the forms of energy is stored energy. Energy can be stored in several ways. Your body stores energy as sugar in your body's cells. Some heating systems store energy in hot-water tanks. A battery stores chemical energy released when a circuit is completed. In the last chapter, you saw that a compressed spring can store energy.

By converting stored energy into mechanical energy, an object, such as this racing car, can gain kinetic energy. However, it takes more than the elastic potential energy of a compressed spring to accelerate this high-tech racing car. It takes the energy stored in the fuel combined with the oxygen in the air to get the car to move faster than 350 km/h (217 mph) in a few seconds.

In Chapter 10, you learned that work could transfer energy from the environment of a system. In this chapter, you'll learn about another way to transfer energy—the way that has transformed the lives of people everywhere.

The steam engine, the first invention that produced mechanical energy from fuel, transformed the United States from a society of farms to one with many factories in the 1800s. The steam engine can be called a heat engine because it converts heat into work. In the case of the steam engine, the fuel source is outside the engine. However, you are probably much more familiar with heat engines that burn fuel inside the engine—internal combustion engines. One of these, the gasoline engine, precipitated tremendous changes in the way people traveled, worked, and lived. Now, more than a century after its invention in 1876 by Nikolaus Otto in Germany, the gasoline engine is firmly entrenched in our society.

WHAT YOU'LL LEARN

- You will define temperature.
- You will calculate heat transfer.
- You will distinguish heat from work.

WHY IT'S IMPORTANT

- Thermal energy provides the energy to keep you warm, to prepare and preserve food and to manufacture many of the objects you use on a daily basis.

PHYSICS Online

To find out more about thermal energy, visit the Glencoe Science Web site at **science.glencoe.com**

CLICK HERE

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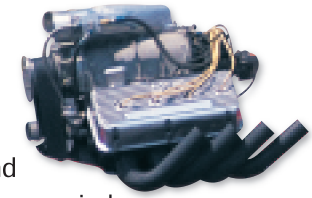
12.1 Temperature and Thermal Energy



OBJECTIVES

- **Describe** the nature of thermal energy.
- **Define** temperature and **distinguish** it from thermal energy.
- **Use** the Celsius and Kelvin temperature scales and **convert** one to the other.
- **Define** specific heat and **calculate** heat transfer.

Europe went through a "Little Ice Age" in the 1600s and 1700s, when temperatures were lower than any other period during the previous one thousand years. Keeping warm was vitally important, and many people devoted themselves to the study of heat. One result was the invention of machines that used the energy produced by burning fuel to do useful work. These machines freed society from its dependence on the energy provided solely by people and animals. As inventors tried to make these machines more powerful and more efficient, they developed the science of **thermodynamics,** the study of heat.

What makes a hot body hot?

Internal combustion engines require very high temperatures to operate. These high temperatures are usually produced by burning fuel. Although the effects of fire have been known since ancient times, only in the eighteenth century did scientists begin to understand how a hot body differs from a cold body. They proposed that when a body is heated, an invisible fluid called "caloric" is added to the body. Hot bodies contain more caloric than cold bodies. The caloric theory could explain observations such as the expansion of objects when heated, but it could not explain why hands get warm when they are rubbed together.

In the mid-nineteenth century, scientists developed a new theory to replace the caloric theory. This theory is based on the assumption that matter is made up of many tiny particles that are always in motion. In a hot body, the particles move faster, and thus have greater kinetic energy than particles in a cooler body. This theory is called the **kinetic-molecular theory.**

The model of a solid shown in **Figure 12–1** can help you understand the kinetic-molecular theory. This model is a solid made up of tiny spherical particles held together by massless springs. The springs represent the electromagnetic forces that bind the solid together. The particles vibrate back and forth and thus have kinetic energy. The vibrations compress and extend the springs, so the solid has potential energy as well. The overall energy of motion of the particles that make up an object is called the **thermal energy** of that object.

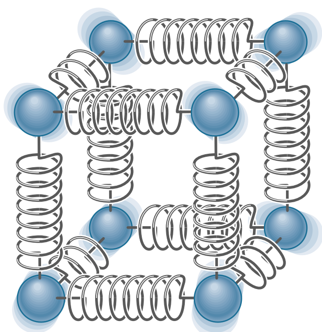


FIGURE 12–1 Molecules of a solid behave in some ways as if they were held together by springs.

Thermal Energy and Temperature

According to the kinetic-molecular theory, a hot body has more thermal energy than a similar cold body, shown in **Figure 12–2.** This means that, as a whole, the particles in a hot body have greater thermal energy than the particles in a cold body. It does not mean that all the particles in a body have exactly the same energy. The particles have a range of energies, some high, others low. It is the *average* energy of particles in a hot body that is higher than that of particles in a cold body. To help you understand this, consider the heights of students in a twelfth-grade class. The heights vary, but you can calculate the average height. This average is likely to be larger than the average height of students in a ninth-grade class, even though some ninth-graders might be taller than some twelfth-graders.

How can you determine the hotness of an object? "Hotness" is a property of an object called its **temperature,** and is measured on a definite scale. Consider two objects. In the one you call hotter, the particles are moving faster. That is, they have a greater average kinetic energy. Because the temperature is a property of matter, the temperature does not depend on the number of particles in the object. Temperature only depends on the average kinetic energy of the particles in the object. To illustrate this, consider two blocks of steel. The first block has a mass of 1 kg and the second block has twice the mass of the first at 2 kg. If the 1 kg mass of steel is at the same temperature as a 2 kg mass of steel, the average kinetic energy of the particles in the both blocks is the same. But the 2 kg mass of steel has twice the mass and the total amount of kinetic energy of particles in the 2 kg mass is twice the amount in the 1 kg mass. Total kinetic energy is divided over all the particles in an object to get the average. Therefore, the thermal energy in an object is proportional to the number of particles in it, but its temperature is not, as is shown in **Figure 12–3.**

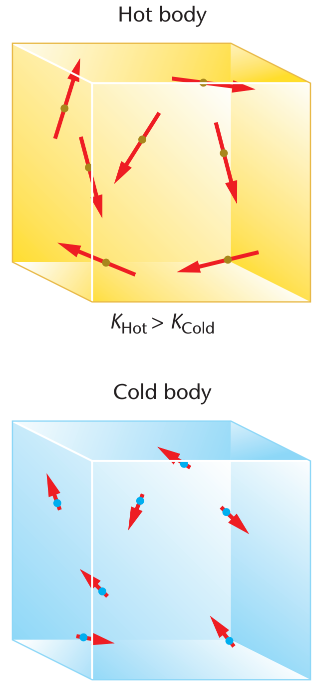


FIGURE 12–2 Particles in a hot body have greater kinetic and potential energies than particles in a cold body.

Equilibrium and Thermometry

How do you measure your temperature? If you suspect that you have a fever, you may place a thermometer in your mouth and wait two or three minutes. The thermometer then provides a measure of the temperature of your body.

You are probably less familiar with the microscopic process involved in measuring temperature. Your body is hot compared to the thermometer, which means that the particles in your body have greater thermal energy and are moving faster. The thermometer is made of a glass tube. When the cold glass touches your hotter body, the faster-moving particles in your skin collide with the slower-moving particles in the glass. Energy is transferred from your skin to the glass particles by the process of **conduction,** the transfer of kinetic energy when particles collide. The thermal energy of the particles that make up the thermometer increases and, at the same time, the thermal energy of the particles in your skin decreases. As the particles in the glass become more energetic, they begin to transfer energy back to the particles in your body. At some point, the rate of transfer of energy back and forth between the glass and your body will become equal and both will be at the same temperature. Your body and the thermometer are then in **thermal equilibrium.** That is, the rate at which energy flows from your body to the glass is equal to the rate at which energy flows

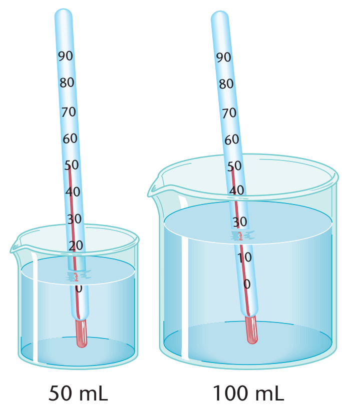
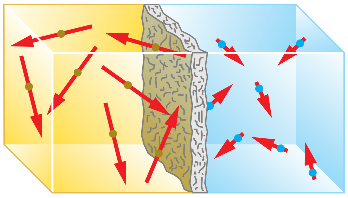


FIGURE 12–3 Temperature does not depend on the number of particles in a body.



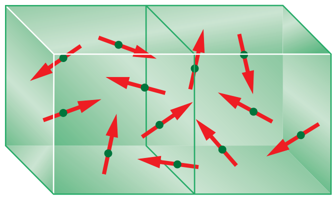


FIGURE 12–4 Thermal energy is transferred from a hot body to a cold body. When thermal equilibrium is reached, the transfer of energy between bodies is equal.

from the glass to your body. Objects that are in thermal equilibrium are at the same temperature, as is shown in **Figure 12–4.**

A **thermometer** is a device that measures temperature. It is placed in contact with an object and allowed to come to thermal equilibrium with that object. The operation of a thermometer depends on some property, such as volume, that 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. Mercury is another liquid commonly used in thermometers. In liquid crystal thermometers, such as the one shown in **Figure 12–5,** 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 creates an instrument that indicates the temperature by color.

Temperature Scales: Celsius and Kelvin

Temperature scales were developed by scientists to allow them to compare their temperature measurements with those of other scientists. A scale based on the properties of water was devised in 1741 by the Swedish astronomer and physicist Anders Celsius. On this scale, now called the Celsius scale, the freezing point of pure water is 0 degrees (0°C). The boiling point of pure water at sea level is 100 degrees (100°C). On the Celsius scale, the average temperature of the human body is 37°C. **Figure 12–6** shows representative temperatures on the three most common scales: Fahrenheit, Celsius, and Kelvin.

The wide range of temperatures in the universe is shown in **Figure 12–7.** Temperatures do not appear to have an upper limit. The interior of the sun is at least 1.5×10^7°C and other stars are even hotter. Temperatures do, however, have a lower limit. Generally, materials contract as they cool. If you cool an "ideal" gas, one in which the particles occupy a tremendously large volume compared to their own size and which don't interact, it contracts in such a way that it occupies a volume that is only the size of the molecules at −273.15°C. At this temperature, all the thermal energy that can 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. This is called **absolute zero,** and is usually rounded to −273°C.

FIGURE 12–5 Thermometers use a change in physical properties to measure temperature. A liquid crystal thermometer changes color with temperature change.

The Kelvin temperature scale is based on absolute zero. Absolute zero is the zero point of the Kelvin scale. On the Kelvin scale, the freezing point of water (0°C) is 273 K and the boiling point of water is 373 K.

Each interval on this scale, called a **kelvin,** is equal to the size of one Celsius degree. Thus, $T_C + 273 = T_K$. 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, or atoms and lasers. Using these techniques, physicists have reached temperatures as low as 2.0×10^{-9} K.

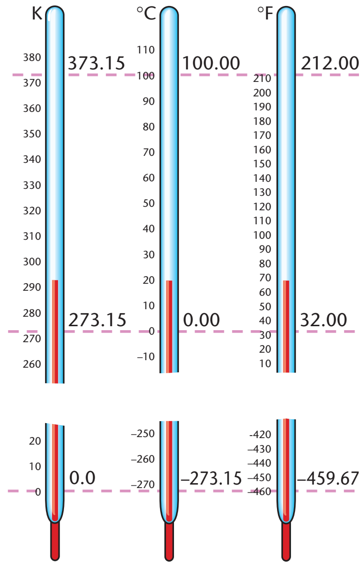


FIGURE 12–6 The three most common temperature scales are Kelvin, Celsius, and Fahrenheit.

Example Problem

Temperature Conversion

Convert 25°C to kelvins.

Calculate Your Answer

Known:

Celsius temperature = 25°C

Unknown:

$T_K = ?$

Strategy:

Change Celsius temperatures to Kelvin by using the relationship $T_K = T_C + 273$.

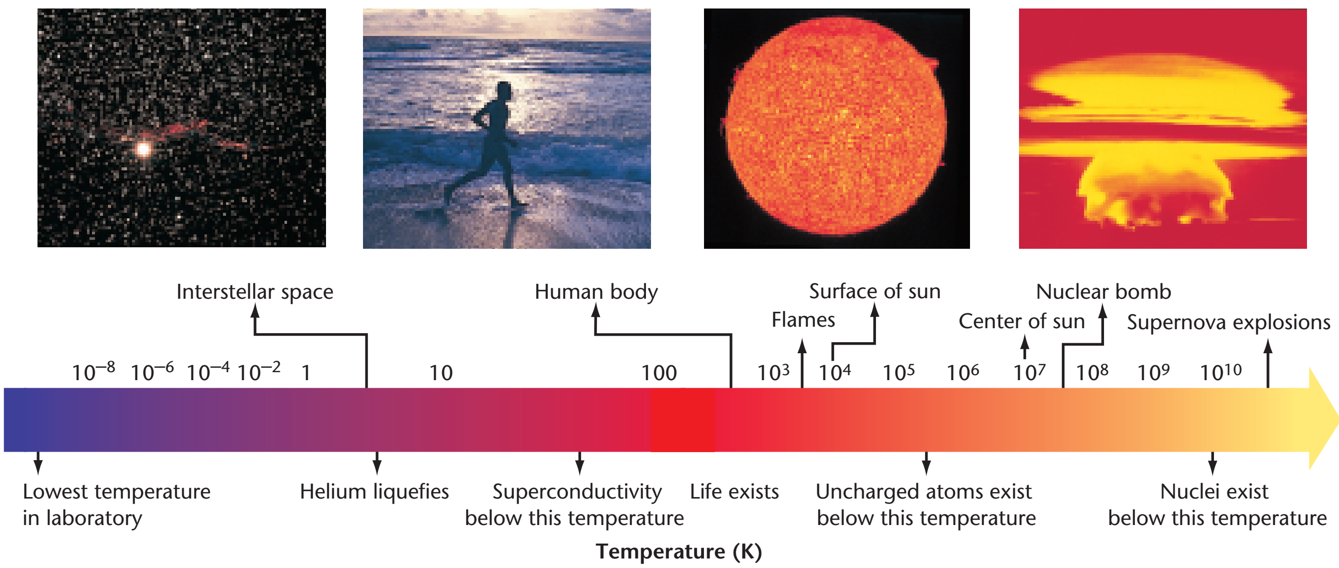
Calculations:

$$T_K = T_C + 273$$
$$= 25 + 273$$
$$= 298 \text{ K}$$

Check Your Answer

- Are your units correct? Temperature is measured in kelvins.
- Does the magnitude make sense? Temperatures on the Kelvin scale are all larger than those on the Celsius scale.

FIGURE 12–7 There is an extremely wide range of temperatures throughout the universe. Note the scale has been expanded in areas of particular interest.



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Practice Problems

1. Make the following conversions.
   - **a.** 0°C to kelvins
   - **b.** 0 K to degrees Celsius
   - **c.** 273°C to kelvins
   - **d.** 273 K to degrees Celsius
2. Convert the following Celsius temperatures to Kelvin temperatures.
   - **a.** 27°C
   - **b.** 150°C
   - **c.** 560°C
   - **d.** −50°C
   - **e.** −184°C
   - **f.** −300°C
3. Convert the following Kelvin temperatures to Celsius temperatures.
   - **a.** 110 K
   - **b.** 70 K
   - **c.** 22 K
   - **d.** 402 K
   - **e.** 323 K
   - **f.** 212 K
4. Find the Celsius and Kelvin temperatures for the following.
   - **a.** room temperature
   - **b.** refrigerator temperature
   - **c.** typical hot summer day
   - **d.** typical winter night

F.Y.I.

Why are degrees used to measure temperature? Fahrenheit set up his temperature scale so that there were 180 degrees separating the temperature where water freezes to where water boils. This was because there are 180 degrees in a straight angle.

Heat and Thermal Energy

One way to increase the temperature of an object is to place it in contact with a hotter object. The thermal energy of the hotter object is decreased, and the thermal energy of the cooler object is increased. Energy always flows from the hotter object to the cooler object. Energy never flows from a colder object to a hotter object. **Heat** is the energy that flows between two objects as a result of a difference in temperature. The symbol Q is used to represent the amount of heat. If Q has a negative value, heat has left the object; if Q has a positive value, heat has been absorbed by the object. Heat, like other forms of energy, is measured in joules.

Thermal energy transfer You have already learned one way that heat flows from a warmer body to a colder one. If you place one end of a metal rod in a flame, it becomes hot. The other end also becomes warm very quickly. Heat is conducted because the particles in the rod are in direct contact.

A second means of thermal transfer involves particles that are not in direct contact. Have you ever looked in a pot of water just about to boil? You can see motion of water, as water heated by conduction at the bottom of the pot flows up and the colder water at the top sinks. Heat flows between the rising hotter water and the descending colder water. This motion of fluid, whether liquid or gas, caused by temperature differences, is **convection.**

The third method of thermal transfer, unlike the first two, does not depend on the presence of matter. The sun warms us from over 150 million kilometers via **radiation,** the transfer of energy by electromagnetic waves. These waves carry the energy from the hot sun to the much cooler Earth.

Specific heat When heat flows into an object, its thermal energy increases, and so does its temperature. The amount of the increase depends on the size of the object. It also depends on the material from which the object is made. 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 one temperature unit. In SI units, specific heat, represented by C (not to be confused with °C), is measured in J/kg·K. **Table 12–1** provides values of specific heat for some common substances. For example, 903 J must be added to one kilogram of aluminum to raise the temperature one kelvin. The specific heat of aluminum is 903 J/kg·K.

Note that water has a high specific heat compared to those of other substances, even ice and steam. One kilogram of water requires the addition of 4180 J of energy to increase its temperature by one kelvin. The same mass of copper requires only 385 J to increase its temperature by one kelvin. The 4180 J of energy needed to raise the temperature of one kilogram of water by one kelvin would increase the temperature of the same mass of copper by 11 K. The high specific heat of water is the reason water is used in car radiators to remove thermal energy from the engine block.

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. The amount of heat transferred can be determined using the following equation.

Heat Transfer $Q = mC\Delta T = mC(T_{\text{final}} - T_{\text{initial}})$

where Q is the heat gained or lost, m is the mass of the object, C is the specific heat of the substance, and ΔT is the change in temperature. When the temperature of 10.0 kg of water is increased by 5.0 K, the heat absorbed, Q, is

$$Q = (10.0\ \text{kg})(4180\ \text{J/kg·K})(5.0\ \text{K}) = 2.1 \times 10^5\ \text{J}.$$

Because one Celsius degree is equal in magnitude to one kelvin, temperature changes can be measured in either kelvins or Celsius degrees.

F.Y.I.

Calorie and calorimeter are derived from *calor*, the Latin word for heat.

TABLE 12–1

Specific Heat of Common Substances

Material	Specific heat J/kg • K	Material	Specific heat J/kg • K
aluminum	903	lead	130
brass	376	methanol	2450
carbon	710	silver	235
copper	385	steam	2020
glass	664	water	4180
ice	2060	zinc	388
iron	450		

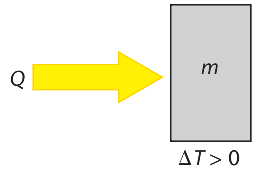
Example Problem

Heat Transfer

A 0.400-kg block of iron is heated from 295K to 325K. How much heat had to be transferred to the iron?

Sketch the Problem

- Sketch the flow of heat into the block of iron



Calculate Your Answer

Known:

$m = 0.400$ kg

$C = 450$ J/kg·K

$T_i = 295$ K

$T_f = 325$ K

Strategy:

The heat transferred is a product of the mass, specific heat, and the temperature change.

Calculations:

$$Q = mC(T_f - T_i)$$
$$= (0.400 \text{ kg})(450 \text{ J/kg·K})(325 - 295 \text{ K})$$
$$= 5.4 \times 10^3 \text{ J}$$

Unknown:

$Q = ?$

Check Your Answer

- Are your units correct? Heat is measured in joules.
- Does the sign make sense? Temperature increased so Q is positive.
- Is the magnitude realistic? Magnitudes of thousands of joules are typical of solids with masses around 1 kg and temperature changes of tens of kelvins.

Practice Problems

5. How much heat is absorbed by 60.0 g of copper when its temperature is raised from 20.0°C to 80.0°C?

6. 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 are added?

b. Suppose it is winter and the 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.

Heating Up

Problem

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

Materials

hot plate (or Bunsen burner)
250-mL ovenproof glass beaker
water
thermometer
stopwatch
goggles
apron

Procedure

1. Turn your hot plate to a medium setting (or as recommended by your teacher). Allow a few minutes for the plate to heat up. Wear goggles.
2. Pour 150 mL of room temperature water into the 250-mL beaker.
3. Make a data and observations table.
4. Record the initial temperature of the water. The thermometer must not touch the bottom or sides of the beaker.
5. Place the beaker on the hot plate and record the temperature every 1.0 minute. Carefully stir the water before taking a temperature reading.
6. Record the time when the water starts to boil. Continue recording the temperature for an additional 4.0 minutes.
7. Carefully remove the beaker from the hot plate. Record the temperature of the remaining water.
8. When you have completed the lab, dispose of the water as instructed by your teacher. Allow equipment to cool before putting it away.

Data and Observations

Time	Temperature

Analyze and Conclude

1. **Analyzing Data** Make a graph of temperature (vertical axis) versus time (horizontal axis). Use a computer or calculator to construct the graph, if possible. What is the relationship between variables?
2. **Interpreting Graphs** What is the slope of the graph for the first 3.0 minutes? Be sure to include units.
3. **Relating Concepts** What is the thermal energy given to the water in the first 3.0 minutes? **Hint:** $Q = mC\Delta T$.
4. **Making Predictions** Use a dotted line on the same graph to predict what the graph would look like if the same procedure was followed with only half as much water.

Apply

1. Would you expect that the hot plate transferred energy to the water at a steady rate?
2. Where is the energy going when the water is boiling?

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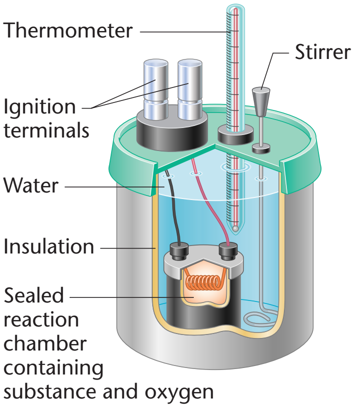


FIGURE 12–8 A calorimeter provides a closed, isolated system in which to measure energy transfer.

Calorimetry: Measuring Specific Heat

A **calorimeter,** shown in **Figure 12–8,** is a device used to measure changes in thermal energy. A calorimeter is carefully insulated so that heat transfer is very small. A measured mass of a substance is placed in the calorimeter and heated to a high temperature. The calorimeter contains a known mass of cold water at a measured temperature. The heat released by the substance is transferred to the cooler water. From the resulting increase in water temperature, the change in thermal energy of the substance is calculated.

The operation of a calorimeter depends on the conservation of energy in isolated, closed systems. Energy can neither enter nor leave an isolated system. As a result of the isolation, if the energy of one part of the system increases, the energy of another part must decrease by the same amount. Consider a system composed of two blocks of metal, block A and block B, as in **Figure 12–9a.** The total energy of the system is constant.

Conservation of Energy in a Calorimeter $E_A + E_B = constant$

Suppose that the two blocks are initially separated but can be placed in contact. 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 following equation.

$$\Delta E_A + \Delta E_B = 0$$

Which means that the following relationship is true.

$$\Delta E_A = -\Delta E_B$$

The change in energy of one block is positive, while the change in energy of the other block is negative. If the thermal energy change is positive, the temperature of that block rises. If the 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 12–9b.** The flow continues until the blocks are in thermal equilibrium. The blocks then have the same temperature.

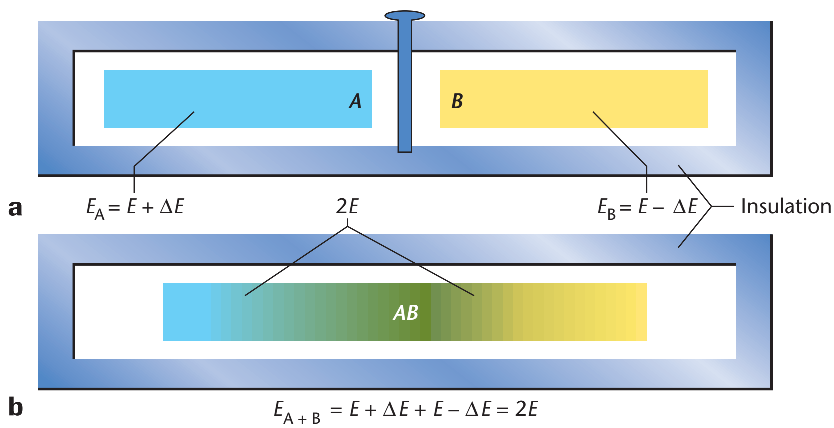


FIGURE 12–9 The total energy for this system is constant.

In a calorimeter, the change in thermal energy is equal to the heat transferred because no work is done. Therefore, the change in energy 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 equation for the transfer of energy is

$$m_A C_A (T_f - T_{Ai}) + m_B C_B (T_f - T_{Bi}) = 0.$$

To solve for T_f, expand the equation.

$$m_A C_A T_f - m_A C_A T_{Ai} + m_B C_B T_f - m_B C_B T_{Bi} = 0$$

$$T_f (m_A C_A + m_B C_B) = m_A C_A T_{Ai} + m_B C_B T_{Bi}.$$

$$T_f = \frac{m_A C_A T_{Ai} + m_B C_B T_{Bi}}{m_A C_A + m_B C_B}$$

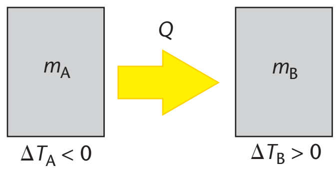
F.Y.I.

For many years, thermal energy was measured in calories. Calories are not part of the SI system of measurements, so today, thermal energy is measured in joules. One calorie is equal to 4.18 joules.

Example Problem

Heat Transfer 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?



Sketch the Problem

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

Calculate Your Answer

Known:

Zinc $m_A = 0.040$ kg, $C_A = 388$ J/kg·°C, $T_{Ai} = 115$°C

Water $m_B = 0.50$ kg, $C_B = 4180$ J/kg·°C, $T_{Bi} = 15$°C

Unknown:

$T_f = ?$

Strategy: Determine final temperature using $T_f = \frac{m_A C_A T_{Ai} + m_B C_B T_{Bi}}{m_A C_A + m_B C_B}$.

Calculations: $$T_f = \frac{(0.040\text{ kg})(388\text{ J/kg·°C})(115\text{°C}) + (0.50\text{ kg})(4180\text{ J/kg·°C})(15\text{°C})}{(0.040\text{ kg})(388\text{ J/kg·°C}) + (0.50\text{ kg})(4180\text{ J/kg·°C})}$$

$$= \frac{(1.78 \times 10^3 + 3.14 \times 10^4)\text{J}}{(15.5 + 2.09 \times 10^3)\text{J/°C}} = 16\text{°C}$$

Continued on next page

Check Your Answer

- Are the units correct? Temperature is measured in °C.
- Is the magnitude realistic? The answer is between the initial temperatures of the two samples and closer to water.

Pocket Lab

Melting

Label two foam cups A and B. Measure 75 mL of room-temperature water into each of the two cups. Add an ice cube to cup A. Add ice water to cup B until the water levels are equal. Measure the temperature of each cup at one minute intervals until the ice has melted.

Analyze and Conclude Do the samples reach the same final temperature? Why?

Practice Problems

7. 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 no heat loss to the surroundings. What is the final temperature of the mixture?

8. 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?

9. A 1.00×10^2-g brass block at 90.0°C is placed in a plastic foam cup containing 2.00×10^2 g of water at 20.0°C. No heat is lost to the cup or the surroundings. Find the final temperature of the mixture.

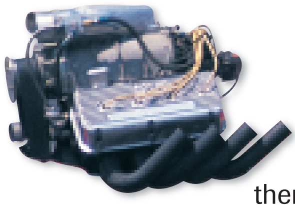
10. A 1.00×10^2-g aluminum block at 100.0°C is placed in 1.00×10^2 g of water at 10.0°C. The final temperature of the mixture is 25.0°C. What is the specific heat of the aluminum?

12.1 Section Review

1. Could the thermal energy of a bowl of hot water equal that of a bowl of cold water? Explain.
2. On cold winter nights before central heating existed, people often placed hot water bottles in their beds. Why would this be more efficient than warmed bricks?
3. If you take a spoon out of a cup of hot coffee and put it in your mouth, you are not likely to burn your tongue. But, you could very easily burn your tongue if you put the hot coffee in your mouth. Why?
4. **Critical Thinking** You use an aluminum cup instead of a plastic foam cup as a calorimeter, allowing heat to flow between the water and the environment. You measure the specific heat of a sample by putting the hot object into room temperature water. How might your experiment be affected? Would your result be too large or too small?

Change of State and Laws of Thermodynamics

12.2



If you rub your hands together, you exert a force and move your hands over a distance. You do work against friction. Your hands start and end at rest, so there is no net change in kinetic energy. They remain the same distance above Earth, so there is no change in potential energy. Yet, if the law of conservation of energy is true, then the energy transferred by the work you did must have gone somewhere. You notice that your hands feel warm; their temperature has increased. The energy to do the work against friction has changed form to thermal energy.

OBJECTIVES

- **Define** heats of fusion and vaporization.
- **State** the first and second laws of thermodynamics.
- **Define** heat engine, refrigerator, and heat pump.
- **Define** entropy.

Change of State

The three most common states of matter are solids, liquids, and gases, as shown in **Figure 12–10.** As the temperature of a solid is raised, it will usually change to a liquid. At even higher temperatures, it will become a gas. How can we explain these changes? Consider a material in a solid state. Your simplified model of the solid consists of tiny particles bonded together by massless springs. These massless springs represent the electromagnetic forces between the particles. When the thermal energy of a solid is increased, the motion of the particles is increased and the temperature increases.

FIGURE 12–10 The three states of water are represented in this photograph. The gaseous state, water vapor, is dispersed in the air and is invisible until it condenses.

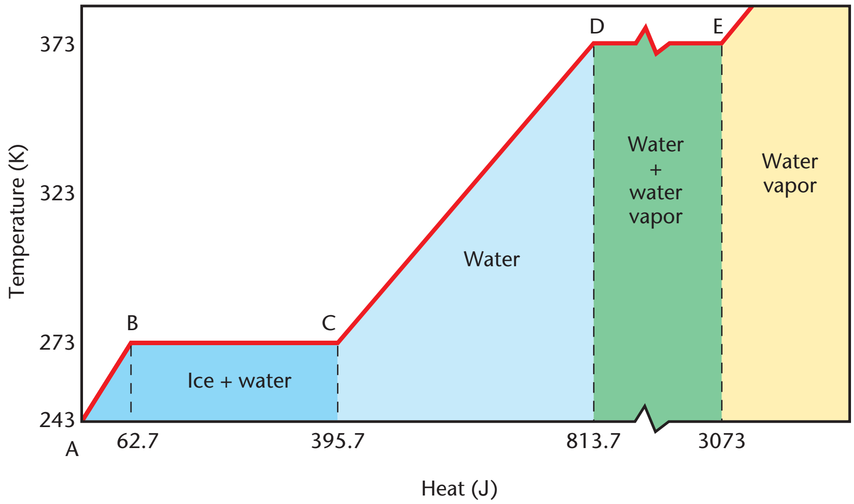


FIGURE 12–11 A plot of temperature versus heat added when 1 g of ice is initially converted to steam.

HELP WANTED

HVAC TECHNICIAN

Knowledge of thermodynamics; ability to read blueprints, specifications and manuals; familiarity with current products, procedures, tools, and test equipment; and ability to work hard in sometimes physically demanding conditions make you an ideal candidate for this position. Knowledge of all kinds of heating, ventilation, and air conditioning systems is required. Work your way up into other positions based on your work performance and completion of further training. For information contact:
Associated Builders and Contractors
1300 N. Seventeenth Street
Suite 800
Rosslyn, VA 22209

Figure 12–11 diagrams this process throughout all the changes of state as thermal energy is added to 1.0 g of H_2O starting at an initial temperature of 243 K and continuing until the temperature is 473 K. 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, but they have more freedom of movement. Eventually, the particles become free enough to slide past each other. At this point, the substance has changed from a solid to a liquid. The temperature at which this change occurs is the **melting point.** 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 the diagram where the added thermal energy melts the ice at a constant 273 K. Since the kinetic energy of the particles does not increase, the temperature does not increase here.

Once the solid is completely melted, there are no more forces holding the particles in the solid state and the added thermal energy again increases the motion of the particles, and the temperature rises. In the diagram, this is between points C and D. As the temperature increases, some particles in the liquid acquire enough energy to break free from other particles. At a specific temperature, known as the **boiling point,** further addition of energy causes another change of state where all the added thermal energy converts the material from the liquid state to the gas state. The motion of the particles does not increase so the temperature is not raised during this transition. In the diagram, this transition is represented between points D and E. After the material is entirely converted to the gas, any added thermal energy again increases the motion of the particles and the temperature rises. Above point E , steam is heated to a temperature of 473 K.

The amount of energy needed to melt one kilogram 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 from point B to point C represents the heat of fusion in **Figure 12–11.**

At normal atmospheric pressure, water boils at 373 K. The thermal energy needed to vaporize one kilogram of a liquid is called the **heat of vaporization.** For water, the heat of vaporization is 2.26×10^6 J/kg. Each substance has a characteristic heat of vaporization. The distance from point D to point E represents the heat of vaporization.

Between points A and B in the diagram, 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 that of ice or steam. 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 equation,

Heat Required to Melt $Q = mH_f$

where H_f is the heat of fusion. Similarly, the heat, Q, required to vaporize a mass, m, of liquid is given by the equation,

Heat Required to Vaporize $Q = mH_v$

where H_v is the heat of vaporization. The values of some heats of fusion, H_f, and heats of vaporization, H_v, can be found in **Table 12–2.**

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 the heat is transferred from the sample to the environment. In the same way, when a vapor condenses to a liquid, an amount of heat, $Q = -mH_v$, must be removed.

TABLE 12–2

Heats of Fusion and Vaporization of Some 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

Pocket Lab

Cool Times

Place a 100-mL beaker in a 250-mL beaker. Put a thermometer in each beaker. Fill the small beaker with hot, colored water. Determine the temperature of the colored water. Slowly pour tap water into the large beaker until the water is at the same height in both beakers. Record the temperature in the large beaker. Record the temperature in both beakers every minute for five minutes. Plot your data for both beakers on a graph of temperature versus time. Measure and record the mass of water in each beaker.

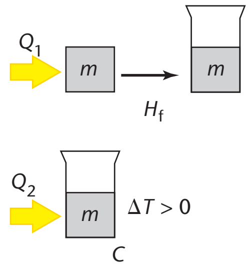
Calculate and Conclude Predict the final temperature. Describe each curve.

Example Problem

Melting a Solid and Warming the Resulting Liquid

You are asked to melt 0.100 kg of ice at its melting point and warm the resulting water to 20.0°C. How much heat is needed?

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 increased.

Calculate Your Answer

Known:

$m = 0.100$ kg

$H_f = 3.34 \times 10^5$ J/kg

$T_i = 0.0°C$

$T_f = 20.0°C$

$C = 4180$ J/kg·°C

Unknown:

$Q_1 + Q_2 = ?$

Strategy:

Calculate heat needed to melt ice.

Calculations:

$$Q_1 = mH_f = (0.100\ \text{kg})(3.34 \times 10^5\ \text{J/kg}) = 3.34 \times 10^4\ \text{J} = 33.4\ \text{kJ}$$

Calculate the temperature change.

$$\Delta T = T_f - T_i = 20.0°\text{C} - 0.0°\text{C} = 20.0°\text{C}$$

Calculate heat needed to raise water temperature.

$$Q_2 = mC\Delta T = (0.100\ \text{kg})(4180\ \text{J/kg}\cdot°\text{C})(20.0°\text{C}) = 8.36 \times 10^3\ \text{J} = 8.36\ \text{kJ}$$

Add all heats together to get total heat needed.

$$Q_1 + Q_2 = 33.4\ \text{kJ} + 8.36\ \text{kJ} = 41.8\ \text{kJ}$$

Check Your 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 about four times greater than the heat needed to increase the water temperature. It takes more energy to overcome the forces holding the particles in the solid state than to raise the temperature of water.

Practice Problems

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

12. 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?

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

14. A 175-g lump of molten lead at its melting point, 327°C, is dropped into 55 g of water at 20.0°C.

a. What is the temperature of the water when the lead becomes solid?

b. When the lead and water are in thermal equilibrium, what is the temperature?

CHEMISTRY CONNECTION

Energy Released or Absorbed Hot packs keep your hands warm, and cold packs keep your soda chilled. In both packs, a thin membrane separates water from a chamber of salt. In a hot pack, the salt might be calcium chloride; in a cold pack, ammonium nitrate. Squeezing the pack breaks the membrane. As the salt dissolves in the water, the solution either releases or absorbs energy. What form does this energy take?

The First Law of Thermodynamics

There are additional means of changing the amount of thermal energy in a system. If you use a hand pump to inflate a bicycle tire, the air and pump become warm. The mechanical energy in the moving piston is converted into thermal energy of the gas. Other forms of energy—light, sound, and electric—can be changed into thermal energy. Some examples include a toaster, which converts electric energy into heat to cook bread, and the sun, which you have already learned warms Earth with light from a distance of over 150 million kilometers.

Thermal energy can be increased either by adding heat or by doing work on a system. Thus, the total increase in the thermal energy of a system is the sum of the work done on it and the heat added to it. This fact is called the **first law of thermodynamics.** Recall that thermodynamics is 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.

The conversion of mechanical energy to thermal energy, as when you rub your hands together, is easy. The reverse process, conversion of thermal to mechanical energy, is more difficult. A device able to convert thermal energy to mechanical energy continuously is called a **heat engine.**

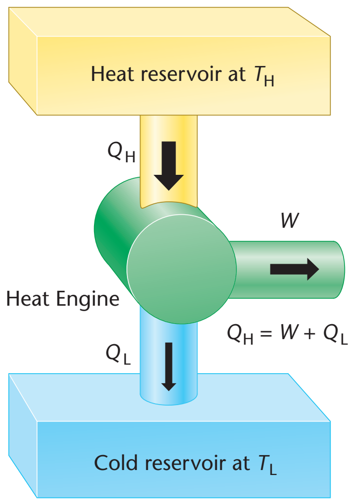
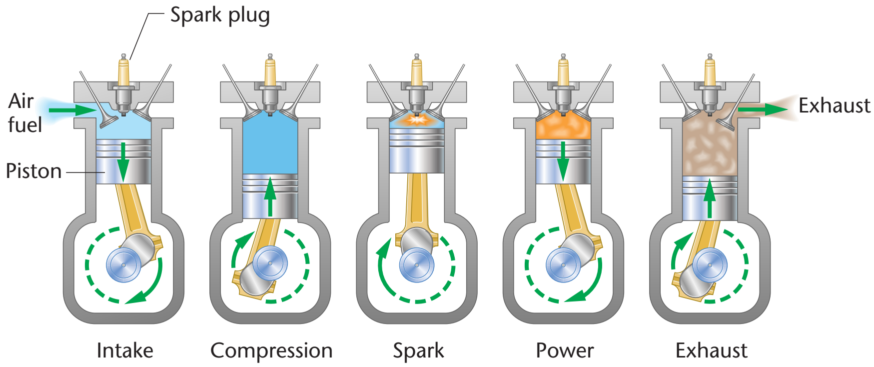


FIGURE 12–12 This diagram represents heat at high temperature transformed into mechanical energy and low-temperature waste heat.

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 is presented in **Figure 12–12.** An automobile internal combustion engine is one

FIGURE 12–13 The heat produced by burning gasoline causes the gases produced to expand and exert force and do work on the piston.



example of a heat engine and is shown in **Figure 12–13.** In the automobile, a mixture of air and gasoline vapor is ignited, producing a high-temperature flame. Heat flows from the flame to the air in the cylinder. The hot air expands and pushes on a piston, changing thermal energy into mechanical energy. In order 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. The entire cycle is repeated many times each minute. The thermal energy from the gasoline is converted into mechanical energy that eventually results in the movement of the car.

Not all the thermal energy from the high-temperature flame is converted into mechanical energy. The exhaust gases and the engine parts become hot. The exhaust comes in contact with outside air, transferring heat to it, and raising its temperature. Heat flow from the hot engine is transferred to a radiator. Outside air passes through the radiator and the

Perpetual Motion?

Answers question from page 272.

the outside of the refrigerator, the gas 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 resulting gas returns to the compressor and repeats the process. The overall change in the thermal energy of the gas is zero. Thus, according

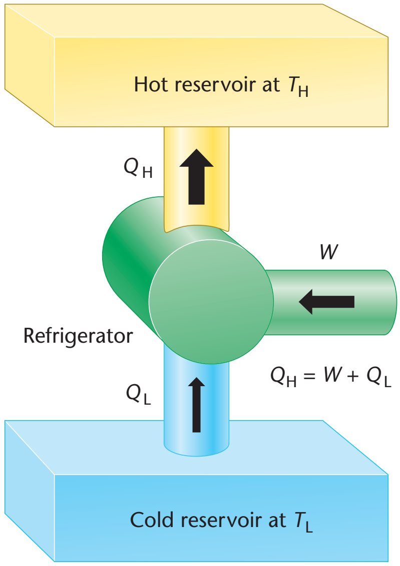


FIGURE 12–14 A refrigerator absorbs heat Q_L from the cold reservoir and gives off heat Q_H to the hot reservoir. Work, W, is done on the refrigerator.

to the first law of thermodynamics, the sum of the heat removed from the refrigerator contents and the work done by the motor is equal to the heat expelled to the outside air at a higher temperature, as is shown in **Figure 12–14.** A heat pump is a refrigerator that can be run in two directions. In summer, heat is removed from the house, and thus cooling the house. The heat is expelled into the warmer air outside. In 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.

The Second Law of Thermodynamics

Many processes that do not violate the first law of thermodynamics have never been observed to occur spontaneously. Three such processes are presented in **Figure 12–15.** For example, the first law of thermodynamics does not prohibit heat flowing from a cold body to a hot body. Still, when a hot body has been placed in contact with a cold body, the hot body has never been observed to become hotter and the cold body colder. Heat flows spontaneously from hot to cold bodies. Another example, heat engines could convert thermal energy completely into mechanical energy with no waste heat and the first law of thermodynamics would still be obeyed. Yet 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, the 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. Consider what happens when heat is added to an object. The particles

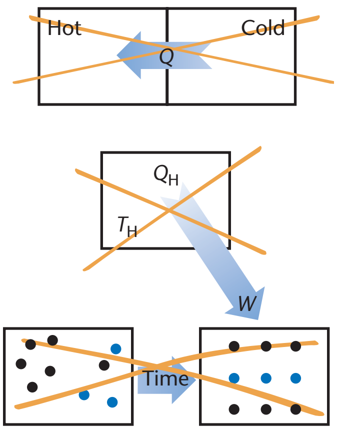


FIGURE 12–15 These three processes are not forbidden by the first law of thermodynamics, yet they do not occur spontaneously.

move in a random way. Some move very quickly, others move slowly, and many move at intermediate speeds. Because the particles are not all traveling at the same speed, disorder ensues. The greater the range of speeds exhibited by the particles, the greater the disorder. The greater the disorder, the larger the entropy. While it is theoretically possible that all the particles could have the same speed, the random collisions and energy exchanges of the particles make the probability of this extremely unlikely.

Entropy, like thermal energy, is contained in an object. If heat is added to a body, entropy is increased. If heat is removed from a body, 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 **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

Physics & Technology

Infrared Detectors

Infrared radiation (IR) is electromagnetic energy that radiates through space. Any object with a temperature above absolute zero gives off infrared energy. The higher the object's temperature, the more IR it emits. Instruments capable of detecting IR can measure the temperature of very cold or very distant objects. They can also be used to "see" objects in total darkness.

Rescue teams can use night-vision goggles to detect the infrared glow given off by people or animals, making them visible in the dark. IR security sensors trigger an alarm when an intruder is detected, and IR-sensitive medical equipment can produce thermal maps of the human body to help locate tumors or circulatory problems. Airborne IR cameras are used to locate the edges of wildfires obscured from view by thick smoke, or to study underground volcanic activity. IR detectors aboard weather satellites map cloud patterns by measuring temperature differences between Earth's surface and high-altitude clouds.

Perhaps the most far-reaching applications for IR detectors are in the fields of astronomy and cosmology. Highly sensitive detectors measure such tiny amounts of IR that they can be used to view objects in the universe that are not hot enough to emit visible light. These instruments can produce images of faraway stars and galaxies, search for planets orbiting nearby stars, and survey the most distant regions of the sky for clues about the origins of the universe. Because Earth's atmosphere absorbs almost all the IR, these studies are usually conducted using arrays of IR detectors aboard orbiting satellites. The detectors are thin layers of electrically conductive materials that absorb IR. This absorption causes changes in conductivity that are monitored with computers.

Thinking Critically Evaluate the impact of IR detector research on society and the environment. Highly sensitive IR detectors used by astronomers are supercooled to extremely low temperatures, sometimes as low as 3 or 4 K. Why is this necessary?

FIGURE 12–16 The spontaneous mixing of the food coloring and water is an example of the second law of thermodynamics.

some action is taken to keep them ordered. Entropy increase and the second law of thermodynamics can be thought of as statements of the probability of events happening. **Figure 12–16** 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 12–17** shows a familiar example of the second law of thermodynamics that most teenagers can readily recognize.

The second law of thermodynamics predicts that heat flows spontaneously only from a hot body to a cold body. 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 move more slowly. When you plunge the bar into the water and allow thermal equilibrium to be reached, the average kinetic energy of the particles in the iron and the water will be the same. More particles now have a more random motion than they did at the beginning of your test. This final state is less ordered than the initial state. No longer are the fast particles confined solely to the iron and the slower particles confined to the water. All speeds are evenly distributed. The entropy of the final state is larger than that of the initial state.

You take for granted many daily events that occur spontaneously, or naturally, in one direction, but that would shock you if they happened in reverse. You are not surprised when a metal spoon, heated at one end,

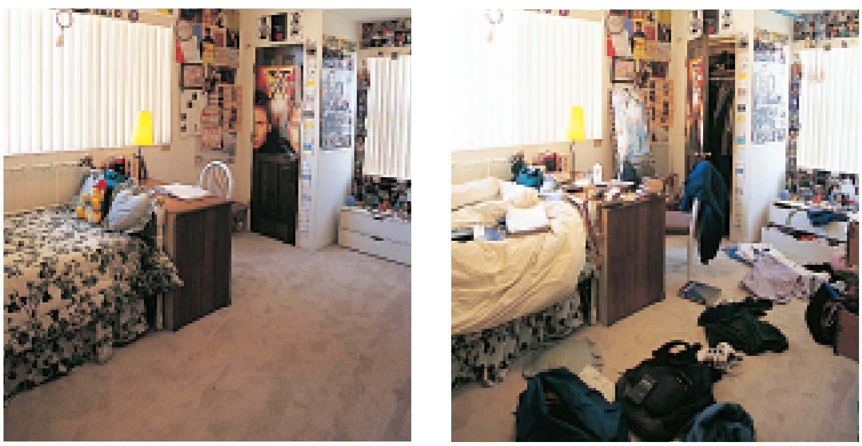


FIGURE 12–17 If no work is done on a system, entropy spontaneously reaches a maximum.

Pocket Lab

Drip, Drip, Drip

Measure equal amounts of very hot and very cold water into two clear glasses (or beakers).

Hypothesize and Test Predict what will happen if you simultaneously put one drop of food coloring in each glass. Try it. What happened? Why? Was the mixing symmetric?

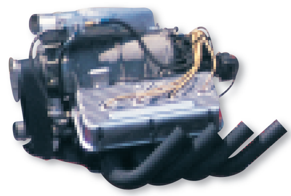
soon becomes uniformly hot, or when smoke from a too-hot frying pan diffuses throughout the kitchen. Consider your reaction, however, if a spoon lying on a table suddenly, on its own, became red hot at one end and icy cold at the other, or if all the smoke from the frying pan collected in a 9-cm^3 cube in the exact center of the kitchen. Neither of the reverse processes violates the first law of thermodynamics. These are simply examples of the countless events that do not occur because their processes would violate the second law of thermodynamics.

The second law of thermodynamics and increase in entropy also give new meaning to what has been commonly called the "energy crisis," 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. The internal chemical energy contained in the molecules of the gas is converted into thermal energy of the flame, which 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 been increased. The chemical structure of natural gas is very ordered. In contrast, the random motion of the warmed air is very disordered. While it is mathematically possible for the original order to be reestablished, the probability of this occurring is essentially zero. For this reason, entropy is often used as a measure of the unavailability of useful energy. The energy in the warmed air in a house is not as available to do mechanical work or to transfer heat to other bodies as the original gas molecules. The lack of usable energy is really a surplus of entropy.

12.2 Section Review

1. Old-fashioned heating systems sent steam into radiators in each room. In a radiator, the steam condenses back to water. Analyze this process and explain how it heats the room.
2. James Joule carefully measured the difference in temperature of water at the top and bottom of a waterfall. Why did he expect a difference?
3. A man uses a 320-kg hammer moving at 5 m/s to smash a 3-kg block of lead against a 450-kg rock. He finds that the temperature of the lead increased by 5°C. Explain how this happens.
4. Evaluate why heating a home with natural gas results in an increased amount of disorder.
5. **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 the original order? Explain. Of what physical law is this an example?

CHAPTER 12 REVIEW



Summary

Key Terms

12.1
- thermodynamics
- kinetic-molecular theory
- thermal energy
- temperature
- conduction
- thermal equilibrium
- thermometer
- absolute zero
- kelvin
- heat
- convection
- radiation
- specific heat
- calorimeter

12.2
- melting point
- boiling point
- heat of fusion
- heat of vaporization
- first law of thermodynamics
- heat engine
- entropy
- second law of thermodynamics

12.1 Temperature and Thermal Energy

- 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 the particles.
- Thermometers reach thermal equilibrium with the objects they contact, then their temperature-dependent property is measured.
- The Celsius and Kelvin temperature scales are used in scientific work. The magnitude of one kelvin is equal to the magnitude of one Celsius degree.
- At absolute zero, no more thermal energy can be removed from a substance.
- Heat is energy transferred because of a difference in temperature.
- Specific heat is the quantity of heat required to raise the temperature of one kilogram of a substance by one kelvin.
- 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.

12.2 Change of State and Laws of Thermodynamics

- The heat of fusion is the quantity of heat needed to change one kilogram of a substance from its solid to liquid state at its melting point.
- The heat of vaporization is the quantity of heat needed to change one kilogram of a substance from its liquid to gaseous state at its boiling point.
- Heat transferred during a change of state doesn't change the temperature.
- The total increase in energy of a system is the sum of the heat added to it and the work done on it.
- A heat engine continuously converts thermal energy to mechanical energy.
- A heat pump or refrigerator uses 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 entropy of the universe always increases, even if the entropy of a system may decrease because of some action taken to increase its order.

Key Equations

12.1

$$Q = mC\Delta T = mC(T_{\text{final}} - T_{\text{initial}})$$

$$E_A + E_B = constant$$

12.2

$$Q = mH_f$$

$$Q = mH_v$$

Reviewing Concepts

Section 12.1

1. Explain the difference between the mechanical energy of a ball, its thermal energy, and its temperature. Give an example.
2. Can temperature be assigned to a vacuum? Explain.
3. Do all of the molecules or atoms in a liquid have the same speed?

4. Your teacher just told your class that the temperature of the core of the sun is 1.5×10^7 degrees.
   a. Sally asks whether this is the Kelvin or Celsius scale. What will be the teacher's answer?
   b. Would it matter whether you use the Celsius or Fahrenheit scale?
5. Is your body a good judge of temperature? On a cold winter day, a metal doorknob feels much colder to your hand than the wooden door. Explain why this is true.
6. Do we ever measure heat transfer directly? Explain.
7. How does heat flow when a warmer object is in contact with a colder object? Do the two have the same temperature changes?

Section 12.2

8. Can you add thermal energy to an object without increasing its temperature? Explain.
9. When wax freezes, is energy absorbed or released by the wax?
10. Analyze and explain why water in a canteen stays cooler if it has a canvas cover that is kept wet.
11. Which process occurs at the coils of a running air conditioner inside the house, vaporization or condensation? Explain.
12. A pot sitting on a warm burner will initially only be hot where it contacts the burner. Over time the entire pot heats up to the same temperature. Evaluate this situation and explain how it results in an increase in disorder.

Applying Concepts

13. 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?
14. On the following pairs of scales, is there one temperature on each scale that has the same value? $T_F = 9/5\ T_C + 32$
    a. Celsius and Fahrenheit
    b. Kelvin and Fahrenheit
    c. Celsius and Kelvin
15. Which liquid would an ice cube cool faster, water or methanol? Explain.
16. Explain why the high specific heat of water makes it desirable for use in hot-water heating systems.
17. 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.
18. Why do easily vaporized liquids, such as acetone and methanol, feel cool to the skin?
19. Explain why fruit growers spray their trees with water, when frost is expected, to protect the fruit from freezing.
20. 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 temperature. Will the two cups of water have equal temperatures after equilibrium is achieved? Explain.

Problems

Section 12.1

21. Liquid nitrogen boils at 77K. Find this temperature in degrees Celsius.
22. The melting point of hydrogen is $-259.14°C$. Find this temperature in kelvins.
23. How much heat is needed to raise the temperature of 50.0 g of water from 4.5°C to 83.0°C?
24. 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.
25. A 4.00×10^2-g glass coffee cup is at room temperature, 20.0°C. It is then plunged into hot dishwater, 80.0°C. If the temperature of the cup reaches that of the dishwater, how much heat does the cup absorb? Assume the mass of the dishwater is large enough so its temperature doesn't change appreciably.
26. A 1.00×10^2-g mass of tungsten at 100.0°C is placed in 2.00×10^2g of water at 20.0°C. The mixture reaches equilibrium at 21.6°C. Calculate the specific heat of tungsten.
27. A 6.0×10^2-g sample of water at 90.0°C is mixed with 4.00×10^2g of water at 22.0°C. Assume no heat loss to the surroundings. What is the final temperature of the mixture?
28. A 10.0-kg piece of zinc at 71.0°C is placed in a container of water. The water has a mass of 20.0 kg and has a temperature of 10.0°C before the zinc is added. What is the final temperature of the water and zinc?

29. To get a feeling for the amount of energy needed to heat water, recall from **Table 11–1** that the kinetic energy of a compact car moving at 100 km/h is 2.9×10^5 J. What volume of water (in liters) would 2.9×10^5 J of energy warm from room temperature (20°C) to boiling (100°C)?

30. A 3.00×10^2-W electric immersion heater is used to heat a cup of water. 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.

31. A 2.50×10^2-kg cast-iron car engine contains water as a coolant. Suppose the engine's temperature is 35.0°C when it is shut off. 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 4.4×10^6 J. What mass of water is used to cool the engine?

Section 12.2

32. 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 delivered. As it melted, how much heat did a block of ice that size absorb?

33. 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?

34. 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?

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

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

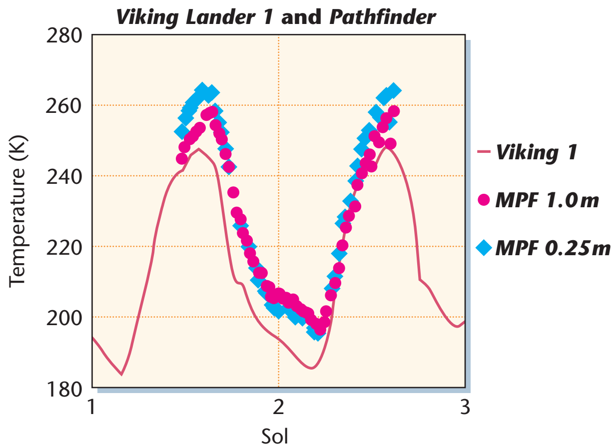
37. A soft drink from Australia is labeled "Low Joule Cola." The label says "100 mL yields 1.7 kJ." The can contains 375 mL. Sally drinks the cola and then wants to offset this input of food energy by climbing stairs. How high would Sally have to climb if she has a mass of 65.0 kg?

Extra Practice For more practice solving problems, go to Extra Practice Problems, Appendix B.

Critical Thinking Problems

38. Your mother demands that you clean your room. Reducing the disorder of your room will reduce its entropy, but the entropy of the universe cannot be decreased. Evaluate how you increase entropy as you clean.

Going Further



Interpreting Graphs The daily cycle of temperatures for the first two robot explorers of Mars is shown plotting the temperature versus time of the solar day starting at midnight. The points labeled "MPF" were recorded by the *Mars Pathfinder* in 1997 at various distances above the ground. Interpret this graph by finding the lowest night temperatures and highest day temperatures recorded by each planetary probe on both the Celsius and Fahrenheit scales.

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