

◀ 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?

What invention has changed our lives more than the internal combustion engine? In the 1700s, the steam engine changed our society from a nation of farms to one of factories. The steam engine burned fuel to change water to steam in a boiler outside the engine. The gasoline engine, invented in Germany by Otto in 1876, burns the fuel inside the engine, and thus is an internal combustion engine. These engines and their descendants, the diesel engines and the turbines used in jet aircraft, change the chemical energy in fuel to thermal energy in hot gases. The thermal energy, in turn, is converted into kinetic energy of motion. Internal combustion engines vary in size from the small engines in lawnmowers and motorbikes, to larger ones in cars, buses, and airplanes, to the giant engines that produce electricity for millions of people. Consider how often we make use of this invention in our everyday lives.

Chapter Outline

12.1 TEMPERATURE AND THERMAL ENERGY

- What Makes a Hot Body Hot?
- Thermal Energy Transfer
- Thermal Energy and Temperature
- Equilibrium and Thermometry
- Temperature Scales: Celsius and Kelvin
- Heat and Thermal Energy
- Calorimetry: Measuring Specific Heat

12.2 CHANGE OF STATE AND LAWS OF THERMODYNAMICS

- Change of State
- The First Law of Thermodynamics
- The Second Law of Thermodynamics

✓ Concept Check

The following terms or concepts from earlier chapters are important for a good understanding of this chapter. If you are not familiar with them, you should review them before studying this chapter.

- work and energy, Chapter 10
- energy conservation, Chapter 11

Objectives

- understand the nature of thermal energy as explained by the kinetic theory.
- define temperature and distinguish it from thermal energy.
- understand the process of reaching equilibrium and its application to the measurement of temperature.
- describe and use the Celsius and Kelvin temperature scales; demonstrate the ability to convert between Celsius and Kelvin.
- distinguish heat from thermal energy.
- define and display an understanding of specific heat; be able to calculate heat transfer.
- display an understanding of the application of conservation of energy to heat transfers; be able to calculate temperature changes due to heat transfer.

12.1 TEMPERATURE AND THERMAL ENERGY

In the 1600s and 1700s, Europe went through a “Little Ice Age” when temperatures were lower than any other period in the last one thousand years. Keeping warm was vitally important. As a result, 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 produce useful work. Although not as convenient as the internal combustion engine, these machines freed society from its dependence on the energy of 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?

To operate, internal combustion engines require very high temperatures—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 was heated, an invisible fluid called “caloric” was added to the body. Hot bodies contained more caloric than cold bodies. The caloric theory could explain observations such as the expansion of objects when heated, but it could not easily explain why hands get warm when they are rubbed together.

In the mid-nineteenth century, scientists developed a new theory to replace 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 a higher energy than particles in a cooler body. The theory is called the **kinetic-molecular theory**.

The kinetic theory is difficult to visualize because the individual particles are too tiny to be seen. A thrown baseball, Figure 12–1, has a kinetic energy that depends on its velocity and a potential energy that is proportional to its height above the ground. These external properties can be seen. The tiny particles that make up the baseball, however, are

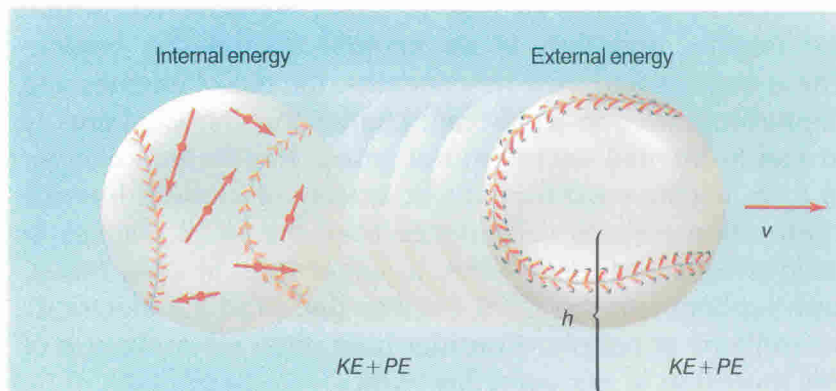


FIGURE 12–1. A baseball in flight has both internal and external energy. The internal energy is the result of the kinetic and potential energies of its particles. The external energy is the result of the position and motion of the baseball in flight.

in constant motion within the ball. This internal motion is invisible under most circumstances.

The model of a solid, Figure 12–2, can help you understand the kinetic theory. This model pictures 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 sum of the kinetic and potential energies of the internal motion of particles that make up an object is called the internal energy or **thermal energy** of that object.

Thermal Energy Transfer

Thermal energy is transferred in three ways. **Conduction**, most common in solids, involves transfer of kinetic energy when the particles of an object collide. It is the principle on which the common household thermometer operates. The movement of fluids caused by their different densities at different temperatures transfers heat by **convection**. Convection currents in the atmosphere are responsible for much of Earth’s weather. Both conduction and convection depend on the presence of matter. The third method of transfer, **radiation**, does not. Thermal energy can be transferred through space in the form of electromagnetic waves. Solar energy is transmitted to Earth by radiation.

Thermal Energy and Temperature

According to the kinetic-molecular theory, a hot body has more thermal energy than a similar cold body, Figure 12–3. This means that, as a whole, the particles in a hot body have larger kinetic and potential energies 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 sixth-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 fourth-grade class, even though some fourth-graders might be taller than some sixth-graders.

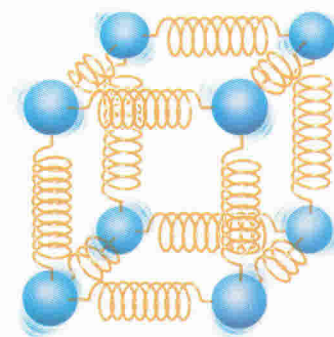
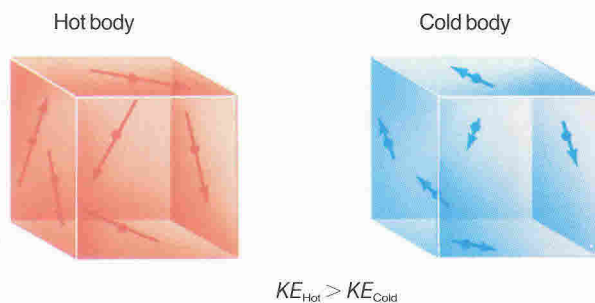


FIGURE 12–2. Molecules of a solid behave in some ways as if they were held together by springs.

FIGURE 12–3. Particles in a hot body have larger kinetic and potential energies than particles in a cold body.

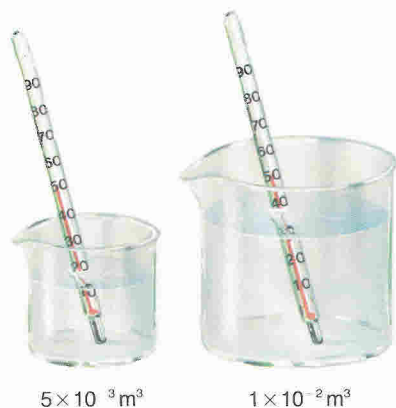


FIGURE 12-4. Temperature does not depend on the number of particles in a body.

Temperature is hotness measured on some definite scale.

The temperature of a gas is proportional to the average kinetic energy of the particles.

Temperature does not depend on the mass of the object; thermal energy does.

How can we measure the “hotness” of an object? Hotness, measured on some definite scale, is a property of an object called its **temperature**. In a hotter object, the particles are moving faster; they have a larger average kinetic energy. For gases, the temperature is proportional to the average kinetic energy of the particles. For solids and liquids, this is only approximately true. For any form of matter, the temperature does not depend on the number of particles in the body. If a one-kilogram mass of steel is at the same temperature as a two-kilogram mass, the average kinetic energy of the particles in both masses is the same. The total amount of kinetic energy of particles in the two-kilogram mass, however, is twice the amount in the one-kilogram mass. The thermal energy in an object is proportional to the number of particles in it, but its temperature is not, Figure 12-4.

Equilibrium and Thermometry

You are familiar with the idea of measuring 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 the particles in your body have higher thermal energy. The thermometer is made of a glass tube. When the cold glass touches your hotter body, the particles in your body hit the particles in the glass. These collisions transfer energy to the glass particles by conduction. The thermal energy of the particles that make up the thermometer increases. 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 is equal. Your body and the thermometer are in **thermal equilibrium**. That is, the rate at which energy that flows from your body to the glass is equal to the rate of flow from the glass to your body. The thermometer and your body are at the same temperature. Objects that are in thermal equilibrium are at the same temperature, Figure 12-5. Note, however, that if the masses of the objects are different, they may not have the same thermal energy.

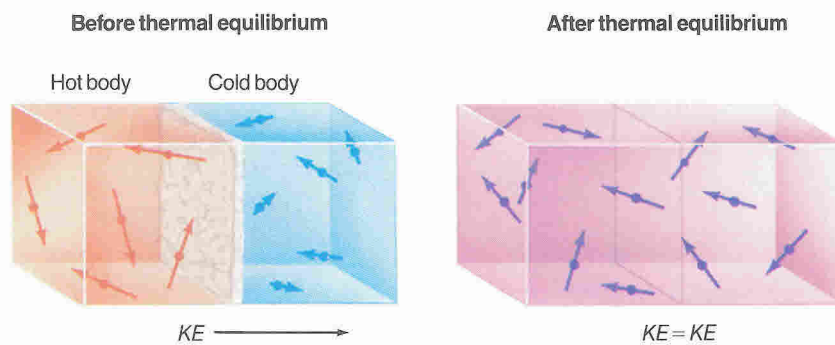
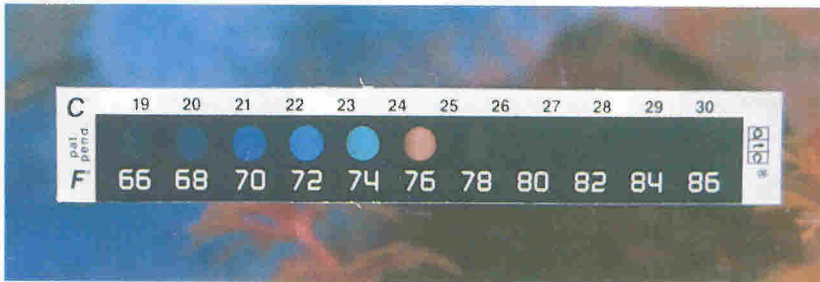


FIGURE 12-5. Thermal energy is transferred from a hot body to a cold body. When thermal equilibrium is reached, the transfer of energy between bodies is equal.



a



b

FIGURE 12–6. Thermometers use a change in physical properties to measure temperature. A liquid crystal thermometer changes color with temperature change (a). Some modern thermometers determine temperature electronically (b).

A **thermometer** is a device to measure 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 larger the volume of the alcohol in it and the higher it rises. Mercury is another liquid commonly used in thermometers.

Other properties of materials change with temperature, allowing them to be used as thermometers. In liquid crystal thermometers, Figure 12–6, the arrangement of the molecules changes at a specific temperature, changing the color of the crystal. As a result, the color depends on temperature. A set of different kinds of liquid crystals is used. Each changes color at a different temperature, creating an instrument that can indicate 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 (1704–1744). On this scale, now called the Celsius scale, Figure 12–7, 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 temperature of the human body is 37°C . Figure 12–7 shows representative temperatures on the three most common scales.

A thermometer measures the temperature of an object with which it is in thermal equilibrium.

Three Temperature Scales

$$\text{Celsius to Kelvin } ^{\circ}\text{C} + 273$$

$$\text{Celsius to Fahrenheit } (^{\circ}\text{C} \times 1.8) + 32$$

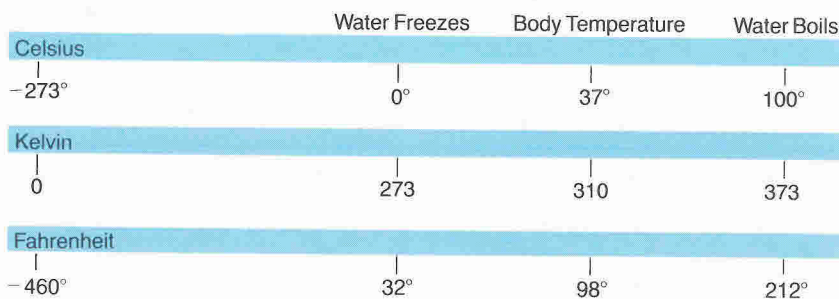


FIGURE 12–7. Some familiar reference temperatures on three temperature scales.

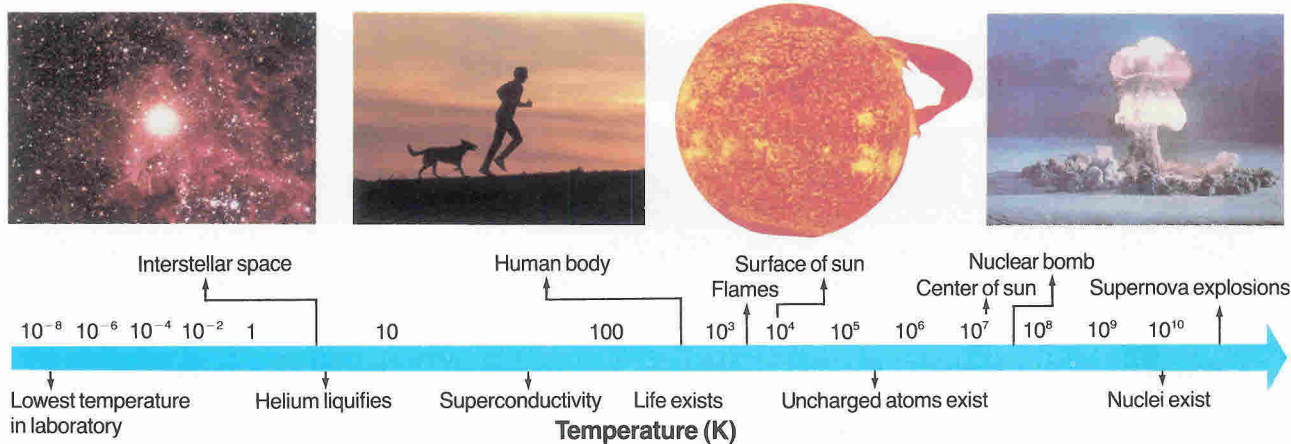


FIGURE 12–8. There is an extremely wide range of temperatures throughout the universe. Note that the scale has been expanded in areas of particular interest.

No temperature can be colder than absolute zero, -273.15°C or 0 K .

The wide range of temperatures in the universe is shown in Figure 12–8. Temperatures do not appear to have an upper limit. The interior of the sun is at least $1.5 \times 10^7^{\circ}\text{C}$. Other stars are even hotter. Temperatures do, however, have a lower limit. Generally, materials contract as they cool. If you cooled an “ideal” gas, one in which the particles have no volume and don’t interact, it would contract in such a way that it would have zero volume at -273.15°C . At this temperature, all the thermal energy would be removed from the gas. It would be impossible to reduce the thermal energy any further. Therefore, there can be no lower temperature than -273.15°C . This is called **absolute zero**.

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 273.15 K and the boiling point of water is 373.15 K . Each interval on this scale, called a **kelvin**, is equal to one Celsius degree. Thus, $^{\circ}\text{C} + 273.15 = \text{K}$.

Very cold temperatures are reached by liquefying gases. Helium liquefies at 4.2 K , or -269.0°C . Even colder temperatures can be reached by using the properties of special substances placed in the fields of large magnets. By using these techniques, physicists have reached temperatures of only $2.0 \times 10^{-9}\text{ K}$.

Example Problem

Converting Celsius to Kelvin Temperature

Convert 25°C to kelvins.

Solution: $\text{K} = ^{\circ}\text{C} + 273.15 = 25^{\circ} + 273.15 = 298\text{ K}$

F. Y. I.

For many years, thermal energy was measured in calories. One calorie is equal to 4.18 J . The calorie is not used in the SI system.

Example Problem

Converting Kelvin to Celsius Temperature

Convert the boiling point of helium, 4.22 K , to degrees Celsius.

Solution: $^{\circ}\text{C} = \text{K} - 273.15 = 4.22 - 273.15 = -268.93^{\circ}\text{C}$

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 these Celsius temperatures to Kelvin temperatures.
 - a. 27°C
 - b. 560°C
 - c. -184°C
 - d. -300°C
3. Convert these Kelvin temperatures to Celsius temperatures.
 - a. 110 K
 - b. 22 K
 - c. 402 K
 - d. 323 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

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 flows from the hotter object to the cooler object. **Heat is the energy that flows as a result of a difference in temperature.** We will use the symbol Q for heat. Heat, like any other form of energy, is measured in joules.

Note that this definition of heat is different from the one in everyday use. We commonly speak of a body containing “heat”. This description is left over from the caloric theory. As we now know, a hot body contains a larger amount of thermal energy than a colder body of the same size. Heat is the energy transferred because of a difference in temperature.

When heat flows into an object, its thermal energy increases, and so does its temperature. The amount of 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 raise the temperature of a unit mass one temperature unit. In SI units, specific heat, C , is measured in $\text{J}/\text{kg} \cdot \text{K}$. For example, 903 J must be added to one kilogram of aluminum to raise the temperature one kelvin. The specific heat of aluminum is $903\text{ J}/\text{kg} \cdot \text{K}$.

Note that water has a high specific heat compared to other substances, even ice and steam. One kilogram of water requires the addition of 4180 J of energy to increase its temperature one kelvin. By comparison, the same mass of copper requires only 385 J . The energy needed to raise the temperature of one kilogram of water 1 K would increase the temperature of the same mass of copper 11 K . The high specific heat of water is the reason water is used in car radiators to remove waste heat.

Specific heat can be used to find the amount of heat that must be transferred to change the temperature of a given mass by any amount. The specific heat of water is $4180\text{ J}/\text{kg} \cdot \text{K}$. When the temperature of one kilogram of water is increased by one kelvin, the heat absorbed by the water is 4180 J . When the temperature of 10 kilograms of water is increased by 5.0 K , the heat absorbed, Q , is

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

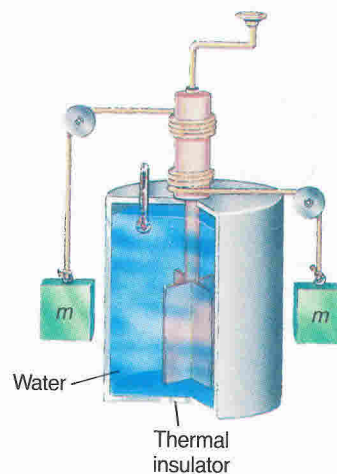


FIGURE 12–9. James Joule’s experiment for measuring the mechanical equivalent of heat. The falling weights rotate the paddles, causing the temperature of the water to rise.

Heat is thermal energy transferred because of a difference in temperature. Heat flows spontaneously from a warmer to a cooler body.

Specific heat is the increase in thermal energy that raises one kilogram of a substance one kelvin.

Table 12–1

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

The heat gained or lost by an object as its temperature changes depends on the mass, change in temperature, and specific heat of the substance. The relationship can be written

$$Q = mC\Delta T,$$

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

Example Problem

Heat Transfer

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

Given: mass, $m = 0.400$ kg

Unknown: Q

specific heat,

Basic equations: $\Delta T = T_f - T_i$

$C = 450$ J/kg · K

$Q = mC\Delta T$

$T_i = 295$ K

$T_f = 325$ K

Solution: $Q = mC\Delta T$

$= (0.400 \text{ kg})(450 \text{ J/kg} \cdot \text{K})(325 \text{ K} - 295 \text{ K})$

$= 5.4 \times 10^3 \text{ J}$



FIGURE 12–10. Brine solutions are evaporated in solar ponds to obtain potassium compounds. The blue dye added to the brine solutions absorbs solar radiation faster and speeds up evaporation.

Practice Problems

- How much heat is absorbed by 60.0 g of copper when it is heated from 20.0°C to 80.0°C?
- A 38-kg block of lead is heated from -26°C to 180°C . How much heat does it absorb during the heating?
- The cooling system of a car engine contains 20.0 L of water. (1 L of water has a mass of 1 kg.)
 - What is the change in the temperature of the water if the engine operates until 836.0 kJ of heat are added?
 - 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?
 - Which is the better coolant, water or methanol? Explain.
- A 565-g cube of iron is cooled from the temperature of boiling water to room temperature (20°C).
 - How much heat must be absorbed from the cube?
 - If the iron is cooled by dunking it into water at 0°C that rises in temperature to 20°C , how much water is needed?

PHYSICS LAB : Heating Up

Purpose

To discover how the temperature increases with a constant supply of thermal energy.

Materials

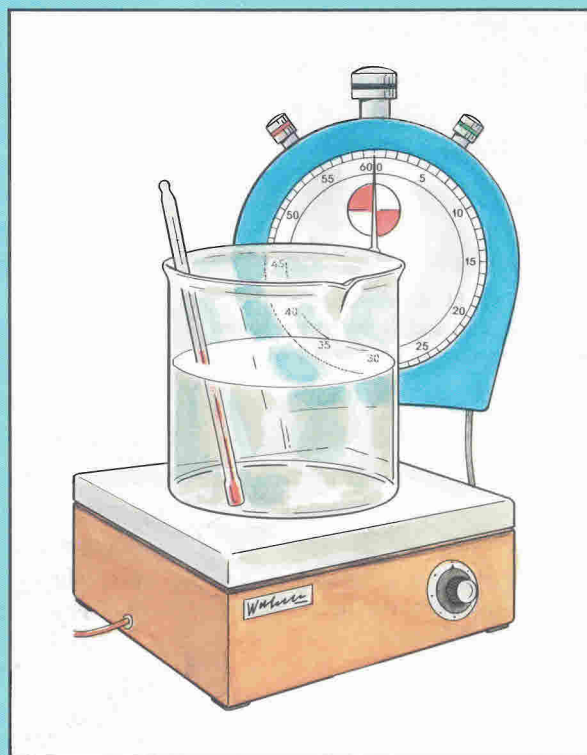
- hot plate (or bunsen burner)
- 250-mL pyrex beaker
- water
- thermometer
- stopwatch
- goggles

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. While heating water wear goggles.
2. Pour 150 ml of room temperature water into the 250-mL beaker.
3. Measure the initial temperature of the water. Keep the thermometer in the water.
4. Place the beaker on the hot plate and record the temperature every 1.0 minute. Carefully stir the water before taking a temperature reading.
5. Record the time when the water starts to boil. Continue recording the temperature for an additional 4.0 minutes.
6. Carefully remove the beaker from the hot plate. Record the remaining water.

Observations and Data

1. Make a graph of temperature (vertical axis) vs time (horizontal axis).
2. Make a generalization about the graph during the first few minutes of the experiment.



Analysis

1. What is the slope of the graph for the first 3.0 minutes? Be sure to include units.
2. What is the thermal energy given to the water in the first 3.0 minutes? **Hint:** $Q = mC\Delta T$.
3. Use a dotted line on the same graph to predict what the graph would look like if the same procedure was made with only half as much water.

Applications

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



FIGURE 12–11. A calorimeter provides a closed, isolated system in which to measure energy transfer.

The total energy of an isolated, closed system is constant. Energy lost by one part is gained by another.

Heat flows until all parts of the system are at the same temperature.

Calorimetry: Measuring Specific Heat

A **calorimeter**, Figure 12–11, is a device used to measure changes in thermal energy. A measured mass of a substance is heated to a known temperature and placed in the calorimeter. The calorimeter contains a known mass of cold water at a measured temperature. From the resulting increase in water temperature, the change in thermal energy of the substance is calculated.

The calorimeter depends on the conservation of energy in isolated, closed systems. Energy can neither enter nor leave an isolated system. A calorimeter is carefully insulated so that heat transfer in or out is very small. Often a covered Styrofoam cup is used. 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, Figure 12–12. The total energy of the system is constant;

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

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 equation

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

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 to the colder block, Figure 12–12. The flow continues until the blocks are in thermal equilibrium. The blocks then have the same temperature.

The change in thermal energy is equal to the heat transferred:

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

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

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

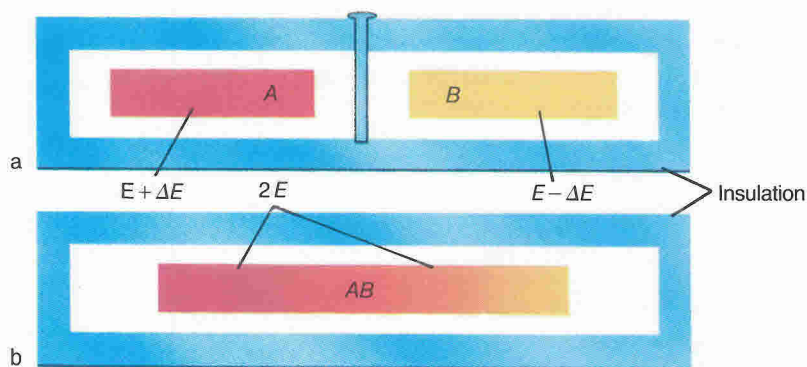


FIGURE 12–12. The total energy for this system is constant.

The change in temperature is the difference between the final and initial temperatures, $\Delta T = T_f - T_i$. 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

$$m_A C_A (T_f - T_{A,i}) + m_B C_B (T_f - T_{B,i}) = 0.$$

To solve for T_f , expand the equation:

$$m_A C_A T_f - m_A C_A T_{A,i} + m_B C_B T_f - m_B C_B T_{B,i} = 0.$$

Isolate T_f and solve,

$$T_f (m_A C_A + m_B C_B) = m_A C_A T_{A,i} + m_B C_B T_{B,i}.$$

$$T_f = \frac{m_A C_A T_{A,i} + m_B C_B T_{B,i}}{m_A C_A + m_B C_B}$$

Note that either the Celsius or Kelvin temperature scale may be used with this equation.

Example Problem

Conservation in Energy Transfer

A 0.500-kg sample of water is at 15.0°C in a calorimeter. A 0.0400-kg block of zinc at 115°C is placed in the water. Find the final temperature of the system? The specific heat of zinc is 388 J/kg · K.

Given: zinc

$$m_A = 0.0400 \text{ kg}$$

$$T_{A,i} = 115^\circ\text{C}$$

$$C_A = 388 \text{ J/kg} \cdot \text{K}$$

water

$$m_B = 0.500 \text{ kg}$$

$$T_{B,i} = 15.0^\circ\text{C}$$

$$C_B = 4180 \text{ J/kg} \cdot \text{K}$$

$$1 \text{ J/kg} \cdot \text{K} = 1 \text{ J/kg} \cdot ^\circ\text{C}$$

Unknown: T_f

Basic equations: $\Delta E_A + \Delta E_B = 0$

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

Solution: $m_A C_A (T_f - T_{A,i}) + m_B C_B (T_f - T_{B,i}) = 0$

$$T_f = \frac{m_A C_A T_{A,i} + m_B C_B T_{B,i}}{m_A C_A + m_B C_B} =$$

$$\frac{(0.0400 \text{ kg})(388 \text{ J/kg} \cdot ^\circ\text{C})(115^\circ\text{C}) + (0.500 \text{ kg})(4180 \text{ J/kg} \cdot ^\circ\text{C})(15.0^\circ\text{C})}{(0.0400 \text{ kg})(388 \text{ J/kg} \cdot ^\circ\text{C}) + (0.500 \text{ kg})(4180 \text{ J/kg} \cdot ^\circ\text{C})}$$

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

$$T_f = \frac{3.32 \times 10^4 \text{ J}}{2.11 \times 10^3 \text{ J/}^\circ\text{C}}$$

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

POCKET LAB

COOL TIMES

Place a 100-mL beaker in a 250-mL beaker. Put a thermometer into 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 5 minutes. Plot your data for both beakers on the same graph of temperature vs time. Measure and record the mass of water in each beaker. Predict the final temperature. Describe each curve.

Using Your CALCULATOR

Use the parentheses () function on your calculator to evaluate complex equations. Note the use of two levels of parentheses in the 'denominator'.

$$T_f = \frac{m_A C_A T_{A,i} + m_B C_B T_{B,i}}{m_A C_A + m_B C_B}$$

$$\left(\left(0.0400 \right) \times 388 \right) \times 115 \quad 1784.8$$

$$+ \left(\left(0.500 \right) \times 4180 \right) \times 15 \quad 31350$$

$$= \div \quad 33134.8$$

$$\left(\left(0.0400 \right) \times 388 \right) + 15.52$$

$$\left(0.500 \times 4180 \right) \quad 2090$$

$$\left) \quad 2105.52$$

$$= \quad 15.73711$$

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

F. Y. I.

Laws of Thermodynamics:

1. You cannot win.
2. You cannot break even.
3. You cannot get out of the game.

Anon

Thermodynamics is the study of the properties of thermal energy and its changes.

Objectives

- define heats of fusion and vaporization; understand the microscopic basis of changes of state; calculate heat transfers needed to effect changes of state.
- distinguish heat from work.
- state the first law of thermodynamics.
- define a heat engine, refrigerator, and heat pump.
- state the second law of thermodynamics; define entropy.

Practice Problems

9. 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?
10. 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 no heat loss to the surroundings. What is the final temperature of the mixture?
11. A 1.00×10^2 -g brass block at 90.0°C is placed in a styrofoam 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.
- ▶ 12. A 1.0×10^2 -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°C . What is the specific heat of the aluminum?

CONCEPT REVIEW

- 1.1 Could the thermal energy of a bowl of hot water equal that of a bowl of cold water? Explain.
- 1.2 On cold winter nights before central heating, people often placed hot water bottles in their beds. Why would this be better than, say warmed bricks?
- 1.3 If you take a spoon out of a cup of hot coffee and put it in your mouth, you won't burn your tongue. But, you could very easily burn your tongue if you put the liquid in your mouth. Why?
- 1.4 **Critical Thinking:** You use an aluminum cup instead of a styrofoam 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?

12.2 CHANGE OF STATE AND LAWS OF THERMODYNAMICS

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 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 is increased. The energy is now in the form of thermal energy. The branch of physics called **thermodynamics** explores the properties of thermal energy.

Change of State

The three most common states of matter are solids, liquids, and gases, Figure 12–13. As the temperature of a solid is raised, it first changes to a liquid. At even higher temperatures, it will become a gas. How can we explain these changes? Our simplified model of a solid consists of

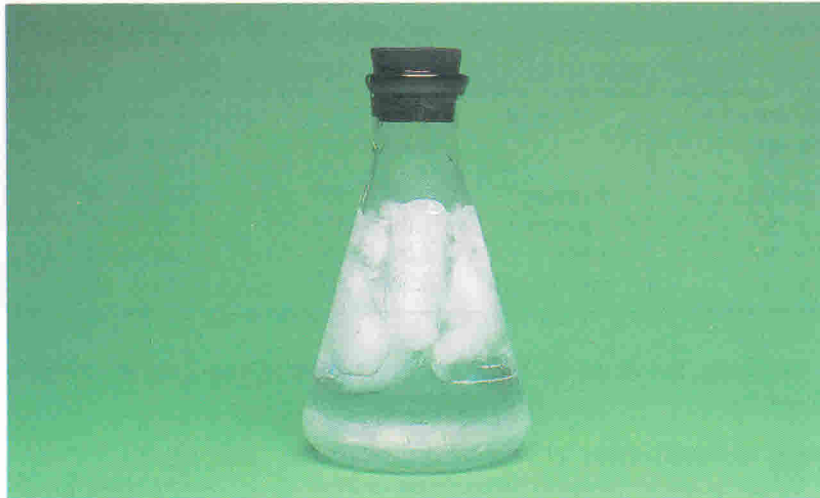


FIGURE 12–13. The three states of water are represented in this photograph. The gaseous state, water vapor, is dispersed in the air and is invisible until it condenses.

tiny particles bonded together by springs. The springs represent the electromagnetic forces between the particles. When the thermal energy of a solid is increased, both the potential and kinetic energies of the particles increase.

At sufficiently high temperatures, the forces between the particles are no longer strong enough to hold them in fixed locations. 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 called the **melting point**.

When a substance is melting, added thermal energy increases the potential energy of particles, breaking the bonds holding them together. The added thermal energy does not increase the temperature.

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.

After the substance is totally melted, a further increase in thermal energy once again increases the temperature. Added thermal energy increases the kinetic and potential energies. 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. It does not raise the temperature; it converts particles in the liquid state to particles in the vapor or gaseous state. At normal atmospheric pressure, water boils at 373 K. The amount of 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 heat, Q , required to melt a solid of mass m is given by

$$Q = mH_f$$

The value of some heats of fusion, H_f , can be found in Table 12–2.

While a substance is melting, added energy does not increase the temperature.

The heat of fusion is the energy needed to melt one kilogram of a substance.

The heat of vaporization is the energy needed to vaporize one kilogram of a substance.

**POCKET
LAB**

MELTING

Label two foam cups A and B. Measure 75 ml of room temperature water into 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 1 minute intervals until the ice has melted. Do the samples reach the same final temperature? Why?

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)	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	1.04×10^4	2.36×10^6
lead	2.04×10^4	8.64×10^5	water (ice)	3.34×10^5	2.26×10^6

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

$$Q = mH_v.$$

Heats of vaporization, H_v , can also 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.

Example Problem

Heat of Fusion—1

If 5.00×10^3 J is added to ice at 273 K, how much ice is melted?

Given: heat added,

$$Q = 5.00 \times 10^3 \text{ J}$$

heat of fusion,

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

Unknown: mass, m

Basic equation: $Q = mH_f$

Solution: $m = \frac{5.00 \times 10^3 \text{ J}}{3.34 \times 10^5 \text{ J/kg}} = 0.0150 \text{ kg}$

Figure 12–14 shows the changes in temperature as thermal energy is added to 1.0 g of H_2O at 243 K. Between points A and B, the ice is warmed to 273 K. Between points B and C, the added thermal energy melts the ice at a constant 273 K. The horizontal distance from point B to point C represents the heat of fusion. Between points C and D, the

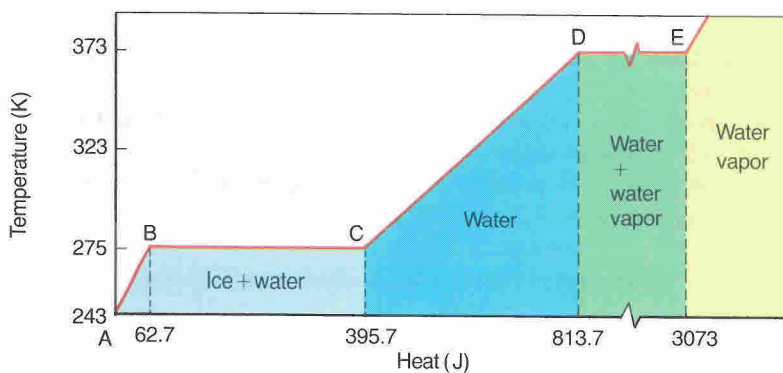


FIGURE 12–14. A plot of temperature versus heat added when 1 g of ice is initially converted to steam.

water temperature rises. The slope is smaller here than between points **A** and **B**, showing that the specific heat of water is higher than that of ice. Between points **D** and **E**, the water boils, becoming water vapor. The distance from point **D** to point **E** represents the heat of vaporization. Between points **E** and **F**, the steam is heated to 473 K. The slope is larger than that from point **C** to point **D**, indicating that the specific heat of steam is less than that of water.

Example Problem

Heat of Fusion—2

How much heat must be transferred to 100.0 g of ice at 0.0°C until the ice melts and the temperature of the resulting water rises to 20.0°C?

Given: $m = 100.0 \text{ g}$

$$T_i = 0.0^\circ\text{C}$$

$$T_f = 20.0^\circ\text{C}$$

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

$$C = 4180 \text{ J/kg} \cdot ^\circ\text{C}$$

Unknown: Q_{total}

Basic equations: $Q = mH_f$

$$Q = mC\Delta T$$

Solution:

First, find the amount of heat the ice absorbs as it changes from solid to liquid.

$$Q = mH_f = (0.100 \text{ kg})(3.34 \times 10^5 \text{ J/kg}) = 33\,400 \text{ J}$$

Second, calculate the amount of heat the water absorbs as its temperature rises from 0.0°C to 20.0°C.

$$Q = mC\Delta T = (0.100 \text{ kg})(4180 \text{ J/kg} \cdot ^\circ\text{C})(20.0^\circ\text{C} - 0.0^\circ\text{C}) = 8360 \text{ J}$$

Finally, add the two quantities of heat.

$$Q_{\text{total}} = 33\,400 \text{ J} + 8360 \text{ J} = 41\,760 \text{ J}$$

Thus, 41.8 kJ of heat is transferred from the water to the ice.

Practice Problems

13. How much heat is absorbed by $1.00 \times 10^2 \text{ g}$ of ice at -20.0°C to become water at 0.0°C ?
14. A $2.00 \times 10^2\text{-g}$ sample of water at 60.0°C is heated to steam at 140.0°C . How much heat is absorbed?
15. How much heat is needed to change $3.00 \times 10^2 \text{ g}$ of ice at -30.0°C to steam at 130.0°C ?
- ▶ 16. 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?

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. This organization offers excellent opportunity to work your way up into other positions based on your work performance and completion of further training. For information contact:

Association Builders and Contractors, 729 15th Street, N.W., Washington, DC 20005

F. Y. I.

Evaporation of perspiration from the skin is an effective way of cooling your body. Over two million joules of thermal energy are carried away for each liter of liquid lost. If the perspiration runs down your face, however, its ability to cool is almost all lost.

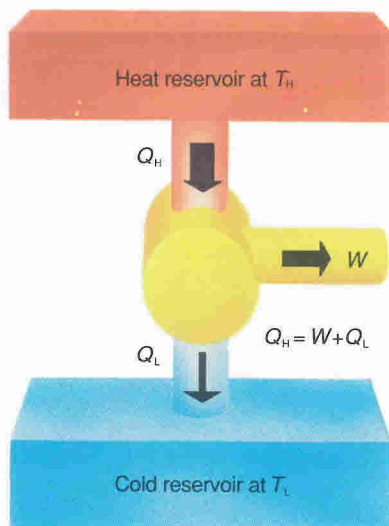


FIGURE 12–15. The flow chart diagram represents heat at high temperature transformed into mechanical energy and low temperature waste heat.

Thermal energy is increased by both work done and heat added to a system.

A heat engine accepts heat from a high temperature source, performs work, and transfers heat out at a low temperature.

The First Law of Thermodynamics

You don't have to transfer heat to increase the thermal energy of a body. If you rub your hands together, they are warmed, yet they were not brought into contact with a hotter body. Instead, work was done on your hands by means of friction. The mechanical energy of your moving hands was changed into thermal energy.

There are other means of converting mechanical energy into thermal energy. 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, electrical, as well as mechanical—can be changed into thermal energy.

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**. Thermodynamics is the study of the changes in thermal properties of matter. The first law is merely a restatement of the law of conservation of energy.

All forms of energy are measured in joules. Work, energy transferred by mechanical means, and heat, energy transferred because of a difference in temperature, are also measured in joules.

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**.

Heat engines require a high temperature source from which thermal energy can be removed, and a low temperature sink into which thermal energy can be delivered, Figure 12–15. An automobile engine is an example of a heat engine, Figure 12–16. A mixture of air and gasoline vapor is ignited, producing a very 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

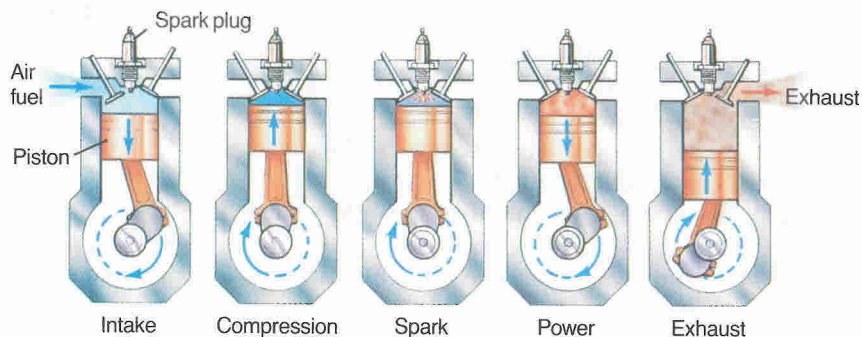


FIGURE 12–16. The heat produced by burning gasoline causes the gases produced to expand and exert force on the cylinder.

air, and the piston is returned to the top of the cylinder. The entire cycle is repeated many times each minute. The heat from the burning gasoline is converted into mechanical energy that eventually results in the movement of the car.

Not all the thermal energy from the very 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. Heat from the hot engine is transferred to a radiator. Outside air passes through the radiator and the air temperature is raised. This heat transferred out of the engine is called waste heat, heat that cannot be converted into work. All heat engines generate waste heat. In a car engine, the waste heat is at a lower temperature than the heat of the gasoline flame. The overall change in total energy of the car-air system is zero. Thus, according to the first law of thermodynamics, the thermal energy in the flame is equal to the sum of the mechanical energy produced and the waste heat expelled.

Heat flows spontaneously from a warm body to a cold body. It is possible to remove thermal energy from a colder body and add it to a warmer body. An external source of energy, usually mechanical energy, however, is required to accomplish this transfer. A refrigerator is a common example of a device that accomplishes this transfer. Electrical energy runs a motor that does work on a gas such as Freon. Heat is transferred from the contents of the refrigerator to the Freon. Food is cooled, usually to 4.0°C , and the Freon is warmed. Outside the refrigerator, heat is transferred from the Freon to room air, cooling the Freon again. The overall change in the thermal energy of the Freon is zero. Thus, according to the first law of thermodynamics, the sum of the heat removed from the food and the work done by the motor is equal to the heat expelled to the outside at a higher temperature, Figure 12–17.

A heat pump is a refrigerator that can be run in two directions. In summer, heat is removed from the house, 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, Figure 12–18. In either case, mechanical energy is required to transfer heat from a cold object to a warmer one.

Perpetual Motion?

FIGURE 12–17. Diagram for a refrigerator. It 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.

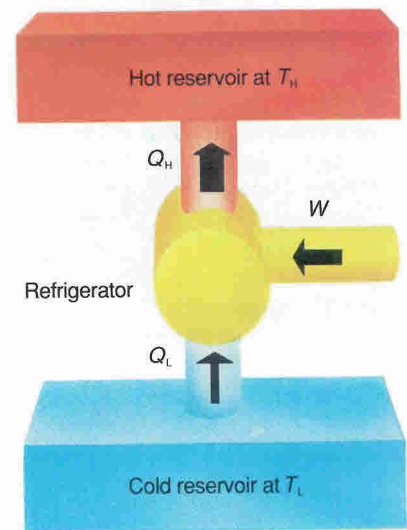
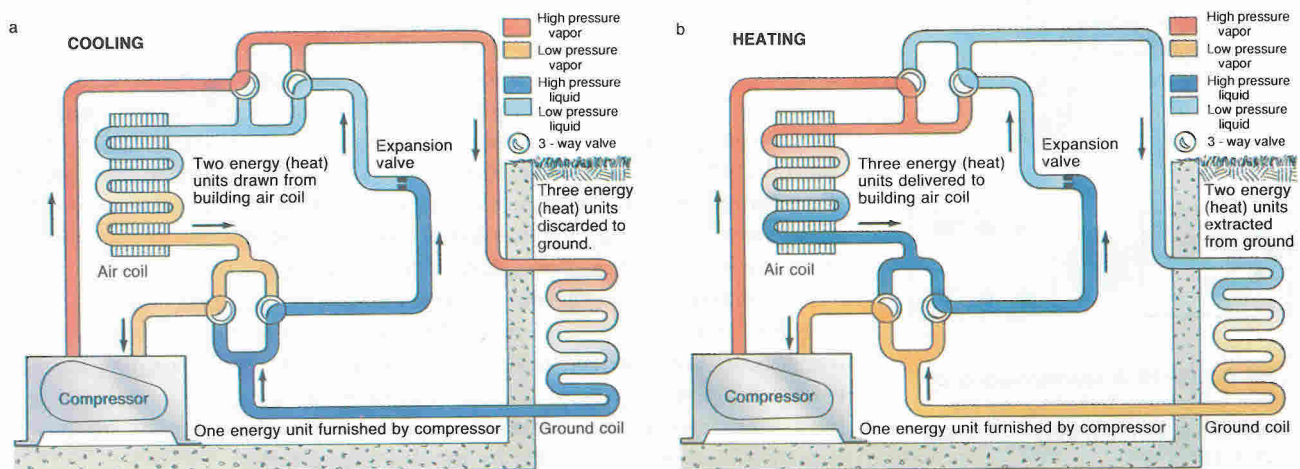


FIGURE 12–18. A heat pump runs in either direction depending on whether it is used in heating or cooling. In cooling, heat is extracted from the air in the house and pumped outside (a). In heating, heat is extracted from the outside and pumped inside (b).



EARTH SCIENCE CONNECTION

POCKET LAB

DRIP, DRIP, DRIP

Measure equal amounts of very hot and very cold water into 2 clear glasses (or beakers). Predict what will happen if you simultaneously put 1 drop of food coloring into each glass. Try it. What happened? Why? Was the mixing symmetric?

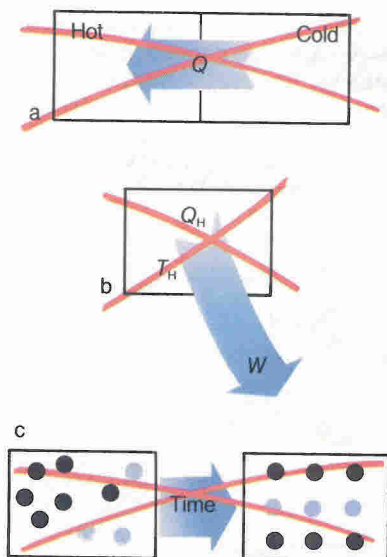


FIGURE 12-19. A representation of three processes forbidden by the second law of thermodynamics, but not the first law.

Freon is the name of a family of chemical compounds invented for use in refrigerators and air conditioners. Each Freon compound was designed by chemists to have a specific boiling point. They were also designed to be very stable, and not to react with materials used in refrigerators. They work very well in a variety of important applications. Unfortunately, these properties also mean that when these gases escape into the atmosphere, they also do not decompose at low altitudes. In the past few years, however, scientists have found that these chemicals do react when they move into the upper atmosphere, miles above Earth. The products of the reaction can destroy Earth's protective ozone layer, increasing ultraviolet radiation that reaches Earth, and increasing incidence of skin cancer. They may also increase the "greenhouse effect" and raise Earth's temperature. Scientists are now trying to create new molecules to replace the Freons that are safe for the ozone and are not greenhouse gases.

The Second Law of Thermodynamics

Many processes that do not violate the first law of thermodynamics have never been observed to occur spontaneously. For example, the first law does not prohibit heat flowing from a cold body to a hot body, Figure 12-19a. Still, when a hot body is 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. As another example, heat engines could convert thermal energy completely into mechanical energy with no waste heat and still obey the first law, Figure 12-19b. Yet waste heat is always generated.

In the nineteenth century, the French engineer Sadi Carnot (1796–1832) studied the ability of engines to convert heat into mechanical energy. He developed a logical proof that even an ideal engine would generate some waste heat. Real engines generate even more waste heat. Carnot's result is best described in terms of a quantity called **entropy** (EN truh pee). 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, however, the entropy does not change, as long as friction is ignored.

On a microscopic level, entropy is described as the disorder in a system. When heat is added to an object, the particles move in a random way. Some move very quickly, others move slowly, many move at intermediate speeds. 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.

The **second law of thermodynamics** states: *natural processes go in a direction that increases the total entropy of the universe.* Entropy and the second law can be thought of as statements of the probability of events happening. Figure 12-20 illustrates an increase in entropy as food color molecules, originally separate from clear water, are thoroughly mixed with the water molecules after a time.

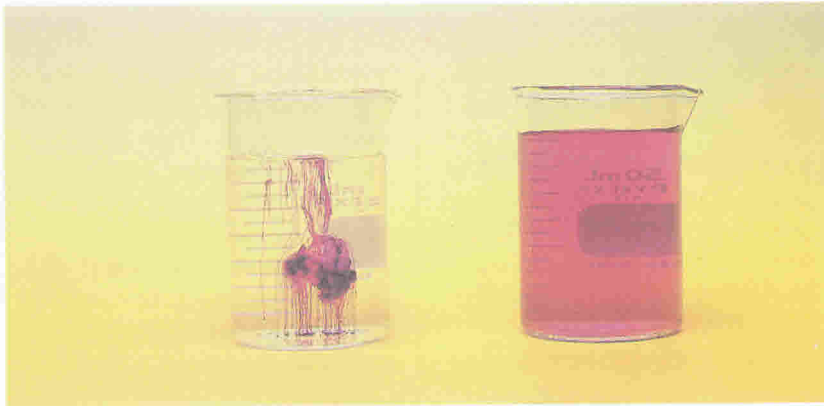


FIGURE 12–20. An example of the second law of thermodynamics.

The second law 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. The bar is plunged into the water. When thermal equilibrium is reached, the average kinetic energy of the particles in the iron and the water will be the same. This final state is less ordered than the first situation. No longer are the fast particles confined mainly in the iron and the slow particles in the water. All speeds are evenly distributed. The entropy of the final state is larger than that of the initial state.

We take for granted many daily events that occur spontaneously, or naturally, in one direction, but that would really shock us if they happened in reverse. You are not surprised when a metal spoon, heated at one end, 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 skillet collected in a 9-cm³ cube in the exact center of the kitchen. Neither of the reverse processes violates the first law of thermodynamics. The events are simply examples of the countless events that are not spontaneously reversible because the reverse process would violate the second law of thermodynamics.

The second law and entropy also give new meaning to what is commonly called the “energy crisis.” When you use a resource such as natural gas to heat your home, you do not use up the energy in the gas. The potential 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 thermal 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 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.

When fuel is burned, energy is not lost; entropy is increased.

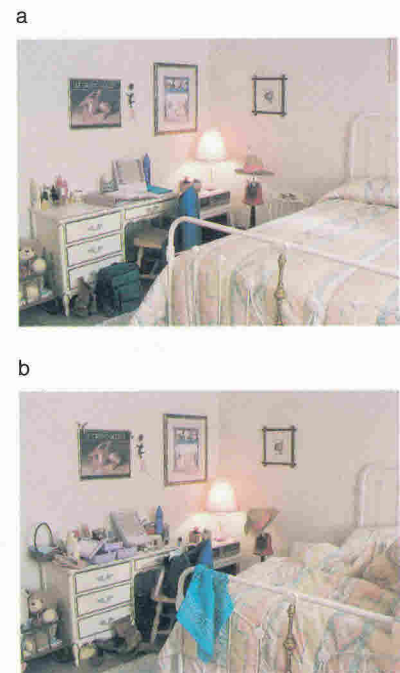


FIGURE 12–21. A familiar example of the second law of thermodynamics. If no work is done on a system, entropy spontaneously reaches a maximum.

Physics and technology

RANGE-TOP COOKING

Top-of-the-range cooking is bringing new technology into the kitchen. In addition to the electric coil and gas burner, consumers can select solid disk heating elements, glass-ceramic cooktops, and sealed gas burners, among others.

The familiar electric coil cooks food with radiant heat and conduction. It heats up and cools down quickly. A disadvantage is that the heating elements and drip bowls can be hard to clean.

Solid disk heating elements were introduced in the United States in the 1980s and have grown in popularity. The disks consist of cast iron which has

electric heating wires embedded in the underside. Solid disks are easier to clean but take longer to heat up and cool down than coil-type elements. They also require cookware with extra-flat bottoms in order to maximize heat transfer from the disk to the utensil.

Glass-ceramic cooktops, or

smoothtops, contain their heating elements under a sheet of ceramic glass. They may use electric resistance heating, a halogen gas “bulb”, or magnetic induction heating.

• Why would the solid disk heating elements take longer to heat up and cool down than the coil-type elements?



CONCEPT REVIEW

- 2.1 Old-fashioned heating systems sent steam into radiators in each room. In the radiator, the steam condensed back to water. How did this heat the room?
- 2.2 James Joule carefully measured the difference in temperature of water at the top and bottom of a waterfall. Why would he have expected a difference?
- 2.3 In 1861, the French scientist Hirn used a 320-kg hammer moving at 5 m/s to smash a 3-kg block of lead against a 450-kg rock. He found that the temperature of the lead increased 5°C. Explain.
- 2.4 **Critical Thinking:** A new deck of cards has all the suits (clubs, diamonds, hearts, and spades) in order, and the cards ordered by number within the suits. If you shuffle the cards many times, are you likely to return the cards to the original order? Of what physical law is this an example?

CHAPTER 12 REVIEW

SUMMARY

12.1 Temperature and Thermal Energy

- The thermal energy of an object is the sum of the kinetic and potential energies of the internal motion of the particles.
- The temperature of a gas is proportional to the average kinetic energy of the particles.
- Thermometers use some property of a substance, such as thermal expansion, that depends on temperature.
- Thermometers reach thermal equilibrium with the objects they contact, and then some temperature-dependent property of the thermometer is measured.
- The Celsius and Kelvin temperature scales are widely used in scientific work. One kelvin is equal to one degree Celsius.
- At absolute zero, 0 K or -273.15°C , matter has no thermal energy.
- Heat is the energy transferred because of a difference in temperature. Heat flows naturally from a hot to a cold body.
- Specific heat is the quantity of heat required to raise the temperature of one kilogram of a substance one kelvin.
- In an isolated system, the thermal energy of one part may change 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 required to change one kilogram of a substance from its solid state to its liquid state at its melting point.
- The heat of vaporization is the quantity of heat required to change one kilogram of a substance from the liquid state to the vapor state at its boiling point.
- The heat transferred during a change of state does not produce a change in temperature.
- The first law of thermodynamics states that the total increase in thermal energy of a system is

equal to 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 an area of lower temperature to an area of higher temperature.
- Entropy, a measure of disorder, never decreases in natural processes.

KEY TERMS

kinetic-molecular theory	calorimeter
thermal energy	thermodynamics
conduction	melting point
convection	heat of fusion
radiation	boiling point
temperature	heat of vaporization
thermal equilibrium	first law of thermodynamics
thermometer	heat engine
absolute zero	entropy
kelvin	second law of thermodynamics
heat	
specific heat	

REVIEWING CONCEPTS

1. Explain the difference between a ball's external energy and its thermal energy. Give an example.
2. Explain the difference between a ball's thermal energy and temperature.
3. Can temperature be assigned to a vacuum? Explain.
4. Do all of the molecules or atoms in a liquid have about the same speed?
5. Your teacher just told your class that the temperature of the sun is 1.5×10^7 degrees.
 - a. Sally asks whether this is the Kelvin or Celsius scale. What is the teacher's answer?
 - b. Would it matter, between the Celsius and Fahrenheit scales, which one you use?
6. Is your body a good judge of temperature? On a cold winter day, a metal door knob feels

much colder to your hand than the wooden door. Is this true? Explain.

7. A hot steel ball is dropped into a cup of cool water. Explain the difference between heat and the ball's thermal energy.
8. Do we ever measure heat transfer directly? Explain.
9. When a warmer object is in contact with a colder object, does temperature flow from one to the other? Do the two have the same temperature changes?
10. Can you add thermal energy to an object without increasing its temperature? Explain.
11. When wax freezes, is energy absorbed or released by the wax?
12. Why does water in a canteen stay cooler if it has a canvas cover that is kept wet?
13. Are the coils of an air conditioner that are inside the house the location of vaporization or condensation of the Freon? Explain.
14. Which situation has more entropy, an unbroken egg or a scrambled egg?

APPLYING CONCEPTS

1. Sally is cooking pasta in a pot of boiling water. Will the pasta cook faster if the water is boiling vigorously than if it is boiling gently?
2. What temperatures on the following pairs of scales are the same? ($T_F = 9/5 T_C + 32$)
 - a. Celsius and Fahrenheit
 - b. Kelvin and Fahrenheit
 - c. Celsius and Kelvin
3. Which liquid would an ice cube cool faster, water or methanol? Explain.
4. Explain why the high specific heat of water makes it desirable for use in hot water heating systems.
5. 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.
6. Two blocks of lead are heated to the same temperature. Block A has twice the mass of block B. They are dropped into identical cups of water and both systems come to thermal equilibrium. If the cups started with water at the same temperature, will the water have the same temperature after the blocks are added? Explain.
7. Why do easily vaporized liquids, such as acetone or methanol, feel cool to the skin?
8. Explain why fruit growers spray their trees with water, when frost is expected, to protect the fruit from freezing.
9. Would opening the refrigerator door on a warm day help cool the kitchen? Explain.

PROBLEMS

12.1 Temperature and Thermal Energy

1. Liquid nitrogen boils at 77 K. Find this temperature in degrees Celsius.
2. The melting point of hydrogen is -259.14°C . Find this temperature in kelvin.
3. Sadi Carnot showed that no real heat engine can have an efficiency greater than

$$\text{efficiency} = \frac{\text{work output}}{\text{heat input}} = \frac{T_{\text{hot}} - T_{\text{cold}}}{T_{\text{hot}}}$$

where T_{hot} and T_{cold} are the temperatures of the input and waste thermal energy reservoirs. Note: Kelvin temperatures must be used in this equation.

- a. What is the efficiency of an ideal steam engine that uses superheated steam at 685 K to drive the engine and ejects waste steam at 298 K?
 - b. If the steam generator produces 1.00×10^8 J each second, how much work can the ideal engine do each second?
4. How much heat is needed to raise the temperature of 50.0 g of water from 4.5°C to 83.0°C ?
 5. How much heat must be added to 50.0 g of aluminum at 25°C to raise its temperature to 125°C ?
 6. 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.
 7. 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.
 8. A copper wire has a mass of 165 g. An electric current runs through the wire for a short

time and its temperature rises from 21°C to 39°C. What minimum quantity of energy is converted by the electric current?

9. 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 21.6°C. Calculate the specific heat of tungsten.
10. 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°C. Assume no heat loss to the surroundings. What is the final temperature of the mixture?
11. 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)?
12. A 10.0-kg piece of zinc at 71°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?
13. A 2.00×10^2 -g sample of brass at 100.0°C is placed in a calorimeter cup that contains 261 g of water at 20.0°C. Disregard the absorption of heat by the cup and calculate the final temperature of the brass and water.
- ▶ 14. 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 for the heater to bring the water to the boiling point? Assume the temperature of the cup to be the same as the temperature of the water at all times and no heat is lost to the air.
15. A 2.50×10^2 -kg cast-iron car engine contains water as a coolant. Suppose the engine's temperature is 35°C when it is shut off. The air temperature is 10°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?

12.2 Change of State and Laws of Thermodynamics

16. 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 deliv-

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

17. A person who eats 2400 food calories each day consumes 1.0×10^7 J of energy in a day. How much water at 100°C could that much energy vaporize?
18. 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?
- ▶ 19. How much heat is removed from 60.0 g of steam at 100.0°C to change it to 60.0 g of water at 20.0°C?
20. A 750-kg car moving at 23 m/s brakes to a stop. The brakes contain about 15 kg of iron that absorb the energy. What is the increase in temperature of the brakes?
- ▶ 21. How much heat is added to 10.0 g of ice at -20.0°C to convert it to steam at 120.0°C ?
- ▶ 22. A 50.0-g sample of ice at 0.00°C is placed in a glass beaker containing 4.00×10^2 g of water at 50.0°C. All the ice melts. What is the final temperature of the mixture? Disregard any heat loss to the glass.
23. 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?
24. 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 offsets this input of food energy by climbing stairs. How high would she have to climb if Sally has a mass of 65.0 kg?
25. When air is compressed in a bicycle pump, an average force of 45 N is exerted as the pump handle moves 0.24 m. During this time, 2.0 J of heat leave the cylinder through the walls. What is the net change in thermal energy of the air in the cylinder?

THINKING PHYSIC-LY

1. Picture a cup of hot coffee and an iceberg.
 - a. Which has a greater amount of internal energy?
 - b. Which has a higher temperature?
2. Why can't you tell if you have a fever by touching your own forehead?