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التواصل الاجتماعي بحسب الصف الحادي عشر المتقدم

					صفحة المناهج الإماراتية على فيسبوك
الرياضيات	اللغة الانجليزية	اللغة العربية	التربية الاسلامية	المواد على تلغرام	

المزيد من الملفات بحسب الصف الحادي عشر المتقدم والمادة فيزياء في الفصل الثاني

حل أسئلة الاختبار الوزاري للمنهجين الجزء الالكتروني	1
أسئلة الاختبار الوزاري للمنهجين الجزء الالكتروني بدون الحل	2
أسئلة الاختبار التجريبي القسم الالكتروني والورقي	3
حل أسئلة الامتحان النهائي القسم الالكتروني منهج بريدج Bridge	4
ملف مراجعة نهائية وحدة Energy Thermal ووحدة Matter of States وفق الهيكل منهج انسباير Inspire	5

Grade 11 ADV

Physics

End of term Examination

Term 2 EOT – (2025-2026)

Prepared by

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Thermal Energy

You have studied how objects collide and trade kinetic energies. Every material is made of microscopic particles. The many particles present in a gas have linear and rotational kinetic energies. The particles also might have potential energy due to their internal bonds and interactions with each other. As gas particles collide with each other and with the walls of the container, as shown in **Figure 1**, they transfer energy. There are numerous molecules that make up the gas, resulting in many collisions. The energies of the particles become randomly distributed.

Thus, it is convenient to discuss the total energy of the particles that compose the gas and the average energy per particle in the gas. Recall that the sum of the particles' energies is the object's thermal energy. The average kinetic energy per particle is related to the temperature of the gas. The relationship between the particles' random motions and the bulk property of the material is described by kinetic theory.

Thermal equilibrium The thermometer's particles also transfer energy to your body's particles. As the thermometer's particles gain more energy, the amount of energy they give back to the skin increases. At some point, the rate of energy transfer from the thermometer to your body is equal to the rate of transfer in the other direction. At this point, your body and the thermometer have reached thermal equilibrium. **Thermal equilibrium** is the state in which the rates of thermal energy transfer between two objects are equal and the objects are at the same temperature. In other words, uncontrolled systems always evolve toward more stable states—that is, toward more uniform energy distribution (objects hotter than their surrounding environment cool down until thermal equilibrium is reached by the object and its surroundings). **Figure 4** on the next page shows two blocks reaching equilibrium.
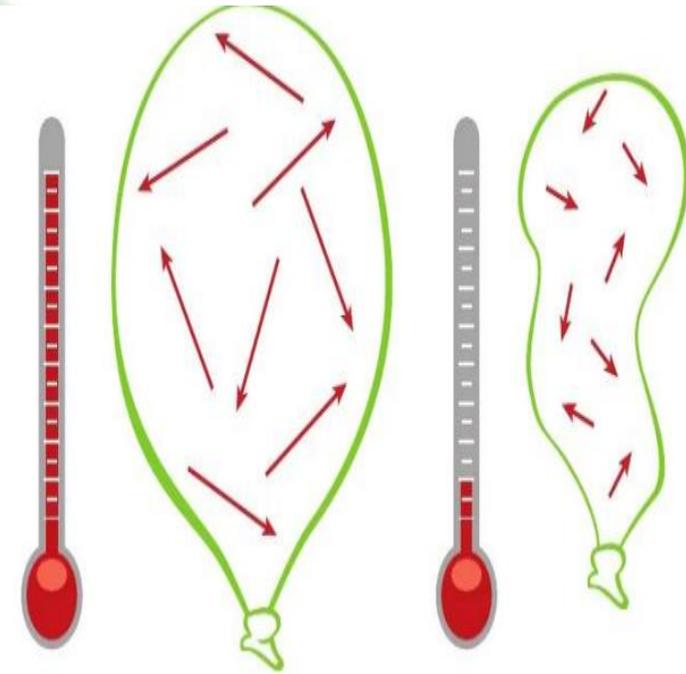
Reaching Thermal Equilibrium

Before Thermal Equilibrium

Hot object (A)   Cold object (B)

Thermal Equilibrium

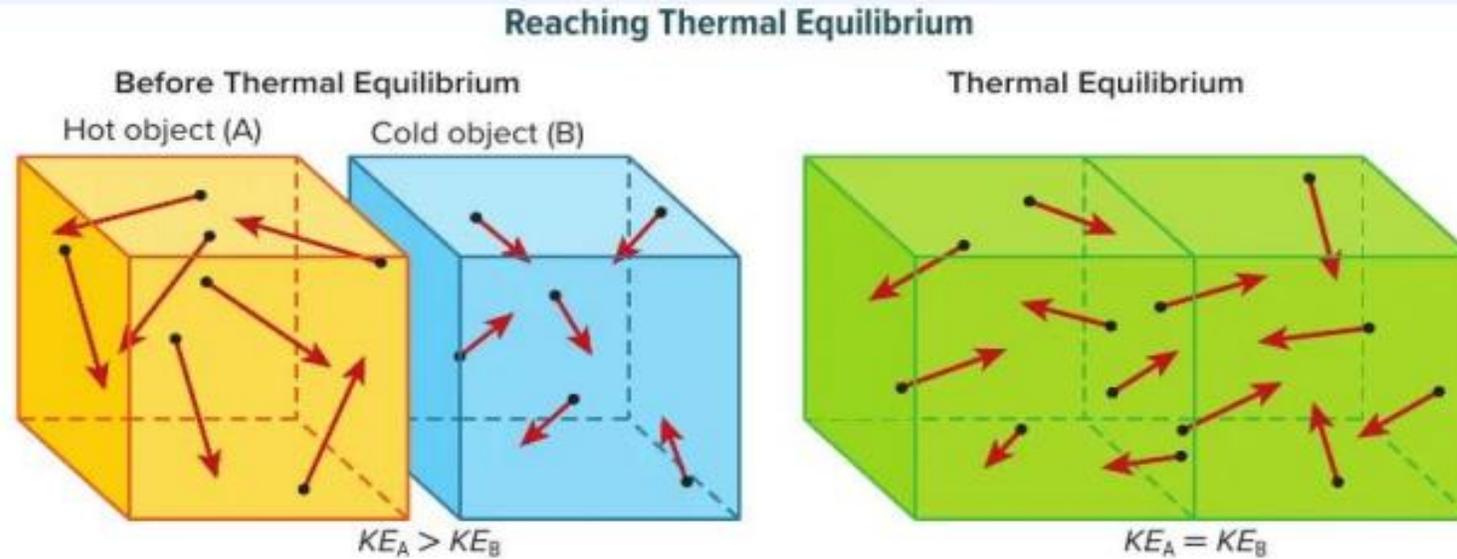$KE_A > KE_B$ $KE_A = KE_B$

Figure 4 When a hot object and a cold object are in contact, there is a net transfer of thermal energy from the hot object to the cold object. When the two objects reach thermal equilibrium, the transfer of energy between the objects is equal, and the objects are at the same temperature.

Thermometers Every thermometer has some useful property that changes with temperature. Household thermometers often contain colored alcohol that expands when heated. The hotter the thermometer, the more the alcohol expands and the higher it rises in the tube. The liquid crystal thermometer in **Figure 5** uses a variety of long molecules that rearrange and cause a color change at specific temperatures. Medical thermometers and the thermometers that monitor automobile engines use very small, temperature-sensitive electronic circuits to take rapid temperature measurements.

Differentiate between the three temperature scales: Celsius, Fahrenheit, and Kelvin scales. Explain why the term absolute zero is appropriate for the coldest temperature possible.

Temperature scales In the United States, weather agencies report temperatures in degrees Fahrenheit. Scientists, however, use the Celsius and Kelvin scales. The Celsius scale is based on the properties of water and was devised in 1741 by Swedish physicist Anders Celsius. On the Celsius scale, the freezing point of pure water at sea level is defined to be 0°C. The boiling point of pure water at sea level is defined to be 100°C. The Celsius scale is useful for day-to-day measurements of temperature.

The Kelvin scale On the Celsius scale, temperatures can be negative. Negative temperatures suggest a particle could have negative kinetic energy. Because temperature represents average kinetic energy of the object's particles, it makes more sense to use a temperature scale whose zero temperature is where the particles' kinetic energy is also zero. Therefore the zero point of the Kelvin scale is defined to be absolute zero. On the Kelvin scale, the freezing point of water (0°C) is about 273 K and the boiling point of water is about 373 K. Each interval on this scale, called a kelvin, is equal to 1°C. Thus, $T_C + 273 = T_K$. **Figure 7** compares the Fahrenheit, Celsius, and Kelvin scales.

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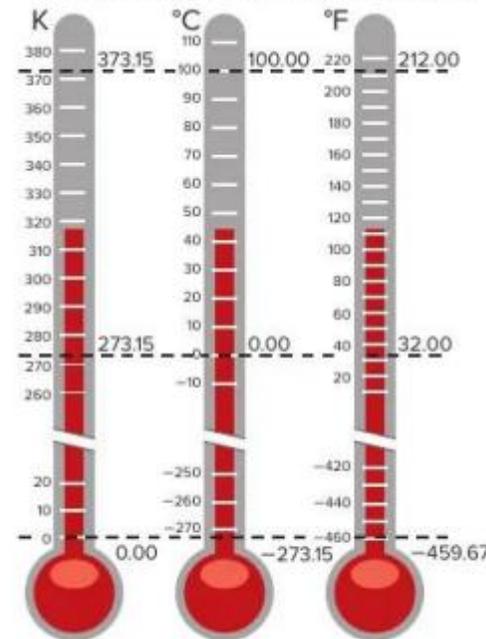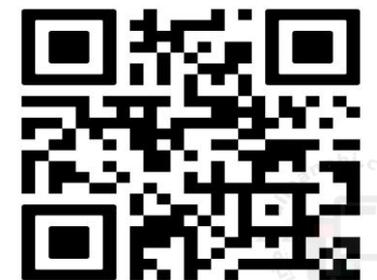

Comparing Temperature Scales

Figure 7 The Kelvin and Celsius scales are used by scientists. In the United States, the Fahrenheit scale is often used for weather reports and cooking.

Define the term specific heat (C) and specify its unit (J/(kg.K))

Specific Heat

Some objects are easier to warm up than others. On a bright summer day, the Sun radiates thermal energy to the sand on a beach and to the ocean water. The sand on the beach becomes quite hot, while the ocean water stays relatively cool. When an object is heated, its thermal energy increases and its temperature can increase. The amount of the increase in temperature depends on the size of the object and its composition.

The **specific heat** of a material is the amount of energy that must be added to a unit mass of the material to raise its temperature by one temperature unit. In SI units, specific heat (C) is measured in J/(kg·K).

Figure 9 These pans are made of stainless steel and have copper bottoms and plastic handles.

Explain *how the selection of these materials is related to their specific heats.*

Table 1 provides values of specific heat for some common substances. For example, 897 J must be added to 1 kg of aluminum to raise its temperature by 1 K. The specific heat of aluminum is therefore 897 J/(kg·K). Materials with different specific heats are used for different purposes. Metals, such as those used to make the pans in **Figure 9,** have low specific heats and are good thermal conductors. Notice that liquid water has a high specific heat compared to other substances. Ice and water vapor also have relatively high specific heats. These high specific heats have had significant effects on our climate and our bodies.

Pressure

When considering fluids (as well as solids), it is often useful to think about pressure as well as force. You have probably heard people talk about water pressure and air pressure, and you might already have a sense that pressure and force are related. Pressure and force are not the same, however. **Pressure** is the perpendicular component of a force on a surface divided by the area of the surface. Since pressure is force exerted on a surface, anything that exerts pressure is capable of producing change and doing work. In **Figure 2,** both the astronaut and the legs of the landing module are exerting pressure on the Moon's surface.

Pressure

Pressure equals the perpendicular component of the force divided by the surface area to which it is applied.

$$P = \frac{F}{A}$$

Pressure is a scalar. In the SI system, the unit of pressure is the **pascal** (Pa), which is 1 N/m². One pascal is a very small amount of pressure, about equal to the pressure that a flat dollar bill exerts on a tabletop. Thus the kilopascal (kPa), equal to 1000 Pa, is usually used. **Table 1** shows pressures in various locations.

1. The atmospheric pressure at sea level is about 1.0×10^5 Pa. What is the force at sea level that air exerts on the top of a desk that is 152 cm long and 76 cm wide?

2. A car tire makes contact with the ground on a rectangular area of 12 cm by 18 cm. If the car's mass is 925 kg, what pressure does the car exert on the ground as it rests on all four tires?

3. A lead brick, 5.0 cm × 10.0 cm × 20.0 cm, rests on the ground on its smallest face. Lead has a density of 11.8 g/cm³. What pressure does the brick exert on the ground?

EXAMPLE Problem 1

CALCULATING PRESSURE A child weighs 364 N and sits on a three-legged stool, which weighs 41 N. The bottoms of the stool's legs touch the ground over a total area of 19.3 cm².

a. What is the average pressure that the child and the stool exert on the ground?

b. How does the pressure change when the child leans over so that only two legs of the stool touch the floor?

1 ANALYZE AND SKETCH THE PROBLEM

- Sketch the child and the stool, labeling the total force that they exert on the ground.
- List the variables, including the force that the child and the stool exert on the ground and the areas for parts **a** and **b**.

Known

$F_{g\ child} = 364\ N$

$F_{g\ stool} = 41\ N$

$F_{g\ total} = F_{g\ child} + F_{g\ stool}$

$\quad = 364\ N + 41\ N$

$\quad = 405\ N$

$A_a = 19.3\ cm^2$

$A_b = \frac{2}{3} \times 19.3\ cm^2$

$\quad = 12.9\ cm^2$

Unknown

$P_a = ?$

$P_b = ?$

2 SOLVE FOR THE UNKNOWN

Find each pressure.

$$P = \frac{F}{A}$$

a. $P_a = \left(\frac{405\ N}{19.3\ cm^2}\right)\left(\frac{(100\ cm)^2}{(1\ m)^2}\right)$ Substitute $F = F_{g\ total} = 405\ N$, $A = A_a = 19.3\ cm^2$.

$\quad = 2.10 \times 10^2\ kPa$

b. $P_b = \left(\frac{405\ N}{12.9\ cm^2}\right)\left(\frac{(100\ cm)^2}{(1\ m)^2}\right)$ Substitute $F = F_{g\ total} = 405\ N$, $A = A_b = 12.9\ cm^2$.

$\quad = 3.14 \times 10^2\ kPa$

$F_g = 405\ N$

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4. Suppose that during a storm, the atmospheric pressure suddenly drops by 15 percent outside. What net force would be exerted on a front door to a house that is 195 cm high and 91 cm wide? In what direction would this force be exerted?

5. CHALLENGE Large pieces of industrial equipment are placed on wide steel plates that spread the weight of the equipment over larger areas. If an engineer plans to install a 454-kg device on a floor that is rated to withstand additional pressure of 5.0×10^4 Pa, how large should the steel support plate be?

The First Law of Thermodynamics

The first steam engines were built in the eighteenth century and used to power trains and factories. Steam engines change thermal energy into mechanical energy. The invention of the steam engine contributed greatly to the Industrial Revolution and to the study of how heat is related to work. The study of how thermal energy is transformed into other forms of energy is called thermodynamics.

Before 1900, scientists did not realize that the concepts of thermodynamics were linked to the motions of particles in an object, and they considered thermodynamics to be a separate topic from mechanics. Today, engineers routinely apply the concepts of thermodynamics to develop higher performance refrigerators, automobile engines, aircraft engines, and numerous other kinds of machines.

The first law developed for thermodynamics was a statement about what thermal energy is and where it can go. As you know, you can raise the temperature of a glass of cold water by placing it on a hot plate or by stirring it. That is, you can increase the water's thermal energy by heating or by doing work on it. If we consider the system to be the water, the work the system does on you is equal to the negative of the work you do on the system. The **first law of thermodynamics** states that the change in thermal energy (ΔU) of an object is equal to the heat (Q) that is added to the object minus the work (W) done by the object. Note that ΔU, Q, and W are all measured in joules, the unit of energy.

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

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

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

State the first law of thermodynamics. Infer that the first law of thermodynamics is merely a restatement of the law of conservation of energy

The First Law of Thermodynamics

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

$$\Delta U = Q - W$$

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 transported from one place to another or transformed into another kind of energy. An example of transformed thermal energy in a system is a hand pump used to inflate a bicycle tire. As a person pumps, the air and the hand pump become warm. The mechanical energy in the moving piston is transformed into thermal energy of the gas. Similarly, other forms of energy, such as light, sound, and electrical energy, can be transformed into thermal energy. For example, a toaster transforms electrical energy into radiant energy when it toasts bread.

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

28. **CHALLENGE** An expansion valve does work on 100 g of water. The system is isolated, and all of the work is used to convert the 90°C water into water vapor at 110°C. How much work does the expansion valve do on the water?

Change of State

The three most common states of matter on Earth are solid, liquid, and gas. As the temperature of a solid rises, that solid usually changes to a liquid. At even higher temperatures, it becomes a gas. If the gas cools, it returns to the liquid state. If the cooling continues, the liquid will return to the solid state. How can these changes be explained? Recall that when the thermal energy of a material changes, the motion of its particles also changes, as does the temperature.

Figure 12 diagrams the changes of state as thermal energy is added to 1.0 kg of water starting at 243 K (ice) and continuing until that water reaches 473 K (water vapor). Between points A and B, the ice is warmed to 273 K. At this point, the added thermal energy gives the water molecules enough energy to partially overcome the forces holding them together. The particles are still touching each other, but they have more freedom of movement. Eventually, the molecules become free enough to slide past each other.
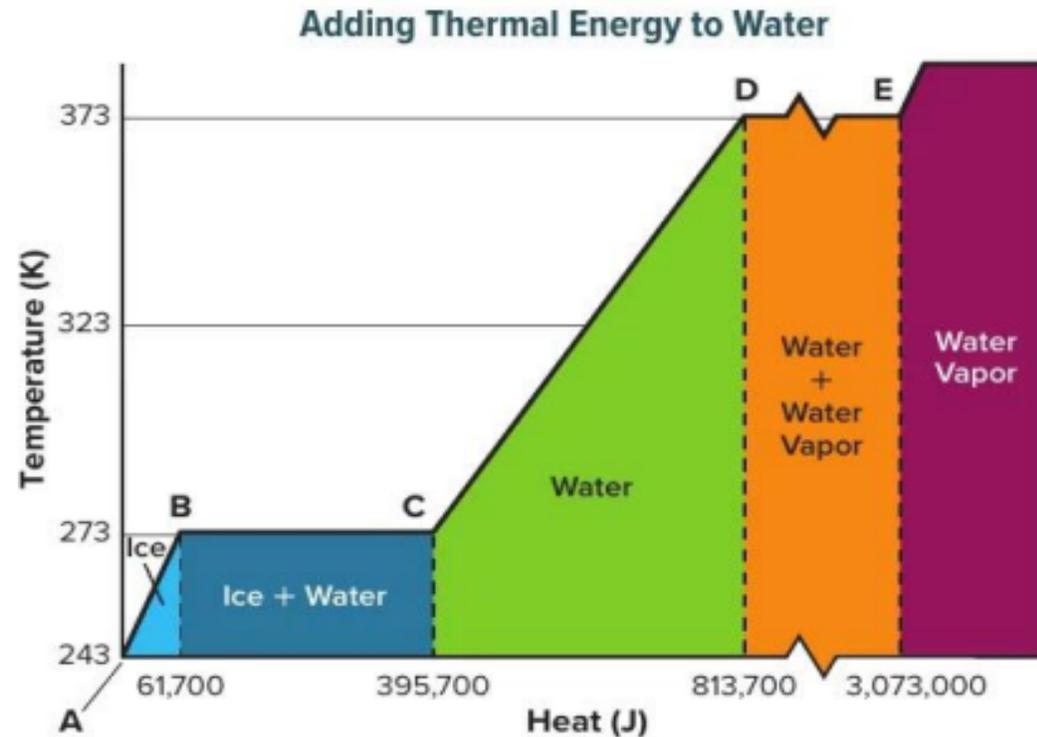

Adding Thermal Energy to Water

Figure 12 Thermal energy added to a substance can raise the temperature or cause a change in state. Note that the scale is broken between points D and E.

Melting point

At this point, water changes from a solid to a liquid, just as the snowman in **Figure 13** does. The temperature at which this change occurs is the melting point of a substance. When a substance melts, the addition of thermal energy allows the particles to move, rotate, and vibrate in ways not available in the solid. Each of these new motions might add new modes of kinetic or potential energy. The added thermal energy does not change the temperature of the material. This can be observed between points B and C in **Figure 12**, where the added thermal energy melts all the ice at a constant 273 K.

Boiling point

Once the ice is completely melted, adding more thermal energy increases the motion of the water molecules again. The temperature rises as shown between points C and D in **Figure 12**. As the temperature increases further, some of the particles that make up the liquid acquire enough energy to break free from the other particles.

At a specific temperature, known as the boiling point, adding more energy to a substance causes it to undergo another change of state. All the added thermal energy converts it from the liquid state to the gas state. As in melting, the temperature does not rise while a liquid boils, as shown between points D and E in **Figure 12**. After the water is entirely converted to gas, any added thermal energy again increases the motion of the molecules, and the temperature rises. After point E, the water vapor is heated to temperatures greater than 373 K.

Heat of fusion and heat of vaporization

The amount of thermal energy needed to melt 1 kg of a substance is called the substance's **heat of fusion** (H_f). For ice, the heat of fusion 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. The added energy causes a change in state but not in tempera-ture. The horizontal distance from point B to point C in **Figure 12** represents the heat of fusion.

The thermal energy needed to vaporize 1 kg of a liquid is called the **heat of vaporization** (H_v). Water's heat of vaporization is 2.26×10^6 J/kg. The horizontal distance from point D to point E in **Figure 12** represents the heat of vaporization. Every material has a characteristic heat of fusion (H_f) and heat of vaporization (H_v). Some values are shown in **Table 2**.

Figure 13 Thermal energy is transferred from the warmer air to the snowman, causing the snowman to melt.

Describe what happens during a phase change in terms of the heat energy absorbed or released. Define the heat of fusion and heat of vaporization of a substance. Relate the changes of state to the heats of fusion and vaporization.

Energy and changes of state There is a definite slope to the graph in **Figure 14** between about 300 s and 800 s. Heat is added at a constant rate, so this slope is proportional to the reciprocal of the specific heat of water. The slope between points A and B in **Figure 12** is proportional to the reciprocal of the specific heat of ice, and the slope above point E is proportional to the reciprocal of the specific heat of water vapor. The slope for water is less than those of both ice and water vapor. This is because water has a greater specific heat than does ice or water vapor. The heat (Q) required to melt a solid of mass (m) is given by the following equation.

Heat Required to Melt a Solid

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

$$Q = mH_f$$

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

Heat Required to Vaporize a Liquid

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

$$Q = mH_v$$

When a liquid freezes, an amount of thermal energy $(Q = -mH_f)$ must be removed from the liquid to turn it into a solid. The negative sign indicates that the thermal energy is transferred from the sample to the external world. In the same way, when a vapor condenses to a liquid, an amount of thermal energy $(Q = -mH_v)$ must be removed from the vapor.

Water absorbs significant amounts of thermal energy when it melts or evaporates. Every day you use the large heats of fusion and vaporization of water. Each gram of sweat that evaporates from your skin carries off about 2.3 kJ of thermal energy. This is one cooling process that many warm-blooded animals use to regulate their body temperatures. Similarly the melting of a 24-g cube of ice absorbs enough thermal energy, 8 kJ, to cool a glass of water by about 30°C.

Table 2 Heats of Fusion and Vaporization of Common Substances

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

EXAMPLE Problem 3

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

1 ANALYZE AND SKETCH THE PROBLEM

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

$T_i = 0.0°C$ $T_f = 70.0°C$

Known	Unknown
$m = 1.50$ kg $H_f = 3.34 \times 10^5$ J/kg	$Q_{melt\ ice} = ?$
$T_i = 0.0°C$ $T_f = 70.0°C$	$Q_{heat\ liquid} = ?$
$C = 4180$ J/(kg·K)	$Q_{total} = ?$

2 SOLVE FOR THE UNKNOWN

Calculate the heat needed to melt ice.

$Q_{melt\ ice} = mH_f$
$= (1.50\ \text{kg})(3.34 \times 10^5\ \text{J/kg})$ Substitute $m = 1.50$ kg, $H_f = 3.34 \times 10^5$ J/kg.
$= 5.01 \times 10^5\ \text{J} = 5.01 \times 10^2\ \text{kJ}$

Calculate the temperature change.

$\Delta T = T_f - T_i$
$= 70.0°C - 0.0°C = 70.0°C = 70.0K$ Substitute $T_f = 70.0°C$, $T_i = 0.0°C$. $\Delta°C = \Delta K$

Calculate the heat needed to raise the water temperature.

$Q_{heat\ liquid} = mC\Delta T$
$= (1.50\ \text{kg})(4180\ \text{J/(kg·K)})(70.0K)$ Substitute $m = 1.50$ kg, $C = 4180$ J/(kg·K), $\Delta T = 70.0K$.
$= 4.39 \times 10^5\ \text{J} = 4.39 \times 10^2\ \text{kJ}$

Calculate the total amount of heat needed.

$Q_{total} = Q_{melt\ ice} + Q_{heat\ liquid}$
$= 5.01 \times 10^2\ \text{kJ} + 4.39 \times 10^2\ \text{kJ}$ Substitute $Q_{melt\ ice} = 5.01 \times 10^2$ kJ, $Q_{heat\ liquid} = 4.39 \times 10^2$ kJ.
$= 9.40 \times 10^2\ \text{kJ}$

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

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

21. Use **Figure 15** to calculate the heat of fusion and heat of vaporization of water in J/kg.

22. A steel plant worker will change 100.0 kg of 25°C iron into molten iron (melting point = 1538°C). How much thermal energy must be added?

23. **CHALLENGE** How much thermal energy is needed to change 3.00×10^2 g of ice at $-30.0°C$ to water vapor at 130.0°C?
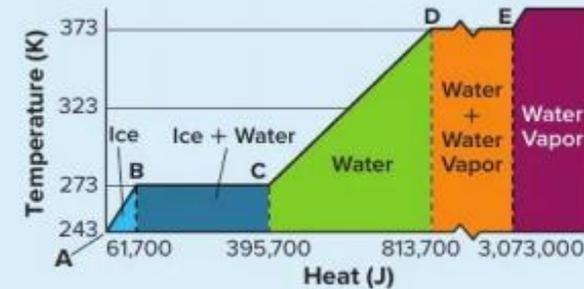

Figure 15

The change in length of a solid is proportional to the change in temperature, as shown in **Figure 21.** A solid will expand in length twice as much when its temperature is increased by 20°C than when it is increased by 10°C. The expansion also is proportional to its length. This means that a 2-m bar will expand twice as much as a 1-m bar with the same change in temperature. The length (L_2) of a solid at temperature T_2 can be found using the following relationship, where L_1 is the length at temperature T_1 and alpha (α) is the coefficient of linear expansion. The **coefficient of linear expansion** is equal to the change in length divided by the original length and the change in temperature.

$$L_2 = L_1 + \alpha L_1(T_2 - T_1)$$

With some algebra, you can solve for α.

$$L_2 - L_1 = \alpha L_1(T_2 - T_1)$$
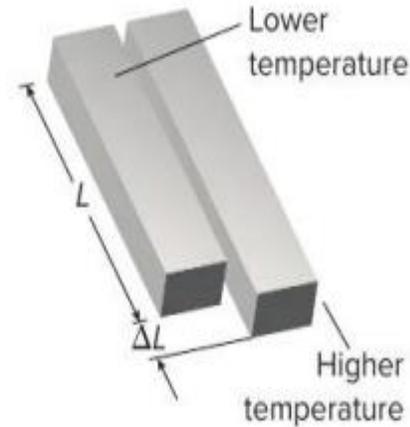
$$\Delta L = \alpha L_1 \Delta T$$

Figure 21 The change in length of a material is proportional to its original length and the change in temperature.

Coefficient of Linear Expansion

The coefficient of linear expansion is equal to the change in length divided by the product of the original length and the change in temperature.

$$\alpha = \frac{\Delta L}{L_1 \Delta T}$$

EXAMPLE Problem 4

LINEAR EXPANSION A metal bar is 1.60 m long at room temperature (21°C). The bar is put into an oven and heated to a temperature of 84°C. It is then measured and found to be 1.7 mm longer. What is the coefficient of linear expansion of this material?

1 ANALYZE AND SKETCH THE PROBLEM

- Sketch the bar, which is 1.7 mm longer at 84°C than at 21°C.
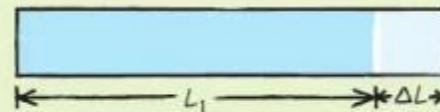- Identify the initial length of the bar (L_1) and the change in length (ΔL).

Known	Unknown
$L_1 = 1.60$ m	$\alpha = ?$
$\Delta L = 1.7 \times 10^{-3}$ m	
$T_1 = 21°C$	
$T_2 = 84°C$	

2 SOLVE FOR THE UNKNOWN

Calculate the coefficient of linear expansion using the relationship among known length, change in length, and change in temperature.

$$\alpha = \frac{\Delta L}{L_1 \Delta T}$$

$$= \frac{1.7 \times 10^{-3} \text{ m}}{(1.60 \text{ m})(84°C - 21°C)}$$

Substitute $\Delta L = 1.7 \times 10^{-3}$ m, $L_1 = 1.60$ m, $\Delta T = (T_2 - T_1) = 84°C - 21°C$

$$= 1.7 \times 10^{-5} \text{ °C}^{-1}$$

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39. A piece of aluminum house siding is 3.66 m long on a cold winter day of −28°C. How much longer is it on the hot summer day shown in **Figure 22?**
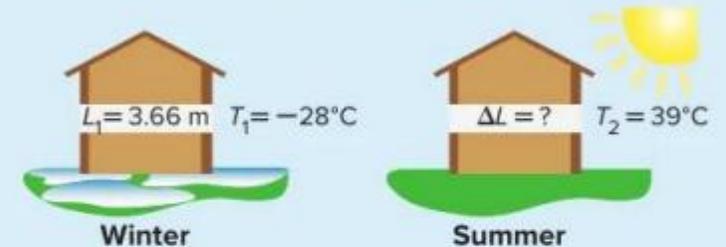

$L_1 = 3.66$ m $T_1 = -28°C$ $\Delta L = ?$ $T_2 = 39°C$

Winter   Summer

Figure 22

40. A piece of steel is 11.5 cm long at 22°C. It is heated to 1221°C, close to its melting temperature. How long is it?

41. A 400-mL glass beaker at room temperature is filled to the brim with cold water at 4.4°C. When the water warms up to 30.0°C, how much water will spill from the beaker?

The unit for the coefficient of linear expansion is the reciprocal of degrees Celsius (which can be written as 1/°C or °C⁻¹). The **coefficient of volume expansion** is equal to the change in volume divided by the product of the original volume and the change in temperature. The coefficient of volume expansion (β) is generally about three times the coefficient of linear expansion.

Coefficient of Volume Expansion

The coefficient of volume expansion is equal to the change in volume divided by the product of the original volume and the change in temperature.

$$\beta = \frac{\Delta V}{V_1 \Delta T}$$

The unit for β is also 1/°C (°C⁻¹). The two coefficients of thermal expansion for a variety of materials are given in **Table 2**. (See reference table in Student Resources for additional materials.)

42. A tank truck takes on a load of 45,725 L of gasoline in Houston, where the temperature is 28.0°C. The truck delivers its load in Minneapolis, where the temperature is −12.0°C.
    a. How many liters of gasoline does the truck deliver?
    b. What happened to the gasoline?

43. A hole with a diameter of 0.85 cm is drilled into a steel plate. At 30.0°C, the hole exactly accommodates an aluminum rod of the same diameter. What is the spacing between the plate and the rod when they are cooled to 0.0°C?

44. **CHALLENGE** A steel ruler is marked in millimeters so that the ruler is absolutely correct at 30.0°C. By what percentage would the ruler be incorrect at −30.0°C?

Material	Coefficient of Linear Expansion, α(°C⁻¹)	Coefficient of Volume Expansion, β(°C⁻¹)
Solids		
Aluminum	23×10⁻⁶	69×10⁻⁶
Glass (soft)	9×10⁻⁶	27×10⁻⁶
Glass (ovenproof)	3×10⁻⁶	9×10⁻⁶
Concrete	12×10⁻⁶	36×10⁻⁶
Copper	17×10⁻⁶	51×10⁻⁶
Liquids		
Methanol	Not Applicable	1200×10⁻⁶
Gasoline	Not Applicable	950×10⁻⁶
Water	Not Applicable	210×10⁻⁶

Heat engines A **heat engine** is a device that is able to continuously convert thermal energy to mechanical energy. A heat engine requires a high-temperature source, a low-temperature receptacle, called a sink, and a way to convert the thermal energy into work, as shown in **Figure 16.**

Internal combustion engines An automobile's internal combustion engine, shown in **Figure 17,** is one example of a heat engine. In the engine, input heat (Q_H) is transferred from a high-temperature flame to a mixture of air and gas vapor in the cylinder. The hot air expands and pushes on a piston, thereby changing thermal energy into mechanical energy. The heated air is expelled, the piston returns to the top of the cylinder, and the cycle repeats. Car engines repeat this cycle many times each minute. The thermal energy from the flame is converted into mechanical energy, which propels the car.
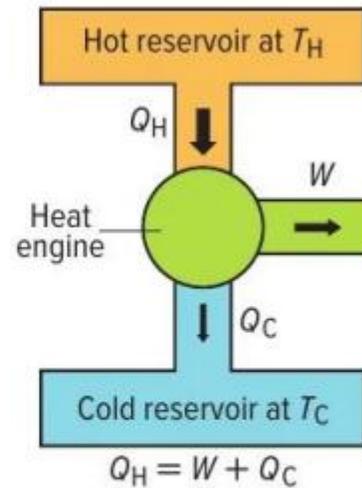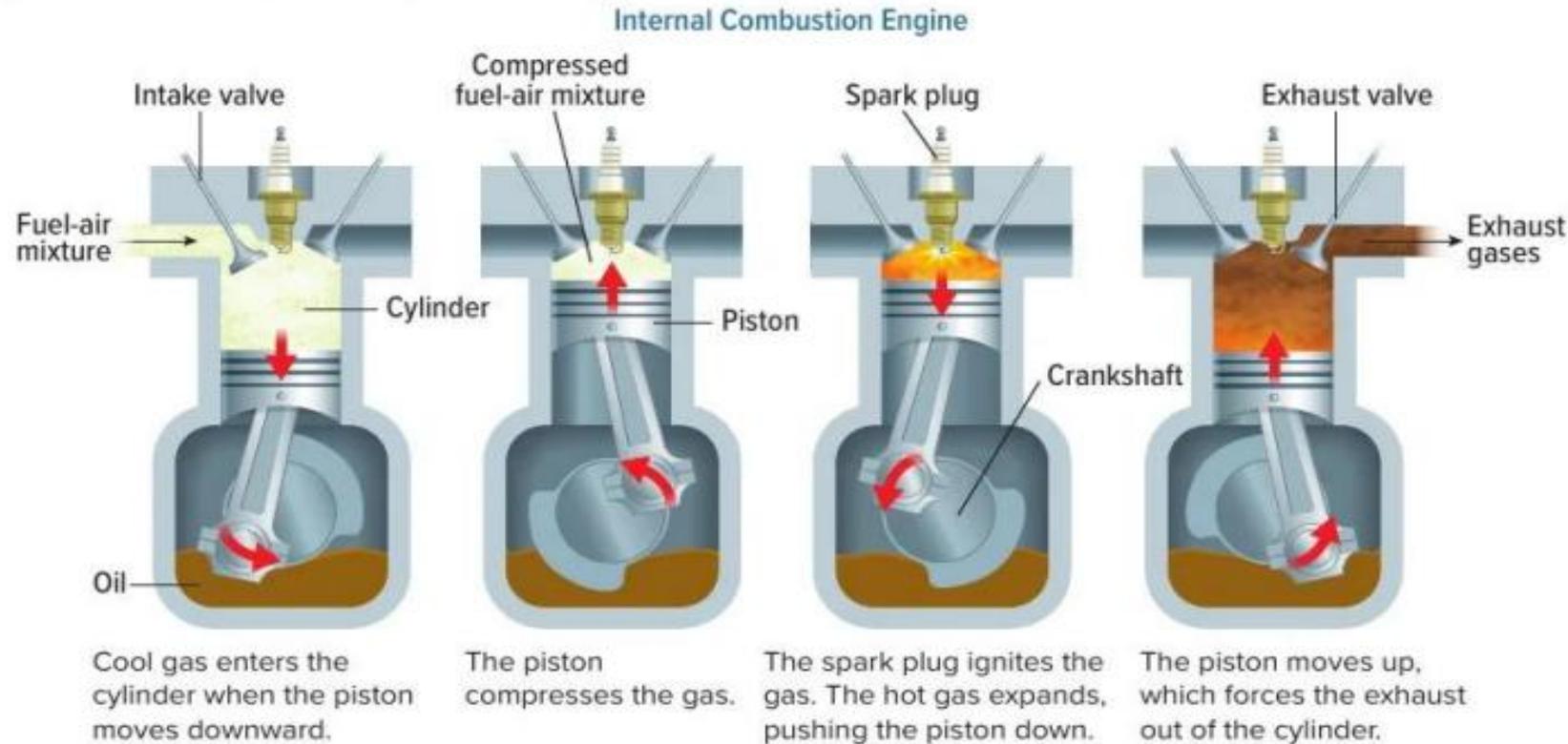

Energy Diagram of a Heat Engine

$$Q_H = W + Q_C$$

Figure 16 Heat engines transform thermal energy into mechanical energy and waste heat. This schematic shows the energy transfers and transformations.

Waste heat Not all thermal energy from the flame is converted into mechanical energy. When the engine is running, the gases and the engine parts become hot. The exhaust comes in contact with outside air and heats the air. In addition, thermal energy from the engine is transferred to a radiator. Outside air passes through the radiator and becomes warmer. All of this energy (Q_C) transferred out of the automobile engine is called waste heat. When the engine is working continuously, its internal energy does not change. That is, $\Delta U = 0 = Q - W$. The net heat going into the engine is $Q = Q_H - Q_C$. Thus, the work done by the engine is $W = Q_H - Q_C$. All heat engines generate waste heat, and no engine can convert all the energy into useful kinetic energy.

Use energy diagrams to show the energy transfers and transformations in a heat engine

Figure 17 Internal combustion engines are one type of heat engine. They are used in automobiles.

Internal Combustion Engine

Intake valve

Compressed fuel-air mixture

Spark plug

Exhaust valve

Fuel-air mixture

Exhaust gases

Cylinder

Piston

Crankshaft

Oil

Cool gas enters the cylinder when the piston moves downward.

The piston compresses the gas.

The spark plug ignites the gas. The hot gas expands, pushing the piston down.

The piston moves up, which forces the exhaust out of the cylinder.

State and explain the second law of thermodynamics. Define entropy

Energy spreads out Examine the melting ice pop and the cooling pizza in **Figure 20**. The first law of thermodynamics does not prohibit net thermal energy transfers from the cold ice pop to the air or from the air to the hot pizza. This does not occur, however, because of the second law of thermodynamics. When a hot object is placed in contact with cooler surroundings, the thermal energy in the hot object has the opportunity to disperse, or spread out more. Some of the thermal energy moves into the cold object, warming it and therefore cooling the originally hotter object. The **second law of thermodynamics** states that in an uncontrolled system whenever there is an opportunity for energy dispersal, the energy always spreads out toward more uniform energy distribution and to a more stable state.

Consider the cooling pizza. The particles in the pizza have a greater average kinetic energy than the particles in the air. Some of the pizza's original thermal energy disperses into the air. As a result, the pizza's temperature decreases and the air temperature increases a small amount. When the pizza and the air reach the same temperature, the average kinetic energy of the particles in the pizza and the air will be the same. That is, the energy spreads out among the particles. Similarly, if you leave the ice pop sitting on the counter, thermal energy from the air will be dispersed to the ice pop. The ice pop will heat up and melt, while the air will experience a small temperature decrease.

State and explain the second law of thermodynamics. Define entropy

Entropy The measure of this dispersal of energy is known as **entropy** (S). A system in which the thermal energy is concentrated in one place is referred to as a system with low entropy. A system in which the thermal energy is spread throughout the system has high entropy.

Another way of stating the second law of thermodynamics is that natural processes go in a direction that maintains or increases the total entropy of the universe. That is, energy will naturally disperse unless some action is taken to localize it. Once a system is in a high-entropy state, it is highly unlikely that it will return to a lower entropy state on its own. Events that occur spontaneously, such as the melting ice pop or the cooling pizza, are events in which the entropy of the system increases. Processes that would decrease the entropy of a system do not tend to occur spontaneously but require work done by an external agent.

Entropy and heat engines How does entropy relate to heat engines? If heat engines completely converted thermal energy into mechanical energy with no waste heat, energy would still be conserved, and so the first law of thermodynamics would be obeyed. However, waste heat is always generated, dispersing thermal energy beyond the engine. In the nineteenth century, French engineer Sadi Carnot studied the ability of engines to convert thermal energy into mechanical energy. He developed a logical proof that even an ideal engine would generate some waste heat. Carnot's result was one of the first formal analyses leading to the development of the concept of entropy.

Changes in entropy Like energy, entropy is a property of a system. If thermal energy is added to a system, the entropy increases. If thermal energy is removed from a system, its entropy decreases. If a system does work on its surroundings without any transfer of thermal energy, the entropy does not change. For a reversible process, the change in entropy (ΔS) is expressed by the following equation. Entropy has units of J/K. The temperature, in kelvins, is constant.

Change in Entropy

For a reversible process, the change in entropy of a system is equal to the heat added to the system divided by the temperature of the system in kelvins.

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

Its a Phys

Apply the conservation of energy for solving problems on specific heat = capacity involving calorimeters. State Boyle's laws as PV = constant or P1V1 P2V2 State Charles's law as V/T = constant or (V1/T1) = (V2/T2), at constant

Combined Gas Law

For a fixed amount of an ideal gas, the pressure times the volume, divided by the Kelvin temperature equals a constant.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} = \text{constant}$$

Ideal Gas Law

For an ideal gas, the pressure times the volume is equal to the number of moles times the constant R times the temperature.

$$PV = nRT$$

8. A tank containing 200.0 L of hydrogen gas at 0.0°C is kept at 156 kPa. The temperature is raised to 95°C, and the volume is decreased to 175 L. What is the new pressure of the gas?

9. **CHALLENGE** The average molar mass of the components of air (mainly diatomic nitrogen gas and diatomic oxygen gas) is about 29 g/mol. What is the volume of 1.0 kg of air at atmospheric pressure and 20.0°C?

6. A tank of helium gas used to inflate toy balloons is at a pressure of 15.5×10^6 Pa and a temperature of 293 K. The tank's volume is 0.020 m³. How large a balloon would it fill at 1.00 atmosphere and 323 K?

7. What is the mass of the helium gas in the previous problem? The molar mass of helium gas is 4.00 g/mol.

Pert

Its a Physics Adventure

Cohesive Forces

Figure 7 shows a water strider walking across the surface of a pond. This lightweight insect can do this because of surface tension—the tendency of the surface of a liquid to contract to the smallest possible area. Surface tension results from the cohesive forces among the particles of a liquid. **Cohesive forces** are the forces of attraction that like particles exert on one another. Notice that beneath the liquid's surface in **Figure 7,** each particle of the liquid is attracted equally in all directions by neighboring particles. As a result, no net force acts on any of the particles beneath the surface. At the surface, however, the particles are attracted downward and to the sides but not upward. There is a net downward force, which acts on the top layers and causes the surface layer to be slightly compressed. The surface layer acts like a tightly stretched sheet that is strong enough to support the weight of very light objects, such as the water strider.
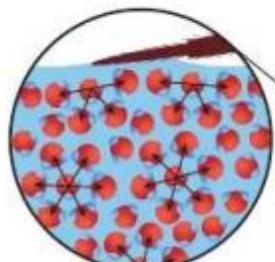
Adhesive Forces

Similar to cohesive forces, **adhesive forces** are attractive forces that act between particles of different substances. When a glass tube is placed in a beaker of water, the surface of the water climbs the outside of the tube, as shown in **Figure 8.** The adhesive forces between the particles that make up the glass and the water molecules are greater than the cohesive forces between the water molecules. In contrast, the cohesive forces between mercury atoms are greater than the adhesive forces between the mercury and the glass, so the liquid does not climb the tube. These forces also cause the mercury's surface to depress, as shown in **Figure 8.**


Side view

Water   Mercury

1

Conduction, convection, and radiation

Figure 8 shows the three types of heat—conduction, convection, and radiation. If you place one end of a metal rod in a flame, the hot gas conducts heat to the rod. The other end of the rod also becomes warm because the particles that make up the rod conduct thermal energy to their neighbors.

Convection   Conduction   Radiation
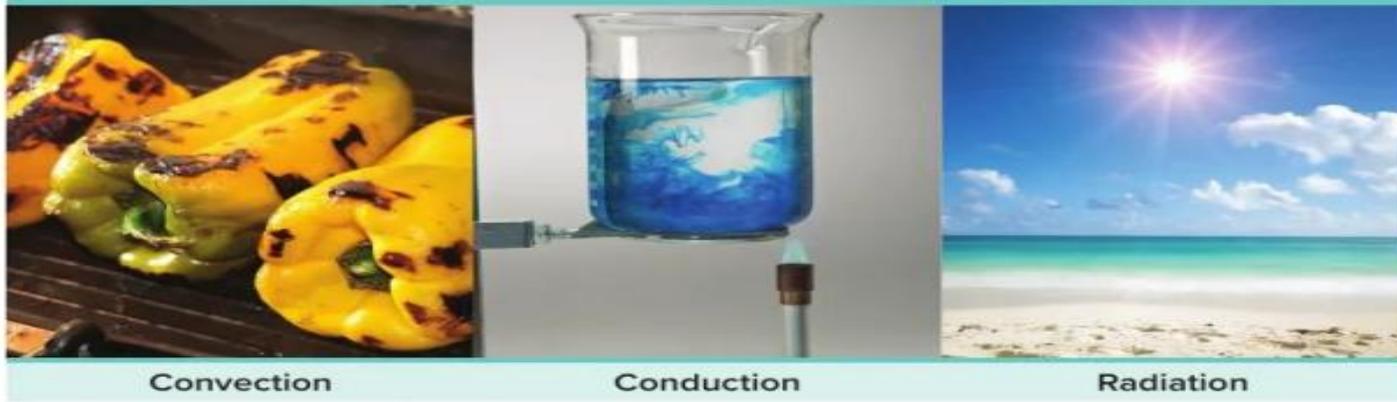
Have you ever noticed the motion on the surface of a pot of water just about to boil? The water at the bottom of the pot is heated by conduction, expands, and floats to the top, while the colder, denser water at the top sinks to the bottom. The motion of the hot water rapidly carries heat from the bottom of the pot to the top surface of the water. Heating caused by the motion of fluid in a liquid or gas due to temperature differences is called **convection.** Atmospheric turbulence is caused by convection of gases in the atmosphere. Thunderstorms and hurricanes are excellent examples of large-scale atmospheric convection. Convection also contributes to ocean currents that move water and materials over large distances.

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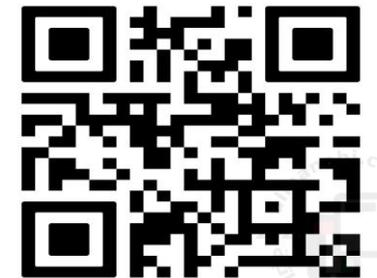
Solids, liquids, and pressure

Imagine you are standing on the surface of a frozen lake. The forces your feet exert on the ice are spread over the area of your shoes, resulting in pressure on the ice. Ice is a solid that is made up of vibrating water molecules, and the forces that hold the water molecules in place cause the ice to exert upward forces on your feet that equal your weight. If the ice melted, most of the bonds between the water molecules would be weakened. Although the molecules would continue to vibrate and remain close to each other, they also would slide past one another, and you would break through the surface. The moving water molecules would continue to exert forces on your body.

Gases and pressure

The pressure exerted by a gas can be understood by applying the kinetic-molecular theory of gases, which explains the properties of an ideal gas. In this model, particles are treated as taking up no space and having no intermolecular attractive forces. In spite of the fact that particles of a real gas take up space and exert attractive forces, an ideal gas is an accurate model of a real gas under most conditions.

Table 1 Some Typical Pressures

Location	Pressure (kPa)
The center of Earth	4×10^{8}
The deepest ocean trench	1.1×10^{5}
Standard atmosphere	1.01325×10^{2}
Blood pressure	1.6×10^{1}
Air pressure on top of Mt. Everest	3×10^{1}
The best vacuum	1×10^{-10}

You might have seen beaded water droplets on a freshly washed and waxed car. Why do these spherical drops form? The force pulling the surface particles into a liquid causes the surface to become as small as possible, and the shape that has the least surface for a given volume is a sphere. The higher the surface tension of the liquid, the more resistant the liquid is to having its surface broken. For example, liquid mercury has much stronger cohesive forces than water does. Liquid mercury forms spherical drops, even when it is placed on a smooth surface. A drop of water flattens out on a smooth surface.
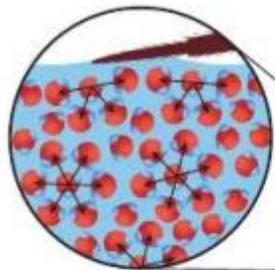
Side view

Figure 7 A water strider can walk on water because molecules at the surface experience a net downward force. Below the surface, each particle of liquid is equally attracted in all directions.

Water   Mercury

Figure 8 Due to adhesive forces, water climbs the outside wall of the glass tube. In the mercury, however, the forces of attraction between mercury atoms are stronger than any adhesive forces between the mercury and the glass. Therefore, the mercury is depressed by the tube.

To analyse first law of thermodynamics, the mathematical equation and application in solving problems,, Working of a heat engine, compare the energy diagrams of heat engine and refrigerator, explain the working of a heat engine,. Explain how second law of thermodynamics and entropy is connected to heat engine.

Energy Diagram of a Heat Engine

Hot reservoir at T_H

Q_H

Heat engine

W

Q_C

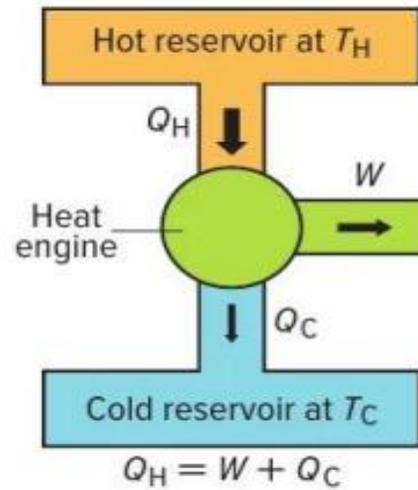Cold reservoir at T_C

$$Q_H = W + Q_C$$

Figure 16 Heat engines transform thermal energy into mechanical energy and waste heat. This schematic shows the energy transfers and transformations.

Energy Diagram of a Refrigerator

Hot reservoir at T_H

Q_H

Refrigerator

W

Q_C

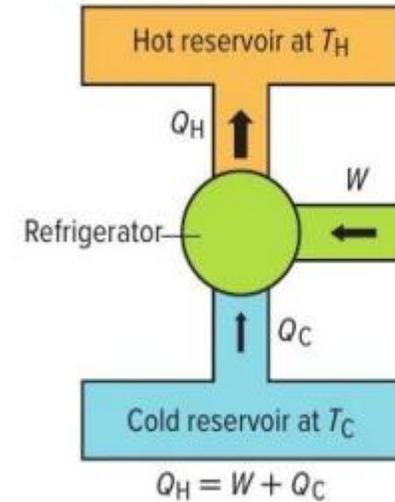Cold reservoir at T_C

$$Q_H = W + Q_C$$

Figure 19 When work is done on the refrigerator, thermal energy is transferred from the cold reservoir to the hot reservoir.

Identify that heat Q is positive if thermal energy is absorbed by an object and negative if thermal energy is transferred from an object

Student Textbook

121

When a liquid freezes, an amount of thermal energy ($Q = -mH_f$) must be removed from the liquid to turn it into a solid. The negative sign indicates that the thermal energy is transferred from the sample to the external world. In the same way, when a vapor condenses to a liquid, an amount of thermal energy ($Q = -mH_v$) must be removed from the vapor.

Water absorbs significant amounts of thermal energy when it melts or evaporates. Every day you use the large heats of fusion and vaporization of water. Each gram of sweat that evaporates from your skin carries off about 2.3 kJ of thermal energy. This is one cooling process that many warm-blooded animals use to regulate their body temperatures. Similarly the melting of a 24-g cube of ice absorbs enough thermal energy, 8 kJ, to cool a glass of water by about 30°C.

Measuring heat When a substance is heated, the substance's temperature can change. The change of temperature (ΔT) depends on heat (Q), the mass of the substance (m), and the specific heat of the substance (C). By using the following equation, you can calculate the heat (Q) required to change the temperature of an object.

Heat

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

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

For example, when the temperature of 10.0 kg of water is increased from 80 K to 85 K, the heat is $Q = (10.0 \text{ kg})(4180 \text{ J/(kg·K)})(85 \text{ K} - 80 \text{ K}) = 2.1 \times 10^5$ J. Remember that the temperature interval for the Kelvin scale is the same as that for the Celsius scale. For this reason, you can calculate ΔT on the Kelvin scale or on the Celsius scale.

Identify that heat Q is positive if thermal energy is absorbed by an object and negative if thermal energy is transferred from an object

Student Textbook

121

EXAMPLE Problem 3

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

1.5 kg
Snow
H_f
$T_i = 0.0°C$
$T_f = 70.0°C$

1 ANALYZE AND SKETCH THE PROBLEM

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

Known

$m = 1.50$ kg
$T_i = 0.0°C$
$C = 4180$ J/(kg·K)

$H_f = 3.34×10^5$ J/kg
$T_f = 70.0°C$

Unknown

$Q_{melt\ ice} = ?$
$Q_{heat\ liquid} = ?$
$Q_{total} = ?$

2 SOLVE FOR THE UNKNOWN

Calculate the heat needed to melt ice.

$Q_{melt\ ice} = mH_f$
$= (1.50$ kg$)(3.34×10^5$ J/kg$)$ Substitute $m = 1.50$ kg, $H_f = 3.34×10^5$ J/kg.
$= 5.01×10^5$ J $= 5.01×10^2$ kJ

Calculate the temperature change.

$\Delta T = T_f - T_i$
$= 70.0°C - 0.0°C = 70.0°C = 70.0K$ Substitute $T_f = 70.0°C$, $T_i = 0.0°C$. $\Delta°C = \Delta K$

Calculate the heat needed to raise the water temperature.

$Q_{heat\ liquid} = mC\Delta T$
$= (1.50$ kg$)(4180$ J/(kg·K$))(70.0K)$ Substitute $m = 1.50$ kg, $C = 4180$ J/(kg·K), $\Delta T = 70.0K$.
$= 4.39×10^5$ J $= 4.39×10^2$ kJ

Calculate the total amount of heat needed.

$Q_{total} = Q_{melt\ ice} + Q_{heat\ liquid}$
$= 5.01×10^2$ kJ $+ 4.39×10^2$ kJ   Substitute $Q_{melt\ ice} = 5.01×10^2$ kJ, $Q_{heat\ liquid} = 4.39×10^2$ kJ.
$= 9.40×10^2$ kJ

Plasma

If you heat a solid, it melts to form a liquid. Further heating results in a gas. What happens if you increase the temperature still further? Collisions between the particles become violent enough to tear the electrons off the atoms, thereby producing positively charged ions. The gaslike state of negatively charged electrons and positively charged ions is called **plasma.** Plasma is considered to be another state of matter.

The plasma state may seem to be uncommon, but plasma is actually the most common state of matter in the universe. Stars consist mostly of plasma at extremely high temperatures. Much of the matter between stars and galaxies consists of energetic hydrogen that has no electrons. This hydrogen is in the plasma state. The primary difference between gas and plasma is that plasmas can conduct an electric current, whereas gases cannot. Lightning bolts are in the plasma state. Neon signs, such as the one shown in **Figure 6,** contain plasma. The fluorescent bulbs that probably light your school also contain plasma.
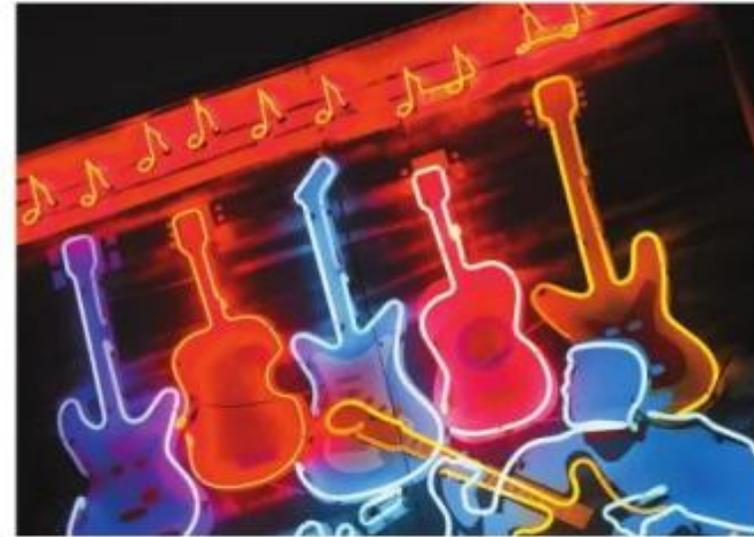
Figure 6 Plasma emits light as it conducts electricity. The color produced by glowing plasma depends on the gas inside the tube.

Solid Bodies

Solids are stiff, can be cut, and retain their shapes. Liquids flow and do not retain their shapes. When the temperature of a liquid is lowered, the average kinetic energy of the particles decreases. As the particles slow down, the cohesive forces have more effect, and for many solids, the particles become frozen into a fixed pattern called a **crystal lattice,** shown in **Figure 19.** Although the cohesive forces hold the particles in place, the particles in a crystalline solid do not stop moving completely. Rather, they vibrate around their fixed positions. In other materials, such as butter and glass, the particles do not form a fixed crystalline pattern. A substance that has no regular crystal structure but does have a definite volume and shape is called an **amorphous solid.**

Pressure and freezing As a liquid becomes a solid, its particles usually fit more closely together than in the liquid state, making solids more dense than liquids. Also, an increase in the pressure on the surface of the liquid increases its freezing point. However, water is an exception. It is most dense at 4°C. Because water expands as it freezes, an increase in pressure forces the molecules closer together and opposes the freezing. Therefore, higher pressure lowers the freezing point of water very slightly.

Solid Ice   **Crystal Lattice**

- = O
- = H

Figure 19 As the temperature of water is lowered and it changes from a liquid to a solid, the particles are frozen in a pattern called a crystal lattice.

Elasticity of solids External forces applied to a solid object may twist, stretch, or bend it out of shape. The ability of a solid object to return to its original form when the external forces are removed is called the elasticity of the solid. If too much deformation occurs, the object will not return to its original shape because its elastic limit has been exceeded. Elasticity is a property of each substance and depends on the forces holding its particles together. Malleability and ductility are two properties that depend on the structure and elasticity of a substance. Because gold can be flattened and shaped into thin sheets, it is said to be malleable. Copper is a ductile metal because it can be pulled into thin strands of wire.

Thermal Expansion of Solids

It is standard practice for engineers to design small gaps, called expansion joints, into concrete-and-steel highway bridges to allow for the expansion of parts in the heat of summer. Expansion joints are shown in **Figure 20**. Objects expand only a small amount when they are heated, but that small amount could be several centimeters in a 100-m-long bridge. If expansion gaps were not present, the bridge could buckle, or parts of it could break.

Some materials, such as the ovenproof glass used for laboratory experiments and cooking, are designed to have the least possible thermal expansion. Large telescope mirrors are made of a ceramic material that is designed to undergo essentially no thermal expansion.

To understand the expansion of heated solids, picture a solid as a collection of particles connected by springs that represent the attractive forces between the particles. When the particles get too close, the springs push them apart. When a solid is heated, the kinetic energy of the particles increases, and they vibrate rapidly and move farther apart, weakening the attractive forces between the particles. As a result, when the particles vibrate more violently with increased temperature, their average separation increases and the solid expands.

Its a Physics Adventure

Define efficiency of an engine as the ratio of useful work to the amount of heat energy ($e = W/QH \times 100$)

Calculate the efficiency of an internal combustion engine.

Efficiency Engineers and car salespeople often discuss the fuel efficiencies of automobile engines. They are referring to the amount of the input heat (Q_H) that is turned into useful work (W). The actual efficiency of an engine is given by the ratio W/Q_H. If all the input heat could be turned into useful work by the engine, the engine would have an efficiency of 100 percent. Because there is always waste heat (Q_C), even the most efficient engines fall short of 100 percent efficiency.

In fact, most heat engines are significantly less than 100 percent efficient. For example, even the most efficient automobile gasoline engines have an efficiency of less than 40 percent. A typical gasoline engine in an automobile has an efficiency that is closer to 20 percent. Energy is not destroyed, but a considerable amount of thermal energy transfers from a hot automobile engine to the surrounding environment.
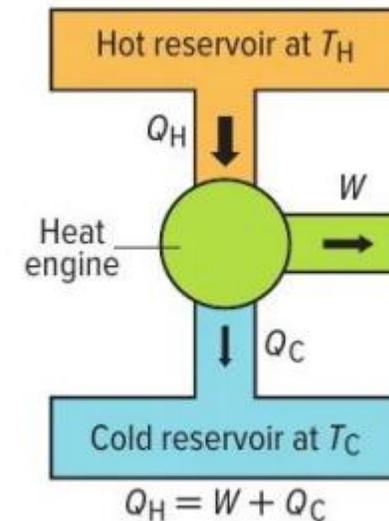

Energy Diagram of a Heat Engine

Hot reservoir at T_H

Q_H

W

Heat engine

Q_C

Cold reservoir at T_C

$$Q_H = W + Q_C$$

Figure 16 Heat engines transform thermal energy into mechanical energy and waste heat. This schematic shows the energy transfers and transformations.

29. **Heat Engine** Describe the energy transformations and transfers made by a heat engine, and explain why operating a heat engine causes an increase in entropy.

30. **Heat of Vaporization** Old heating systems sent water vapor into radiators in each room of a house. In the radiators, the water vapor condensed to water. Analyze this process and explain how it heated a room.

31. **Heat of Fusion** How much thermal energy is needed to change 50.0 g of ice at $-20.0°C$ to water at $10.0°C$?

32. **Heat of Vaporization** How much energy is needed to heat 1.0 kg of mercury metal from $10.0°C$ to its boiling point ($357°C$) and vaporize it completely? For mercury, $C = 140$ J/kg·°C and $H_v = 3.06 \times 10^5$ J/kg.

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

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

35. **Mechanical Energy and Thermal Energy** For the waterfall in **Figure 22,** calculate the temperature difference between the water at the top and the bottom of the fall. Assume that the potential energy of the water is all converted to thermal energy.

125.0 m

Figure 22

36. **Entropy** Evaluate why heating a home with natural gas results in increased entropy.

37. **Critical Thinking** Many outdoor amusement parks and zoos have systems that spray a fine mist of water, which evaporates quickly. Explain why this process cools the surrounding air.

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

$T_i = 0.0°C$ $T_f = 70.0°C$

1 ANALYZE AND SKETCH THE PROBLEM

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

Known

$m = 1.50$ kg

$T_i = 0.0°C$

$C = 4180$ J/(kg·K)

$H_f = 3.34×10^5$ J/kg

$T_f = 70.0°C$

Unknown

$Q_{melt\ ice} = ?$

$Q_{heat\ liquid} = ?$

$Q_{total} = ?$

2 SOLVE FOR THE UNKNOWN

Calculate the heat needed to melt ice.

$$Q_{melt\ ice} = mH_f$$
$$= (1.50\ kg)(3.34×10^5\ J/kg) \qquad \text{Substitute } m = 1.50\ kg,\ H_f = 3.34×10^5\ J/kg.$$
$$= 5.01×10^5\ J = 5.01×10^2\ kJ$$

Calculate the temperature change.

$$\Delta T = T_f - T_i$$
$$= 70.0°C - 0.0°C = 70.0°C = 70.0K \qquad \text{Substitute } T_f = 70.0°C,\ T_i = 0.0°C.\ \Delta°C = \Delta K$$

Calculate the heat needed to raise the water temperature.

$$Q_{heat\ liquid} = mC\Delta T$$
$$= (1.50\ kg)(4180\ J/(kg\text{-}K))(70.0K) \qquad \text{Substitute } m = 1.50\ kg,\ C = 4180\ J/(kg\cdot K),\ \Delta T = 70.0K.$$
$$= 4.39×10^5\ J = 4.39×10^2\ kJ$$

Calculate the total amount of heat needed.

$$Q_{total} = Q_{melt\ ice} + Q_{heat\ liquid}$$
$$= 5.01×10^2\ kJ + 4.39×10^2\ kJ \qquad \text{Substitute } Q_{melt\ ice} = 5.01×10^2\ kJ,\ Q_{heat\ liquid} = 4.39×10^2\ kJ.$$
$$= 9.40×10^2\ kJ$$

TRANSFERRING HEAT IN A CALORIMETER A calorimeter contains 0.50 kg of water at 15°C. A 0.10-kg block of an unknown substance at 62°C is placed in the water. The final temperature of the system is 16°C. What is the substance?

1 ANALYZE AND SKETCH THE PROBLEM

- Let the unknown be sample A and water be sample B.

- Sketch the transfer of thermal energy from the hotter unknown sample to the cooler water.
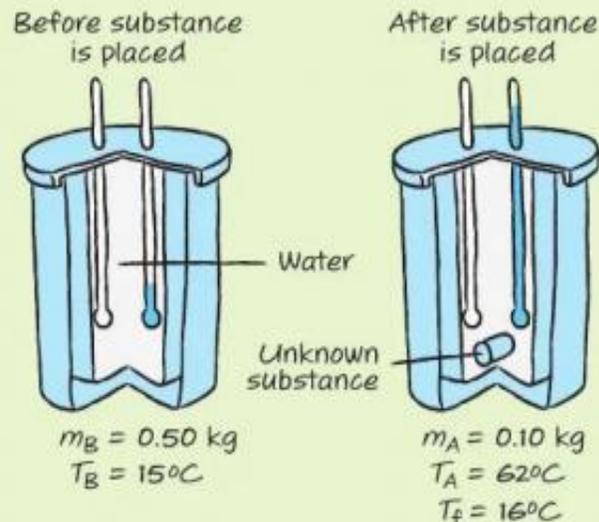
Known

$m_A = 0.10$ kg

$T_A = 62°C$

$m_B = 0.50$ kg

$C_B = 4180$ J/(kg·K)

$T_B = 15°C$

$T_f = 16°C$

Unknown

$C_A = ?$

Before substance is placed

After substance is placed

Water

Unknown substance

$m_B = 0.50$ kg
$T_B = 15°C$

$m_A = 0.10$ kg
$T_A = 62°C$
$T_f = 16°C$

2 SOLVE FOR THE UNKNOWN

Determine the final temperature using the following equation. Beware of the minus signs.

$$C_A = \frac{-m_B C_B \Delta T_B}{m_A \Delta T_A}$$

$$= \frac{-(0.50 \text{ kg})(4180 \text{ J/(kg·K)})(16°C - 15°C)}{(0.10 \text{ kg})(16°C - 62°C)}$$

Substitute $m_A = 0.10$ kg, $T_A = 62°C$, $m_B = 0.50$ kg, $C_B = 4180$ J/(kg·K), $T_B = 15°C$, $T_f = 16°C$.

$= 450$ J/(kg·K)

According to **Table 1,** the specific heat of the unknown substance equals that of iron.

4. A 1.00×10^2-g aluminum block at 100.0°C is placed in 1.00×10^2g of water at 10.0°C. The final temperature of the mixture is 26.0°C. What is the specific heat of the aluminum?

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

6. A 2.00×10^2-g sample of water at 80.0°C is mixed with 2.00×10^2 g of water at 10.0°C in a calorimeter. What is the final temperature of the mixture?

7. A 1.50×10^2-g piece of glass at a temperature of 70.0°C is placed in a container with 1.00×10^2 g of water initially at a temperature of 16.0°C. What is the equilibrium temperature of the water?

8. **CHALLENGE** A 4.00×10^2-g sample of water at 15.0°C is mixed with 4.00×10^2 g of water at 85.0°C. After the system reaches thermal equilibrium, 4.00×10^2 g of methanol at 15°C is added. Assume there is no thermal energy lost to the surroundings. What is the final temperature of the mixture?

PhysixsPert

Its a Physics Adventure

Discuss some engineering solutions and applications related to thermal expansion in solids, such as expansion joints, filling materials, the use of bimetallic strips in home thermostats, or others

practices problems

189

Applications of thermal expansion

Engineers take the thermal expansion of materials into consideration as they design structures. You've already read about the expansion joints that are installed on concrete highways and bridges. The regular gaps between slabs of concrete in sidewalks also help keep sidewalks from buckling when the concrete expands during hot weather. Different materials expand at different rates, as indicated by the different coefficients of expansion given in **Table 2**. Engineers also consider different expansion rates when designing systems. Steel bars are often used to reinforce concrete. The steel and concrete must have the same expansion coefficient. Otherwise, the structure could crack on a hot day. Similarly, filling materials used to repair teeth must expand and contract at the same rate as tooth enamel.

Different rates of expansion have useful applications. For example, engineers have taken advantage of these differences to construct a useful device called a bimetallic strip, which is used in thermostats. A bimetallic strip consists of two strips of different metals welded or riveted together. Usually, one strip is brass, and the other is iron. When heated, brass expands more than iron does. Thus, when the bimetallic strip of brass and iron is heated, the brass part of the strip becomes longer than the iron part. As a result, the bimetallic strip bends with the brass on the outside of the curve. If the bimetallic strip is cooled, it bends in the opposite direction. The brass is then on the inside of the curve.

In a home thermostat, the bimetallic strip is installed so that it bends toward an electric contact as the room cools. When the room cools below the setting on the thermostat, the bimetallic strip bends enough to make electric contact with the switch, which turns on the heater. As the room warms, the bimetallic strip bends in the other direction. When the room's temperature reaches the setting on the thermostat, the electric circuit is broken, and the heater switches off.

Thermal Expansion

As you applied the combined gas law, you discovered how gases expand as their temperatures increase. **Thermal expansion** is a property of all forms of matter that causes the matter to expand, becoming less dense, when heated. Thermal expansion has many useful applications, such as circulating air in a room.

Convection currents **Figure 5** shows that when the air near the radiator of a room is heated, it becomes less dense and, therefore, rises. Gravity pulls the denser, colder air near the ceiling down. The cold air is subsequently warmed by the radiator, and air continues to circulate. This circulation of air within a room is called a convection current. Convection currents also occur in a pot of hot, but not boiling, water on a stove. When the pot is heated from the bottom, the colder and denser water sinks to the bottom where it is warmed and then pushed up by the continuous flow of cooler water from the top.

This thermal expansion occurs in most fluids. A good model for all liquids does not exist, but it is useful to think of a liquid as a finely ground solid. Groups of two, three, or more particles move together as if they were tiny pieces of a solid. When a liquid is heated, particle motion causes these groups to expand in the same way that particles in a solid are pushed apart. The spaces between groups increase. As a result, the whole liquid expands. With an equal change in temperature, liquids expand considerably more than solids but not as much as gases.
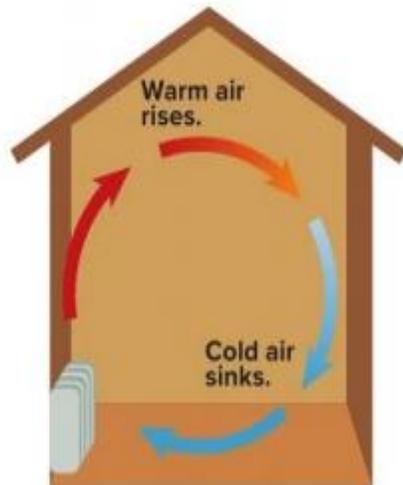

Warm air rises.

Cold air sinks.

Figure 5 Convection currents occur as warmer, less dense air rises and cooler, denser air sinks.

Why ice floats Because matter expands as it is heated, you might predict that ice would be more dense than water, and therefore, it should sink. However, when water is heated from 0°C to 4°C, instead of expanding, it contracts as the forces between particles increase and the ice crystals collapse. These forces between water molecules are strong, and the crystals that make up ice have a very open structure. Even when ice melts, tiny crystals remain. These remaining crystals are melting, and the volume of the water decreases until the temperature reaches 4°C. However, once the temperature of water moves above 4°C, its volume increases because of greater molecular motion. The practical result is that water is most dense at 4°C, and ice floats. This unique property of water is very important to our lives and environment. If ice sank, lakes would freeze from the bottom each winter and many would never melt completely in the summer.
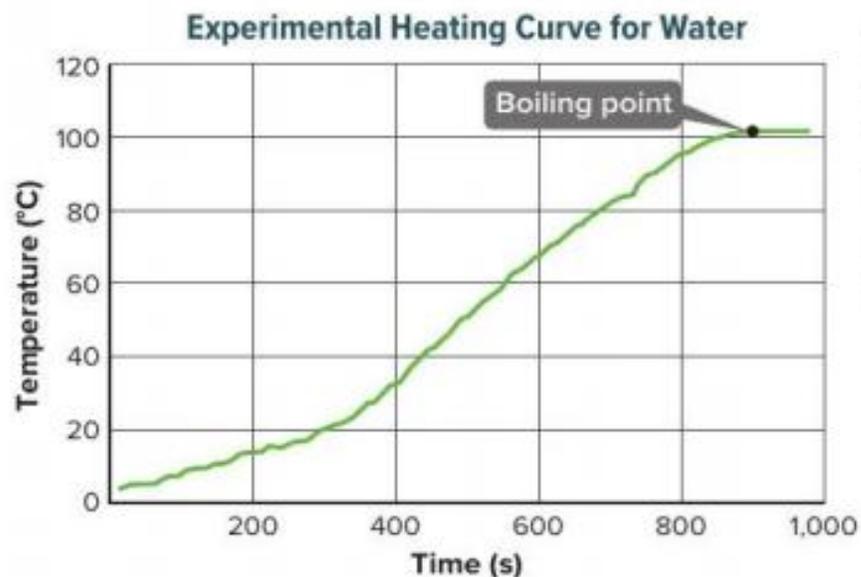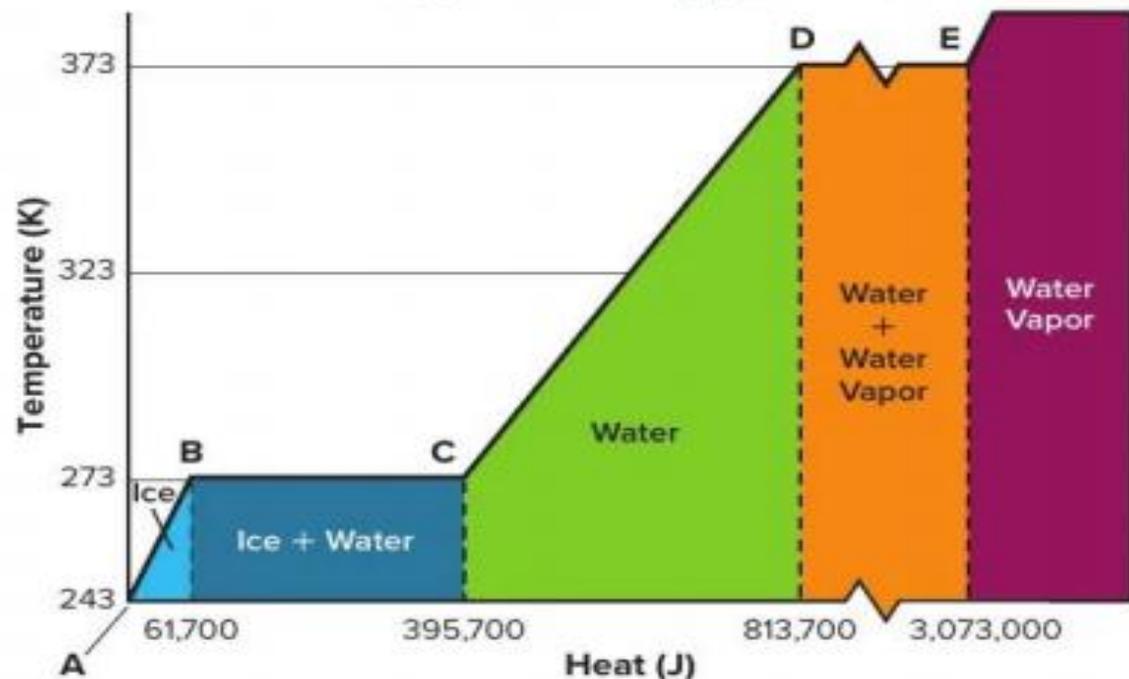
Experimental Heating Curve for Water

Figure 14 One way to measure the absorption of energy by a material is to add energy from a constant source of thermal energy and measure the change in temperature over time. The plot of temperature v. time is called a heating curve. For this figure, a beaker of cold water was placed on a hot plate. The resulting heating curve is graphed.

Explain *why thermal energy must be added at a constant rate to calculate the specific heat of water from this graph.*

Energy and changes of state There is a definite slope to the graph in **Figure 14** between about 300 s and 800 s. Heat is added at a constant rate, so this slope is proportional to the reciprocal of the specific heat of water. The slope between points A and B in **Figure 12** is proportional to the reciprocal of the specific heat of ice, and the slope above point E is proportional to the reciprocal of the specific heat of water vapor. The slope for water is less than those of both ice and water vapor. This is because water has a greater specific heat than does ice or water vapor. The heat (Q) required to melt a solid of mass (m) is given by the following equation.

Adding Thermal Energy to Water

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

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

21. Use **Figure 15** to calculate the heat of fusion and heat of vaporization of water in J/kg.

22. A steel plant worker will change 100.0 kg of 25°C iron into molten iron (melting point = 1538°C). How much thermal energy must be added?

23. **CHALLENGE** How much thermal energy is needed to change 3.00×10^2 g of ice at $-30.0°C$ to water vapor at $130.0°C$?
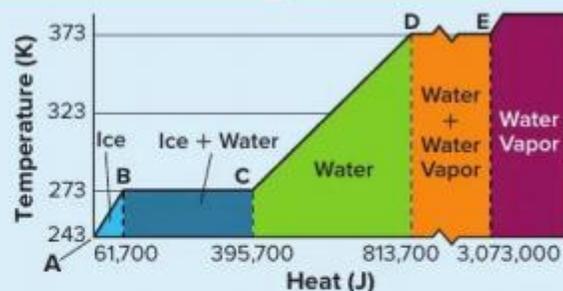

Figure 15