13-1 Temperature Scales

In the next chapter, we'll come to a fundamental understanding of what temperature is. Let's begin our investigation of thermal physics, however, by discussing various temperature scales and how to measure temperature.

Temperature is something that has an important impact on our daily lives. In the United States, temperatures in a weather forecast are generally specified in Fahrenheit. In most of the rest of the world, however, such temperatures are given in Celsius. It is useful to know how to convert a temperature from one unit to another. The picture of the thermometer in Figure 13.1 helps us to understand the conversion process to go from Fahrenheit to Celsius, or vice versa, as well as the conversion between Celsius and Kelvin. Note that a temperature of –40°F is the same as a temperature of –40°C. Starting there, every change by 5°C is equivalent to a change of 9°F. This is where the factor of 9/5, or 5/9, in the conversion equations comes from.

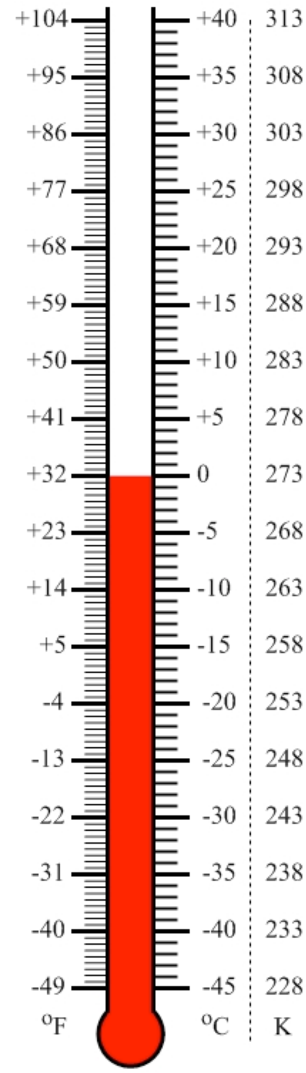

Figure 13.1: A thermometer that is immersed in a cup of ice water (not shown). The thermometer is calibrated in Fahrenheit (°F), Celsius (°C), and Kelvin (K).

$$T_C = \left(\frac{5°\text{C}}{9°\text{F}}\right)(T_F - 32°\text{F}).$$ (Equation 13.1: **Converting from Fahrenheit to Celsius**)

$$T_F = 32°\text{F} + \left(\frac{9°\text{F}}{5°\text{C}}\right)T_C.$$ (Equation 13.2: **Converting from Celsius to Fahrenheit**)

The Celsius scale is convenient for measuring everyday temperatures. The scale is based on the properties of water, with 0°C corresponding to the freezing point of water, and 100°C corresponding to its boiling point (at standard atmospheric pressure, at least). In scientific work temperatures are usually measured in Celsius, or in Kelvin. The Kelvin scale is an absolute temperature scale, because its zero corresponds to absolute zero. Because an increase by 1°C corresponds to an increase of 1K, it is easy to convert between the two scales.

$$T_C = T_K - 273.16°.$$ (Equation 13.3: **Converting from Kelvin to Celsius**)

In this three-chapter sequence on thermal physics, we will use several equations that involve either the temperature T or the change in temperature ΔT. When an equation involves T the temperature is an absolute temperature – use a temperature in Kelvin. When an equation has ΔT, we can use Celsius or Kelvin, because the change in temperature is the same on the two scales.

Measuring Temperature

There are many devices that can be used to measure temperature, such as the glass thermometer illustrated in Figure 13.1. These used to be filled with mercury but, with mercury now known to have negative health and environmental effects, such thermometers are now usually filled with alcohol. Almost any property of a material that changes with temperature can be exploited to make a temperature-measuring device. Examples include:

- A thermocouple – Two different kinds of metal are bonded together at two junctions. A temperature difference between the two junctions gives rise to a voltage (we'll explore that concept later in the book) that corresponds to that temperature difference.
- A gas-law thermometer – The pressure of a gas in a bulb of fixed volume can be directly converted to temperature via the ideal gas law (see Chapter 14).
- A thermopile – The higher its temperature, the more energy an object gives off. A thermopile picks up such radiated energy, which can be converted to temperature.

How does an alcohol-filled thermometer measure temperature? The thermometer, initially at room temperature (about 20°C), is placed in some hot water. To measure the water temperature, several things happen, which we will explore in the next three chapters:

- Energy is transferred from the water to the thermometer, raising the temperature of the thermometer (and lowering the temperature of the water by a small amount). This transferred energy is known as **heat**. Energy is always spontaneously transferred from the higher-temperature object to the lower-temperature object. Energy can only be transferred in the reverse direction with the aid of something else, such as a heat pump.

- The energy is transferred through the glass wall of the thermometer into the alcohol, through a process known as **thermal conduction**.

- The transfer of energy continues until **thermal equilibrium** is reached (when the thermometer reaches the same temperature as the water).

- As the temperature of the alcohol increases, the alcohol expands, causing the level of alcohol to rise in the thermometer. We exploit this expansion to measure temperature.

We will explore this last process, known as **thermal expansion**, in the next section.

Related End-of-Chapter Exercises: 18 and 20

A note about "heat" and "temperature"

We will discuss the ideas of heat and temperature in more detail in this chapter and in Chapters 14 and 15, but it is not too early to begin distinguishing between these two concepts. The word heat, as a noun, is reserved for a transfer of energy because of a temperature difference. Thus, it is incorrect to say that a hot object contains heat. A hot object has a high temperature, and thus has more **internal energy** (energy associated with the motion of the object's atoms and molecules) when it is warmer than when it is cooler. Temperature, as we will see, is a direct measure of this internal energy.

To add to the confusion, we also use "heat" as a verb, such as "I put the kettle on the stove to heat water so I could make tea." However, note that the meaning of heat as a transfer of energy is preserved here. In this situation, the water's internal energy, and therefore its temperature, increases because of heat (energy transferred to it) from, for instance, a high-temperature element on the stove.

Essential Question 13.1: Consider an alcohol thermometer of the kind drawn in Figure 13.1. Increasing the temperature causes both the alcohol and the glass shell of the thermometer to expand. Predict whether you think the cavity inside the glass (the volume occupied by the alcohol, and the space above the alcohol) increases or decreases because of the expansion of the glass. How does this affect the level of the alcohol? Is the level higher or lower because of the glass expansion than would be achieved by the expansion of the alcohol alone?

Answer to Essential Question 13.1: We'll address this issue further in the next section, but a prediction that the cavity volume decreases because of the expansion of the glass is consistent with an alcohol level higher than it would be if only the alcohol expanded. Increasing the cavity volume corresponds to an alcohol level lower than it would be if only the alcohol expanded.

13-2 Thermal Expansion

When an object's temperature changes, we assume the change in length ΔL experienced by each dimension of the object is proportional to its change in temperature ΔT. This model is valid as long as ΔT is not too large. We can express this linear relationship as an equation:

$\Delta L = L_i \alpha \Delta T$, (Equation 13.4: **Length change from thermal expansion**)

where L_i is the original length and α is known as the linear thermal expansion coefficient, which depends on the material (see Table 13.1).

Material	**α (× 10^{-6} /°C)**
Aluminum	23
Brass	19
Copper	17
Diamond	1
Glass	8.5
Gold	14
Iron or steel	12
Lead	29
Stainless steel	17

Table 13.1: Values of α, the linear thermal expansion coefficient, for various materials at 20°C.

Another way to express the linear nature of the model is to relate the initial length L_i to the new length L. Because $L = L_i + \Delta L$ we can express the new length as:

$L = L_i + L_i \alpha \Delta T = L_i (1 + \alpha \Delta T)$. (Equation 13.5)

Consider now a rectangular object with an area A_i that is the product of its height H_i and its width W_i. Both the height and width change with temperature, so the new area is:

$$A = HW = H_i (1 + \alpha \Delta T) W_i (1 + \alpha \Delta T) = H_i W_i (1 + \alpha \Delta T)(1 + \alpha \Delta T) = A_i \left[1 + 2\alpha \Delta T + (\alpha \Delta T)^2\right]$$

Because values of α are small, for small temperature changes $\alpha \Delta T$ is significantly less than 1. Squaring a number less than 1 makes it even smaller, so the $(\alpha \Delta T)^2$ term in the previous equation is negligible compared to the $2\alpha \Delta T$ term. Thus we can write the area equation as:

$A = A_i (1 + 2\alpha \Delta T)$. (Equation 13.6: **Area thermal expansion**)

A similar process can be followed to show that the equation for the volume V resulting from imposing a temperature change ΔT on an object of original volume V_i is:

$V = V_i (1 + 3\alpha \Delta T)$. (Equation 13.7: **Volume thermal expansion**)

EXAMPLE 13.2 – Shrink to fit

You are building a plane, but the stainless steel rivets you are using will not fit through the holes in the skin of the plane. At 20°C, each rivet has a diameter of 12.020 mm, while the diameter of a hole is 12.000 mm. To what temperature should you cool a rivet so it fits in a hole?

SOLUTION

We need to find the temperature change required so the diameter of the rivet equals the diameter of the hole. To find this we can use Equation 13.5, $L = L_0 (1 + \alpha\,\Delta T)$, with the thermal expansion coefficient from Table 13.1 for stainless steel, $\alpha = 17 \times 10^{-6}\,/°\text{C}$. Solving for ΔT gives:

$$\Delta T = \frac{L - L_0}{L_0 \alpha} = \frac{12.000\,\text{mm} - 12.020\,\text{mm}}{(12.020\,\text{mm})(17 \times 10^{-6}\,/°\text{C})} = \frac{-0.020\ \text{mm}}{204.34 \times 10^{-6}\ \text{mm}/°\text{C}} = -98°\text{C}\,.$$

Because the initial temperature is 20°C, we get a final temperature of –78°C. Note that the equation involves ΔT, so we can work entirely in Celsius. Also, there is no reason to convert the lengths to meters because the length units cancel out. One way to cool the rivets this much would be to immerse them in liquid nitrogen, which has a temperature of about –196°C, or 77K.

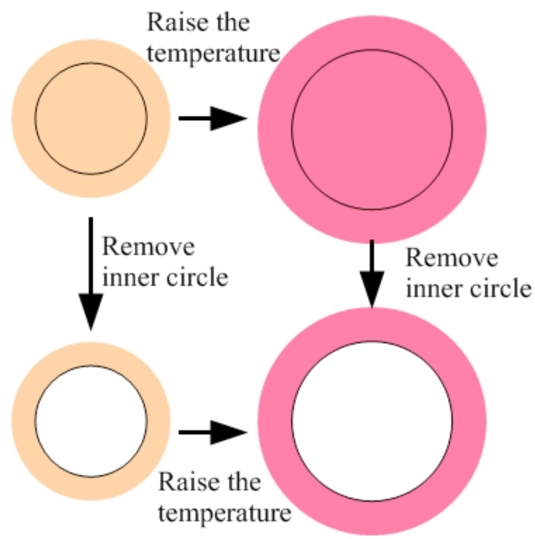

Figure 13.2: When the temperature of a disk increases, the entire disk expands, including the circle that was drawn on the disk. Removing the area within the circle leaves a ring with an inside diameter larger than the original diameter of the circle. Doing this in reverse order, removing the inside of the circle and then raising the temperature, has the same result. Holes expand as if they are filled with the surrounding material.

EXPLORATION 13.2 – What do holes do?

Imagine trying to fit a brass ball through a brass ring that has an inner radius just a little smaller than the ring. Should we heat the ring or cool it so that the ball can pass through?

Step 1 – ***Take a solid disk with a radius equal to the outer radius of the ring. Draw a circle on this disk so the radius of the circle matches the inner radius of the ring. What happens when the temperature of the disk increases?*** A picture of this situation is shown in Figure 13.2. The disk expands when its temperature increases, and so does the circle drawn on the disk. Removing the material inside the circle leaves a ring.

Step 2 – ***Reverse the order of the actions. First remove the inner part of the circle and then increase the temperature. What is the result?*** Figure 13.2 shows that changing the order of the actions makes no difference. The ring is the same size in both cases – a hole expands or contracts as if it were solid.

Key idea for holes: Holes and cavities in materials expand and contract as if they were made from the surrounding material.
Related End-of-Chapter Exercises: 17 – 19.

Increasing the temperature of a solid makes its atoms vibrate more energetically, so each atom effectively needs more space. If we imagine a circle of atoms around the inside diameter of the ring, as in Figure 13.3, the atoms spread out when heated, making the hole in the ring larger.

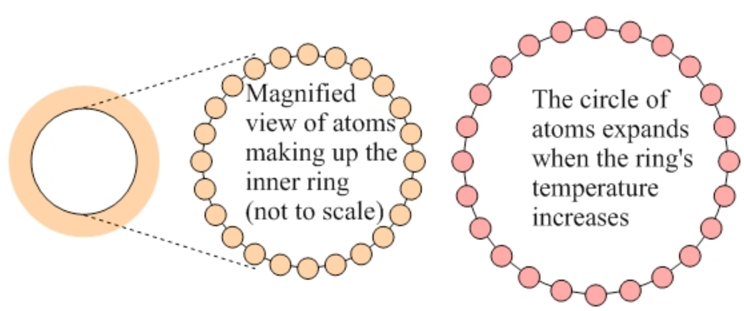

Figure 13.3: An object expands when its temperature increases because its atoms vibrate more energetically. This causes the circle of atoms along the inside diameter of a ring to expand when the temperature increases.

Essential Question 13.2: An iron ring will not quite fit over an aluminum cylinder. Can you fit the ring over the cylinder by increasing or decreasing the temperature of both objects at the same time? Explain.

Answer to Essential Question 13.2: The key here is that aluminum has a larger expansion coefficient than iron (see Table 13.1), so the aluminum expands or contracts more than does the iron ring for the same change in temperature. Increasing the temperature would make the mismatch worse but, if we decrease the temperature enough, both objects will contract, with the aluminum contracting more so the iron ring will slide onto the cylinder.

13-3 Specific Heat

Heat is energy transferred between objects that are at different temperatures, transferred from the higher-temperature object to the lower-temperature object. Although this is a form of energy we have not discussed in previous chapters, we will deal with it in much the same way we dealt with other forms of energy. Conservation of energy, for instance, will be a very useful tool, as it was previously.

The type of question we'll deal with next is the following: what is the equilibrium temperature when a 500-gram lead ball at 100°C is added to 400 g of water that is at room temperature, 20°C? To answer this kind of question we will assume that no energy is transferred into or out of the system, and we will also use a simple model that says that when energy in the form of heat is added to or removed from a substance, the substance's change in temperature is proportional to the amount of heat added or removed. The equation that goes with this statement is the following, where the symbol Q is used to represent heat, with units of energy:

$Q = mc\Delta T$, (Eq. 13.8: **The heat required to change an object's temperature**)

where m is the mass of the substance, c is known as the specific heat capacity (a constant that depends on the material), and ΔT is the temperature change. Note that, if heat is added to a substance, Q and ΔT are both positive; if heat is removed, Q and ΔT are both negative. Table 13.2 shows values of specific heat capacity for various materials. Equation 13.8 is generally valid as long as the substance does not change phase (such as from solid to liquid). We will learn how to deal with phase changes in the next section.

Material	**c, specific heat capacity (J / kg °C)**
Aluminum	900
Brass	377
Copper	385
Gold	128
Iron	449
Lead	129
Water (gas)	1850
Water (liquid)	4186
Water (solid)	2060

Table 13.2: Specific heat capacities for several solid metals, and for three common forms of water. The value for gaseous water is valid at standard atmospheric pressure.

Specific heat capacities are sometimes given in calories / (g °C). One calorie is the amount of heat needed to change the temperature of 1 g of water by 1°C. Thus 1000 cal is needed to change the temperature of 1 kg (1000 g) of water by 1°C. Looking at the specific heat capacity of liquid water in Table 13.2, we can see that 4186 J is equivalent to 1000 cal, giving us a conversion factor of 4.186 J/cal.

EXPLORATION 13.3 – The large heat capacity of water

A 500-gram lead ball that is initially at 100°C is added to an unknown mass of water that is initially at room temperature, 20°C, in a Styrofoam cup. After allowing the system to come to thermal equilibrium (i.e., waiting until the lead ball and the water have the same temperature), the final temperature of the system is measured to be 60°C, exactly halfway between the initial temperatures of the lead ball and the water. Based on this, determine the mass of water in the system.

Step 1 – ***Using Table 13.2, use a qualitative argument to predict how the mass of water compares to the mass of the lead ball.*** Both the lead and the water experience a temperature change of the same magnitude, with the temperature of the lead falling 40°C and the water temperature rising 40°C. If their heat capacities were equal, therefore, their masses would also be equal. From Table 13.2, however, we see that their heat capacities differ by a large factor. The heat capacity of lead (chemical symbol Pb) is $c_{Pb} = 129 \text{ J/kg°C}$ while the heat capacity of water (H_2O), is much larger, at $c_{H_2O} = 4186 \text{ J/kg°C}$. Refer to equation 13.8, $Q = m c \Delta T$. The heat Q has the same magnitude for both, because it is the energy transferred from the lead to the water. Therefore, the larger the specific heat capacity, the smaller the mass. In this case, then, the mass of water must be considerably smaller than the mass of the lead ball.

Step 2 – ***Use equation 13.8, and energy conservation, to find the mass of water in the system.*** In the static equilibrium problems we analyzed in Chapter 10 we used the equations $\sum \vec{F} = 0$ and $\sum \vec{\tau} = 0$. We can apply a similar equation in this thermal equilibrium situation:

$$\sum Q = 0 .$$ (Equation 13.9: **An equation for thermal equilibrium**)

This is really a simple statement of conservation of energy. In the case of the lead ball and the water, it means that all the heat transferred out of the lead ball (this Q is negative) is transferred to the water (that Q is positive). Combining equations 13.8 and 13.9 gives:

$$Q_{Pb} + Q_{H_2O} = 0 ;$$

$$m_{Pb} \, c_{Pb} \, \Delta T_{Pb} + m_{H_2O} \, c_{H_2O} \, \Delta T_{H_2O} = 0 .$$

Solving for the unknown mass of the water gives:

$$m_{H_2O} = \frac{-m_{Pb} \, c_{Pb} \, \Delta T_{Pb}}{c_{H_2O} \, \Delta T_{H_2O}} = \frac{-m_{Pb} \, c_{Pb} \left(T_{f,Pb} - T_{i,Pb}\right)}{c_{H_2O} \left(T_{f,H_2O} - T_{i,H_2O}\right)} ,$$

where the subscripts i and f on the temperatures correspond to the initial and final values, respectively. Plugging in the numerical values gives:

$$m_{H_2O} = \frac{-(500\,\text{g})(129\,\text{J/kg°C})(60°\text{C} - 100°\text{C})}{(4186\,\text{J/kg°C})(60°\text{C} - 20°\text{C})} = 15.4\,\text{g} .$$

As we predicted, the mass of water is much smaller than the mass of the lead ball. This is because of the unusually large specific heat capacity of liquid water.

Key ideas regarding thermal equilibrium: In a thermal equilibrium situation we can use a simple energy conservation relationship, $\sum Q = 0$. We also observed that liquid water has a large specific heat capacity. **Related End-of-Chapter Exercises: 22 – 25.**

Note that, in Exploration 13.3, there is no need to convert the temperatures from Celsius to Kelvin, because the equation for Q involves ΔT and not T. In addition, there is no need to convert the given mass of the lead ball from grams to kilograms, even though the specific heat capacities are specified in kg. In the end, the units in the two specific heats cancel out, so the mass of the water comes out in the units the mass of the lead ball is given in.

Essential Question 13.3: Many people learn to do thermal equilibrium problems by setting the heat lost by one or more parts of a system equal to the heat gained by other parts. Compare and contrast that method to the $\sum Q = 0$ used in Exploration 13.3.

Answer to Essential Question 13.3: First, the two methods are completely equivalent, and they should give the same answer. There is a difference, however, in how signs are handled for the ΔT 's. In the $\sum Q = 0$ method, each ΔT is specified as a final temperature minus an initial temperature, so a ΔT can be positive (if the temperature increases) or negative (if the temperature decreases). In the heat lost vs. heat gained method, all the ΔT 's are taken to be positive. In the case of the lead ball and water, the analysis would proceed like this:

$$Q_{gained} = Q_{lost} \qquad \text{so} \qquad m_{H_2O}\, c_{H_2O} \left|\Delta T_{H_2O}\right| = m_{Pb}\, c_{Pb} \left|\Delta T_{Pb}\right| .$$

Using the absolute values of the ΔT 's ensures that every term in the equation is positive.

13-4 Latent Heat

Let's expand our knowledge of thermal equilibrium problems by learning to handle phase changes. In general, a change of phase is associated with a relatively large amount of energy, and occurs at a particular temperature. To melt ice, for instance, first heat is added to raise the ice's temperature to 0°C, the melting point. Adding more heat breaks bonds and gradually transforms the solid water into liquid water. The transformation takes place at a constant temperature of 0°C.

Because there is no change in temperature associated with a phase change, the equation we use has a different form than the $Q = mc\Delta T$ equation we use for the heat required to change a substance's temperature. The amount of heat associated with a phase change is given by:

$Q = m L_f$, (Equation 13.10: **Heat for a liquid-solid phase change**)

where L_f is known as the **latent heat of fusion,** or,

$Q = m L_v$, (Equation 13.11: **Heat for a gas-liquid phase change**)

where L_v is known as the **latent heat of vaporization**.

As shown in Table 13.3, latent heat values depend on the material. To melt or vaporize a substance requires that heat is added, while heat must be removed from a substance to solidify or condense it. Table 13.4 shows melting points and latent heats of fusion of some common metals.

Material	**Melting point**	**Latent heat of fusion (kJ/kg)**	**Boiling point**	**Latent heat of vaporization (kJ/kg)**
Water	0°C	335	100°C	2272
Nitrogen	-210°C	25.7	-196°C	200
Oxygen	-219°C	13.9	-183°C	213
Ethyl alcohol	-114°C	108	-78.3°C	855

Table 13.3: A table of melting points, boiling points, and latent heats for various materials.

Material	**Melting point**	**Latent heat of fusion (kJ/kg)**
Aluminum	933.5K	396
Copper	1356.6K	209
Gold	1337.58K	64
Iron	1808K	250
Lead	600.65K	23

Table 13.4: A table of melting points and latent heats of fusion for various metals.

EXPLORATION 13.4 – Temperature vs. time

100 grams of ice at –20°C is put in a pot on a burner on the stove. The burner transfers energy to the water at a rate of 400 W. The ice melts, and eventually all the water boils away.

Step 1 – ***How many different heat terms do we need to keep track of in this process? Describe them in words.*** There are four heat terms we need to keep track of. These include:

1. The heat needed to increase the temperature of the ice to the melting point.
2. The heat needed to melt the ice at 0°C.
3. The heat needed to raise the liquid water to the boiling point.
4. The heat needed to boil the water at 100°C.

Step 2 – ***Plot a graph of the temperature of the water as a function of time, starting at t = 0 when the temperature is at –20°C.*** To draw the graph it's helpful to find the time each of the four steps takes in the process. Because energy is power multiplied by time, the times can be found by dividing the heat in each case by the power.

Raising the temperature of the ice takes $t_1 = \frac{mc\Delta T}{P} = \frac{(0.100\,\text{kg})(2060\,\text{J/kg°C})(+20°\text{C})}{400\,\text{J/s}} = 10.3\,\text{s}$.

Melting the ice takes $t_2 = \frac{mL_f}{P} = \frac{(0.100\,\text{kg})(3.33\times10^5\,\text{J/kg})}{400\,\text{J/s}} = 83.25\,\text{s}$.

Raising the water by 100°C takes $t_3 = \frac{mc\Delta T}{P} = \frac{(0.100\,\text{kg})(4186\,\text{J/kg°C})(+100°\text{C})}{400\,\text{J/s}} = 105\,\text{s}$.

Boiling the water takes $t_4 = \frac{mL_V}{P} = \frac{(0.100\,\text{kg})(2.256\times10^6\,\text{J/kg})}{400\,\text{J/s}} = 564\,\text{s}$.

The total time required to vaporize all the water is t = 762 s, which is equivalent to 12.7 minutes. The water spends most of the time changing phase, particularly while it is vaporizing. This tells us that it takes a lot of energy to change phase. We can also see the large fraction of time spent changing phase on the graph of the temperature as a function of time in Figure 13.4.

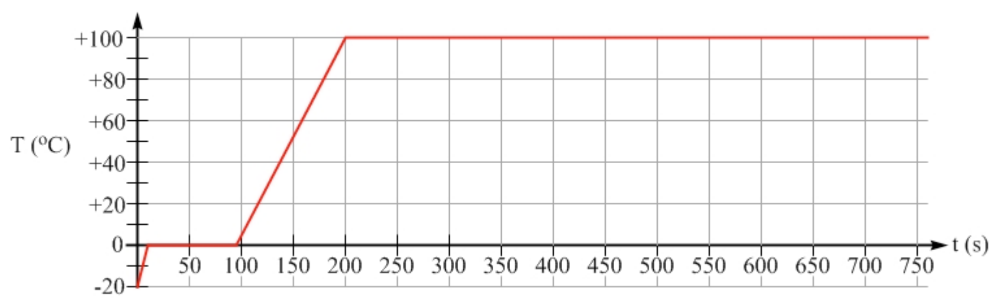

Figure 13.4: A plot of the temperature as a function of time (in seconds) for a sample of water that starts as a solid at –20°C and ends up boiling away completely. Note the large fraction of the time the water spends changing phase, in the two constant-temperature sections of the graph.

Key idea regarding phase changes: Changes of phase are generally associated with a relatively large amount of energy. **Related End-of-Chapter Exercises: 4, 31, 32.**

Essential Question 13.4: Return to the graph in Figure 13.4. Why is the slope of the graph constant when the sample's temperature is changing? What is the slope of the graph equal to?

Answer to Essential Question 13.4: Let's begin by solving for the slope. For a particular time interval Δt, we have $m c \Delta T = P \Delta t$. The slope is equal to the rise (ΔT) over the run (Δt):

$$\frac{\Delta T}{\Delta t} = \frac{P}{m c}.$$

Until the phase changes, the power, mass, and specific heat capacity, *c*, are constant, so the slope is constant. Also, the slope is inversely proportional to *c*. A lower specific heat capacity (which water has in the solid phase) corresponds to a larger slope, so the temperature rises faster. It does not take as much energy to change temperature when the specific heat is lower.

13-5 Solving Thermal Equilibrium Problems

Let's summarize the various steps involved in solving a thermal equilibrium problem.

A General Method for Solving a Thermal Equilibrium Problem

1. Write out in words a brief description of the various heats involved.
2. Use the equation $\sum Q = 0$ to write out an equation of the form $Q_1 + Q_2 + \ldots = 0$, where each *Q* corresponds to one of the brief descriptions from step 1.
3. Use $Q = m c \Delta T$ to write expressions for each *Q* associated with a change in temperature, ΔT. Express each ΔT as $T_{final} - T_{initial}$ and the sign of that term will be built into the ΔT.
4. Use the equation $Q = m L_f$ or $Q = m L_v$ to write an expression for the heat associated with a phase change. Use a plus sign if heat is added to produce the phase change (melting or vaporizing) and a negative sign if heat is removed (solidifying or condensing).
5. Solve the resulting equation for the unknown variable.

This method applies when no heat is transferred between a system and its surroundings. If a heat transfer is involved, we can generalize the method to account for that. In that case, $\sum Q = q$, where *q* is the heat added to (positive *q*) or removed from (negative *q*) the system.

Let's do an example in which we combine changes of temperature and changes of phase.

EXPLORATION 13.5 – Freezing the punch

You have 2.00 liters of fruit punch at 20.0°C that you are trying to cool to get ready for a party. Because fruit punch is mostly water let's assume that the relevant specific heat capacities and latent heats for water can also be used for the fruit punch, and that the freezing point is 0°C.

Step 1 - ***To cool the fruit punch quickly, you pour it into a bowl of ice that is initially –15.0°C. There is so much ice, however, that all the punch freezes and the whole mixture comes to a final temperature of –5.0°C. Find the mass of ice in the bowl originally, assuming no energy is transferred between the ice-fruit punch system and the bowl or the surrounding environment.***

This is a thermal equilibrium problem, so we apply Equation 13.9, $\sum Q = 0$, to solve it.

Let's first just write out in words the different *Q*'s we need to include in the analysis:

1. The heat associated with increasing the temperature of the ice to –5.0 °C.
2. The heat associated with decreasing the temperature of the fruit punch to the freezing point, 0 °C.
3. The heat associated with freezing the fruit punch.
4. The heat associated with decreasing the solid fruit punch from 0 °C to the final temperature of –5.0 °C.

Let's use m_i to represent the mass of the ice, and m_{fp} to represent the mass of the fruit punch. Use the information given in the problem to determine the mass of the fruit punch, assuming again that fruit punch is basically water. Water's density is such that 1 ml of water has a mass of 1 g. With 2.00 liters of punch we have 2000 ml, which is m_{fp} = 2000 g = 2.00 kg.

Setting up the $\sum Q = 0$ equation gives:

$$Q_1 + Q_2 + Q_3 + Q_4 = 0\,;$$

$$m_i c_{solid}\left(-5°\text{C} - -15°\text{C}\right) + m_{fp}c_{liquid}\left(0°\text{C} - 20°\text{C}\right) - m_{fp}L_f + m_{fp}c_{solid}\left(-5°\text{C} - 0°\text{C}\right) = 0\,.$$

In the last term, we use the specific heat capacity for solid water, because that term is the heat required to bring the solid punch from the freezing point to the final temperature of –5°C.

When using $\sum Q = 0$ the signs on the ΔT terms come naturally from the relationship $\Delta T = T_{final} - T_{initial}$. The sign on a heat associated with a phase change, however, needs to be put in explicitly. In this case the fruit punch is solidifying, so heat is removed from the punch. This is why the $m_{fp}L_f$ term has a negative sign. Solving the equation for the mass of the ice gives:

$$m_i = \frac{m_{fp}\left[c_{liquid}\left(20°\text{C}\right) + L_f + c_{solid}\left(5°\text{C}\right)\right]}{c_{solid}\left(10°\text{C}\right)}\,.$$

$$m_i = \frac{2.00\text{ kg}\left[(4186\text{ J / kg °C})(20°\text{C}) + (3.33\times10^5\text{ J/kg}) + (2060\text{ J / kg °C})(5°\text{C})\right]}{(2060\text{ J / kg °C})(10°\text{C})} = 41.5\text{ kg}.$$

The fact that the mass required is so large reflects the fact that it takes a great deal of energy to freeze the punch.

Step 2 - ***Because nobody can drink the solid fruit punch, you start again with a new 2.00-liter batch of punch at 20.0°C. Calculate the mass of ice, at –15.0°C, you should add to the punch so that the final temperature is 0°C and all the ice melts.***

This time we need three terms in the heat equation. They are:

1. The heat to increase the temperature of the ice to the melting point, 0°C.
2. The heat needed to melt all the ice at 0°C.
3. The heat associated with decreasing the temperature of the fruit punch to 0°C.

Write out these terms in the $\sum Q = 0$ analysis to solve for m_i', the new mass of the ice:

$$Q_1 + Q_2 + Q_3 = 0\,;$$

$$m_i' c_{solid}\left(0°\text{C} - -15°\text{C}\right) + m_i' L_f + m_{fp}c_{liquid}\left(0°\text{C} - 20°\text{C}\right) = 0\,.$$

Note that the term involving the phase change is positive, because heat is added to the ice to cause it to melt. Solving for the mass of ice this time gives:

$$m_i' = \frac{m_{fp}c_{liquid}\left(20°\text{C}\right)}{c_{solid}\left(15°\text{C}\right) + L_f} = \frac{\left(2.00\text{kg}\right)\left(4186\,\text{J/kg°C}\right)\left(20°\text{C}\right)}{\left(2060\,\text{J/kg°C}\right)\left(20°\text{C}\right) + 3.33\times10^5\,\text{J / kg}} = 0.447\text{ kg}\,.$$

This is a more reasonable amount of ice than the more than 20 kg we found in step 1!

Key idea for solving thermal equilibrium problems: The $\sum Q = 0$ method can be used in cases involving changes of phase. **Related End-of-Chapter Exercises**: Exercises 43 – 45.

Essential Question 13.5: Return to step 2 of Exploration 13.5. Is 0.447 kg of ice the only mass of ice that gives a final temperature of 0°C? Explain your answer qualitatively.

Answer to Essential Question 13.5: A whole range of masses of ice can produce a mixture with a final temperature of 0°C. The mixture could be all liquid, as in Exploration 13.5, but by increasing the mass of the ice we could have some ice left, all the ice left, some liquid solidifying, or even all the liquid solidifying. We get such a range because the final temperature is 0°C, the temperature of the phase change. This issue is explored further in end-of-chapter exercise 45.

13-6 Energy-Transfer Mechanisms

Let's investigate in more detail the mechanisms by which energy is transferred when there is a temperature difference. In Exploration 13.5, we did not concern ourselves with exactly how energy is transferred from the burner through the pot to the water. That process is known as thermal conduction. Two other important energy-transfer processes are convection and radiation. Let's first discuss those two processes qualitatively, and then discuss conduction in more detail.

Convection

Energy transfer in fluids generally takes place via convection, in which flowing fluid carries energy from one place to another. Convection currents are produced by temperature differences. Hotter (less dense) parts of the fluid rise, while cooler (more dense) areas sink. Birds and gliders make use of upward convection currents to rise, and we also rely on convection to remove ground-level pollution. Forced convection, where the fluid is forced to flow, is often used for heating (e.g., forced-air furnaces) or cooling (e.g., fans, automobile cooling systems).

Thermal Radiation

Thermal radiation involves energy transferred via electromagnetic waves. Often this is infrared radiation (which we can detect with the backs of our hands), but it can also be visible light or radiation of higher energy.

All objects continually absorb energy and radiate it away again. When everything is at the same temperature, the amount of energy received is equal to the amount given off and no changes in temperature occur. If an object emits more than it absorbs, though, it tends to cool down unless some other process replenishes its energy. For an object with a temperature *T* (in Kelvin) and a surface area *A*, the net rate of radiated energy P_{net} depends strongly on temperature:

$$P_{net} = P_{radiated} - P_{absorbed} = A\varepsilon\sigma\left(T^4 - T_{env}^4\right),$$ (Equation 13.12: **Net radiated power**)

where T_{env} is the temperature of the surrounding environment; *A* is the object's surface area; the Stefan-Boltzmann constant has a value $\sigma = 5.67\times10^{-8}\ \mathrm{W/m^2}$; and ε is the emissivity, which depends on the object. If an object readily absorbs or emits radiation, its emissivity is close to 1 (if it is exactly 1 the object is a **perfect blackbody**); if an object is highly reflective and does not absorb (or emit) radiation readily, its emissivity is close to 0. The best absorbers are also the best emitters. Black objects heat up faster than shiny ones, but they cool down faster too.

Thermal Conduction

Thermal conduction involves energy being transferred from a hot region to a cooler region through a material, without a net flow of the material. At the hotter end, the atoms, molecules, and electrons vibrate with more energy than they do at the cooler end. The energy flows through the material, passed along by these vibrations. One thing the rate at which energy is conducted through a uniform slab of material depends on is the temperature difference between the two faces of the slab. The rate of energy transfer is directly proportional to the temperature difference. Let's examine other contributing factors.

EXPLORATION 13.6 – The house beside the ice factory

You're planning to build a house, so you buy a small plot of land that is right next to a factory that makes ice. To get maximum use out of your land, one wall of your house will be built against the wall of the ice factory, where the ice is stored at –20°C. This might help keep your house cool in the summer, but you're more worried about staying warm in the winter.

Step 1 - ***You plan to build a rectangular house. To minimize the rate at which energy is transferred from inside your house, at +20°C, to the ice factory in the winter, should you place a large-area wall or a small-area wall of your house against the wall of the ice factory? Why?*** Imagine drawing a grid of equal-area squares on the wall. The more squares you have (that is, the larger the area of the wall), the more heat is transferred. To minimize your heat transfer, you need to minimize the wall area, because the rate of heat transfer is proportional to the area, A.

Step 2 – ***You put a layer of insulating material in the walls of your house. To be most effective, should the material be thin or thick?*** The thicker the insulation, the lower the rate of heat transfer. The rate of heat transfer is inversely proportional to L, the thickness of the insulation.

Step 3 – ***While shopping at Insulation World you notice that they have slabs of copper on sale for 10% less than the slabs of Styrofoam insulation you intended to buy. Check the thermal conductivity of these materials in Table 13.5 below. Should you save 10% and buy the copper?*** Absolutely not! The larger the thermal conductivity, the more effectively the material transfers heat. This is why copper is used in the base of pots and pans, efficiently transferring heat through the pot to the food. In insulating your house, however, you want to minimize the rate of heat transfer, so choose the material with the lowest thermal conductivity you can find.

Key idea: The rate at which heat is transferred (i.e., the power, P) through a slab of area A, thermal conductivity k, and thickness L is $P = \frac{k\,A}{L}(T_H - T_L)$. (Equation 13.13)

T_H is the temperature at the higher-temperature face of the slab, while T_L is the temperature at the lower-temperature face. **Related End-of-Chapter Exercises: 61, 62.**

R-values of insulation

Insulating materials are rated in terms of their resistance to conduction, or their R-value, also known as their thermal resistance. The larger the R-value, the better the insulating properties. R-values also make some calculations easier, because the total R-value of two materials placed back to back is the sum of their individual R-values. The R-value for a layer of material is found by dividing the material's thickness, L, by its thermal conductivity, k:

$$R = \frac{L}{k}. \quad \text{(Eq. 13.14: \textbf{R-values for insulation})}$$

Material	Thermal Conductivity, k [W/(m K)]
Air	0.0262
Aluminum	237
Brass	120
Copper	401
Diamond	895-2300
Ice	2.2
Glass	1.35
Styrofoam	0.033

Table 13.5: Thermal conductivities for various materials. These values are valid at 25°C, except for the value for ice which applies to ice at 0°C.

Essential Question 13.6: In the United States, R-values for insulation are specified in units of °F h ft^2 / BTU. BTU stands for British Thermal Unit, where 1 BTU = 1055 J. You have some insulation rated at R-15, so its R-value is 15 °F h ft^2 / BTU. What is this in units of K m^2 / W?

Answer to Essential Question 13.6: The conversion factors we need here include 1 BTU = 1055 J; 1 h = 3600 s; 1 m = 3.281 ft; and a change of 1 K is equivalent to a change of 1.8°F. Putting all this together gives us an overall conversion factor of:

$$1\frac{°\text{F h ft}^2}{\text{BTU}} \times \frac{1\text{ K}}{1.8°\text{F}} \times \frac{3600\text{ s}}{1\text{ h}} \times \left(\frac{1\text{ m}}{3.281\text{ ft}}\right)^2 \times \frac{1\text{ BTU}}{1055\text{ J}} = 0.1761\frac{\text{K m}^2}{\text{W}}.$$

Converting 15 °F h ft^2 / BTU requires multiplying the 15 by the conversion factor above, giving an R-value of 2.64 K m^2 / W.

Chapter Summary

Essential Idea: Understanding thermal expansion and energy transfer

In this chapter, we applied simple models to help us understand thermal expansion (the change in size an object experiences when its temperature changes) and energy transfer, especially the process of thermal conduction (energy transferred through a solid object because of a temperature difference).

Temperature scales

Temperature can be measured on various scales. Among the more common temperature scales are the Fahrenheit scale, used in daily life in the United States; the Celsius scale used in daily life in the rest of the world; and the Kelvin scale, which is used in scientific work.

$$T_C = \left(\frac{5°\text{C}}{9°\text{F}}\right)(T_F - 32°\text{F}).$$ (Equation 13.1: **Converting from Fahrenheit to Celsius**)

$$T_F = 32°\text{F} + \left(\frac{9°\text{F}}{5°\text{C}}\right)T_C .$$ (Equation 13.2: **Converting from Celsius to Fahrenheit**)

$$T_C = T_K - 273.16° .$$ (Equation 13.3: **Converting from Kelvin to Celsius**)

Thermal Expansion

When an object's temperature changes, its dimensions also generally change, with most materials expanding as the temperature increases. The model we applied to understand this assumes that each dimension of the object experiences a change in length that depends on the change in temperature, the initial length L_i, and the material the object is made from.

$$\Delta L = L_i \,\alpha\, \Delta T \quad ,$$ (Equation 13.4: **Length change from thermal expansion**)

where α is the thermal expansion coefficient, which depends on the material.

Equations giving the new length, area, or volume are:

$$L = L_i + L_i\alpha\,\Delta T = L_i\left(1 + \alpha\,\Delta T\right).$$ (Equation 13.5: **Length thermal expansion**)

$$A = A_0\left(1 + 2\alpha\,\Delta T\right).$$ (Equation 13.6: **Area thermal expansion**)

$V = V_0\left(1 + 3\alpha\,\Delta T\right)$. (Equation 13.7: **Volume thermal expansion**)

A General Method for Solving a Thermal Equilibrium Problem

1. Write out in words a brief description of the various heats involved. This helps to keep track of all the terms.
2. Use the equation $\sum Q = 0$ to write out an equation of the form $Q_1 + Q_2 + \ldots = 0$, where each Q corresponds to one of the brief descriptions you wrote in step 1.
3. Use the equation $Q = m\,c\,\Delta T$ to write expressions for each Q associated with a change in temperature. Express each ΔT as $T_{final} - T_{initial}$ and the sign of that heat term will be built into that temperature change.
4. Use the equation $Q = m\,L_f$ or $Q = m\,L_v$ to write an expression for the heat associated with a change of phase. Use a plus sign if heat must be added to produce the phase change (melting or vaporizing) and a negative sign if heat must be removed (solidifying or condensing).
5. Solve the resulting equation for the unknown.

The method above applies when no heat is transferred between a system and its surroundings. If heat is transferred into or out of a system, we can say that $\sum Q = q$, where q represents heat added to (q is positive) or removed from (q is negative) the system.

Energy-Transfer Mechanisms

Three mechanisms of energy transfer, driven by temperature differences, include:

1. Convection - energy is carried by a flowing fluid.

2. Thermal radiation – energy is given off in the form of electromagnetic radiation. Power radiated by an object is strongly dependent on how its temperature compares to the temperature of its surroundings.

$P_{net} = P_{radiated} - P_{absorbed} = A\varepsilon\sigma\left(T^4 - T_{env}^4\right)$, (Equation 13.12: **Net radiated power**)

Where T_{env} is the temperature of the surrounding environment; A is the surface area of the object; the Stefan-Boltzmann constant σ has a value $\sigma = 5.67\times10^{-8}\ \mathrm{W/m^2}$; and ε is known as the emissivity, which depends on the object. If an object readily absorbs or emits radiation then its emissivity is close to 1.

3. Thermal conduction – energy is transferred through a material by the vibrations of atoms and molecules. The rate at which energy is transferred (i.e., the power, P) through a slab of area A, thermal conductivity k, and thickness L is:

$P = \dfrac{k\,A}{L}\left(T_H - T_L\right)$. (Equation 13.13: **Thermal conduction**)

T_H is the temperature at the higher-temperature face of the slab, while T_L is the temperature at the lower-temperature face.

End-of-Chapter Exercises

Exercises 1 – 12 are conceptual questions that are designed to see if you have understood the main concepts of the chapter.

1. An American named Bill is visiting London, England. Listening to the weather forecast, Bill hears that the temperature will be 25°. Back home in New York, this would be a relatively cold day, so Bill puts on warm clothing. When Bill goes out of his hotel on his way to Trafalgar Square, however, he realizes that the temperature is much warmer than he thought. (a) What happened? (b) On the temperature scale that Bill is used to, what is the temperature?

2. A density ball is a ball that is weighted so its density is very similar to that of water. The ball floats in cold water but sinks in warm water. Explain why.

3. A solid iron disk is rotating without friction about an axle through its center. The Sun then comes out from behind a cloud and increases the temperature of the disk. You notice that the disk's angular velocity changes a little. What is responsible for the change, and does the disk speed up or slow down?

4. The graph in Figure 13.5 shows the temperature of a sample of an unknown material as a function of time. Heat is either being transferred into or out of this sample at a steady rate. (a) Is heat being transferred into or out of the sample? How do you know? (b) Does it look like the material undergoes a phase change? How can you tell? (c) If, over the period shown in the graph, the sample is in the solid phase for part of the time and is in the liquid phase for another part of the time, how do its two specific heat capacities compare? If possible, find the ratio of the specific heat capacity in the solid phase to the specific heat capacity in the liquid phase.

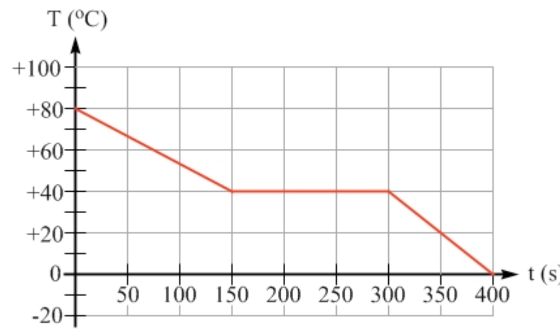

Figure 13.5: A graph of temperature as a function of time for a sample to which heat is being transferred at a constant rate, for Exercise 4.

5. A bimetallic strip is actually made from strips of two different metals that are bonded together back-to-back, as shown in Figure 13.6. The two strips have the same length at room temperature, 20°C. If the temperature changes, the two strips expand or contract different amounts, causing the bimetallic strip to bend into a circular arc, as shown in Figure 13.6, with the longer strip on the outside of the arc. (a) If the bimetallic strip in Figure 13.6 is composed of a strip of brass bonded to a strip of iron, which side is which? Why? (b) Bimetallic strips are often used as switches in thermostats, turning a furnace on if the temperature falls below a certain pre-set minimum value and turning the furnace off again when the temperature has risen sufficiently. Briefly explain how this process works.

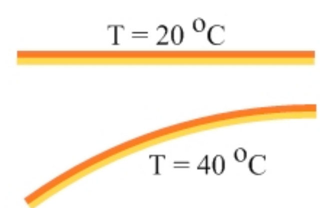

Figure 13.6: A bimetallic strip, for Exercise 5.

6. You have three blocks of equal mass. Block A is made of aluminum; block B is made of gold; and block C is made of copper. Each block starts at room temperature, 20°C. 100 J of energy is then added to each of the blocks. Rank the blocks based on their final temperatures, from largest to smallest.

7. You have three blocks. Block A is made of aluminum, block B is made of gold, and block C is made of copper. The blocks start at the same temperature. When 50 J of heat is added to each block, the blocks also have the same final temperature. Rank the blocks based on their masses, from largest to smallest.

8. You have three blocks of equal mass. Block A is made of aluminum; block B is made of gold; and block C is made of copper. Each block is initially at 80°C. You also have three identical Styrofoam cups, each containing the same amount of water at 10°C. You add one block to each of the cups and measure the final temperature. Assuming no heat is transferred to the cup or the surroundings, rank the final temperatures from highest to lowest.

9. You have three blocks. Block A is made of aluminum; block B is made of gold; and block C is made of copper. Each block is initially at 80°C. You also have three identical Styrofoam cups, each containing the same amount of water at 10°C. You add one block to each of the cups and measure the final temperature, which is exactly the same in each case. Assuming no heat is transferred to the cup or the surroundings, rank the blocks from largest to smallest based on their masses.

10. You decide to have spaghetti for dinner, so you fill a large pot with water and place it on the stove. While you are waiting for the water to boil, you decide to do an experiment, so you place a couple of drops of food coloring in the water to observe what the water is doing. You observe that the water is moving, with some water rising from the bottom of the pot to the top, moving sideways at the top, and then gradually falling back down to the bottom of the pot again. Which of the three heat transfer mechanisms are you observing here? Explain how it works.

11. You take two identical shiny metal cans and paint one black. You then place them out in the sun on a hot summer's day, and measure their temperatures as a function of time. (a) If the cans are initially at 20°C, which can reaches 30°C first? Why? (b) Later on, you take the cans inside and fill them both with hot water so both cans are initially at 95°C. Again you measure their temperatures as a function of time. Which can reaches 85°C first? Why?

12. Thin films of diamond are used on computer chips to ensure that the chips do not get too hot. Why is diamond such a good material for this application?

Exercises 13 – 16 deal with temperature scales and conversions between temperature scales.

13. What is the conversion equation to transform from (a) Fahrenheit to Kelvin? (b) Kelvin to Fahrenheit?

14. (a) What is the Rankine temperature scale? What is the equation for converting from the Rankine scale to (b) the Fahrenheit scale? (c) the Kelvin scale?

15. While visiting Toronto, Canada, James buys a cake mix at the grocery store. When James gets home to Los Angeles he tries baking the cake. He carefully follows the instructions on the package, including baking the cake at 250° for 45 minutes, but the cake is a disaster. (a) Explain what happened. (b) Did James burn the cake or was it underdone? (c) To what temperature should James have set his oven?

16. Liquid nitrogen boils at a temperature of 77K. What is this in (a) Celsius? (b) Fahrenheit?

Questions 17 – 21 deal with thermal expansion.

17. An iron ring has an inner radius of 2.5000 cm and an outer radius of 3.5000 cm, giving the ring a thickness of 1.0000 cm. If the temperature of the ring is increased from 20°C to 80°C, what is the thickness of the ring?

18. Do you think it is important for bridge designers to worry about thermal expansion when they design bridges? Why? Carry out the following two calculations to support your answer. At 10°C, a particular steel bridge has a length of precisely 500 m. Find the length of the bridge on a hot summer day when the temperature reaches 40°C. Then find the length in the middle of winter when the temperature drops to –30°C.

19. Liquid water has a linear thermal expansion coefficient of 70 x 10^{-6} /°C. You absent-mindedly fill an aluminum pot to the brim with water. Both the pot and the water are at 20°C. You then want to bring the water to a boil, so you put the pot on the stove to heat the water. (a) Assuming the temperature of the water and the pot are always equal, what happens as the temperature increases? Does the water spill out of the pot or does the water level fall relative to the top of the pot? Explain. Neglect any loss of water because of evaporation. (b) When the temperature reaches 80°C, determine either what fraction of the original water has overflowed the pot, or what fraction of the volume inside the pot is no longer occupied by water.

20. In section 13.2, we derived the following expression for the area A resulting from imposing a temperature change ΔT on an object of original area A_0 : $A = A_0\left[1+2\alpha\,\Delta T+(\alpha\,\Delta T)^2\right]$. We then argued that the $(\alpha\,\Delta T)^2$ inside the bracket is negligible in comparison to the $2\alpha\,\Delta T$ term. Let's check that to see the effect of neglecting the last term. Start with an aluminum cylinder with a cross-sectional area of 5.000000 cm^2. Compute the new cross-sectional area when the temperature is increased by 100°C by (a) using the complete area equation above; (b) using the approximation $A = A_0\left[1+2\alpha\,\Delta T\right]$. (c) What is the percentage difference between your two answers? (d) Based on this, do you think it is reasonable to neglect the $(\alpha\,\Delta T)^2$ term?

21. A solid iron disk is rotating without friction about an axle through its center. The Sun then comes out from behind a cloud and increases the temperature of the disk. You notice that the disk's angular velocity changes a little. When the temperature is 20°C, the disk's radius is 15.00 cm, and the angular speed of the disk is 20.00 rad/s. What is the disk's angular speed when its temperature is 60°C?

Exercises 22 – 27 are designed to give you practice applying the general method for solving a thermal equilibrium problem. For each of these exercises, begin with the following steps:

(a) Write out in words a brief description of the various heats involved.
(b) Apply $\sum Q = 0$ to obtain an equation of the form $Q_1 + Q_2 + \ldots = 0$. Each Q corresponds to one of the brief descriptions you wrote in part (a).
(c) For any temperature changes, apply the equation $Q = mc\,\Delta T$. Express each ΔT as $T_{final} - T_{initial}$. For any changes of phase, apply the equation $Q = m\,L_f$ or $Q = m\,L_v$. Use a positive sign if heat must be added to produce the phase change, and a negative sign if heat must be removed.

22. An aluminum block with a mass of 300 g and a temperature of 80°C is placed in a Styrofoam cup that contains 500 g of water at 10°C. Ignore any temperature change associated with the Styrofoam cup. Start by doing parts (a) – (c) as described above. (d) Find the equilibrium temperature.

23. Repeat Exercise 22, but now the water is in an aluminum container that has a mass of 200 g and is initially at the temperature of the water. The block, container, and water all come to the same final temperature.

24. A copper block with a mass of 500 g is cooled to 77K by being immersed in liquid nitrogen. The block is then placed in a Styrofoam cup containing some water that is initially at +50°C. Assume no heat is transferred to the cup or the surroundings. The goal of the exercise is to determine the mass of water in the cup, if the final temperature is +20°C. Start by doing parts (a) – (c) as described above. (d) Find the mass of the water.

25. Repeat Exercise 24, but this time the final temperature is –20°C.

26. Repeat Exercise 24, but this time the final temperature is 0°C. Start by doing parts (a) – (c) as described above. (d) Find the maximum mass of water in the cup. (e) Find the minimum mass of water in the cup.

27. Repeat Exercise 24, but this time the water is in an aluminum cup that has a mass of 300 g. Assume the temperature of the cup is equal to the temperature of the water at all times.

Exercises 28 – 32 are designed to give you practice solving problems in which heat (*q*) is added or removed from a system. Applying $\sum Q = q$ should help you answer these exercises.

28. A Styrofoam cup contains 200 g of water at 20°C. The cup is then placed in the freezer. The freezer can remove heat from the water at a steady rate of 50 W. (a) If we neglect any heat transfer involving the Styrofoam cup, how long does it take until the cup contains ice at –5°C? (b) Plot a graph of the temperature of the water as a function of time as the water cools from 20°C to –5°C.

29. Return to the situation described in Exercise 28, except now the water is placed in a 300 g aluminum container. The temperature of the container matches the water at all times and, because aluminum has a larger thermal conductivity than Styrofoam, the freezer removes heat from the aluminum-water system at a rate of 90 W. How long does it take the temperature of the system to drop from 20°C to –5°C now?

30. 500 g of water at 20°C is in a pot on the stove. An unknown mass of ice that is originally at –10°C is placed in an identical pot on the stove. Heat is then added to the two samples of water at precisely the same constant rate. You observe that both samples of water reach 80°C at the same time. (a) How does the mass of the ice in the second pot compare to the mass of the water in the first pot? (b) Which system reaches 90°C first? (c) Solve for the mass of the ice.

31. You have a 100 g block of lead that you intend to melt and then pour into a mold to form a bell. The lead is initially at room temperature, 20°C. You then add heat to the lead at a steady rate of 200 W. (a) How long does it take for the lead to reach its melting point? (b) How much additional time is required to completely melt the lead block? (c) Graph the temperature of the sample versus time, ending the graph when the lead is completely melted.

32. The graph in Figure 13.7 shows the temperature of a sample of an unknown material as a function of time. Heat is either being transferred into or out of this sample at a steady rate. Over the period shown in the graph, the sample is in the solid phase for part of the time and is in the liquid phase for another part of the time. (a) Is heat being transferred into or out of this sample? (b) What is the material's melting point? (c) What is the ratio of the material's specific heat when liquid to the specific heat when solid?

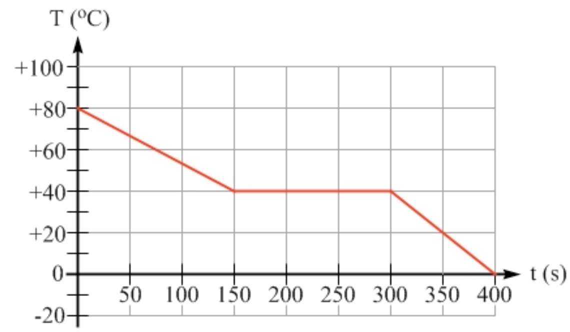

Figure 13.7: A graph of temperature as a function of time for a sample to which heat is being transferred at a constant rate, for Exercise 32.

Exercises 33 – 35 involve the energy-transfer mechanisms of convection, conduction, or radiation.

33. To keep yourself warm in the winter, you heat a solid metal ball and place it on a stand in the center of your room. The ball has a radius of 10.0 cm and an emissivity of 0.82. If your room has a temperature of 15°C, what is the net power radiated by the ball initially, when the ball's temperature is 200°C?

34. You have two rods that have the same dimensions but which are made from different materials. One rod is made of brass while the other is made of copper, and each rod is 1.00 m long. The rods are joined end-to-end, as shown in Figure 13.8. The far end of the brass rod is maintained at a temperature of 0°C, while the far end of the copper rod is maintained at a temperature of 90°C. The system is allowed to come to equilibrium (defined as each point on the rods reaching a constant temperature). What is the temperature at the point where the rods meet? Neglect thermal expansion.

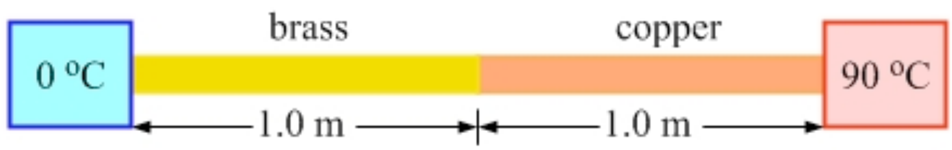

Figure 13.8: A brass rod and a copper rod are joined together end-to-end. The other end of the brass rod is kept at 0°C, while the far end of the copper rod is maintained at 90°C. For Exercises 34 and 35.

35. Return to the situation described in Exercise 34. Plot a graph of the temperature as a function of position along the rods, taking the cooler end of the brass rod to be the origin and the hotter end of the copper rod to be *x* = +2.00 m.

General Exercises and Conceptual Questions

36. The Fahrenheit and Celsius temperature scales are named after particular individuals. Do a little research about these people and write a paragraph or two describing each one.

Figure 13.9: A photograph of a Galileo thermometer, for Exercise 37. Photo credit: Johanna Goodyear / iStockPhoto.

37. A photograph of a Galileo thermometer is shown in Figure 13.9. Explain how such a thermometer works. What property of the liquid inside the thermometer is being exploited in this thermometer?

38. The "Mpemba effect" is the name for an interesting phenomenon, namely that in some circumstances warmer water can end up freezing before cooler water. Do some research on the Mpemba effect and write two or three paragraphs describing how this might be possible.

39. To make a cup of tea, you put 1000 g of water at 15°C into a kettle and bring the water to a boil. However, you only need 300 g of hot water to make your cup of tea. How much energy did you waste bringing the extra water to the boiling point?

40. In the second paragraph of Section 13.3, we ask the following question: what is the equilibrium temperature when a 500-gram lead ball at 100°C is added to 400 g of water that is at room temperature, 20°C? What is the answer?

41. A 500-gram lead ball that is initially at 100°C is added to 500 g of water that is initially at room temperature, 20°C, in a Styrofoam cup. The system is allowed to come to thermal equilibrium. Note that 60°C is exactly halfway between the initial temperatures of the lead ball and the water. (a) Come up with a qualitative argument for whether the final temperature is more than 60°C, less than 60°C, or equal to 60°C. (b) Find the final temperature.

42. A lead bullet with a mass of 25 g is fired into a target. The bullet completely melts upon impact. Assuming all the kinetic energy of the bullet goes into raising the bullet's temperature and then melting it, what is the speed of the bullet when it strikes the target?

43. You have 2.00 liters of fruit punch at 20.0°C that you are trying to cool to get ready for a party. Assume that the relevant specific heat capacities and latent heats for water can also be used for the fruit punch, and that the freezing point is 0°C. To cool the fruit punch quickly, you pour it over a large bowl of ice that is initially at –15.0°C. The mixture comes to a final temperature of +5.0°C. Find the mass of ice that was originally in the bowl, assuming no energy is transferred between the ice-fruit punch system and the bowl or the surrounding environment.

44. Repeat Exercise 43, but now account for the bowl. Assume the bowl is made from 300 g of aluminum and that the bowl is also initially at –15.0°C.

45. In part (b) of Exploration 13.3, we found that when 0.447 kg of ice at –15.0°C is added to 2.00 liters of fruit punch at 20.0°C and allowed to come to thermal equilibrium, the result is that all the ice melts and the final temperature of the mixture is 0°C. There is a whole range of masses of ice, however, that could have been added to the punch to achieve that same final temperature. (a) Does 0.447 kg represent the minimum or the maximum amount of ice at –15.0°C we can add to the punch to produce a final temperature of 0°C? Explain. (b) Determine the other end of the range, the other extreme in the amount of ice we can add to the punch and yet still achieve a final temperature of 0°C.

46. A block of ice has an unknown initial temperature. Heat is transferred to the ice, first bringing the ice to 0°C, then melting it, and then bringing the resulting water to 50°C. The total heat required for the two changes in temperature is equal to the heat associated with the melting. What was the block's initial temperature?

47. A copper block, with a mass of 1500 g, is cooled to 77K by being immersed in liquid nitrogen. The block is then transferred to a Styrofoam cup containing 1.20 liters of water at 50°C. Assuming no energy is transferred to the cup, determine the final temperature of the system.

48. Repeat Exercise 47, but this time the block is aluminum instead of copper.

49. Repeat Exercise 47, but this time assume the water is in an aluminum container that has a mass of 400 g, and that the temperature of the container is equal to the temperature of the water at all times.

50. You have three blocks of equal mass. Block A is made of aluminum; block B is made of gold; and block C is made of copper. Each block is initially at 80°C. The blocks are added, one at a time, to a Styrofoam cup containing 500 g of water at 10°C. The final temperature is 40°C. Assuming no heat is transferred to the cup or the environment, what is the mass of each block?

51. You have three balls of equal mass. Ball A is made of aluminum; ball B is made of gold; and ball C is made of copper. Each ball is initially at –50°C. You also have three identical Styrofoam cups, each containing equal amounts of water at 10°C. You add one ball to each of the cups and measure the final temperature. Assuming no heat is transferred to the cup or the surroundings, is there enough information provided to rank the final temperatures from highest to lowest? If so, provide the ranking. If not, explain why not.

52. Return to the situation described in Exercise 51. Is it possible for the final temperature in one of the cups to be below 0°C, the final temperature in another to be 0°C, and the final temperature in the remaining cup to be above 0°C? If so, come up with an example specifying the mass of the balls and the mass of the water in the cup. If not, explain why not.

53. James Prescott Joule carried out an experiment known as the mechanical equivalent of heat. Write a few paragraphs about Joule and the experiment.

54. The water at Niagara Falls drops through a height of 52 m. (a) If the water's loss of gravitational potential energy shows up as an increase in temperature of the water, what is the temperature difference between the water at the top of the falls and the water at the bottom?

Figure 13.10: A photograph of Niagara Falls, for Exercise 54. Photo credit: Robert Glusic/ PhotoDisc/Getty Images.

55. As part of an experiment, you fill a cardboard tube that has a length of 1.2 m with 200 g of lead shot (small lead balls) and seal the ends of the tube. Aligning the axis of the tube vertically, you then invert the tube 100 times. Predict what you observe for the temperature difference of the lead balls at the end of the experiment compared to what it was at the beginning.

56. Return to Exercise 55. Doing the experiment with lead balls can be something of a health risk, because you can breathe in lead dust if you open the tube to measure the temperature. You try the experiment with small copper balls instead of lead. (a) In which case would you observe a larger temperature change, when you used the copper balls or when you used the lead balls? Explain your answer. (b) If you take the ratio of the two temperature changes, what would you expect to find?

57. The Sun has a radius of 6.96×10^5 km, and an average temperature at its surface of 5780 K. (a) Calculate the power radiated by the Sun. (b) The distance from the Sun to the Earth is about 150 million km. Estimate the power from the incident sunlight on a 1.0 m^2 solar cell that is part of an array being used to provide energy for a satellite in orbit around the Earth.

58. The base of a copper-bottomed pot has a radius of 15 cm and a thickness of 3.0 mm. Its bottom surface is maintained at a temperature of 250°C by being placed on a hot burner on the stove. (a) If the pot contains 2.00 liters of water that is initially at 20°C determine the initial rate at which energy is transferred through the base of the pot to the water. (b) As the water temperature increases, does the rate at which energy is transferred through the base of the pot increase, decrease, or stay the same? Explain. (c) Calculate the rate of energy transfer when the water temperature is 95°C.

59. Return to the situation described in Exercise 58. Let's say it takes a time T to raise the water temperature to the boiling point. The process is repeated with a second pot in which everything is the same except for the fact that the base of the second pot is aluminum instead of copper. (a) Is the time it takes to bring the water to the boiling point in the second pot greater than, less than, or the same as T? Justify your answer. (b) How much time, in terms of T, does the process take in the second pot?

60. On a chilly November day, you go out for a hike to the top of a local mountain with your friend. The two of you dress in layers, but for your inner-most layer you are wearing a high-tech fabric that wicks moisture away from your skin while your friend is wearing a long-sleeved cotton shirt that is damp with perspiration by the time you reach the top of the mountain. Coming back down, you are quite comfortable, while your friend is feeling colder with each passing minute. Fortunately, you reach the lodge at the base of the mountain before your friend becomes hypothermic, and your friend is able to warm up again in front of a roaring fire. Explain what happened, given that the fabric you were wearing has a thermal conductivity of 0.06 W/(m K), dry cotton has a thermal conductivity of 0.04 W/(m K), and water has a thermal conductivity of 0.6 W/(m K).

61. The outer walls of your house have an R-value of 5.0 K m^2 / W, and a total area of 2000 m^2. Let's assume that, from the beginning of December to the end of February, the temperature outside the house is 0°C while you set your thermostat to maintain a constant temperature of 22°C inside the house. (a) How much energy is conducted through the walls of your house in this three-month period? Assume it is not a leap year. (b) How much energy would be conducted through the walls if you lowered the thermostat so as to maintain a constant temperature of 20°? (c) Every kW-h of energy costs about 20 cents. First, convert a kW-h to joules, and then determine how much money you would save by keeping your thermostat at 20°C for the three months.

62. You have four square aluminum sheets, each with an area of 0.25 m^2 and a thickness of 5.0 mm, and four copper sheets, having exactly the same dimensions as the aluminum sheets. You plan to create a square piece of insulation, with an area of 1.00 m^2, by placing four of your sheets together in one layer, and layering the remaining four sheets on top of the first four. (a) To minimize the rate of energy transfer through your arrangement, should you place the four sheets of copper over the four sheets of aluminum, or should you stack sheets of aluminum together and stack the copper together? (b) Assuming a temperature difference of 20°C between the two faces of the arrangement, support your answer to part (a) by calculating the rate of energy transfer in the two cases.

63. What thickness of aluminum has the same R-value as 5.0 cm of Styrofoam?

64. You overhear two of your classmates discussing Essential Question 13.1. Comment on each of their statements.

Liam: Did you notice that we never got the answer to Essential Question 13.1? Does the space inside the glass thermometer increase or decrease when the temperature goes up? Well, the glass expands, so it must fill in some of that space – that would cause the alcohol level to go up even more than it would if the glass did not change size.

Sherry: Except, in Section 13.2 we looked at how holes expand when they're heated. Can't we apply that to the cavity inside the glass? That would mean the cavity volume increases when the temperature goes up, so the alcohol level is less than if the glass doesn't change size. Except, how do we know the level goes up at all when the temperature increases? Couldn't the level even go down, or stay the same? That doesn't sound like a very good thermometer!

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