1.3.5: Energy and Heat Capacity Calculations

Learning Objectives

- To relate heat transfer to temperature change.

Heat is a familiar manifestation of transferring energy. When we touch a hot object, energy flows from the hot object into our fingers, and we perceive that incoming energy as the object being "hot." Conversely, when we hold an ice cube in our palms, energy flows from our hand into the ice cube, and we perceive that loss of energy as "cold." In both cases, the temperature of the object is different from the temperature of our hand, so we can conclude that differences in temperatures are the ultimate cause of heat transfer.

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = c_p \times m \times \Delta T$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Every substance has a characteristic specific heat, which is reported in units of cal/g•°C or cal/g•K, depending on the units used to express Δ*T*. The specific heat of a substance is the amount of energy that must be transferred to or from 1 g of that substance to change its temperature by 1°. Table $1.3.5.1$lists the specific heats for various materials.

Table 1.3.5.1*: Specific Heats of Some Common Substances*

Substance	**Specific Heat** $(\mathrm{J/g^oC})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

The *direction* of heat flow is not shown in heat = *mc*Δ*T*. If energy goes into an object, the total energy of the object increases, and the values of heat Δ*T* are positive. If energy is coming out of an object, the total energy of the object decreases, and the values of heat and Δ*T* are negative.

Example 1.3.5.1

A 15.0 g piece of cadmium metal absorbs 134 J of heat while rising from 24.0^oC to 62.7^oC. Calculate the specific heat of cadmium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Heat} = q = 134 \text{ J}$
- $\text{Mass} = m = 15.0 \text{ g}$
- $\Delta T = 62.7^{\text{o}}\text{C} - 24.0^{\text{o}}\text{C} = 38.7^{\text{o}}\text{C}$

Unknown

- c_p of cadmium $= ? \text{ J/g}^{\text{o}}\text{C}$

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134 \text{ J}}{15.0 \text{ g} \times 38.7^{\text{o}}\text{C}} = 0.231 \text{ J/g}^{\text{o}}\text{C}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals. The result has three significant figures.

Since most specific heats are known (Table $1.3.5.1$), they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that a 60.0 g of water at 23.52^{o}C was cooled by the removal of 813 J of heat. The change in temperature can be calculated using the specific heat equation:

$$\Delta T = \frac{q}{c_p \times m} = \frac{813 \text{ J}}{4.18 \text{ J/g}^{\text{o}}\text{C} \times 60.0 \text{ g}} = 3.24^{\text{o}}\text{C}$$

Since the water was being cooled, the temperature decreases. The final temperature is:

$$T_f = 23.52^{\text{o}}\text{C} - 3.24^{\text{o}}\text{C} = 20.28^{\text{o}}\text{C}$$

Example $1.3.5.2$

What quantity of heat is transferred when a 150.0 g block of iron metal is heated from 25.0°C to 73.3°C? What is the direction of heat flow?

Solution

We can use heat = $mc\Delta T$ to determine the amount of heat, but first we need to determine ΔT. Because the final temperature of the iron is 73.3°C and the initial temperature is 25.0°C, ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 73.3°\text{C} - 25.0°\text{C} = 48.3°\text{C}$$

The mass is given as 150.0 g, and Table 7.3 gives the specific heat of iron as 0.108 cal/g•°C. Substitute the known values into heat = $mc\Delta T$ and solve for amount of heat:

$$\text{heat} = (150.0 \text{ g}) \left(0.108 \, \frac{\text{cal}}{\text{g} \cdot {}^\circ\text{C}}\right) (48.3^\circ\text{C}) = 782 \text{ cal}$$

Note how the gram and °C units cancel algebraically, leaving only the calorie unit, which is a unit of heat. Because the temperature of the iron increases, energy (as heat) must be flowing *into* the metal.

Exercise $1.3.5.1$

What quantity of heat is transferred when a 295.5 g block of aluminum metal is cooled from 128.0°C to 22.5°C? What is the direction of heat flow?

Answer

Heat leaves the aluminum block.

Example 1.3.5.2

A 10.3 g sample of a reddish-brown metal gave off 71.7 cal of heat as its temperature decreased from 97.5°C to 22.0°C. What is the specific heat of the metal? Can you identify the metal from the data in Table 1.3.5.1?

Solution

The question gives us the heat, the final and initial temperatures, and the mass of the sample. The value of ΔT is as follows:

$$\Delta T = T_{\text{final}} - T_{\text{initial}} = 22.0°\text{C} - 97.5°\text{C} = -75.5°\text{C}$$

If the sample gives off 71.7 cal, it loses energy (as heat), so the value of heat is written as a negative number, −71.7 cal. Substitute the known values into heat = $mc\Delta T$ and solve for c:

$$-71.7\text{ cal} = (10.3\text{ g})(c)(-75.5°\text{C})$$

$$c = \frac{-71.7\text{ cal}}{(10.3\text{ g})(-75.5°\text{C})}$$

$$c = 0.0923\text{ cal/g}\bullet°\text{C}$$

This value for specific heat is very close to that given for copper in Table 7.3.

Exercise 1.3.5.2

A 10.7 g crystal of sodium chloride (NaCl) has an initial temperature of 37.0°C. What is the final temperature of the crystal if 147 cal of heat were supplied to it?

Answer

Summary

Specific heat calculations are illustrated.