Temperature Change and Heat capacity

Table 14.1 Specific Heats[1] of Various Substances

Substances	Specific heat (c)	
Solids	J/kg·ºC	kcal/kg·ºC[2]
Aluminum	900	0.215
Asbestos	800	0.19
Concrete, granite (average)	840	0.20
Copper	387	0.0924
Glass	840	0.20
Gold	129	0.0308
Human body (average at 37 °C)	3500	0.83
Ice (average, -50°C to 0°C)	2090	0.50
Iron, steel	452	0.108
Lead	128	0.0305
Silver	235	0.0562
Wood	1700	0.4
Liquids		
Benzene	1740	0.415
Ethanol	2450	0.586
Glycerin	2410	0.576
Mercury	139	0.0333
Water (15.0 °C)	4186	1.000
Gases [3]		
Air (dry)	721 (1015)	0.172 (0.242)
Ammonia	1670 (2190)	0.399 (0.523)
Carbon dioxide	638 (833)	0.152 (0.199)
Nitrogen	739 (1040)	0.177 (0.248)
Oxygen	651 (913)	0.156 (0.218)
Steam (100°C)	1520 (2020)	0.363 (0.482)

Table 14.2 Heats of Fusion and Vaporization [4]

		L_f			L_v	
Substance	Melting point (°C)	kJ/kg	kcal/kg	Boiling point (°C)	kJ/kg	kcal/kg
Helium	−269.7	5.23	1.25	−268.9	20.9	4.99
Hydrogen	−259.3	58.6	14.0	−252.9	452	108
Nitrogen	−210.0	25.5	6.09	−195.8	201	48.0
Oxygen	−218.8	13.8	3.30	−183.0	213	50.9
Ethanol	−114	104	24.9	78.3	854	204
Ammonia	−75		108	−33.4	1370	327
Mercury	−38.9	11.8	2.82	357	272	65.0
Water	0.00	334	79.8	100.0	2256[5]	539[6]
Sulfur	119	38.1	9.10	444.6	326	77.9
Lead	327	24.5	5.85	1750	871	208
Antimony	631	165	39.4	1440	561	134
Aluminum	660	380	90	2450	11400	2720
Silver	961	88.3	21.1	2193	2336	558
Gold	1063	64.5	15.4	2660	1578	377
Copper	1083	134	32.0	2595	5069	1211
Uranium	1133	84	20	3900	1900	454
Tungsten	3410	184	44	5900	4810	1150

1. On a hot day, the temperature of an 80,000-L swimming pool increases by 1.50ºC . What is the net heat transfer during this heating? Ignore any complications, such as loss of water by evaporation.

2. Show that 1 cal/g · ºC = 1 kcal/kg · ºC .

3. To sterilize a 50.0-g glass baby bottle, we must raise its temperature from 22.0ºC to 95.0ºC . How much heat transfer is required?

4. The same heat transfer into identical masses of different substances produce different temperature changes. Calculate the final temperature when 1.00 kcal of heat transfers into 1.00 kg of the following, originally at 20.0ºC : (a) water; (b) concrete; (c) steel; and (d) mercury.

5. Rubbing your hands together warms them by converting work into thermal energy. If a woman rubs her hands back and forth for a total of 20 rubs, at a distance of 7.50 cm per rub, and with an average frictional force of 40.0 N, what is the temperature increase? The mass of tissues warmed is only 0.100 kg, mostly in the palms and fingers.

6. A 0.250-kg block of a pure material is heated from 20.0ºC to 65.0ºC by the addition of 4.35 kJ of energy. Calculate its specific heat and identify the substance of which it is most likely composed.

7. Suppose identical amounts of heat transfer into different masses of copper and water, causing identical changes in temperature. What is the ratio of the mass of copper to water?

Phase change and latent heat

11. How much heat transfer (in kilocalories) is required to thaw a 0.450-kg package of frozen vegetables originally at 0ºC if their heat of fusion is the same as that of water?

12. A bag containing 0ºC ice is much more effective in absorbing energy than one containing the same amount of 0ºC water.

a. How much heat transfer is necessary to raise the temperature of 0.800 kg of water from 0ºC to 30.0ºC?

b. How much heat transfer is required to first melt 0.800 kg of 0ºC ice and then raise its temperature?

c. Explain how your answer supports the contention that the ice is more effective.

13. (a) How much heat transfer is required to raise the temperature of a 0.750-kg aluminum pot containing 2.50 kg of water from 30.0ºC to the boiling point and then boil away 0.750 kg of water? (b) How long does this take if the rate of heat transfer is 500 W

1 watt = 1 joule/second (1 W = 1 J/s) ?

14. The formation of condensation on a glass of ice water causes the ice to melt faster than it would otherwise. If 8.00 g of condensation forms on a glass containing both water and 200 g of ice, how many grams of the ice will melt as a result? Assume no other heat transfer occurs.

15. On a trip, you notice that a 3.50-kg bag of ice lasts an average of one day in your cooler. What is the average power in watts entering the ice if it starts at 0ºC and completely melts to 0ºC water in exactly one day

1 watt = 1 joule/second (1 W = 1 J/s) ?

16. On a certain dry sunny day, a swimming pool's temperature would rise by $1.50^\circ C$ if not for evaporation. What fraction of the water must evaporate to carry away precisely enough energy to keep the temperature constant?

17. (a) How much heat transfer is necessary to raise the temperature of a 0.200-kg piece of ice from $-20.0^\circ C$ to $130^\circ C$, including the energy needed for phase changes?

(b) How much time is required for each stage, assuming a constant 20.0 kJ/s rate of heat transfer?

(c) Make a graph of temperature versus time for this process.

18. In 1986, a gargantuan iceberg broke away from the Ross Ice Shelf in Antarctica. It was approximately a rectangle 160 km long, 40.0 km wide, and 250 m thick.

(a) What is the mass of this iceberg, given that the density of ice is 917 kg/m^3 ?

(b) How much heat transfer (in joules) is needed to melt it?

(c) How many years would it take sunlight alone to melt ice this thick, if the ice absorbs an average of 100 W/m^2 , 12.00 h per day?

19. How many grams of coffee must evaporate from 350 g of coffee in a 100-g glass cup to cool the coffee from $95.0^\circ C$ to $45.0^\circ C$? You may assume the coffee has the same thermal properties as water and that the average heat of vaporization is 2340 kJ/kg (560 cal/g). (You may neglect the change in mass of the coffee as it cools, which will give you an answer that is slightly larger than correct.)

20. (a) It is difficult to extinguish a fire on a crude oil tanker, because each liter of crude oil releases $2.80\times10^7 \text{ J}$ of energy when burned. To illustrate this difficulty, calculate the number of liters of water that must be expended to absorb the energy released by burning 1.00 L of crude oil, if the water has its temperature raised from $20.0^\circ C$ to $100^\circ C$, it boils, and the resulting steam is raised to $300^\circ C$. (b) Discuss additional complications caused by the fact that crude oil has a smaller density than water.

21. The energy released from condensation in thunderstorms can be very large. Calculate the energy released into the atmosphere for a small storm of radius 1 km, assuming that 1.0 cm of rain is precipitated uniformly over this area.

22. To help prevent frost damage, 4.00 kg of $0^\circ C$ water is sprayed onto a fruit tree.

(a) How much heat transfer occurs as the water freezes?
(b) How much would the temperature of the 200-kg tree decrease if this amount of heat transferred from the tree?
Take the specific heat to be $3.35\ \text{kJ/kg}\cdot{}^{\circ}\text{C}$, and assume that no phase change occurs.

23. A 0.250-kg aluminum bowl holding 0.800 kg of soup at 25.0°C is placed in a freezer. What is the final temperature if 377 kJ of energy is transferred from the bowl and soup, assuming the soup's thermal properties are the same as that of water? Explicitly show how you follow the steps in **Problem-Solving Strategies for the Effects of Heat Transfer**.

24. A 0.0500-kg ice cube at -30.0°C is placed in 0.400 kg of 35.0°C water in a very well-insulated container. What is the final temperature?

25. If you pour 0.0100 kg of 20.0°C water onto a 1.20-kg block of ice (which is initially at -15.0°C), what is the final temperature? You may assume that the water cools so rapidly that effects of the surroundings are negligible.

26. Indigenous people sometimes cook in watertight baskets by placing hot rocks into water to bring it to a boil. What mass of 500°C rock must be placed in 4.00 kg of 15.0°C water to bring its temperature to 100°C, if 0.0250 kg of water escapes as vapor from the initial sizzle? You may neglect the effects of the surroundings and take the average specific heat of the rocks to be that of granite.

1. *On a hot day, the temperature of an 80,000-L swimming pool increases by* 1.50°C. *What is the net heat transfer during this heating? Ignore any complications, such as loss of water by evaporation.*

Solution

$$m = \rho V = (1.00\times10^{3}\ \text{kg/m}^{3})(80{,}000\ \text{L})\times\frac{1\ \text{m}^{3}}{1000\ \text{L}} = 8.00\times10^{4}\ \text{kg}.$$

Therefore, $Q = mc\Delta T = (8.00\times10^{4}\ \text{kg})(4186\ \text{J/kg}\cdot{}^{\circ}\text{C})(1.50^{\circ}\text{C}) = \underline{5.02\times10^{8}\ \text{J}}$

2. *Show that* $1\ \text{cal/g}\cdot{}^{\circ}\text{C} = 1\ \text{kcal/kg}\cdot{}^{\circ}\text{C}$.

Solution $\frac{1\text{ cal}}{\text{g}\cdot°\text{C}}\times\frac{1\text{ kcal}}{1000\text{ cal}}\times\frac{1000\text{ g}}{1\text{ kg}}=\underline{1\text{ kcal/kg}\cdot°\text{C}}$

3. *To sterilize a 50.0-g glass baby bottle, we must raise its temperature from* $22.0°\text{C}$ *to* $95.0°\text{C}$ *. How much heat transfer is required?*

Solution $Q=mc\Delta T=(50.0\times10^{-3}\text{ kg})(840\text{ J/kg}\cdot°\text{C})(73.0°\text{C})=3066\text{ J}=\underline{3.07\times10^{3}\text{ J}}$

4. *The same heat transfer into identical masses of different substances produces different temperature changes. Calculate the final temperature when 1.00 kcal of heat transfers into 1.00 kg of the following, originally at* $20.0°\text{C}$ *: (a) water; (b) concrete; (c) steel; and (d) mercury.*

Solution $Q=mc\Delta T\Rightarrow\Delta T=\frac{Q}{mc}$

(a) $\Delta T=\frac{1.00\text{ kcal}}{(1.00\text{ kg})(1.00\text{ kcal/kg}\cdot°\text{C})}=1.00\text{ °C}\Rightarrow T=20.0\text{ °C}+1.00\text{ °C}=\underline{21.0\text{ °C}}$

(b) $\Delta T=\frac{1.00\text{ kcal}}{(1.00\text{ kg})(0.20\text{ kcal/kg}\cdot°\text{C})}=5.0°\text{C}\Rightarrow T=20.0°\text{C}+5.0°\text{C}=\underline{25.0°\text{C}}$

(c) $\Delta T=\frac{1.00\text{ kcal}}{(1.00\text{ kg})(0.108\text{ kcal/kg}\cdot°\text{C})}=9.26°\text{C}\Rightarrow T=20.0°\text{C}+9.26°\text{C}=\underline{29.3°\text{C}}$

(d) $\Delta T=\frac{1.00\text{ kcal}}{(1.00\text{ kg})(0.0333\text{ kcal/kg}\cdot°\text{C})}=30.0°\text{C}\Rightarrow T=20.0°\text{C}+30.0°\text{C}=\underline{50.0°\text{C}}$

5. *Rubbing your hands together warms them by converting work into thermal energy. If a woman rubs her hands back and forth for a total of 20 rubs, at a distance of 7.50 cm per rub, and with an average frictional force of 40.0 N, what is the temperature increase? The mass of tissues warmed is only 0.100 kg, mostly in the palms and fingers.*

Solution Let N be the number of hand rubs and F be the average frictional force of a hand rub:

$$Q=NFd=mc\Delta T\Rightarrow\Delta T=\frac{NFd}{mc}=\frac{20(40.0\text{ N})(7.50\times10^{-2}\text{ m})}{(0.100\text{ kg})(3500\text{ J/kg}\cdot°\text{C})}=\underline{0.171°\text{C}}$$

6. *A 0.250-kg block of a pure material is heated from* $20.0°\text{C}$ *to* $65.0°\text{C}$ *by the addition of 4.35 kJ of energy. Calculate its specific heat and identify the substance of which it is most likely composed.*

Solution

$$Q = mc\Delta T \Rightarrow c = \frac{Q}{m\Delta T} = \frac{1.04\ \text{kcal}}{(0.250\ \text{kg})(45.0°\text{C})} = \underline{0.0924\ \text{kcal/kg}\cdot°\text{C}}$$

It is copper.

7. *Suppose identical amounts of heat transfer into different masses of copper and water, causing identical changes in temperature. What is the ratio of the mass of copper to water?*

Solution

$$m_\text{w}c_\text{w}\Delta T = Q = m_\text{c}c_\text{c}\Delta T$$

$$\frac{m_\text{c}}{m_\text{w}} = \frac{c_\text{w}}{c_\text{c}} = \frac{1\ \text{kcal/kg}\cdot°\text{C}}{0.0924\ \text{kcal/kg}\cdot°\text{C}} = \underline{10.8}$$

8. *(a) The number of kilocalories in food is determined by calorimetry techniques in which the food is burned and the amount of heat transfer is measured. How many kilocalories per gram are there in a 5.00-g peanut if the energy from burning it is transferred to 0.500 kg of water held in a 0.100-kg aluminum cup, causing a* $54.9°\text{C}$ *temperature increase? (b) Compare your answer to labeling information found on a package of peanuts and comment on whether the values are consistent.*

Solution

$$Q = m_\text{w}c_\text{w}\Delta T + m_\text{Al}c_\text{Al}\Delta T = \left(m_\text{w}c_\text{w} + m_\text{Al}c_\text{Al}\right)\Delta T$$

$$Q = \begin{bmatrix}(0.500\ \text{kg})(1.00\ \text{kcal/kg}\cdot°\text{C}) + \\ (0.100\ \text{kg})(0.215\ \text{kcal/kg}\cdot°\text{C})\end{bmatrix}(54.9°\text{C}) = 28.63\ \text{kcal}$$

(a) $\dfrac{Q}{m_\text{p}} = \dfrac{28.63\ \text{kcal}}{5.00\ \text{g}} = \underline{5.73\ \text{kcal/g}}$

(b) A label for unsalted dry roasted peanuts says that 33 g contains 200 calories (kcal), which is $\dfrac{Q}{m_\text{p}} = \dfrac{200\ \text{kcal}}{33\ \text{g}} = \underline{6\ \text{kcal/g},}$ which is consistent with our results to part (a), to one significant figure.

9. *Following vigorous exercise, the body temperature of an 80.0-kg person is* $40.0°\text{C}$. *At what rate in watts must the person transfer thermal energy to reduce the body*

temperature to $37.0°\text{C}$ *in 30.0 min, assuming the body continues to produce energy at the rate of 150 W?* $(1 \text{ watt} = 1 \text{ joule/second or } 1 \text{ W} = 1 \text{ J/s})$

Solution

$$Q = mc_{\text{human body}}\Delta T = (80.0 \text{ kg})(3500 \text{ J/kg}\cdot°\text{C})(40°\text{C - }37°\text{C}) = 8.40\times10^5 \text{ J}$$

$$P_{\text{cooling}} = \frac{Q}{t} = \frac{8.40\times10^5 \text{ J}}{(30 \text{ min})(60 \text{ s/1 min})} = 4.67\times10^2 \text{ W}$$

Thus, $P_{\text{required}} = P_{\text{cooling}} + P_{\text{body}} = 467 \text{ W} + 150 \text{ W} = \underline{617 \text{ W}}$.

10. *Even when shut down after a period of normal use, a large commercial nuclear reactor transfers thermal energy at the rate of 150 MW by the radioactive decay of fission products. This heat transfer causes a rapid increase in temperature if the cooling system fails* $(1 \text{ watt} = 1 \text{ joule/second or } 1 \text{ W} = 1 \text{ J/s and } 1 \text{ MW} = 1 \text{ megawatt})$ *. (a) Calculate the rate of temperature increase in degrees Celsius per second* $(°\text{C/s})$ *if the mass of the reactor core is* $1.60\times10^5 \text{ kg}$ *and it has an average specific heat of* $0.3349 \text{ kJ/kg}\cdot°\text{C}$ *. (b) How long would it take to obtain a temperature increase of* $2000°\text{C}$ *, which could cause some metals holding the radioactive materials to melt? (The initial rate of temperature increase would be greater than that calculated here because the heat transfer is concentrated in a smaller mass. Later, however, the temperature increase would slow down because the* $5\times10^5\text{ - kg}$ *steel containment vessel would also begin to heat up.)*

Solution

(a) $Q = mc\Delta T \Rightarrow \Delta T = \dfrac{Q}{mc}$

Recall that 1 W = 1 J/s. Thus ΔT for 1 s is given by

$$\Delta T = \frac{(150\times10^6 \text{ J})(1 \text{ kcal/4186J})}{(1.60\times10^5 \text{ kg})(0.0800 \text{ kcal/kg}\cdot°\text{C})} = 2.80°\text{C} \Rightarrow \text{Rate} = \underline{2.80°\text{C/s}}$$

(b) $t = \dfrac{2000°\text{C}}{2.80°\text{C/s}} = 714 \text{ s} = \underline{11.9 \text{ min}}$

14.3 Phase Change and Latent Heat

11. *How much heat transfer (in kilocalories) is required to thaw a 0.450-kg package of frozen vegetables originally at* $0°\text{C}$ *if their heat of fusion is the same as that of water?*

Solution $Q = mL_\text{f} = (0.450\ \text{kg})(79.8\ \text{kcal/kg}) = \underline{35.9\ \text{kcal}}$

12. *A bag containing* $0°\text{C}$ *ice is much more effective in absorbing energy than one containing the same amount of* $0°\text{C}$ *water. (a) How much heat transfer is necessary to raise the temperature of 0.800 kg of water from* $0°\text{C}$ *to* $30.0°\text{C}$ *? (b) How much heat transfer is required to first melt 0.800 kg of* $0°\text{C}$ *ice and then raise its temperature? (c) Explain how your answer supports the contention that the ice is more effective.*

Solution

(a) $Q = mc\Delta T = (0.800\ \text{kg})(4186\ \text{J/kg}\cdot°\text{C})(30.0°\text{C}) = \underline{1.00\times10^5\ \text{J}}$

(b) $Q = mL_\text{f} + mc\Delta T = (0.800\ \text{kg})(334\times10^3\ \text{J/kg}) + 1.005\times10^5\ \text{J} = \underline{3.68\times10^5\ \text{J}}$

(c) The ice is much more effective in absorbing heat because it first must be melted, which requires a lot of energy, then it gains the same amount of heat as the bag that started with water. The first $2.67\times10^5\ \text{J}$ of heat is used to melt the ice, then it absorbs the $1.00\times10^5\ \text{J}$ of heat as water.

13. *(a) How much heat transfer is required to raise the temperature of a 0.750-kg aluminum pot containing 2.50 kg of water from* $30.0°\text{C}$ *to the boiling point and then boil away 0.750 kg of water? (b) How long does this take if the rate of heat transfer is 500 W [* $1\ \text{watt} = 1\ \text{joule/second}\ (1\ \text{W} = 1\ \text{J/s})$ *]?*

Solution

$$Q = m_\text{w}c_\text{w}\Delta T + m_\text{Al}c_\text{Al}\Delta T + m'_\text{w}L_\text{v}$$

$$Q = (2.50\ \text{kg})(1.00\ \text{kcal/kg}\cdot°\text{C})(70.0°\text{C}) + (0.750\ \text{kg})(0.215\ \text{kcal/kg}\cdot°\text{C})(70.0°\text{C}) + (0.750\ \text{kg})(539\ \text{kcal/kg})$$

(a) $\text{Q} = 590.5\ \text{kcal} = \underline{591\ \text{kcal}}$

$$Q = Pt \Rightarrow t = \frac{Q}{P},\ \text{where}\ P = \text{power and}\ t = \text{time.}$$

(b) $t = (590.5\ \text{kcal})\left(\dfrac{4186\ \text{J/kcal}}{500\ \text{W}}\right) = \underline{4.94\times10^3\ \text{s}}$

14. *The formation of condensation on a glass of ice water causes the ice to melt faster than it would otherwise. If 8.00 g of condensation forms on a glass containing both water and 200 g of ice, how many grams of the ice will melt as a result? Assume no other heat transfer occurs.*

Solution

$$m_{\text{w}}L_{\text{v}} = m_{\text{ice}}L_{\text{f}} \Rightarrow m_{\text{ice}} = m_{\text{w}}\frac{L_{\text{v}}}{L_{\text{f}}} = (8.00\text{ g})\left(\frac{580\text{ kcal/kg}}{79.8\text{ kcal/kg}}\right) = \underline{58.1\text{ g}}$$

(Note that L_{v} for water at $37°\text{C}$ is used here as a better approximation than L_{v} for $100°\text{C}$ water.)

15. *On a trip, you notice that a 3.50-kg bag of ice lasts an average of one day in your cooler. What is the average power in watts entering the ice if it starts at* $0°\text{C}$ *and completely melts to* $0°\text{C}$ *water in exactly one day [* $1\text{ watt} = 1\text{ joule/second } (1\text{ W} = 1\text{ J/s})$ *]?*

Solution

$$P = \frac{Q}{t} = \frac{mL_{\text{f}}}{t} = \frac{(3.50\text{ kg})(334\ \text{kJ/kg}\cdot°\text{C})}{86400\text{ s}} = 13.53\text{ W} = \underline{13.5\ \text{W}}$$

16. *On a certain dry sunny day, a swimming pool's temperature would rise by* $1.50°\text{C}$ *if not for evaporation. What fraction of the water must evaporate to carry away precisely enough energy to keep the temperature constant?*

Solution

Let M be the mass of pool water and m be the mass of pool water that evaporates.

$$Mc\Delta T = mL_{\text{v}(37°\text{C})} \Rightarrow \frac{m}{M} = \frac{c\Delta T}{L_{\text{v}(37°\text{C})}} = \frac{(1.00\text{ kcal/kg}\cdot°\text{C})(1.50°\text{C})}{580\text{ kcal/kg}} = \underline{2.59\times10^{-3}}$$

(Note that L_{v} for water at $37°\text{C}$ is used here as a better approximation than L_{v} for $100°\text{C}$ water.)

17. *(a) How much heat transfer is necessary to raise the temperature of a 0.200-kg piece of ice from* $-20.0°\text{C}$ *to* $130°\text{C}$*, including the energy needed for phase changes? (b) How much time is required for each stage, assuming a constant 20.0 kJ/s rate of heat transfer? (c) Make a graph of temperature versus time for this process.*

Solution

(a) (i) Heat needed to warm ice to $0°C$

$$Q_1 = m_i c_i \Delta T = (0.200\text{ kg})(2.090\text{ kJ/kg}\cdot°\text{C})(20°\text{C}) = 8.36\text{ kJ}$$

(ii) Heat needed to melt ice at $0°C$

$$Q_2 = m_i L_f = (0.200\text{ kg})(334\text{ kJ/kg}) = 66.8\text{ kJ}$$

(iii) Heat required to warm $0°C$ water to $100°C$

$$Q_3 = m_i c_w \Delta T = (0.200\text{ kg})(4.186\text{ kJ/kg}\cdot^{\circ}\text{C})(100^{\circ}\text{C}) = 83.73\text{ kJ} = 83.7\text{ kJ}$$

(iv) Heat required to vaporize water at $100°C$

$$Q_4 = m_i L_v = (0.200\text{ kg})(2256\text{ kJ/kg}) = 451.2\text{ kJ} = 451\text{ kJ}$$

(v) Heat required to warm $100°C$ vapor to $130°C$

$$Q_5 = m_i c_v \Delta T = (0.200\text{ kg})(1.520\text{ kJ/kg}\cdot°\text{C})(30°\text{C}) = 9.12\text{ kJ}$$

Total heat required $Q = Q_1 + Q_2 + Q_3 + Q_4 + Q_5 = 619.0\text{ kJ} = 147.9\text{ kcal} = \underline{148\text{ kcal}}$

.

(b) $P = \frac{Q}{t} \Rightarrow t = \frac{Q}{P}$

(i) $t_1 = \frac{Q_1}{P} = \frac{8.36\text{ kJ}}{20\text{ kJ/s}} = 0.418\text{ s}$

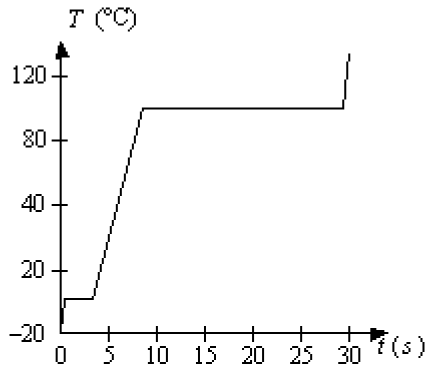
(ii) $t_2 = \frac{Q_2}{P} = \frac{66.8\text{ kJ}}{20\text{ kJ/s}} = 3.34\text{ s}$

(iii) $t_3 = \frac{Q_3}{P} = \frac{83.7\text{ kJ}}{20\text{ kJ/s}} = 4.185\text{ s} = 4.19\text{ s}$

(iv) $t_4 = \frac{Q_4}{P} = \frac{451\text{ kJ}}{20\text{ kJ/s}} = 22.6\text{ s}$

(v) $t_5 = \frac{Q_5}{P} = \frac{9.12\text{ kJ}}{20\text{ kJ/s}} = 0.456\text{ s}$

Total time $t = t_1 + t_2 + t_3 + t_4 + t_5 = 31.00\text{ s} = \underline{31\text{ s}}$

(c) 

18. *In 1986, a gargantuan iceberg broke away from the Ross Ice Shelf in Antarctica. It was approximately a rectangle 160 km long, 40.0 km wide, and 250 m thick. (a) What is the mass of this iceberg, given that the density of ice is* 917 kg/m^3*? (b) How much heat transfer (in joules) is needed to melt it? (c) How many years would it take sunlight alone to melt ice this thick, if the ice absorbs an average of* 100 W/m^2*, 12.00 h per day?*

Solution

$$m = \rho V = \rho(lwh) = (917\text{ kg/m}^3)(160\times10^3\text{ m})(40.0\times10^3\text{ m})(250\text{ m})$$

(a) $$= 1.467\times10^{15}\text{ kg} = \underline{1.47\times10^{15}\text{ kg}}$$

(b) $$Q = mL_\text{f} = (1.467\times10^{15}\text{ kg})(79.8\text{ kcal/kg})(4186\text{ J/kcal}) = \underline{4.90\times10^{20}\text{ J}}$$

$$Q_\text{day} = (100\text{ W/m}^2)(160\times10^3\text{ m})(40.0\times10^3\text{ m})(12\text{ h})\left(\frac{3600\text{ s}}{1\text{ h}}\right) = 2.765\times10^{16}\text{ J}$$

(c) $$n = \frac{Q}{Q_\text{day}} = \frac{(1.171\times10^{17}\text{ kcal})(4186\text{ J/kcal})}{2.765\times10^{16}\text{ J}} = 1.773\times10^4\text{ d}\times\frac{1\text{ y}}{365.25\text{ d}} = \underline{48.5\text{ y}}$$

19. *How many grams of coffee must evaporate from 350 g of coffee in a 100-g glass cup to cool the coffee from* $95.0°\text{C}$ *to* $45.0°\text{C}$*? You may assume the coffee has the same thermal properties as water and that the average heat of vaporization is 2340 kJ/kg (560 cal/g). (You may neglect the change in mass of the coffee as it cools, which will give you an answer that is slightly larger than correct.)*

Solution The heat gained in evaporating the coffee equals the heat leaving the coffee and glass to lower its temperature, so that $ML_{\mathrm{v}} = m_{\mathrm{c}}c_{\mathrm{c}}\Delta T + m_{\mathrm{g}}c_{\mathrm{g}}\Delta T,$ where M is the mass of coffee that evaporates. Solving for the evaporated coffee gives:

$$\begin{aligned} M &= \frac{\Delta T(m_{\mathrm{c}}c_{\mathrm{c}} + m_{\mathrm{g}}c_{\mathrm{g}})}{L_{\mathrm{v}}} \\ &= \frac{(95.0°\mathrm{C} - 45.0°\mathrm{C})}{560\ \mathrm{cal/g}} \cdot \left[(350\ \mathrm{g})(1.00\ \mathrm{cal/g}\cdot°\mathrm{C}) + (100\ \mathrm{g})(0.20\ \mathrm{cal/g}\cdot°\mathrm{C})\right] = \underline{33.0\ \mathrm{g}} \end{aligned}$$

20. *(a) It is difficult to extinguish a fire on a crude oil tanker, because each liter of crude oil releases* $2.80\times10^{7}\ \mathrm{J}$ *of energy when burned. To illustrate this difficulty, calculate the number of liters of water that must be expended to absorb the energy released by burning 1.00 L of crude oil, if the water has its temperature raised from* $20.0°\mathrm{C}$ *to* $100\,°\mathrm{C}$*, it boils, and the resulting steam is raised to* $300\,°\mathrm{C}$*. (b) Discuss additional complications caused by the fact that crude oil has a smaller density than water.*

Solution (a) $Q = mc_{\mathrm{w}}\Delta T_{\mathrm{w}} + mL_{\mathrm{v}} + mc_{\mathrm{s}}\Delta T_{\mathrm{s}}$

$$\begin{aligned} m &= \frac{Q}{c_{\mathrm{w}}\Delta T_{\mathrm{w}} + L_{\mathrm{v}} + c_{\mathrm{s}}\Delta T_{\mathrm{s}}} \\ &= \frac{2.80\times10^{7}\ \mathrm{J}}{(4186\ \mathrm{J/kg}\cdot°\mathrm{C})(80.0°\mathrm{C}) + 2256\times10^{3}\ \mathrm{J/kg} + (1520\ \mathrm{J/kg}\cdot°\mathrm{C})(200°\mathrm{C})} = 9.67\ \mathrm{kg} \\ \Rightarrow V &= 9.67\ \mathrm{kg}\times\frac{1\,\mathrm{m}^{3}}{1.00\times10^{3}\ \mathrm{kg}}\times\frac{1\,\mathrm{L}}{10^{-3}\ \mathrm{m}^{3}} = \underline{9.67\ \mathrm{L}} \end{aligned}$$

(b) Crude oil is less dense than water, so it floats on top of the water, thereby exposing it to the oxygen in the air, which it uses to burn. Also, if the water is under the oil, it is less able to absorb the heat generated by the oil.

21. *The energy released from condensation in thunderstorms can be very large. Calculate the energy released into the atmosphere for a small storm of radius 1 km, assuming that 1.0 cm of rain is precipitated uniformly over this area.*

Solution We have a phase change $Q = mL_{\mathrm{v}}$. We need to find mass of rain in a cloud of radius 1 km. $m = \rho V = (1000\ \mathrm{kg/m^{3}})(0.01\,\mathrm{m})(\pi\times10^{6}\ \mathrm{m}^{2}) = \pi\times10^{7}\ \mathrm{kg}$. With $Q = mL_{\mathrm{v}}$ and $L_{\mathrm{v}} = 2256\ \mathrm{kJ/kg}$, we find $Q = \underline{7\times10^{13}\ \mathrm{J}}$ – about the energy released in the first atomic bomb explosion.

22. *To help prevent frost damage, 4.00 kg of* $0°C$ *water is sprayed onto a fruit tree. (a) How much heat transfer occurs as the water freezes? (b) How much would the temperature of the 200-kg tree decrease if this amount of heat transferred from the tree? Take the specific heat to be* $3.35\ \text{kJ/kg}\cdot°\text{C}$*, and assume that no phase change occurs.*

Solution

(a) $Q = mL_\text{f} = (4.00\ \text{kg})(79.8\ \text{kcal/kg}\cdot°\text{C}) = 319.2\ \text{kcal} = \underline{319\ \text{kcal}}$

(b) $$Q = mc\Delta T \Rightarrow \Delta T = \frac{Q}{mc} = \frac{319.2\ \text{kcal}}{(200\ \text{kg})(0.800\ \text{kcal/kg}\cdot°\text{C})} = \underline{2.00°\text{C}}$$

23. *A 0.250-kg aluminum bowl holding 0.800 kg of soup at* $25.0°C$ *is placed in a freezer. What is the final temperature if 377 kJ of energy is transferred from the bowl and soup, assuming the soup's thermal properties are the same as that of water? Explicitly show how you follow the steps in the Problem-Solving Strategies for the Effects of Heat Transfer.*

Solution

To bring the system to $0°C$ requires heat, Q, of:

$$\begin{aligned} Q &= m_\text{Al}c_\text{Al}\Delta T + m_\text{s}c_\text{s}\Delta T \\ &= \left[(0.250\ \text{kg})(0.215\ \text{kcal/kg}\cdot°\text{C}) + (0.800\ \text{kg})(1.00\ \text{kcal/kg}\cdot°\text{C})\right](25.0°\text{C}) = 21.34\ \text{kcal} \end{aligned}$$

This leaves $(90.0 - 21.34)\ \text{kcal} = 68.66\ \text{kcal}$ to freeze all the soup, leaving

$Q'' = (68.66 - 63.84)\ \text{kcal} = 4.82\ \text{kcal}$ to be removed. So, we can now determine the final temperature of the frozen soup:

$$\begin{aligned} Q'' &= (m_\text{Al}c_\text{Al} + m_\text{s}c_\text{s})\Delta T = (m_\text{Al}c_\text{Al} + m_\text{s}c_\text{s})(0°\text{C} - T_\text{f}). \\ T_\text{f} &= \frac{-Q''}{m_\text{Al}c_\text{Al} + m_\text{s}c_\text{s}} \\ &= \frac{-4.82\ \text{kcal}}{(0.250\ \text{kg})(0.215\ \text{kcal/kg}\cdot°\text{C}) + (0.800\ \text{kg})(0.500\ \text{kcal/kg}\cdot°\text{C})} = \underline{-10.6°\text{C}} \end{aligned}$$

24. *A 0.0500-kg ice cube at* $-30.0°C$ *is placed in 0.400 kg of* $35.0°C$ *water in a very well-insulated container. What is the final temperature?*

Solution

First bring the ice up to $0°C$ and melt it with heat Q_1:

$$Q_1 = mc\Delta T_1 + mL_\text{f} = (0.0500\text{ kg})\left[(0.500\text{ kcal/kg}\cdot°\text{C})(30.0°\text{C}) + (79.8\text{ kcal/kg})\right]$$
$$= 4.74\text{ kcal}$$

This lowers the temperature of water by ΔT_2 :

$$Q_1 = mc\Delta T_2 \Rightarrow \Delta T_2 = \frac{Q_1}{mc} = \frac{4.74\text{ kcal}}{(0.400\text{ kg})(1.00\text{ kcal/kg}\cdot°\text{C})} = 11.85°\text{C}$$
$$\text{New } T_\text{w} = 35.0°\text{C} - 11.85°\text{C} = 23.15°\text{C}$$

Now, the heat lost by the hot water equals that gained by the cold water (T_f is the final temperature):

$$m_\text{c}c_\text{w}(T_\text{f} - T_\text{c}) = m_\text{h}c_\text{w}(T_\text{h} - T_\text{f})$$
$$T_\text{f} = \frac{m_\text{h}T_\text{h} + m_\text{c}T_\text{c}}{m_\text{c} + m_\text{h}} = \frac{(0.400\text{ kg})(296.3\text{ K}) + (0.0500\text{ kg})(273.15\text{ K})}{0.450\text{ kg}} = 293.7\text{ K} = \underline{20.6°\text{C}}$$

25. *If you pour 0.0100 kg of* $20.0°\text{C}$ *water onto a 1.20-kg block of ice (which is initially at* $-15.0°\text{C}$ *), what is the final temperature? You may assume that the water cools so rapidly that effects of the surroundings are negligible.*

Solution First, we need to calculate how much heat would be required to raise the temperature of the ice to $0°\text{C}$:

$$Q_\text{ice} = mc\Delta T = (1.20\text{ kg})(2090\text{ J/kg}\cdot°\text{C})(15°\text{C}) = 3.762\times10^4\text{ J}$$

Now, we need to calculate how much heat is given off to lower the water to $0°\text{C}$:

$$Q_1 = mc\Delta T_1 = (0.0100\text{ kg})(4186\text{ J/kg}\cdot°\text{C})(20.0°\text{C}) = 837.2\text{ J}$$

Since this is less than the heat required to heat the ice, we need to calculate how much heat is given off to convert the water to ice:

$$Q_2 = mL_\text{f} = (0.0100\text{ kg})(334\times10^3\text{ J/kg}) = 3.340\times10^3\text{ J}$$

Thus, the total amount of heat given off to turn the water to ice at $0°\text{C}$:

$$Q_\text{water} = 4.177\times10^3\text{ J}.$$

Since $Q_\text{ice} > Q_\text{water}$, we have determined that the final state of the water/ice is ice at some temperature below $0°\text{C}$. Now, we need to calculate the final temperature. We set the heat lost from the water equal to the heat gained by the ice, where we now know that the final state is ice at $T_f < 0°C$:

$$Q_{\text{lost by water}} = Q_{\text{gained by ice,}} \text{ or } m_{\text{water}} c_{\text{water}} \Delta T_{20\to 0} + m_{\text{water}} L_f + m_{\text{water}} c_{\text{ice}} \Delta T_{0\to ?} = m_{\text{ice}} c_{\text{ice}} \Delta T_{-15\to ?}$$

Substituting for the change in temperatures (being careful that ΔT is always positive) and simplifying gives:

$$m_{\text{water}}[c_{\text{water}}(20°C) + L_f + (c_{\text{ice}})(0 - T_f)] = m_{\text{ice}} c_{\text{ice}}[T_f - (-15°C)].$$

Solving for the final temperature gives $T_f = \dfrac{m_{\text{water}}[c_{\text{water}}(20°C) + L_f] - m_{\text{ice}} c_{\text{ice}}(15°C)}{(m_{\text{water}} + m_{\text{ice}})c_{\text{ice}}}$

and so finally,

$$\begin{aligned} T_f &= \frac{(0.0100\text{ kg})[(4186\text{ J/kg}\cdot°C)(20°C) + 334\times 10^3\text{ J/kg}]}{(0.0100\text{ kg} + 1.20\text{ kg})(2090\text{ J/kg}\cdot°C)} \\ &\quad - \frac{(1.20\text{ kg})(2090\text{ J/kg}\cdot°C)(15°C)}{(0.0100\text{ kg} + 1.20\text{ kg})(2090\text{ J/kg}\cdot°C)} \\ &= \underline{-13.2°C} \end{aligned}$$

26. *Indigenous people sometimes cook in watertight baskets by placing hot rocks into water to bring it to a boil. What mass of* $500°C$ *rock must be placed in 4.00 kg of* $15.0°C$ *water to bring its temperature to* $100°C$*, if 0.0250 kg of water escapes as vapor from the initial sizzle? You may neglect the effects of the surroundings and take the average specific heat of the rocks to be that of granite.*