43 HEAT CALCULATIONS

1. Calculate the quantity of heat required to warm 1.25 L of water from 22.0°C to 98.0°C in an electric kettle.

$$q = vc\Delta t$$
$$= 1.25\ \text{L} \times \frac{4.19\ \text{kJ}}{\text{L}\bullet{}^\circ\text{C}} \times (98.0 - 22.0)^\circ\text{C}$$
$$= 398\ \text{kJ}$$

2. Assuming a volumetric heat capacity the same as pure water, calculate the heat that must be released from a soft drink in a 2.00 L container when the soft drink is cooled from 22.0°C to 10.0°C.

$$q = vc\Delta t$$
$$= 2.00\ \text{L} \times \frac{4.19\ \text{kJ}}{\text{L}\bullet{}^\circ\text{C}} \times (22.0 - 10.0)^\circ\text{C}$$
$$= 101\ \text{kJ}$$

3. What mass of aluminum in a car engine will absorb 1.00 MJ of heat when the temperature rises from 22°C to 102°C after the car is started?

$$q = mc\Delta t$$
$$1.00\ \text{MJ} = m \times \frac{0.900\ \text{J}}{\text{g}\bullet{}^\circ\text{C}} \times (102 - 22)^\circ\text{C}$$
$$m = 14\ \text{kg}$$

4. Assume the liquid coolant in a car engine has a volumetric heat capacity of 3.88 kJ/(L•°C). Determine the volume of coolant that will absorb 1.00 MJ of heat during a temperature rise from 22°C to 102°C.

$$q = vc\Delta t$$
$$1.00\ \text{MJ} = v \times \frac{3.88\ \text{kJ}}{\text{L}\bullet{}^\circ\text{C}} \times (102 - 22)^\circ\text{C}$$
$$v = 3.2\ \text{L}$$

5. In a laboratory experiment, 2.00 kJ of heat flowed to a 100 g sample of a liquid solvent, causing a temperature increase from 15.40°C to 21.37°C. Calculate the specific heat capacity of the liquid.

$$q = mc\Delta t$$
$$2.00\ \text{kJ} = 100\ \text{g} \times c \times (21.37 - 15.40)^\circ\text{C}$$
$$c = 3.35\ \text{J/(g}\bullet{}^\circ\text{C)}$$

6. A human body loses about 360 kJ of heat every hour. Assuming that an average human body is equivalent to about 60 kg of water, what temperature decrease would this heat transfer cause? (Of course, this heat is replaced by body metabolism.)

$$q = mc\Delta t$$
$$360\ \text{kJ} = 60\ \text{kg} \times \frac{4.19\ \text{J}}{\text{g}\bullet{}^\circ\text{C}} \times \Delta t$$
$$\Delta t = 1.4^\circ\text{C}$$

44 ENTHALPY CHANGES

1. Enthalpy changes may be classified into three types — phase, chemical, and nuclear.
   (a) How are these types different empirically?
   Phase changes involve only a change of state. Chemical changes result in new chemical substances being formed. In nuclear changes, new elements or subatomic particles are produced.
   (b) How are they similar theoretically?
   Phase, chemical and nuclear changes all involve a change in potential energy as a result of changes in bonding.

2. Calculate the enthalpy change for the melting of a 30 g ice cube.

$$\begin{aligned}\Delta H_{fus} &= nH_{fus}\\ &= 30\text{ g} \times \frac{1\text{ mol}}{18.02\text{ g}} \times \frac{6.03\text{ kJ}}{\text{mol}}\\ &= 10\text{ kJ}\end{aligned}$$

3. A reference gives a value of +39.23 kJ/mol for the molar enthalpy of vaporization for methanol. What enthalpy change occurs in the evaporation of 10.0 g of methanol?

$$\begin{aligned}\Delta H_{vap} &= nH_{vap}\\ &= 10.0\text{ g} \times \frac{1\text{ mol}}{32.05\text{ g}} \times \frac{39.23\text{ kJ}}{\text{mol}}\\ &= 12.2\text{ kJ}\end{aligned}$$

4. Given H_{vap} is +1.37 kJ/mol for NH_3 (page 293), find the mass of ammonia that can be condensed from vapor to liquid (with no temperature change) during an enthalpy change of 10.0 kJ.

$$\begin{aligned}\Delta H_{cond} &= nH_{cond}\\ 10.0\text{ kJ} &= m \times \frac{1\text{ mol}}{17.04\text{ g}} \times \frac{1.37\text{ kJ}}{\text{mol}}\\ m &= 124\text{ g}\end{aligned}$$

5. An experiment produces evidence that the evaporation of 4.00 g of liquid butane, $C_4H_{10(l)}$, requires a gain in enthalpy of 1.67 kJ. Find the molar enthalpy of vaporization for butane from this evidence.

$$\begin{aligned}\Delta H_{vap} &= nH_{vap}\\ 1.67\text{ kJ} &= 4.00\text{ g} \times \frac{1\text{ mol}}{58.14\text{ g}} \times H_{vap}\\ H_{vap} &= 24.3\text{ kJ/mol}\end{aligned}$$