How much heat, Q, is required to melt 500 g of ice at 0 °C ?

By definition $Q = mL_f$ where the heat of fusion of ice is $L_{ice} = 3.33 \times 10^5$ J/kg

Therefore, $Q = (0.5 \text{ kg}).\ (3.33 \times 10^5) = 166.5$ kJ

45. You mix 18 kg of water at 25 °C with 6 kg of water at 2 °C, what is the final temperature ?

All we can say is that the hotter water changes temperature by: $\Delta T_{\text{hot}} = T_f - 25°\text{C}$, while the colder water changes temperature by: $\Delta T_{\text{cold}} = T_f - 2.0°\text{C}$. We will be able to solve for T_f using Equation 13.2,

$$\text{i.e. } Q = c.m.\ \Delta T$$

and the fact that the heat lost by the hot water is gained by the cold water: $Q_{\text{hot}} = -Q_{\text{cold}}$, assuming of course that no heat is lost to the surroundings.

Known: $m_{\text{hot}} = 18$ kg, $m_{\text{cold}} = 6$ kg. and c_{water}= 4186 J/(kg.°C) (Table 13.1)

SOLVE The equal but opposite heat exchange implies:

$$Q_{\text{hot}} = -Q_{\text{cold}} \Rightarrow m_{\text{hot}} c \Delta T_{\text{hot}} = -m_{\text{cold}} c \Delta T_{\text{cold}}$$

Solving for the final temperature:

$$T_f - 25°\text{C} = -\frac{6 \text{ kg}}{18 \text{ kg}}(T_f - 2.0°\text{C}) \Rightarrow T_f = 19°\text{C}$$

Sample Problem 18-3

(a) How much heat must be absorbed by ice of mass m = 720 g at T_1 = -10°C to take it to liquid state at T_3 = 15°C?

Let T_2 = 0°C . Then

$Q_{12} = c_{ice} m(T_2 - T_1) = (2{,}220\ \text{J/kg K})(0.72\ \text{kg})[0°\text{C} - (-10°\text{C})]$

$= 15{,}984\ \text{J} = 15.98\ \text{kJ}$

Table 13.3

$Q_F = L_F\, m = (333\ \text{kJ/kg})(0.720\ \text{kg}) = 239.8\ \text{kJ}$

Table 13.1

$Q_{23} = c_w\, m\, (T_3 - T_2) = (4{,}190\ \text{J/kg K})(0.720\ \text{kg})(15°\text{C} - 0°\text{C})$

$= 45{,}252\ \text{J} = 45.25\ \text{kJ}$

$Q = Q_{12} + Q_F + Q_{23} = 15.98\ \text{kJ} + 239.8\ \text{kJ} + 45.25\ \text{kJ} = 300\ \text{kJ}$

PHY 130 Quiz Name:

You have 300 g of coffee at 55 °C. How much 10 °C water do you need to add in order to reduce the coffee's temperature to a more bearable 49 °C ? Note the specific heat, c, of coffee and water are the same (c = 4186 J. kg.$^{-1}$.°C^{-1})

Heat is exchanged between the hot coffee and the cold water, but the whole system does not lose or receive heat. Therefore, $Q_{\text{hot}} + Q_{\text{cold}} = 0$. The temperature of the coffee drops, while that of the added water rises:

$$\Delta T_{\text{hot}} = 49^\circ\text{C} - 55^\circ\text{C} = -6^\circ\text{C}$$
$$\Delta T_{\text{cold}} = 49^\circ\text{C} - 10^\circ\text{C} = 39^\circ\text{C}$$

Known: $m_{\text{hot}} = 300 \text{ g}$

SOLVE The heat exchange is written:

$$Q_{\text{hot}} + Q_{\text{cold}} = m_{\text{hot}} c \Delta T_{\text{hot}} + m_{\text{cold}} c \Delta T_{\text{cold}} = 0$$

The specific heat of coffee is the same as water, so the c's will cancel out of the equation. Solving for the cold water mass:

$$m_{\text{cold}} = -m_{\text{hot}} \frac{\Delta T_{\text{hot}}}{\Delta T_{\text{cold}}} = -(300 \text{ g}) \frac{(-6^\circ\text{C})}{(39^\circ\text{C})} = 46 \text{ g}$$

Sample Problem 18-3 (cont)

(b) If we supply the ice with a total energy of only 210 kJ (as heat), what then are the final state and the temperature of the water?

$Q_{12} = 15.98$ kJ, $Q_F = 239.8$ kJ, $Q_{23} = 45.25$ kJ

Final state: ICE and WATER, $T_f = 0^{\circ}$ C

$Q_{rem} = 210 \text{ kJ} - 15.98 \text{ kJ} = 194 \text{ kJ}$

$Q_{rem} = m_w L_F$

$m_w = Q_{rem} / L_F = 194 \text{ kJ}/333 \text{ kJ/kg} = 0.583 \text{ kg}$

$m_{ice} = 0.720 \text{ kg} - 0.583 \text{ kg} = 0.137 \text{ kg}$