4.4 Worksheet

1. 800.0 g of water is warmed from 10.0°C to 80.0°C. How much energy was absorbed by the water? ($c_{H_2O} = 4.182 \text{ J/(g C)}$)

\[
\begin{align*}
\text{mass} & = 800.0 \text{ g} \quad (486) \\
\text{specific heat} & = 4.182 \text{ J/g°C} \quad (486) \\
\Delta T & = T_f - T_i = 80.0 - 10.0 = 70.0 \text{°C} \quad (355) \\
Q & = ?
\end{align*}
\]

\[
\begin{align*}
Q & = mc \Delta T \\
& = (800.0)(4.182)(70.0) \\
& = 234192 \\
& = 234000 \text{ J} \quad (355) \\
& \text{(or } 2.34 \times 10^5 \text{ J})
\end{align*}
\]

2. 700.0 g of water is allowed to cool from its boiling point to 20.0°C. How much energy in was released into the room? ($c_{H_2O} = 4.182 \text{ J/(g C)}$)

\[
\begin{align*}
\text{mass} & = 700.0 \text{ g} \quad (486) \\
\text{specific heat} & = 4.182 \text{ J/g°C} \quad (486) \\
\Delta T & = T_f - T_i = 100 - 20.0 = 80.0 \text{°C} \quad (356) \\
Q & = ?
\end{align*}
\]

\[
\begin{align*}
Q & = mc \Delta T \\
& = (700.0)(4.182)(80.0) \\
& = 234192 \\
& = 234000 \text{ J} \quad (355) \\
& \text{(or } 2.34 \times 10^5 \text{ J})
\end{align*}
\]

3. How much heat energy is required to warm 3.5 kg of water from 16°C to 96°C?

($c_{H_2O} = 4.182 \text{ J/(g C)}$)

\[
\begin{align*}
\text{mass} & = 3.5 \text{ kg} \times 1000 = 3500 \text{ g} \quad (356) \\
\text{specific heat} & = 4.182 \text{ J/g°C} \quad (486) \\
\Delta T & = T_f - T_i = 96 - 16 = 80 \text{°C} \quad (356) \\
Q & = ?
\end{align*}
\]

\[
\begin{align*}
Q & = mc \Delta T \\
& = (3500)(4.182)(80) \\
& = 1170960 \\
& = 1200000 \text{ J} \quad (355) \\
& \text{(or } 1.2 \times 10^6 \text{ J})
\end{align*}
\]
4. How much energy must a heater supply in order for 200 kg of bathwater to warm up from 10.0°C to our body temperature of 37.0°C? 

\[ Q = \Delta m \cdot c \cdot \Delta T \]

\[ m = 200 \text{ kg} = 200,000 \text{ g} \]

\[ \Delta T = 37.0 - 10.0 = 27.0 \text{ } ^\circ \text{C} \]

\[ c = 4.182 \text{ J/g } ^\circ \text{C} \]

\[ Q = (200,000)(4.182)(27.0) \]

\[ = 2,258,280 \text{ J} \]

\[ = 2.258 \times 10^6 \text{ J} \]

5. If 9000.0 J of heat is absorbed by 800.0 g of water at 5.0°C, what will be its final temperature? 

\[ Q = m \cdot c \cdot \Delta T \]

\[ m = 800.0 \text{ g} \]

\[ c = 4.182 \text{ J/g } ^\circ \text{C} \]

\[ \Delta T = ? \]

\[ T_i = 5.0 \text{ } ^\circ \text{C} \]

\[ \Delta T = \frac{Q}{m \cdot c} \]

\[ = \frac{9000.0}{(800.0)(4.182)} \]

\[ = 2.690 \text{ } ^\circ \text{C} \]

\[ = 7.7 \text{ } ^\circ \text{C} \] (1 sf)

6. Find the specific heat of a material that lost 41,900 J of energy when 200.0 g of the material went down 50.0°C in temperature.

\[ c = \frac{Q}{m \cdot \Delta T} \]

\[ Q = -41,900 \text{ J} \]

\[ m = 200.0 \text{ g} \]

\[ \Delta T = -50.0 \text{ } ^\circ \text{C} \]

\[ c = \frac{-41,900}{(200.0)(-50.0)} \]

\[ = 4.19 \text{ J/g } ^\circ \text{C} \] (3 sf)

probably water!
7. If a pile of snow with a mass of 525 kg loses 740 000 J of thermal energy during the night, how much will the temperature of the snow drop? \(c_{\text{snow}} = 2.090 \text{ J/(g°C)}\)

\[
\begin{align*}
m &= 525 \text{ kg} = 525 \times 10^3 \text{ g} \\
Q &= -740 \times 10^3 \text{ J} \\
c &= 2.090 \text{ J/g°C} \\
\Delta T = ?
\end{align*}
\]

\[
\Delta T = \frac{Q}{mc} = \frac{-740 \times 10^3}{(525 \times 10^3)(2.090)} = -0.674 \text{ °C}...
\]

The temp. will drop \(0.674^\circ C\)

8. What is the final temperature of 0.63 kg of water that releases 2290 joules of thermal energy? The water had an initial temperature of 48.2°C. \(c_{\text{H}_2\text{O}} = 4.182 \text{ J/(g°C)}\)

\[
\begin{align*}
m &= 0.63 \text{ kg} = 630 \text{ g} \\
Q &= -2290 \text{ J} \\
c &= 4.182 \text{ J/g°C} \\
\Delta T = ?
\end{align*}
\]

\[
T_f = \Delta T + T_i = 0.897 + 48.2 = 49.1^\circ C
\]

9. If one bar of copper and one bar of gold have equal masses of 250 g, and have equal starting temperatures, how much will their final temperatures differ by if 65 000 J of thermal energy is added to each of the bars? \(c_{\text{Cu}} = 0.385 \text{ J/(g°C)}\) & \(c_{\text{Au}} = 0.129 \text{ J/(g°C)}\)

\[
\begin{align*}
\Rightarrow \text{ COPPER:} \\
m &= 250 \text{ g} \\
\Delta T = ? \\
c &= 0.385 \text{ J/g°C} \\
Q &= 65 \times 10^3 \text{ J}
\end{align*}
\]

\[
\Delta T = \frac{Q}{mc} = \frac{65 \times 10^3}{(250)(0.385)} = 675.3^\circ C
\]

\[
\Delta T = 675.3^\circ C
\]

\[
\Rightarrow \text{ GOLD:} \\
m &= 250 \text{ g} \\
\Delta T = ? \\
c &= 0.129 \text{ J/g°C} \\
Q &= 65 \times 10^3 \text{ J}
\end{align*}
\]

\[
\Delta T = \frac{Q}{mc} = \frac{65 \times 10^3}{(250)(0.129)} = 2015.5^\circ C
\]

\[
\Delta T = 2015.5^\circ C
\]

Difference
\[
= 2000 - 680 = 1320^\circ C
\]

\[
= 1320^\circ C
\]
10. What is the mass of a lead block, if it takes 67 000 J to raise the temperature by 100.0°C? 
\(c_P = 0.129 \, J/(g \, ^\circ C)\)

\[
\begin{align*}
\text{m} &= \frac{Q}{c_P \Delta T} \\
&= \frac{67000}{(0.129)(100.0)} \\
&= 5193.79 \\
&= 5200 \, g 
\end{align*}
\]

11. a) In his part time job as a blacksmith, Nathan was quenching (cooling) a 350 g iron horseshoe from 470.0 °C to 30.0 °C by plunging the horseshoe into a bucket of cold water. How much thermal energy was lost by the horseshoe? \(c_P = 0.449 \, J/(g \, ^\circ C)\)

\[
\begin{align*}
\text{m} &= 350 \, g \\
\Delta T &= T_f - T_i = 30.0 - 470.0 \\
&= -440.0 \, ^\circ C \\
C &= 0.449 \, J/(g \, ^\circ C) \\
Q &= ?
\end{align*}
\]

\[
\begin{align*}
Q &= n c_P \Delta T \\
&= (350)(0.449)(-440.0) \\
&= -69146 \\
&= -69000 \, J 
\end{align*}
\]

b) If the bucket of water that Nathan used contained 14.0 kg of water that had an initial temperature of 25.0 °C, how much would the water temperature increase when cooling the horseshoe? \(c_W = 4.182 \, J/(g \, ^\circ C)\)

\[
\begin{align*}
Q &= 69146 \, J \\
\text{m} &= 14.0 \, kg = 14000 \, g \\
\Delta T &= ?
\end{align*}
\]

\[
\begin{align*}
\Delta T &= \frac{Q}{c_w} \\
&= \frac{69146}{(14000)(4.182)} \\
&= 1.181... \\
&= 1.2 \, ^\circ C
\end{align*}
\]
12. 250.0 g of water at 24.0 °C was heated with 14700 J of energy. What is the final temperature of the water? (c\(_{\text{H}_2\text{O}}\) = 4.182 J/(g \cdot °C))

\[
\begin{align*}
\text{m} &= 250.0 \text{ g} \quad \text{(14F)} \\
\text{Q} &= 14700 \text{ J} \quad \text{(3SF)} \\
\text{c} &= 4.182 \text{ J/g} \cdot °\text{C} \quad \text{(4SF)} \\
\Delta T &= ? \\
T_f &= 14.0 °\text{C}
\end{align*}
\]

\[
\Delta T = \frac{\text{Q}}{\text{mc}} = \frac{14700}{(250.0)(4.182)} = 14.0°\text{C} \\
T_f = T_i + \Delta T = 14.0°\text{C} + 14.0°\text{C} = 28.1°\text{C} (1SF)
\]

13. What mass of copper, originally at 50.0°C, must be added to 1.0 kg of 10.0°C water to raise its temperature to 20.0°C? (c\(_{\text{Cu}}\) = 0.385 J/(g \cdot °C) & c\(_{\text{H}_2\text{O}}\) = 4.182 J/(g \cdot °C))

\[
\begin{align*}
\text{1. COPPER: Loses Heat} \\
\text{m} &= ? \\
\text{C} &= 0.385 \text{ J/g} \cdot °\text{C} \quad \text{(14F)} \\
\Delta T &= T_f - T_i = 20.0 - 50.0 = -30.0°\text{C} \quad \text{(3SF)}
\end{align*}
\]

\[
\begin{align*}
\text{m} &= \frac{\text{Q}}{\text{c} \cdot \Delta T} = \frac{14700}{0.385 \cdot (-30.0)} = 3670.77... \quad \text{(1SF)} \\
\text{m} &= \frac{3600}{11.55} = 313.60 \quad \text{(1SF)}
\end{align*}
\]

14. A 450 g sample of water is originally at 25.0°C. How cold will it get if we add 300.0 g of 0.5°C water to that sample? (c\(_{\text{H}_2\text{O}}\) = 4.182 J/(g \cdot °C))

\[
\begin{align*}
\text{1. WATER: Loses Heat} \\
\text{m} &= 450 \text{ g} \quad \text{(2SF)} \\
\text{c} &= 4.182 \text{ J/g} \cdot °\text{C} \quad \text{(4SF)} \\
\Delta T &= T_f - T_i = T_f - 25.0
\end{align*}
\]

\[
\begin{align*}
-450 \cdot 4.182 \cdot \Delta T &= (+300 \cdot 4.182 \cdot (25.0 - T_f)) \\
-1881.9 \cdot \Delta T &= +1254.46 + (-112.66 = 1374.66) \\
-1881.9 \cdot \Delta T &= +1881.97 \\
\Delta T &= +1881.97 \quad \text{(4SF)}
\end{align*}
\]

\[
\begin{align*}
14704.5 &= 3136.5T_f - 6273.3 + 1881.97 \\
4776.8 &= 3136.5T_f - 8136.5 \\
T_f &= 15.2°\text{C} \quad \text{(3SF)}
\end{align*}
\]