Clicker Questions: Monday, March 25

1) How many kcal (to 3 sig figs) are required to raise the temperature of 35.0 mL of alcohol from 23.0 °C to 45.0 °C? Density of alcohol = 0.789 g/mL and its specific heat capacity = 2.14 J/g°C.
   A) 59.6 kcal  
   B) 1.30 x 10³ kcal  
   C) 0.311 kcal  
   D) 311 kcal  
   E) 0.0596 kcal  
   F) 3.30 x 10³ kcal

2) If we burn 1 packet of oatmeal in a bomb calorimeter containing 5.00 kg water, and the temperature of the water increases from 23.0°C to 42.5°C, how many nutritional Calories does the packet of oatmeal contain?
   A) 104 cal  
   B) 104 Cal  
   C) 408 cal  
   D) 408 Cal  
   E) 9.74 x 10⁴ cal  
   F) 9.74 x 10⁴ Cal  
   G) 97.4 cal  
   H) 97.4 Cal

3) While hiking in the Sierra, you find a shiny piece of metal weighing 415 g. You decide to determine the specific heat of the metal to see if it might be gold. You heat the metal to 164 °C and drop it in 200.0 g of water at 22.0 °C. The final temperature of the water and gold is 46.2 °C. What is the heat capacity of the metal?
   A) \( C_{\text{metal}} = 0.128 \text{ J/g °C} \) 
   B) \( C_{\text{metal}} = 0.414 \text{ J/g °C} \) 
   C) \( C_{\text{metal}} = 0.258 \text{ J/g °C} \) 
   D) \( C_{\text{metal}} = 0.195 \text{ J/g °C} \)

4) A 20.0 g sample of copper (specific heat = 0.384 J/g °C) is heated to 203 °C and dropped into 80.0 g of water (specific heat = 4.18 J/g °C) at 25.0 °C. What is the final temperature of the water (to 3 sig figs)? [Hint: both the copper and the water end up at the same final temperature, so \( T_f \) is the same variable on both sides of the equation.]
   A) 19.9 °C  
   B) 29.0 °C  
   C) 20.9 °C  
   D) 43.5 °C  
   E) 155 °C  
   F) 30.4 °C

Answers: 1) C*  
         2) H*  
         3) B*  
         4) B*  
* See answer worked out on next page
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Selected answers with work shown

1) How many kcal (to 3 sig figs) are required to raise the temperature of 35.0 mL of alcohol from 73.4 °F to 113 °F? Density of alcohol = 0.789 g/mL and its specific heat capacity = 2.14 J/g °C.

Answer: 

\[ q = m \cdot C \cdot \Delta T = m \cdot C \cdot (T_f - T_i) \]

collect the values to plug into our equation and put convert them to the right units:

\[ C = 2.14 \text{ J/g °C} \]
\[ m = (35.0 \text{ mL}) \frac{0.789 \text{ g}}{1 \text{ mL}} = 27.62 \text{ g} \]
\[ \Delta T = 45.0 - 23.0 = 22.0 \text{ °C} \]

perform the calculation:

\[ q = (27.62 \text{ g})(2.14 \text{ J/g °C})(22.0 \text{ °C}) = 1300.3496 \text{ J} \]

\[ = (1300.3496 \text{ J}) \cdot \frac{1 \text{ cal}}{4.184 \text{ J}} \cdot \frac{1 \text{ kcal}}{1000 \text{ cal}} = 0.310791013 \text{ kcal} = 0.311 \text{ kcal} \]

2) If we burn 1 packet of oatmeal in a bomb calorimeter containing 5.00 kg water, and the temperature of the water increases from 23.0°C to 42.5°C, how many nutritional Calories does the packet of oatmeal contain?

Answer:

\[ q_{\text{water}} = (m_{\text{water}})(C_{\text{water}})(\Delta T_{\text{water}}) = (m_{\text{water}})(C_{\text{water}})(T_f - T_i) \]
\[ = (5.00 \times 10^3 \text{ g})(4.18 \text{ J/g °C})(42.5 \text{ °C} - 23.0 \text{ °C}) \]
\[ = 407550 \text{ J} \]

Convert answer to Calories:

\[ q_{\text{water}} = (407550 \text{ J})(1 \text{ cal}/4.184 \text{ J})(1 \text{ Cal}/1000 \text{ cal}) = 97.4 \text{ Cal} \]

This is the heat the water absorbed, so it must be the amount of heat given off by the oatmeal.

3) While hiking in the Sierra, you find a shiny piece of metal weighing 415 g. You decide to determine the specific heat of the metal to see if it might be gold. You heat the metal to 164 °C and drop it in 200.0 g of water at 22.0 °C. The final temperature of the water and gold is 46.2 °C. What is the heat capacity of the metal?

Answer:

\[ q_{\text{metal}} = -q_{\text{water}} \]
\[ (m_{\text{metal}})(C_{\text{metal}})(T_f, \text{ metal} - T_i, \text{ metal}) = - (m_{\text{water}})(C_{\text{water}})(T_f, \text{ water} - T_i, \text{ water}) \]
\[ (415 \text{ g})(C_{\text{metal}})(46.2 \text{ °C} - 164 \text{ °C}) = - (200.0 \text{ g})(4.18 \text{ J/g °C})(46.2 \text{ °C} - 22.0 \text{ °C}) \]
\[ (415 \text{ g})(C_{\text{metal}})(-117.8 \text{ °C}) = - (200.0 \text{ g})(4.18 \text{ J/g °C})(24.2 \text{ °C}) \]
\[ (-48887 \text{ g °C})(C_{\text{metal}}) = -20231.2 \text{ J} \]

\[ C_{\text{metal}} = 0.414 \text{ J/g °C} \]
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4) A 20.0 g sample of copper (specific heat = 0.384 J/g °C) is heated to 203 °C and dropped into 80.0 g of water (specific heat = 4.18 J/g °C) at 25.0 °C. What is the final temperature of the water (to 3 sig figs)?

**Answer:**

\[
q_{\text{copper}} = -q_{\text{water}}
\]

\[
(m_{\text{Cu}})(C_{\text{Cu}})(T_f, \text{Cu} - T_i, \text{Cu}) = - (m_{\text{water}})(C_{\text{water}})(T_f, \text{water} - T_i, \text{water})
\]

\[
(20.0 \text{ g})(0.384 \text{ J/g °C})(T_f, \text{copper} - 203^\circ \text{C}) = - (80.0 \text{ g})(4.18 \text{ J/g °C})(T_f, \text{water} - 25.0^\circ \text{C})
\]

This would leave us with two variables, but we realize that \(T_f\) is the same for both:

\[
(20.0 \text{ g})(0.384 \text{ J/g °C})(T_f - 203^\circ \text{C}) = - (80.0 \text{ g})(4.18 \text{ J/g °C})(T_f - 25.0^\circ \text{C})
\]

\[
(7.68 \text{ J/°C})(T_f - 203^\circ \text{C}) = - (334.4 \text{ J/°C})(T_f - 25.0^\circ \text{C})
\]

Distribute/factor through the parenthesis. Be really careful with the negative sign!

\[
(7.68 \text{ J/°C})(T_f) - (1559.04 \text{ J}) = - (334.4 \text{ J/°C})(T_f) + (8360 \text{ J})
\]

Group like terms. Again, be really careful with negative signs!

\[
(342.08 \text{ J/°C})(T_f) = (9919.04 \text{ J})
\]

Solving for \(T_f\) gives:

\[
T_f = 28.996^\circ \text{C} = 29.0^\circ \text{C}
\]

➢ Does it make sense that the temp of the water changes so little? Yes, it makes sense; there is more of the water (4x mass of copper) and the water has a huge specific heat (almost 10x that of copper) so the water will resist changing temperature.