AP Worksheet 6a (Calorimetry)

1. A 466-gram sample of water is heated from 8.50 °C to 74.60 °C. Calculate the amount of energy absorbed by the water.

$q = mc\Delta T$

$= (466\,g)(4.184\,J/g{\cdot}°C)(74.60°C - 8.50°C)$

$= 1.29 \cdot 10^5\,J = 1.29 \cdot 10^2\,kJ = 129\,kJ$ ✓

2. The specific heat of aluminum is 0.215 cal/g°C. How many calories are given off by an aluminum bar that has a mass of 200.0 grams and the temperature change of the block is 35°C?

$q = mc\Delta T$

$= (200.0\,g)(.215\,cal/g°C)(35°C)$

$= 1500\,J = 1.5\,kJ$

3. The specific heat of gold is 0.031cal/g°C. What is the mass of the water used if a gold ring with a mass of 250 grams raises the temperature of the water from 10°C to 20°C, and the temperature of the ring goes from 150°C to 20°C?

$-q_{Au} = q_{H_2O}$

$-m_{Au}C_{Au}\Delta T_{Au} = m_{H_2O}C_{H_2O}\Delta T_{H_2O}$

$-(250\,g)(.031\,cal/g°C)(20°C - 150°C) = (m_{H_2O})(1.00\,cal/g°C)(20°C - 10°C)$

$1007.5\,cal = (10.)(m_{H_2O})$ $\quad m_{H_2O} = 1.0 \cdot 10^2\,g$

4. How many Joules of heat is needed to increase the temperature of 750 ml of water by 67°C?

$750\,mL\,H_2O \left| \frac{1.00\,g\,H_2O}{1\,mL} \right. = 750\,g$

$q = mc\Delta T$

$= (750\,g)(4.184\,J/g°C)(67°C)$

$= 2.1 \cdot 10^5\,J$

5. If 95 grams a of a metal at a temperature of 26.8 °C is placed in 42.2 grams of water at a temperature of 9.5°C and the final temperature of the system is 12.5°C, what is the specific heat of the sample in J/g°C?

$-q_m = q_{H_2O}$

$-m_m C_m \Delta T_m = m_{H_2O} C_{H_2O} \Delta T_{H_2O}$

$-(95\,g)(C_m)(12.5°C - 26.8°C) = (42.2\,g)(4.184\,J/g°C)(12.5°C - 9.5°C)$

$1358.5(C_m) = 529.7$

$C_m = .39\,J/g°C$

6. Using the data below calculate the specific heat of an unknown metal

Mass of unknown metal m_m	65.0 grams
Start temperature of metal T_{i_m}	96.3°C
Mass of water M_{H_2O}	150 grams
Start temp of water $T_{i\,H_2O}$	23.4°C
End temp of water T_f	26.3°C

$-q_m = q_{H_2O}$

$-m_m C_m \Delta T_m = M_{H_2O} C_{H_2O} \Delta T_{H_2O}$

$-(65.0\,g)(C_m)(26.3°C - 96.3°C) = (150\,g)(4.184\,J/g°C)(26.3°C - 23.4°C)$

$(C_m)(4550) = 1820$

$C_m = .40\,J/g°C$

7. Suppose a Cheeto is burned underneath a soda can containing 200. milliliters of 21.6 °C distilled water. The heat released from the burning of the Cheeto raises the temperature of the water by 24.8 °C.

$\frac{200.\,mL}{1\,mL} \left| \frac{1.00\,g}{} \right.$

a. Calculate the amount of energy released by the burning of the Cheeto.

$-q_{cheeto} = q_{H_2O}$

$= mC\Delta T = (200.\,g)(4.184\,J/g°C)(24.8°C - 21.6°C)$

$= 2680\,J$

b. How many calories of energy did the Cheeto contain?

$\frac{2680\,J}{4.184\,J} \left| \frac{1\,cal}{} \right. = 640.\,cal = .640\,Kcal\ (Cal)$

c. How many Calories are in a bag of Cheetos containing 42 pieces?

1 calorie = 4.18 Joules

1 nutritional Calorie = 1 kilocalorie = 1000 calories

$\frac{640\,cal}{cheeto} \left| \frac{42\,cheeto}{1\,bag} \right| \frac{1\,Cal}{1000\,cal} = 26.9\,Cal$

d. Calculate the percent error in this experiment. (Note: the bag of Cheetos used in this experiment actually contained 300. Calories.)

$\%\ error = \frac{|300.\,Cal - 26.9\,Cal|}{300.\,Cal} = 91.0\%$

e. Why do you think this experiment had such a large percent error?

heat absorbed by can

heat released into air, not just used to heat H_2O

did cheeto burn completely?