Calculations Practice

16.1

1. A fast food item contains 544 nutritional Calories. Convert this energy to calories and Joules.

$544 \text{ Calories} \times \frac{1000 \text{ calories}}{1 \text{ Calories}} = 540{,}000$ calories

$544 \text{ Calories} \times \frac{1000 \text{ calories}}{1 \text{ calorie}} \times \frac{4.184}{1 \text{ calorie}} =$ ~~2,259,360 Joules~~ 2276096 Joules

2. An endothermic process absorbs 138 kJ. How many calories of heat are absorbed?

$138 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ calorie}}{4.184 \text{ J}} = 32982.79$ calories

16.2

1. How many Joules of heat are lost by 3580 kg granite as it cools from 41.2°C to -12.9°C? c =.803J/(g-C)

3580,000

$q = cm\Delta T$

negative!

$q = 0.803 \frac{J}{gC} \times 3580{,}000 g \times -54.1°C = -1{,}55{,}523{,}434$ Joules

2. How much heat is absorbed by a 2,000.kg granite boulder as energy from the sun causes its temperature to change from 10°C to 29°C? c =.803J/(g-C)

$q = cm\Delta T$

$q = 0.803 \frac{J}{g \cdot C} \times 2{,}000{,}000 g \times 19°C = +30{,}514{,}000 J$

3. A swimming pool, 20.0 m x 12.5 m is filled with water to a depth of 3.75 m. If the initial temperature of the water is 18.4°C, how much heat must be added to the water to raise its temperature to 29.0°C? Assume that the density of water is 1.000 g/mL. c = 4.184 J/(g-C)

$937.5 m^3 \times \frac{100 cm^3}{1 m^3}$

Volume = $937.5 m^3$

$D = 1 g/mL = \frac{m}{937.5 m^3} = \frac{m}{937500000}$

$m = 937{,}500{,}000 g$

$q = cm\Delta T$

$q = 4.184 \frac{J}{g \cdot C} \times 937{,}500{,}000 g \times 10.6°C$

$q = 4.16 \times 10^{10}$ Joules

16.3

1. How much heat is required to vaporize 343 g of liquid ethanol, C_2H_5OH, at its boiling point? H_{vap}=38.6 kJ/mol.

$q = H_{vap} \times Mol$

$343 g \ C_2H_5OH \times \frac{1 \text{ mol}}{46.08 g \ C_2H_5OH}$

$7.44 \text{ mol } C_2H_5OH$

$q = 38.6 \frac{kJ}{mol} \times 7.44 Mol$

$q = 287 kJ$

16.3 continued

2. How much heat is evolved when 1255 g of water, H_2O, condenses to a liquid at 100°C? H_{cond} = -40.7 kJ/mol

$q = H_{cond} \times mol$ $\quad 1255g\ H_2O \left| \frac{1\ mol}{18.02g} \right. = 69.64\ mol$

$q = -40.7 \frac{kJ}{mol} \times 69.64\ mol = -2835\ kJ$

3. How much heat is liberated by the combustion of 206 g of hydrogen, H_2? H_{comb} = -286 kJ/mol

$q = H_{comb} \times mol$ $\quad 206g\ H_2 \left| \frac{1\ mol}{2.02g} \right. = 102\ mol\ H_2$

$q = -286 \frac{kJ}{mol} \times 102\ mol\ H_2 = 29166\ kJ$

4. How much heat is required to warm 225g (12.5 mol) of ice from -46.8°C to 0.0°C (1), melt the ice (2), warm the water from 0.0°C to 100.0°C (3), boil the water (4), and heat the steam to 173.0°C (5)?

$H_f = 40.7\ kJ/mol$
$H_v = 6.01\ kJ/mol$
$C = 4.184\ (l)$
$C = 2.03\ (s)$
$C = 2.01\ (g)$

(1) $q = cm\Delta T \quad (2.03)(225g)(46.8) = \underline{21375.9\ J}$

(2) $= H_f \times mol \quad 40.7\ kJ/mol\ (12.5mol) = 508.75\ kJ = \underline{508750\ J}$

(3) $q = cm\Delta T \quad (4.184)(225g)(100) = \underline{94,140\ J}$

(4) $H_{vap} \times mol \quad (6.01)(12.5mol) = 75.125\ kJ = \underline{75125\ J}$

(5) $q = cm\Delta T \quad (2.01)(225)(73) = \underline{33014.25\ J}$

Total = 732,405.15 J

16.4

1. You are given these two equations.

a. $Sn(s) + Cl_2(g) \rightarrow SnCl_2(s)$ H = -325 kJ

b. $SnCl_2(s) + Cl_2(g) \rightarrow SnCl_4(l)$ H=-186 kJ

Calculate the change in enthalpy for the reaction:

$Sn(s) + 2Cl_2(g) \rightarrow SnCl_4(l)$

$Sn + Cl_2 + \cancel{SnCl_2} + Cl_2 \longrightarrow \cancel{SnCl_2} + SnCl_4$

-325
$+ -186$
$-511\ kJ$