Prac. Quiz

AP Chemistry
Chapter 6: Thermochemistry

Name Key

Solve the following problems showing all work and formulas.

1. If ΔU for a system is 217 J in a process in which the system absorbs 185 calories of heat, how much work, in joules must have been involved?

$$\Delta U = q + w$$

$$217\text{ J} = \left(\frac{185\text{ cal} \mid 4.18\text{ J}}{\mid \text{cal}}\right) + w$$

$$w = \boxed{-556\text{ J}}$$

* System does work on surrounding

2. Combustion of 0.144 g of sucrose (table sugar, $C_{12}H_{22}O_{11}$) in the open air results in the release of 2.38 kJ of heat.

$$C_{12}H_{22}O_{11}(s) + 12\,O_2\,(g) \longrightarrow 12\,CO_2\,(g) + 11\,H_2O\,(l)$$

Is the reaction endothermic or exothermic? releases heat
What is the value of q for the reaction, per mole of sucrose?
Is this value equal to ΔU or ΔH? Explain.

$$q_{rxn} = \frac{2.38\text{ kJ} \mid 342.3\text{ g}}{0.144\text{ g} \mid \text{mol}} = \boxed{5.65 \times 10^3\ \frac{\text{kJ}}{\text{mol}}}$$

This value = ΔH

$q_p = \Delta H_{rxn}$ (at constant pressure)

(If constant vol., then ΔU)

3. Calcium carbide (CaC_2) reacts with water to form acetylene (C_2H_2), a gas used as a fuel in welding.

L.R?

$$\underset{3.5\text{ kg}}{CaC_2\,(s)} + \underset{\substack{1.25\text{ L} \\ \text{L.R}}}{2\,H_2O\,(l)} \longrightarrow C_2H_2\,(g) + Ca(OH)_2\,(s) \qquad \Delta H = -128.0\text{ kJ}$$

How many kilojoules of heat are evolved in the reaction of 3.50 kg CaC_2 with 1.25 L H_2O?

$$\frac{3500\text{ g } CaC_2 \mid \text{mol } CaC_2 \mid 1\text{ mol } C_2H_2}{\mid 64.12\text{ g } CaC_2 \mid 1\text{ mol } CaC_2} = 54.6\text{ mol } C_2H_2$$

$$\frac{1.25\text{ L } H_2O \mid 1000\text{ mL} \mid 1.00\text{ g} \mid \text{mol } H_2O \mid 1\text{ mol } C_2H_2}{\mid \text{L} \mid \text{mL} \mid 18.02\text{ g } H_2O \mid 2\text{ mol } H_2O} = 34.7\text{ mol } C_2H_2$$

$$\Delta H = \frac{-128.0\text{ kJ} \mid 34.7\text{ mol } C_2H_2}{\text{mol } C_2H_2 \mid} = \boxed{-4.44 \times 10^3\text{ kJ}}$$

4. How much heat, in kilojoules, is released when the temperature of 47.0 g water from 45.4°C to 10.0°C?

$$q = m \Delta T C_p$$

$$= \frac{47.0\,g}{1} \left| \frac{(10.0-45.4)^\circ C}{} \right| \frac{4.18\,J}{g^\circ C} \left| \frac{1\,kJ}{1000\,J} \right. = \boxed{-6.95\,kJ}$$

5. A 454 gram iron block initially at 16°C absorbs 63.9 kJ of heat. What is the final temperature of the iron?

$C_p = .449 \frac{J}{g^\circ C}$

$$q = m \Delta T C_p$$

$$= \frac{454\,g}{} \left| \frac{X}{} \right| \frac{.449\,J}{g^\circ C} = 63900\,J$$

$$\Delta T = 313^\circ C = T_f - 16^\circ C$$

$$T_f = \boxed{329^\circ C}$$

6. A 65.0 mL sample of 0.600 *M* HI at 18.46°C is mixed with 84.0 mL of a solution containing excess potassium hydroxide, at 18.46°C in a Styrofoam® cup calorimeter. The reaction

0.600M excess

HI (*aq*) + KOH (*aq*) ----------> KI (*aq*) + H_2O (*l*) ΔH = ?

65.0mL 84.0mL

18.46°C → 21.96°C

takes place, and the temperature rises to 21.96°C. Calculate ΔH for the reaction.

Assume:

1. the solution volumes are additive
2. the solution formed is dilute enough that the density and specific heat are about the same as pure water.
3. the system is completely isolated, no heat escapes the calorimeter
4. the heat required to warm any part of the calorimeter other than the final solution is negligible.

$$\Delta T = 21.96 - 18.46 = 3.50^\circ C$$

$$q_{rxn} = -q_{calorimeter\,(H_2O)}$$

$$q_{cal\,(H_2O)} = \frac{149\,mL}{} \left| \frac{1.00\,g}{mL} \right| \frac{3.50^\circ C}{} \left| \frac{4.18\,J}{g^\circ C} \right. = 2180\,J$$

$$mol\,H_2O = \frac{0.600\,mol\,HI}{L\,HI} \left| \frac{.065\,L}{} \right| \frac{1\,mol\,H_2O}{1\,mol\,HI} = 0.039\,mol\,H_2O$$

$$q_{rxn} = \frac{-2180\,J}{0.039\,mol\,H_2O} \left| \frac{1\,kJ}{1000\,J} \right. = \boxed{-55.9\,\frac{kJ}{mol}}$$

7. Use the following equations

$$\checkmark\ N_2H_4\ (l) + O_2\ (g) \text{ ----------> } N_2\ (g) + 2\ H_2O\ (l) \qquad \Delta H = -622.2\text{ kJ}$$

$$2\left(H_2\ (g) + 1/2\ O_2\ (g) \text{ ---------> } H_2O\ (l)\right) \qquad \Delta H = -285.8\text{ kJ}$$

$$-2\left(H_2\ (g) + O_2\ (g) \text{ ---------> } H_2O_2\ (l)\right) \qquad \Delta H = -187.8\text{ kJ}$$

to calculate the enthalpy change for the reaction

$$N_2H_4\ (l) + 2\ H_2O_2\ (l) \text{ ----------> } N_2\ (g) + 4\ H_2O\ (l) \qquad \Delta H = ?$$

$$\checkmark\ N_2H_4 + \cancel{O_2} + 2\cancel{H_2} + \cancel{O_2} + 2H_2O_2 \rightarrow N_2 + 2H_2O + 2H_2O + 2\cancel{H_2} + 2\cancel{O_2}$$

$$\Delta H_{rxn} = -622.2 + 2(-285.8) + 2(187.8) = \boxed{-818.2\text{ kJ}}$$

8. Use the standard enthalpies of formation to calculate the standard enthalpy change for the following reaction.

$$2\ C_2H_5OH\ (l) + O_2\ (g) \text{ ----------> } 2\ CH_3CHO\ (g) + 2\ H_2O\ (l)$$

ΔH^o_f
$C_2H_5OH\ (l) = -277.7$ kJ
$CH_3CHO(g) = -166.1$ kJ
$H_2O\ (l) = -285.8$ kJ

$$\Delta H_{rxn} = \Sigma \Delta H_P - \Sigma \Delta H_R$$

$$\Delta H_{rxn} = [2(-166.1) + 2(-285.8)] - [2(-277.7) + 0]$$

$$= \boxed{-348.4\text{ kJ}}$$