

CHEM 1035 Homework #7

6.18 This reaction is exothermic, the sign on  $q$  is (-) negative because heat is removed from the system and goes to the surroundings.

6.29 Ethanol should have greater heat of combustion per mole because it has more bonds (one additional C-C, and 2 additional C-H bonds)

$$6.36 \quad q = (c)(\text{mass})\Delta T = \left( \frac{2.087 J}{g \cdot K} \right) (0.10 g) (-75 - 10) = -17.74 J$$

$$6.38 \quad q = -688 J, \text{ mass} = 27.7 g, c = 2.42 J/g \cdot K$$

$$-688 J = \left( 2.42 \frac{J}{g \cdot K} \right) (27.7 g) (32.5 - T_i)$$

$$T_i = 42.7^\circ C$$

$$6.44 \quad q_{\text{alloy}} = -q_{\text{calorimeter}}$$

$$c_{\text{alloy}} (30.5 g) (31.1 - 93.0) = -1 \left[ \left( 4.184 \frac{J}{g \cdot K} \right) (50.0) (31.1 - 22) + \left( 9.2 \frac{J}{K} \right) (31.1 - 22) \right]$$

$$-1888 c_{\text{alloy}} = -1987$$

$$c_{\text{alloy}} = 1.05 \frac{J}{g \cdot K}$$

$$6.46 \quad q_{\text{hydrocarbon}} = -q_{\text{calorimeter}}$$

$$q_{\text{hydrocarbon}} = -1 \left[ \left( 4.184 \frac{J}{g \cdot K} \right) \left( 2.5 L \times \frac{1000 g H_2O}{L} \right) (23.55 - 20.00) + \left( 403 \frac{J}{K} \right) (23.55 - 20.00) \right]$$

$$q_{\text{hydrocarbon}} = -38530 J = -38.53 kJ$$

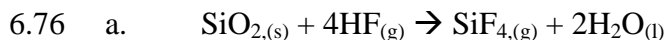
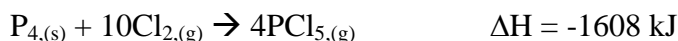
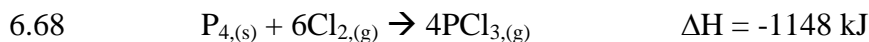
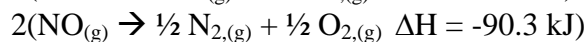
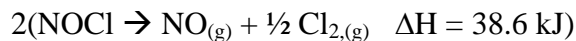
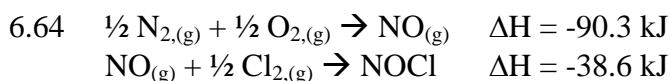
$$q_{\text{hydrocarbon}} / gm = -38.53 kJ / 1.500 g = -25.71 kJ / gm$$

$$6.56 \quad a. \quad 100. \text{ gm Fe} \times \frac{1 \text{ mole Fe}}{55.85 \text{ gm}} \times \frac{-1.65 \times 10^3 \text{ kJ}}{4 \text{ mole Fe}} = -739 \text{ kJ}$$

$$b. \quad -4.93 \times 10^3 \text{ kJ} \times \frac{2 \text{ mole Fe}_2\text{O}_3}{-1.65 \times 10^3 \text{ kJ}} \times \frac{159.7 \text{ gm Fe}_2\text{O}_3}{1 \text{ mole Fe}_2\text{O}_3} = 954 \text{ gm}$$



$$b. \quad 1.0 \text{ gm C}_{12}\text{H}_{22}\text{O}_{11} \times \frac{1 \text{ mole C}_{12}\text{H}_{22}\text{O}_{11}}{342.30 \text{ gm}} \times \frac{-5.64 \times 10^3 \text{ kJ}}{1 \text{ mole C}_{12}\text{H}_{22}\text{O}_{11}} = -16.48 \text{ kJ}$$

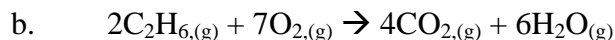


$$\Delta H_{\text{rxn}}^{\circ} = \sum \Delta H_f^{\circ}(\text{prod}) - \sum \Delta H_f^{\circ}(\text{react})$$

$$\Delta H_{\text{rxn}}^{\circ} = \{ [(-1614.9 \text{ kJ}) + 2(-285.84 \text{ kJ})] - [(-910.9 \text{ kJ}) + 4(-273 \text{ kJ})] \}$$

$$= -2186.6 - (-2002.9)$$

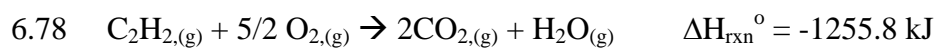
$$\Delta H_{\text{rxn}}^{\circ} = -183.7 \text{ kJ}$$



$$\Delta H_{\text{rxn}}^{\circ} = \{ [4(-393.5 \text{ kJ}) + 6(-241.83 \text{ kJ})] - [2(-84.67 \text{ kJ}) + 7(0)] \}$$

$$= -3025 - (-169.34)$$

$$\Delta H_{\text{rxn}}^{\circ} = -2856 \text{ kJ}$$



$$-1255.8 \text{ kJ} = \{ [2(-393.5 \text{ kJ}) + (-241.83 \text{ kJ})] - [(5/2)(0) + \Delta H_f^\circ(\text{C}_2\text{H}_2)] \}$$

$$-1255.8 \text{ kJ} = -1028.8 - \Delta H_f^\circ(\text{C}_2\text{H}_2)$$

$$\Delta H_f^\circ(\text{C}_2\text{H}_2) = -226.97 \text{ kJ}$$