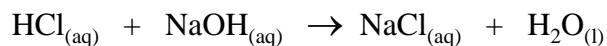# Thermochemistry II: Calorimetry, Enthalpy, and Hess' Law

1. When 100.0 mL of 1.00 M HCl is mixed with 100.0 mL of 1.00 M NaOH, both initially at 21.1°C, are mixed in a two-cup calorimeter the temperature of the mixture rises to 27.9°C. Determine the ΔH of neutralization for the reaction

$$HCl_{(aq)} \ + \ NaOH_{(aq)} \ \rightarrow \ NaCl_{(aq)} \ + \ H_2O_{(l)}$$

By a prior experiment, the heat capacity of the calorimeter was determined to be 125 J/°C. Assume the density of the final solution is 1.0 g/mL and the specific heat of the mixture is 4.18 J/g°C.

$$\Delta T = 27.9°C - 21.1°C = 6.8°C$$

$$n_{HCl} = n_{NaOH} = 0.1000 \text{ L} \times 1.00 \text{ M} = 0.100 \text{ mol}$$

$$C_{cal} = 125 \tfrac{J}{°C}$$

$$n\Delta H + mc\Delta T + C_{cal}\Delta T = 0$$

$$\Delta H = -\frac{(mc + C_{cal})\Delta T}{n} = \frac{\left[(200.0 \text{ g})(4.18\tfrac{J}{g\cdot°C}) + 125\tfrac{J}{°C}\right](6.8°C)}{0.100 \text{ mol}}$$

$$\Delta H = -65,350 \tfrac{J}{mol}$$

2. Consider the reaction

$$C_{12}H_{22}O_{11(s)} \ + \ 12 \text{ O}_{2(g)} \ \rightarrow \ 12 \text{ CO}_{2(g)} \ + \ 11 \text{ H}_2O_{(g)}$$
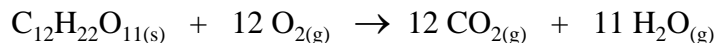
which has a ΔH of -5.65 x 10³ kJ/mol ($C_{12}H_{22}O_{11}$). How much heat (energy) can be produced during the complete combustion of 10.0 g of sucrose?

$$q = n\Delta H_{comb}$$

$$n = \frac{10.0 \text{ g}}{342.2 \tfrac{g}{mol}} = 0.02922 \text{ mol}$$

$$q = 0.2922 \text{ mol}(-5.65 \times 10^3 \tfrac{kJ}{mol}) = -165.1 \text{ kJ}$$

3. If all of the energy in question 2 were used to heat 1.0 L of water at 22.0°C, what would the final temperature of the water be? (Assume 100% energy transfer to the water.)

$$q_{comb} = -q_{H_2O}$$

$$-165.1 \times 10^3 \text{ J} = -1000 \text{ g}(4.184 \tfrac{J}{g \cdot °C})(T_f - 22.0°C)$$

$$T_f - 22.0°C = 39.46°C$$

$$T_f = 61.5°C$$

4. Using standard enthalpies of reaction, calculate the $\Delta H°$ for the following reactions:

$C_2H_{2(g)} + 2 H_{2(g)} \rightarrow C_2H_{6(g)}$

$C_2H_2(g) + \tfrac{5}{2}O_2(g) \rightarrow 2 CO_2(g) + H_2O(l)$	$\Delta H° = -1299.4$ kJ
$2 C(s) + 3 H_2(g) \rightarrow C_2H_6(g)$	$\Delta H° = -84.7$ kJ
$2 CO_2(g) \rightarrow 2 C(s) + 2 O_2(g)$	$\Delta H° = 2(393.5$ kJ$)$
$H_2O(l) \rightarrow H_2(g) + ½ O_2(g)$	$\Delta H° = 285.8$ kJ

$$C_2H_{2(g)} + 2 H_{2(g)} \rightarrow C_2H_{6(g)} \qquad \Delta H°_{rxn} = -311.3 \text{ kJ}$$

$2 CH_{4(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow C_2H_{6(g)} + H_2O_{(l)}$

$2 CH_4(g) + 4 O_2(g) \rightarrow 2 CO_2(g) + 4 H_2O(l)$	$\Delta H° = 2(-890.4$ kJ$)$
$2 C(s) + 3 H_2(g) \rightarrow C_2H_6(g)$	$\Delta H° = -84.7$ kJ
$2 CO_2(g) \rightarrow 2 C(s) + 2 O_2(g)$	$\Delta H° = 2(393.5$ kJ$)$
$3 H_2O(l) \rightarrow 2 H_2(g) + \tfrac{3}{2} O_2(g)$	$\Delta H° = 3(285.8$ kJ$)$

$$2 CH_{4(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow C_2H_{6(g)} + H_2O_{(l)} \qquad \Delta H°_{rxn} = -221.1 \text{ kJ}$$