Thermodynamics
6.4 Heat Capacity and Calorimetry
6.6 Introduction to Enthalpy of Reaction
Worksheet Key

1) A 85.2 g copper bar was heated to 221.32 °C and placed in a coffee cup calorimeter containing 425.0 mL of water at 22.55 °C. The final temperature of the water was recorded to be 26.15 °C.

   a. How much heat was gained by the water?

$$q_w = mc\Delta T$$
$$q_w = (425.0\text{ g})(4.184\text{ J/g}\cdot{}^\circ\text{C})(26.15\ {}^\circ\text{C} - 22.55\ {}^\circ\text{C})$$
$$q_w = +6402\text{ J}$$

   b. How much heat was lost by the copper?

$$q_w = -q_{\text{Cu}}$$
$$q_{\text{Cu}} = -6402\text{ J}$$

   c. What is the specific heat of copper?

$$q_{\text{Cu}} = mc\Delta T$$
$$c = \frac{q_{\text{Cu}}}{m\Delta T}$$
$$c = \frac{-6402\text{ J}}{(85.2\text{ g})(26.15\ {}^\circ\text{C} - 221.32\ {}^\circ\text{C})}$$
$$c = 0.385\text{ J/g}\cdot{}^\circ\text{C}$$

   d. Was energy conserved in the process? Justify your answer.

Energy was conserved in this process. All of the heat that was lost by the copper was gained by the water.

2) A 100.0 mL sample of 0.76 *M* HCl at 23.0 °C was mixed with 100.0 mL of 0.76 *M* NaOH at 23.0°C in a coffee cup calorimeter and the following reaction occurred.

$$H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$$

The temperature of the solution increased and a maximum temperature of 28.2 °C was recorder. Assume that no heat was lost to the surroundings, the volumes were additive, the specific heat capacity of the solution was 4.184 J/gK, and the density of the solution was 1.00 g/mL.

a. Calculate the enthalpy change, $\Delta H_{\text{rxn},}$ for the formation of 1.0 mol H_2O in this reaction.

$$m = 200.0\text{ mL} \times 1.00\text{ g/mL} = 2.00\times10^2\text{g}$$

$$q_{sol} = mc\Delta T$$

$$q_{sol} = (2.00\times10^2\text{g})(4.184\text{ J / g}\cdot^\circ\text{ C})(28.2\ ^\circ\text{C} - 23.0\ ^\circ\text{C})$$

$$q_{sol} = (2.00\times10^2\text{g})(4.184\text{ J / g}\cdot^\circ\text{ C})(5.2\ ^\circ\text{C})$$

$$q_{sol} = +4400\text{ J}$$

$$q_{rxn} = -q_{\text{sol}}$$

$$q_{rxn} = -4400\text{ J}$$

$$0.1000\text{ L} \times \frac{0.76\text{ mol H}^+}{1\text{ L}} = 0.076\text{ mol H}^+ \qquad 0.1000\text{ L} \times \frac{0.76\text{ mol OH}^-}{1\text{ L}} = 0.076\text{ mol OH}^-$$

$$\Delta H_{\text{rxn}} = \frac{q_{\text{rxn}}}{n_{\text{H}_2\text{O}}} = \frac{-4400\text{ J}}{0.076\text{ mol H}_2\text{O}} = -5.8\times10^4\text{ J} = -58\text{ kJ}$$

b. Is the reaction endothermic or exothermic?

Exothermic

c. Was energy conserved in the process? Justify your answer.

Energy was conserved in this process. All of the heat that was lost by the reaction was gained by the solution.

3) A coffee cup calorimeter contains 100.0 mL of 1.50 *M* $Ba(NO_3)_2$ at 25.0°C. A student pours 100.0 mL of 1.50 *M* Na_2SO_4 at 25.0°C into the calorimeter. A precipitate forms and the temperature rises to 29.7°C. Assume that no heat was lost to the surroundings, the volumes were additive, the specific heat capacity of the solution was 4.184 J/gK, and the density of the solution was 1.00 g/mL.

   a. This reaction could be represented as a complete ionic or a net ionic equation. Write the balanced chemical equation that should be used in association with the ΔH_{rxn} value for this reaction. Justify your choice.

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

The net ionic equation was chosen, as the spectator ions do not influence the energy changes associated with this reaction.

   b. Find the amount of heat that was lost or gained by the solution in the calorimeter.

$$m = 200.0 \text{ mL} \times 1.00 \text{ g/mL} = 2.00 \times 10^2 \text{g}$$
$$q_{sol} = mc\Delta T$$
$$q_{sol} = 2.00 \text{ x } 10^2 \text{ g } (4.184 \text{ J/g°C})(29.7\text{°C} - 25.0\text{°C})$$
$$q_{sol} = 2.00 \text{ x } 10^2 \text{ g } (4.184 \text{ J/g°C})(4.7\text{°C})$$
q_{sol} = 3900 J were absorbed by the solution.

   c. Find the heat of reaction, ΔH_{rxn}.

$$0.1000 \text{ L} \times \frac{1.50 \text{ mol Ba}^{2+}}{1 \text{ L}} = 0.150 \text{ mol Ba}^{2+}$$
$$q_{rxn} = - q_{sol}$$
$$q_{rxn} = - 3900 \text{ J}$$
$$\Delta H_{rxn} = \frac{q_{rxn}}{\text{mol BaSO}_4} = \frac{-3900 \text{ J}}{0.150 \text{ mol BaSO}_4} = -26000 \text{ J/mol} = -26 \text{ kJ/mol}$$

   d. Is the reaction endothermic or exothermic?

Exothermic

4) The specific heat capacity of copper is 0.385 J/(gK). If 353 J of heat are added to 3.6 moles of copper at 283 K, what is the final temperature of the sample of copper?

$$C_{m_{\mathrm{Cu}}} = \frac{0.385\mathrm{J}}{\mathrm{gK}} \times \frac{63.55\mathrm{g\ Cu}}{1\ \mathrm{mol\ Cu}} = 24.5\ \mathrm{J/(molK)}$$

$$\begin{aligned} q &= nC_{m_{\mathrm{Au}}}\Delta T \\ q &= nC_{m_{\mathrm{Au}}}(T_f - T_i) \\ T_f &= \frac{q}{nC_{m_{\mathrm{Au}}}} + T_i \\ T_f &= \frac{353\mathrm{J}}{(3.6\ \mathrm{moles})(24.5\ \mathrm{J/(molK)})} + 283\mathrm{K} \\ &= 287\mathrm{K} \end{aligned}$$

5) The specific heat capacity of iron 0.45 J/gK.

a. Find the molar heat capacity for iron.

$$C_{m_{\mathrm{Fe}}} = \frac{0.45\mathrm{J}}{\mathrm{gK}} \times \frac{55.85\mathrm{g\ Fe}}{1\ \mathrm{mol\ Fe}} = 25\ \mathrm{J/(molK)}$$

b. If 1239 J of heat are added to 5.6 mole sample of Fe at 12.2°C, what is the final temperature of the sample?

$$\begin{aligned} q &= nC_{m_{\mathrm{Au}}}\Delta T \\ q &= nC_{m_{\mathrm{Au}}}(T_f - T_i) \\ T_f &= \frac{q}{nC_{m_{\mathrm{Au}}}} + T_i \\ T_f &= \frac{1239\mathrm{J}}{(5.6\ \mathrm{moles})(25\ \mathrm{J/(molK)})} + (12 + 273)\mathrm{K} \\ &= 294\mathrm{K} \end{aligned}$$

6) An experiment was conducted in order to determine the enthalpy change that occurs when 1.0 mole of ice at 0°C melts and becomes 1.0 mole of water at 0°C. The enthalpy change associated with this process is referred to as the heat of fusion, ΔH_{fus}, of ice. In the experiment, a 9.68 g sample of ice at 0°C was added to a coffee cup calorimeter containing 278.25 mL of distilled water. The temperature of the water was 22.485°C before the ice was added. The lowest temperature that was recorded after the ice had melted was 19.050°C.

a. Find the amount of heat lost or gained by the water in the calorimeter.

$q_w = mc\Delta t$

$q_w = 278.25\text{ g }(4.184\text{ J/g°C})(19.050\text{°C} - 22.485\text{°C})$

$q_w = 278.25\text{ g }(4.184\text{ J/g°C})(3.435\text{°C})$

$q_w = -3999$ J were lost by the calorimeter.

b. In this experiment, the ice melted and then the temperature of the water produced by the ice increased from 0°C to 19.050°C. Calculate the amount of heat absorbed by the melted ice ($q_{\text{melted ice}}$) as its temperature increased from 0.000°C to 19.050°C.

$$q_{\text{temp gain}} = mc\Delta T$$

$$q_{\text{temp gain}} = 9.68\text{ g}(4.184\text{ J/g°C})(19.05\text{ °C} - 0.00\text{ °C}) = 772\text{ J}$$

c. Calculate the amount of heat that was gained by the ice during the melting process (q_{melting}). (Hint: $q_{\text{temp gain}} + q_{\text{melting}} = -q_{\text{w}}$)

$$q_{\text{temp gain}} + q_{\text{melting}} = -q_{\text{w}}$$

$$q_{\text{melting}} = -(-3999\text{ J}) - 772\text{ J} = +3227\text{ J}$$

d. Calculate the heat of fusion of ice, ΔH_{fus}.

$$9.68\text{ g H}_2\text{O}\times\frac{1\text{ mol H}_2\text{O}}{18.02\text{ g H}_2\text{O}} = 0.537\text{ mol H}_2\text{O}$$

$$\Delta H_{fus} = \frac{q_{\text{melting}}}{\text{mol H}_2\text{O}} = \frac{+3227\text{ J}}{0.537\text{ mol H}_2\text{O}} = +\,6010\text{ J} = +\,6.01\text{ kJ/mol}$$

e. Is the reaction endothermic or exothermic?

Endothermic

f. Energy was transferred from one system to another during this experiment. Identify the two interacting systems and outline the direction of energy flow.

One system was the ice and the water that was produced by the ice as it melted, and the other system was the water that was in the liquid form for the entire experiment. The heat flowed from the system of water into the system of ice and melted ice.

g. Was energy conserved in the process? Justify your answer.

Energy was conserved in this process. All of the heat that was lost by the water was gained by the ice as it melted and by the melted ice as its temperature increased from 0°C to 19.050°C.

6) How much heat is released when 24.8 g of $CH_4(g)$ is burned in excess oxygen gas to produce carbon dioxide and water?

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H_{rxn} = -\ 802.3$ kJ/mol

$$24.8\text{g CH}_4 \times \frac{1\text{ mol CH}_4}{16.05\text{g CH}_4} \times \frac{-802.3\text{ kJ}}{1\text{ mol CH}_4} = -1240\text{ kJ}$$

7) How much heat is released or absorbed when 45.8 g of hydrogen gas reacts with excess nitrogen gas according to the chemical equation below?

$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ $\Delta H_{rxn} = -\ 92.2$ kJ/mol

$$45.8\text{ g H}_2 \times \frac{1\text{ mol H}_2}{2.02\text{ g H}_2} \times \frac{-92.2\text{ kJ}}{3\text{ mol H}_2} = -697\text{ kJ (released)}$$

8) Suppose 391 kJ of heat is released when carbon monoxide gas reacts with nitrogen monoxide gas according to the thermochemical equation below.

$2\ CO(g) + 2\ NO(g) \rightarrow 2\ CO_2(g) + N_2(g)$ $\Delta H_{rxn} = -\ 746.6$ kJ

How many grams of carbon monoxide were consumed during this reaction?

$$391\text{ kJ} \times \frac{2\text{ mol CO}}{746.6\text{ kJ}} \times \frac{28.01\text{g CO}}{1\text{mol CO}} = 29.3\text{ g CO}$$

The ratio $\frac{2\text{ mol CO}}{746.6\text{ kJ}}$ is used as the balanced thermochemical equation tells us that 746.6 kJ of heat is released when 2 mol $CO(g)$ react with 2 mol $NO(g)$.

9) Use the thermochemical equation for the combustion of methane written below to answer the following question.

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$ $\Delta H_{rxn} = -\ 802.3$ kJ

How much heat is released when 24.8 g of $CH_4(g)$ is burned in excess oxygen gas to produce carbon dioxide and water?

$$24.8\text{g CH}_4 \times \frac{1\text{ mol CH}_4}{16.05\text{g CH}_4} \times \frac{-802.3\text{ kJ}}{1\text{ mol CH}_4} = -1210\text{ kJ}$$

10) Ammonia is produced from a reaction between hydrogen gas and nitrogen gas.

$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ $\Delta H_{rxn} = -\ 92.2$ kJ

a. Does the forward reaction release or absorb heat?

Released

b. How much heat is released or absorbed when 45.8 g of hydrogen gas reacts with excess nitrogen gas according to the chemical equation above?

$$45.8\text{ g H}_2 \times \frac{1\text{ mol H}_2}{2.02\text{ g H}_2} \times \frac{-92.2\text{ kJ}}{3\text{ mol H}_2} = -697\text{ kJ (released)}$$

11) $\Delta H_{rxn} = +\ 80.3$ kJ for the following reaction.

$Ba(OH)_2 \cdot 8H_2O(s) + 2\ NH_4Cl(s) \rightarrow BaCl_2(aq) + 2\ NH_3(aq) + 10\ H_2O(l)$

a. Is the forward reaction endothermic or exothermic?

Endothermic

b. How much heat is released or absorbed when 16.8 g of $NH_4Cl(s)$ reacts with excess $Ba(OH)_2 \cdot 8H_2O(s)$?

$$16.8\text{ g NH}_4\text{Cl} \times \frac{1\text{ mol NH}_4\text{Cl}}{53.50\text{ g NH}_4\text{Cl}} \times \frac{+80.3\text{ kJ}}{2\text{ mol NH}_4\text{Cl}} = +12.6\text{ kJ (absorbed)}$$

c. If this reaction could be reversed in order to produce 16.8 g of $NH_4Cl(s)$, how much heat would be released or absorbed during that process.

$$16.8\text{ g NH}_4\text{Cl} \times \frac{1\text{ mol NH}_4\text{Cl}}{53.50\text{ g NH}_4\text{Cl}} \times \frac{-80.3\text{ kJ}}{2\text{ mol NH}_4\text{Cl}} = -12.6\text{ kJ (released)}$$