During Class Invention

Solution Calorimetry

Name(s) with Lab section in Group

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1. A 192 gram sample of copper metal is heated to 100 ˚C in boiling water and then added to 750 grams of water at 24.0 ˚C in an OSU calorimeter. The heat capacity of the calorimeter is 50.0 J $˚C^{-1}$. Calculate the final temperature of the water and the copper in the OSU calorimeter. (NOTE: the specific heat of copper is 0.385 J g^{-1} $˚C^{-1}$ and for water it is 4.184 J g^{-1} $˚C^{-1}$.)

a) write the first law heat balance equation for this system.

$$q_{\textbf{lost by metal}} = - q_{\textbf{gain by H}_2\textbf{O}}$$

b) Solve for the final temperature.

$$\mathbf{q = m \cdot SH \cdot \Delta T}$$

$$\mathbf{(m \cdot SH \cdot \Delta T)_{lost\ by\ the\ metal} = -(m \cdot SH \cdot \Delta T)_{gained\ by\ the\ H_2O}}$$

$$\mathbf{(192\ g \cdot \left(0.385\ \frac{J}{g \cdot {}^\circ C}\right) \cdot (T_f - 100.0\ {}^\circ C))_{metal} = -(750\ g \cdot \left(4.184\ \frac{J}{g \cdot {}^\circ C}\right) \cdot (T_f - 24.0\ {}^\circ C))_{H_2O}}$$

$$\mathbf{(73.9\ \frac{J}{{}^\circ C} \cdot (T_f - 100.0\ {}^\circ C))_{metal} = -(3138\ \frac{J}{{}^\circ C} \cdot (T_f - 24.0\ {}^\circ C))_{H_2O}}$$

$$\mathbf{\cdot (T_f - 100.0\ {}^\circ C)) = -(42.5 \cdot (T_f - 24.0\ {}^\circ C))}$$

$$\mathbf{\cdot (T_f - 100.0\ {}^\circ C)) = -42.5 \cdot T_f + 1019\ {}^\circ C))}$$

$$\mathbf{\cdot 43.5 T_f = 1119\ {}^\circ C))}$$

$$\mathbf{T_f = 25.7\ {}^\circ C}$$

2. When 100. mLs of 0.200 M CsOH is mixed with 100. mLs of 0.200 M HCl in an OSU calorimeter the following reaction occurs.

$$CsOH(aq) + HCl(aq) \rightarrow CsCl(aq) + H_2O(l)$$

The temperature of both solutions before mixing was 24.30 ˚C. After mixing the temperature was 25.6 ˚C. The heat capacity of the OSU calorimeter is 50 J $˚C^{-1}$, calculate the heat produced in the reaction, and the heat produced per mol of water in the reaction.

$$q_{\textbf{rxn}} = -(q_{\textbf{solution}} + q_{\textbf{calorimeter}})$$

$$q_{\textbf{rxn}} = -(\textbf{mass}_{\textbf{solution}} * \textbf{SH}_{\textbf{ solution}} * \Delta T_{\textbf{solution}} + \textbf{SH}_{\textbf{calorimeter}} * \Delta T_{\textbf{calorimeter}})$$

To determine the mass of the solution we must convert the volume of the solution to grams. The problem does not specifically state what the density of the solution is, but since the concentrations are low we can assumed the density of the solution is equal to

that of water (density = 1.00 g mL^{-1}). The mass of the solution is the sum of the masses of the individual solutions. Remember we can add the two masses because the concentrations of the two solutions are close to each other.

$$q_{rxn} = -((100.\text{ g} + 100.\text{ g})(4.184\ \frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}})(25.6\ {}^\circ\text{C} - 24.30\ {}^\circ\text{C}) + 50.0\ \frac{\text{J}}{{}^\circ\text{C}}(25.6\ {}^\circ\text{C} - 24.30\ {}^\circ\text{C}))$$

$$q_{rxn} = -(1087\text{ J} + 65\text{ J}) = -1152\text{ J}$$

To determine the heat produced per mol of water formed we must calculate how much water is formed. According to the information

$$0.100\text{ L}\left(\frac{0.200\text{ mol CsOH}}{1\text{ L}}\right) = 0.0200\text{ mol CsOH}$$

$$0.100\text{ L}\left(\frac{0.200\text{ mol HCl}}{1\text{ L}}\right) = 0.0200\text{ mol HCl}$$

Since there are equal mol of both CsOH and HCl either can be considered the limiting reagent to calculate the mol of water formed.

$$0.0200\text{ mol HCl}\left(\frac{1\text{ mol }H_2O}{1\text{ mol HCl}}\right) = 0.0200\text{ mol }H_2O$$

$$q_{rxn} = \frac{-1152\text{ J}}{0.0200\text{ mol }H_2O} = -57.6\ \frac{\text{kJ}}{\text{mol}}$$

3a. A 0.692 gram sample of glucose, $C_6H_{12}O_6$, is burned in a constant volume, bomb calorimeter. The temperature change is measured at 1.80 °C. The calorimeter contains 1.05 kg and the ‘dry’ calorimeter has a heat capacity of 650. J °C-1. Calculate the amount heat evolved per mol of glucose.

$$q_{rxn} = -(\ q_{water} + q_{calorimeter})$$

$$\begin{aligned} -q_{rxn} &= -(1.05 \times 10^3\text{ g})(4.184\ \frac{\text{J}}{\text{g}\cdot{}^\circ\text{C}})(1.80^\circ\text{C}) + (650\ \frac{\text{J}}{{}^\circ\text{C}})(1.80^\circ\text{C}) \\ &= -(7.91\text{ kJ} + 1.17\text{ kJ}) = -9.07\text{ kJ} \\ &= \frac{9.07\text{ kJ}}{0.692\text{ g }C_6H_{12}O_6}\left(\frac{180\text{ g}}{1\text{ mol }C_6H_{12}O_6}\right) = -2.36 \times 10^3\ \frac{\text{kJ}}{\text{mol}} \end{aligned}$$