

CH301 Fall 2008 Worksheet 12 Answer Key

1. Assume we want to use a bomb calorimeter to determine the specific heat capacity of an unknown liquid. We use 3 L of the unknown liquid and perform a known reaction that releases 400 kJ of heat. We measure an initial and final temperature of 25 °C and 28.7 °C, respectively. If the heat capacity of the calorimeter is 85 J·K<sup>-1</sup>, and the density of the liquid is 2.34 g·mL<sup>-1</sup>, what is the specific heat capacity of the unknown liquid?

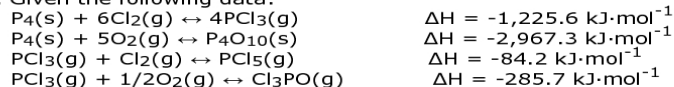
$$\begin{aligned} \Delta H &= 400 \text{ kJ} \\ m &= 3 \text{ L} * 1000 \text{ mL}\cdot\text{L}^{-1} * 2.34 \text{ g}\cdot\text{mL}^{-1} = 7020 \text{ g} \\ \Delta T &= T_f - T_i = 28.7 \text{ }^\circ\text{C} - 25 \text{ }^\circ\text{C} = 3.7 \text{ }^\circ\text{C} = 3.7 \text{ K} \\ c_{\text{cal}} &= 85 \text{ J}\cdot\text{K}^{-1} * .001 \text{ kJ}\cdot\text{J}^{-1} = 0.085 \text{ kJ}\cdot\text{K}^{-1} \\ \Delta H &= m\cdot c\cdot\Delta T + c_{\text{cal}}\cdot\Delta T \\ c &= (\Delta H - c_{\text{cal}}\cdot\Delta T)/(m\cdot\Delta T) \\ &= (400 \text{ kJ} - 0.085 \text{ kJ}\cdot\text{K}^{-1} * 3.7 \text{ K})/(7020 \text{ g} * 3.7 \text{ K}) \\ &= 0.01539 \text{ kJ}\cdot\text{g}^{-1}\cdot\text{K}^{-1} = 15.39 \text{ J}\cdot\text{g}^{-1}\cdot\text{K}^{-1} \end{aligned}$$

2. Lets say we filled the calorimeter above with 3 L of water and performed the same known reaction above. We measured a final temperature of 57.56 °C, but forgot to measure the initial temperature. Considering the density and specific heat capacity of water are 1 g·mL<sup>-1</sup> and 4.184 J·g<sup>-1</sup>·K<sup>-1</sup>, could we calculate what the initial temperature must have been? If so, what was the initial temperature?

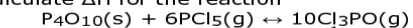
Yes, we can determine T<sub>i</sub>. We don't know the value of either ΔT or T<sub>i</sub>, so we have two unknowns, but we also have two equations, so we can solve:

$$\begin{aligned} \Delta H &= m\cdot c\cdot\Delta T + c_{\text{cal}}\cdot\Delta T \quad \text{and} \quad \Delta T = T_f - T_i \\ \Delta T &= \Delta H/(m\cdot c + c_{\text{cal}}) = 400,000 \text{ J}/(3000 \text{ g} * 4.184 \text{ J}\cdot\text{g}^{-1}\cdot\text{K}^{-1} + 85 \text{ J}\cdot\text{K}^{-1}) \\ &= 32.34 \text{ K} = 32.34 \text{ }^\circ\text{C} \\ T_i &= T_f - \Delta T = 57.56 \text{ }^\circ\text{C} - 32.34 \text{ }^\circ\text{C} = 25.22 \text{ }^\circ\text{C} \end{aligned}$$

3. Given the following data:

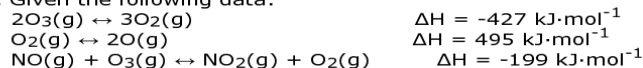


calculate ΔH for the reaction

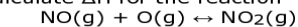


$$\begin{aligned} \Delta H_{\text{rxn}} &= (1 * -1,225.6 \text{ kJ}\cdot\text{mol}^{-1}) + (-1 * -2,967.3 \text{ kJ}\cdot\text{mol}^{-1}) + (-6 * -84.2 \text{ kJ}\cdot\text{mol}^{-1}) + (10 * -285.7 \text{ kJ}\cdot\text{mol}^{-1}) \\ &= -610.1 \text{ kJ}\cdot\text{mol}^{-1} \end{aligned}$$

4. Given the following data:

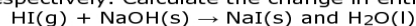


calculate ΔH for the reaction



$$\begin{aligned} \Delta H_{\text{rxn}} &= (-1/2 * -427 \text{ kJ}\cdot\text{mol}^{-1}) + (-1/2 * 495 \text{ kJ}\cdot\text{mol}^{-1}) + (1 * -199 \text{ kJ}\cdot\text{mol}^{-1}) \\ &= -233 \text{ kJ}\cdot\text{mol}^{-1} \end{aligned}$$

5. Hydroiodic acid (HI) and sodium hydroxide (NaOH) are a strong acid and strong base respectively. Calculate the change in enthalpy for their neutralization reaction:



Consult Appendix 2 in your ebook for standard enthalpy of formation values.

$$\begin{aligned} \Delta H_{\text{rxn}} &= \sum H_{\text{f,products}} - \sum H_{\text{f,reactants}} \\ &= (-287.78 \text{ kJ}\cdot\text{mol}^{-1} + -285.83 \text{ kJ}\cdot\text{mol}^{-1}) - (26.48 \text{ kJ}\cdot\text{mol}^{-1} + -425.61 \text{ kJ}\cdot\text{mol}^{-1}) \\ &= -174.48 \text{ kJ}\cdot\text{mol}^{-1} \end{aligned}$$