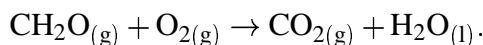


# Chemistry 2720 Fall 2001 Assignment 4 Solutions

1. If we add the four reactions given, we get the reverse of the desired reaction. The enthalpy change for the disproportionation reaction is therefore

$$\Delta\bar{H}^\circ = -(91 + 51.5 + 41.8 + 36 \text{ kJ/mol}) = -220 \text{ kJ/mol}.$$

2. (a) The enthalpy of combustion refers to the reaction



The enthalpy of combustion is therefore

$$\Delta\bar{H}_c^\circ = \Delta\bar{H}_f^\circ(\text{CO}_2) + \Delta\bar{H}_f^\circ(\text{H}_2\text{O}) - \left( \Delta\bar{H}_f^\circ(\text{CH}_2\text{O}) + \Delta\bar{H}_f^\circ(\text{O}_2) \right).$$

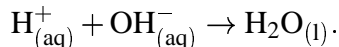
$$\text{Note that } \Delta\bar{H}_f^\circ(\text{O}_2) = 0.$$

$$\begin{aligned} \therefore \Delta\bar{H}_f^\circ(\text{CH}_2\text{O}) &= \Delta\bar{H}_f^\circ(\text{CO}_2) + \Delta\bar{H}_f^\circ(\text{H}_2\text{O}) - \Delta\bar{H}_c^\circ \\ &= -393.51 + (-285.830) - (-570.78) \text{ kJ/mol} \\ &= -108.56 \text{ kJ/mol}. \end{aligned}$$

- (b) For the combustion reaction,

$$\begin{aligned} \Delta\bar{n}_{\text{gas}} &= 1 - 2 = -1. \\ \therefore \Delta\bar{E}^\circ &= \Delta\bar{H}^\circ - RT\Delta\bar{n}_{\text{gas}} \\ &= -570.78 \text{ kJ/mol} - (8.314510 \times 10^{-3} \text{ kJ K}^{-1} \text{ mol}^{-1})(298.15 \text{ K})(-1) \\ &= -568.30 \text{ kJ/mol}. \\ n_{\text{CH}_2\text{O}} &= \frac{1.0045 \text{ g}}{30.03 \text{ g/mol}} = 0.03345 \text{ mol}. \\ \therefore q|_V &= (-568.30 \text{ kJ/mol})(0.03345 \text{ mol}) = -19.01 \text{ kJ}. \end{aligned}$$

3. The overall reaction is



The enthalpy change for this reaction is

$$\Delta\bar{H}^\circ = \Delta\bar{H}_f^\circ(\text{H}_2\text{O}) - \Delta\bar{H}_f^\circ(\text{OH}^-) = -55.815 \text{ kJ/mol}.$$

We have  $(0.080 \text{ L})(0.50 \text{ mol/L}) = 0.040 \text{ mol}$  of hydrogen ions and  $(0.045 \text{ L})(0.85 \text{ mol/L}) = 0.038 \text{ mol}$  of hydroxide ions. The hydroxide ions are therefore limiting. The heat generated is therefore

$$q|_P = (-55.815 \text{ kJ/mol})(0.038 \text{ mol}) = -2.1 \text{ kJ}.$$

Since the reaction produces heat — this is the meaning of the negative sign — the  $80 + 45 = 125$  mL of water gains 2.1 kJ, which will cause the temperature to rise. Since the density of water is about 1 g/mL, we have about 125 g of water. The change in temperature is therefore

$$\Delta T = \frac{q}{m\bar{C}_P} = \frac{2.1 \times 10^3 \text{ J}}{(125 \text{ g})(4.184 \text{ J K}^{-1} \text{ g}^{-1})} = 4.1 \text{ K.}$$

$$\therefore T_{\text{final}} = 4.1 + 22^\circ \text{ C} = 26^\circ \text{ C.}$$

4. (a) We need to warm the iron from  $20^\circ \text{ C}$  and then melt it. The number of moles of iron is

$$n_{\text{Fe}} = \frac{10 \text{ g}}{55.85 \text{ g/mol}} = 0.18 \text{ mol.}$$

Warming the iron to its melting point requires

$$q_1 = (0.18 \text{ mol})(25.1 \text{ J K}^{-1} \text{ mol}^{-1})(1808 - 293 \text{ K}) = 6809 \text{ J.}$$

To melt 10 g of iron, we need

$$q_2 = (10 \text{ g})(267 \text{ J/g}) = 2670 \text{ J.}$$

The total heat required is

$$q_{\text{total}} = q_1 + q_2 = 9479 \text{ J.}$$

The reaction is written per mole of  $\text{Fe}_2\text{O}_3$  so the number of moles of  $\text{Fe}_2\text{O}_3$  required is

$$n_{\text{Fe}_2\text{O}_3} = \frac{9.479 \text{ kJ}}{852 \text{ kJ/mol}} = 0.0111 \text{ mol.}$$

The number of moles of aluminium is twice as large as this, so we have  $n_{\text{Al}} = 0.0223 \text{ mol}$ .

The masses of the two reactants are

$$m_{\text{Fe}_2\text{O}_3} = (0.0111 \text{ mol})(159.70 \text{ g/mol}) = 1.78 \text{ g.}$$

$$m_{\text{Al}} = (0.0223 \text{ mol})(26.98 \text{ g/mol}) = 0.600 \text{ g.}$$

The total mass of thermite reagent is therefore 2.38 g.

- (b) The standard enthalpy of reaction is for a fictitious reaction occurring at  $25^\circ \text{ C}$ . Thus we imagine that the heat released by the reaction goes into heating the products from 25 to  $3000^\circ \text{ C}$ . Heating liquid iron from 25 to  $3000^\circ \text{ C}$  takes

$$\bar{q}_{\text{Fe}} = (46.024 \text{ J K}^{-1} \text{ mol}^{-1})(3000 - 25 \text{ K}) = 136.9 \text{ kJ/mol.}$$

There are two moles of iron in a stoichiometric mixture of the results for the reaction as written, so heating these two moles takes 273.8 kJ/mol of aluminium oxide. From the original 852 kJ/mol produced by the reaction, this leaves 578 kJ/mol. If all of this is used to heat  $\text{Al}_2\text{O}_3(l)$ , then we have

$$\bar{C}_P = \frac{578 \text{ kJ/mol}}{3000 - 25 \text{ K}} = 194 \text{ J K}^{-1} \text{ mol}^{-1}.$$

This is of course a very crude estimate.