

INS - CHEMISTRY WORKSHOP - WEEK 1 - SPRING

①

22.0°C  
90.0g

cold

55.0°C

hot

37°C

final

heat absorbed by cold water = heat given off by hot water

$$m_{\text{cold}} C_{\text{cold}} \Delta T = m_{\text{hot}} C_{\text{hot}} \Delta T$$

$$(90.0\text{g}) (4.184 \text{ J g}^{-1} \text{ K}^{-1}) (15\text{K}) = m_{\text{hot}} (4.184 \text{ J g}^{-1} \text{ K}^{-1}) (18\text{K})$$

$$m_{\text{hot}} = \frac{(90.0\text{g}) (15\text{K})}{(18\text{K})} = \underline{\underline{75.0\text{g}}}$$

② heat given off by (Al + Fe) = heat absorbed by water

$$m_{\text{Al}} C_{\text{Al}} (\Delta T)_{\text{Al}} + m_{\text{Fe}} C_{\text{Fe}} (\Delta T)_{\text{Fe}} = m_{\text{water}} C_{\text{water}} (\Delta T)_{\text{water}}$$

$T_f$  = final temp. of all materials in °C

$$\Rightarrow \left. \begin{aligned} (5.00\text{g}) (0.89 \text{ J } ^\circ\text{C}^{-1} \text{ g}^{-1}) \left(\frac{100}{4} - T_f\right)^\circ\text{C} \\ (10.00\text{g}) (0.45 \text{ J } ^\circ\text{C}^{-1} \text{ g}^{-1}) (100 - T_f)^\circ\text{C} \end{aligned} \right\} = (97.3\text{g}) (4.184 \text{ J } ^\circ\text{C}^{-1} \text{ g}^{-1}) (T_f - 22)^\circ\text{C}$$

$$(4.45 + 4.5) \cancel{\text{J}} [100 - T_f] = 407.10 \cancel{\text{J}} (T_f - 22)$$

$$895 - 8.95 T_f = 407.10 T_f - 8956.27$$

$$9851.27 = 416.05 T_f$$

$$T_f = \underline{\underline{23.68^\circ\text{C}}}$$

$$\textcircled{3} \quad \left. \begin{array}{l} \text{Heat absorbed by} \\ \text{bomb calorimeter} \end{array} \right\} = (1.56 \text{ kJ } ^\circ\text{C}^{-1}) (3.2^\circ\text{C})$$

$$= 4.992 \text{ kJ}$$

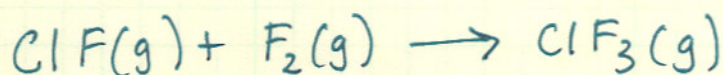
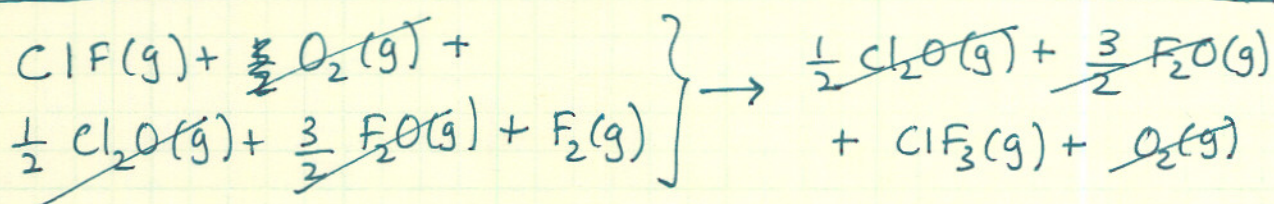
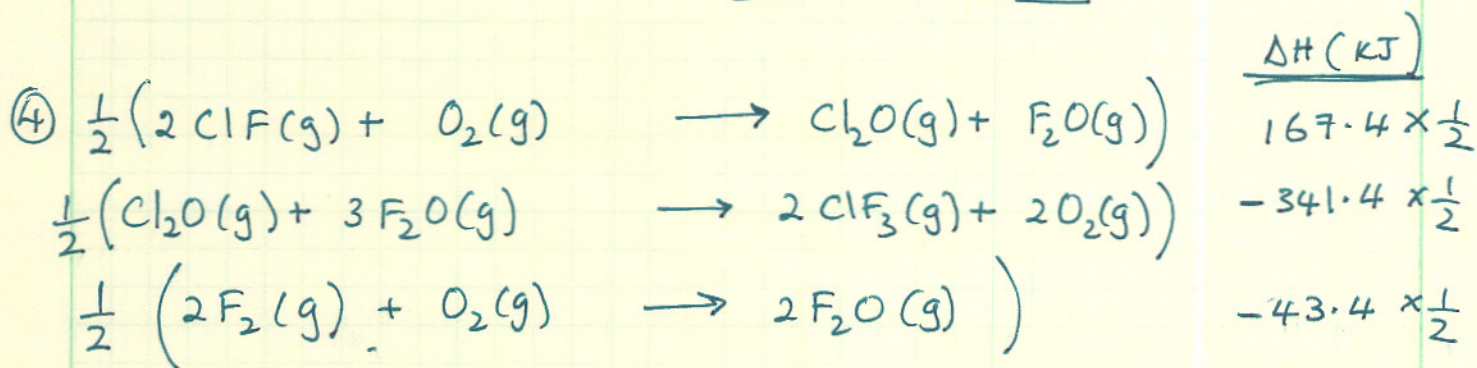
= heat given off by burning  
0.1964 g of quinone

$$\therefore \left. \begin{array}{l} \text{Energy of combustion} \\ \text{per gram of quinone} \end{array} \right\} = \frac{4.992 \text{ kJ}}{0.1964 \text{ g}} = \underline{\underline{25.418 \text{ kJ/g}}}$$

$$\begin{aligned} \text{quinone molar mass} &= [6(12.01) + 4(1.008) + 2(16)] \text{ g/mol} \\ &= 228.192 \text{ g/mol} \end{aligned}$$

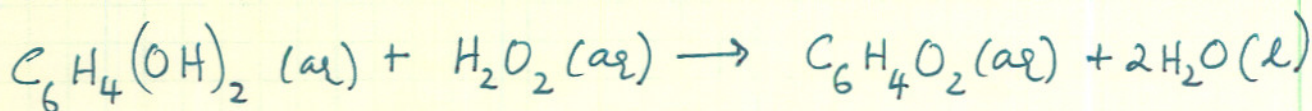
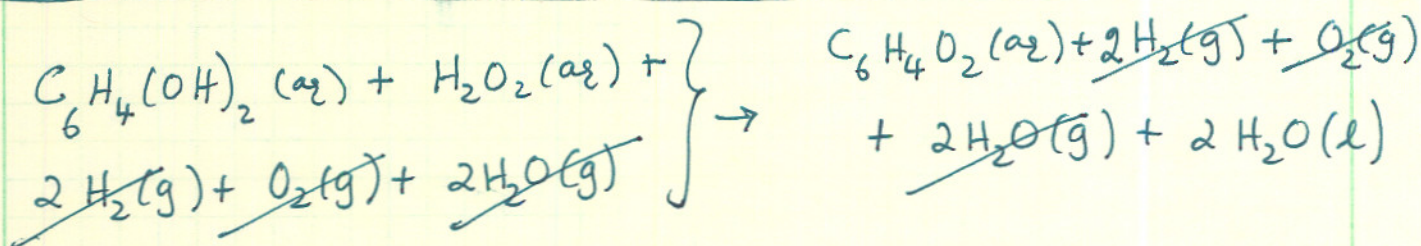
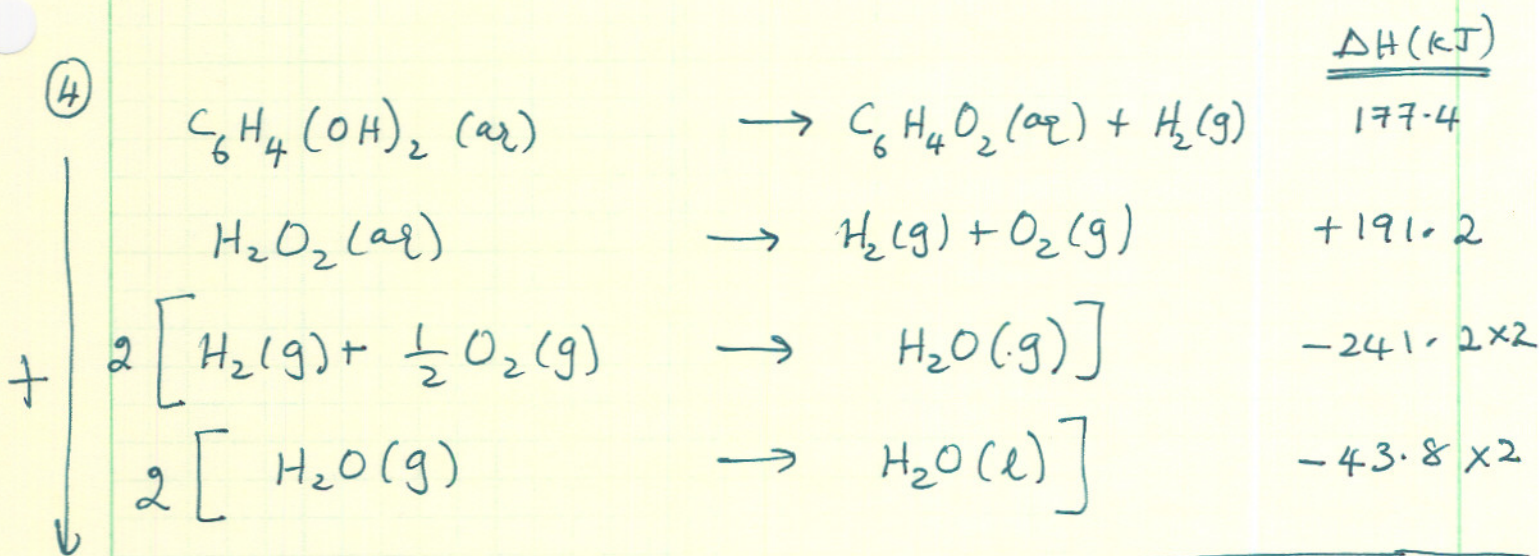
$$\left. \begin{array}{l} \text{Energy of combustion} \\ \text{per mole of quinone} \end{array} \right\} = 4.992 \text{ kJ} \times \left( \frac{228.192 \text{ g}}{0.1964 \text{ g}} \times \frac{1}{\text{mol}} \right)$$

$$= \underline{\underline{5799.27 \text{ kJ/mol}}}$$



$$\Delta H_{\text{in}} = \frac{1}{2} (167.4) + \frac{1}{2} (-341.4) + \frac{1}{2} (-43.4) \text{ kJ}$$

$$\Delta H_{rn} = \underline{\underline{-108.7 \text{ kJ}}}$$



$$\Delta H_{rn} = [177.4 + 191.2 + 2(-241.2) + 2(-43.8)] \text{ kJ}$$

$$= \underline{\underline{-201.4 \text{ kJ}}}$$

$$\begin{aligned}
 \textcircled{5} \quad \Delta H_m^\ominus &= 6\Delta H_f^\ominus [\text{NO}(\text{g})] + 6\Delta H_f^\ominus [\text{H}_2\text{O}(\text{g})] - 4\Delta H_f^\ominus [\text{NH}_3(\text{g})] - \\
 &\quad 5\Delta H_f^\ominus [\text{O}_2(\text{g})] \\
 &= [6(90.29) + 6(-241.83) - 4(-45.90) - 5(0)] \text{ kJ} \\
 &= \underline{\underline{-725.64 \text{ kJ}}}
 \end{aligned}$$

$$\begin{aligned}
 \Delta H_m^\ominus &= 2\Delta H_f^\ominus [\text{NO}_2(\text{g})] - 2\Delta H_f^\ominus [\text{NO}(\text{g})] - \Delta H_f^\ominus [\text{O}_2(\text{g})] \\
 &= [2(33.1) - 2(90.29) - 0] \text{ kJ} \\
 &= \underline{\underline{-114.38 \text{ kJ}}}
 \end{aligned}$$

$$\begin{aligned}
 \Delta H_m^\ominus &= 4\Delta H_f^\ominus [\text{HNO}_3(\text{aq})] + \Delta H_f^\ominus [\text{NO}(\text{g})] - 3\Delta H_f^\ominus [\text{NO}_2(\text{g})] - \\
 &\quad \Delta H_f^\ominus [\text{H}_2\text{O}(\text{l})] \\
 &= 4(-207.36) + 90.29 - 3(33.1) - (-285.83) \text{ kJ} \\
 &= \underline{\underline{-552.62 \text{ kJ}}}
 \end{aligned}$$