

### Objective 3

2.  $\Delta H$  is negative and obtained by taking the enthalpy of the products minus the enthalpy of the reactants. For  $\Delta H$  to be negative, the enthalpy of the reactants (2Cl) must be greater than that of the product.

3. a. exothermic

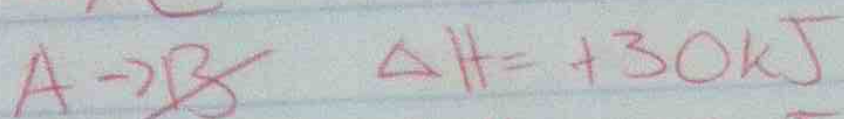
$$b. \frac{2.4 \text{ g Mg}}{1} \times \frac{1 \text{ mol Mg}}{24.31 \text{ g}} \times \frac{-1204 \text{ kJ}}{2 \text{ mol Mg}} = \boxed{-59.4 \text{ kJ}}$$

$$c. \frac{-96.0 \text{ kJ}}{1} \times \frac{2 \text{ mol MgO}}{-1204 \text{ kJ}} \times \frac{40.31 \text{ g}}{1 \text{ mol MgO}} = \boxed{6.43 \text{ g}}$$

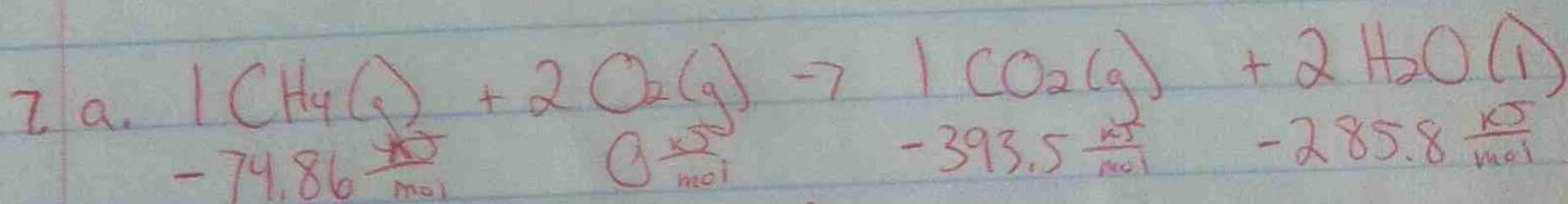
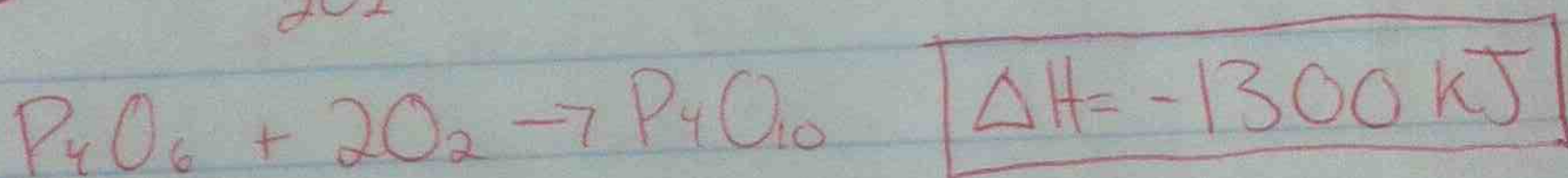
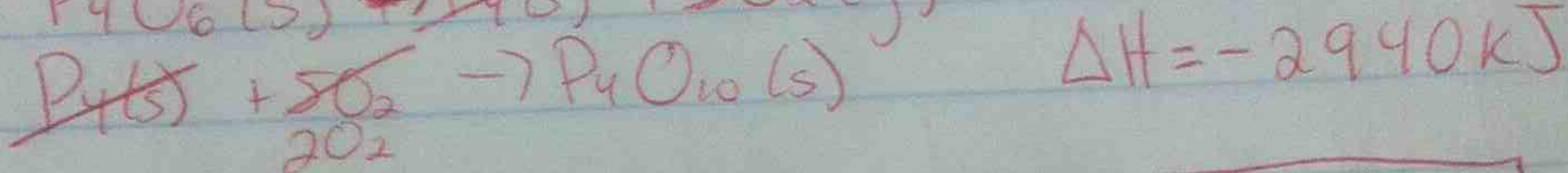
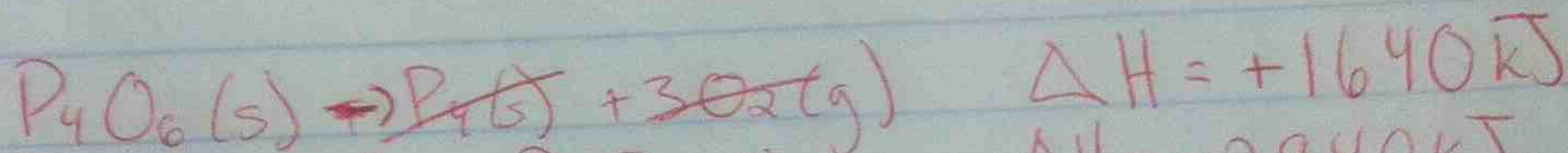
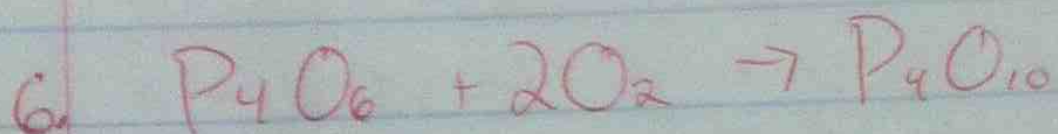
$$4. q = m \cdot C \cdot \Delta T = 60.0 \text{ g} \cdot 4.184 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} (16.9^\circ\text{C} - 22.0^\circ\text{C}) \\ = -1280 \text{ J lost by water}$$

+1280 J endothermic or 1.28 kJ  
need to find mol

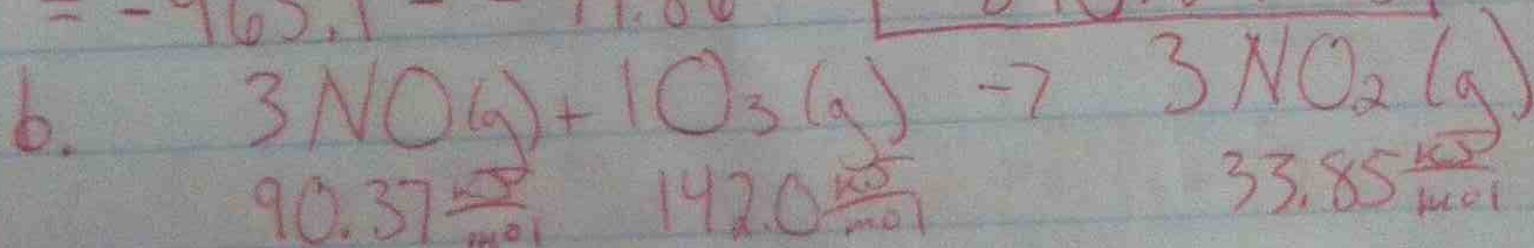
$$\frac{4.25 \text{ g NH}_4\text{NO}_3}{1} \times \frac{1 \text{ mol}}{80 \text{ g}} = 0.0531 \text{ mol}$$
$$\frac{1.28 \text{ kJ}}{0.0531 \text{ mol}} = \boxed{24.1 \frac{\text{kJ}}{\text{mol}}}$$



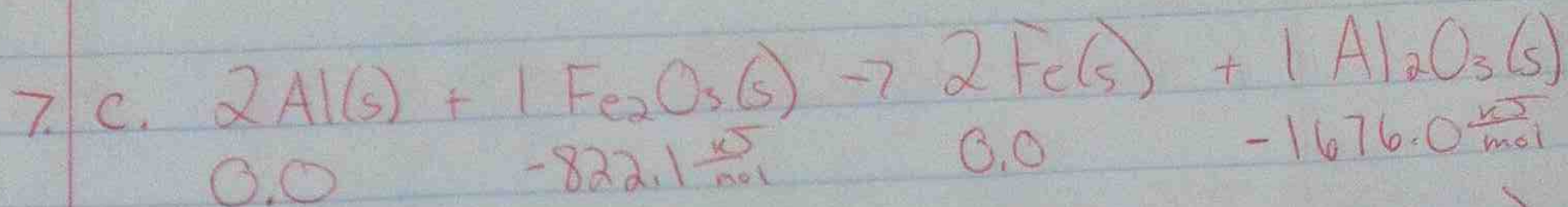
b. Don't worry about enthalpy diagram  
Hess's law is a law of summations  
Partial paths add to the desired path.



$\Delta H = \Delta H(\text{products}) - \Delta H(\text{reactants})$   
 $= \left( (1 \text{ mol} \cdot -393.5 \frac{\text{kJ}}{\text{mol}}) + (2 \text{ mol} \cdot -285.8 \frac{\text{kJ}}{\text{mol}}) \right) - \left( (1 \text{ mol} \cdot -74.86 \frac{\text{kJ}}{\text{mol}}) + (2 \text{ mol} \cdot 0) \right)$   
 $= -965.1 - -74.86 = -890.24 \text{ kJ}$



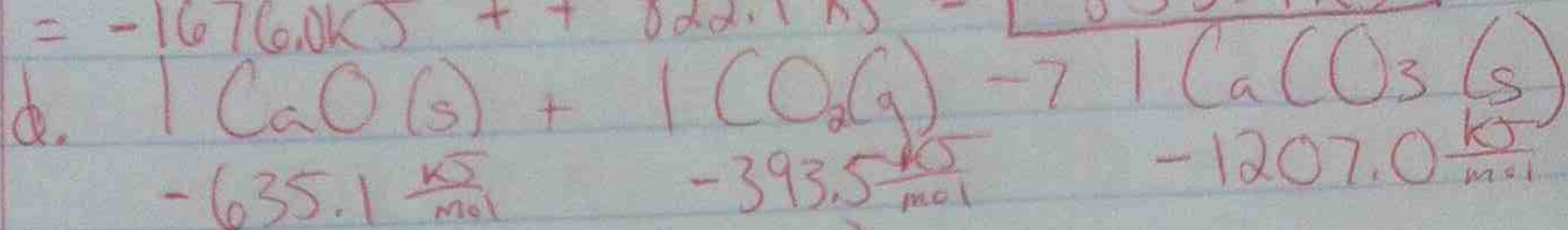
$\Delta H = \Delta H(\text{products}) - \Delta H(\text{reactants})$   
 $= \left( 3 \text{ mol} \cdot 33.85 \frac{\text{kJ}}{\text{mol}} \right) - \left( (3 \text{ mol} \cdot 90.37 \frac{\text{kJ}}{\text{mol}}) + (1 \text{ mol} \cdot 142.0 \frac{\text{kJ}}{\text{mol}}) \right)$   
 $= 101.55 \text{ kJ} - 413.11 \text{ kJ} = -311.56 \text{ kJ}$



$$\Delta H = \Delta H(\text{products}) - \Delta H(\text{reactants})$$

$$= ((1 \text{ mol} \cdot -1676.0 \frac{\text{kJ}}{\text{mol}}) + (2 \text{ mol} \cdot 0)) - ((2 \text{ mol} \cdot 0) + (1 \text{ mol} \cdot -822.1 \frac{\text{kJ}}{\text{mol}}))$$

$$= -1676.0 \text{ kJ} + + 822.1 \text{ kJ} = \boxed{-853.9 \text{ kJ}}$$



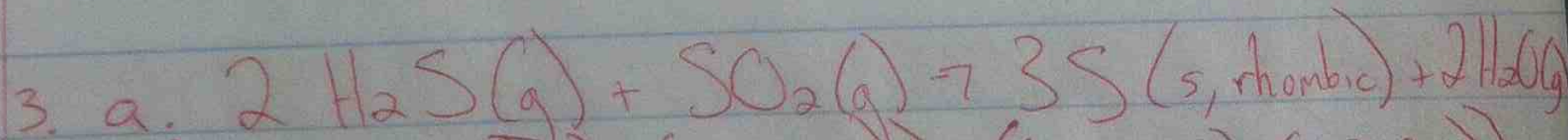
$$\Delta H = \Delta H(\text{products}) - \Delta H(\text{reactants})$$

$$= (1 \text{ mol} \cdot -1207.0 \frac{\text{kJ}}{\text{mol}}) - ((1 \text{ mol} \cdot -635.1 \frac{\text{kJ}}{\text{mol}}) + (1 \text{ mol} \cdot -393.5 \frac{\text{kJ}}{\text{mol}}))$$

$$= -1207.0 \frac{\text{kJ}}{\text{mol}} + + 1028.6 \text{ kJ} = \boxed{-178.4 \text{ kJ}}$$

### Objective 4

2. a. + decreases  
 b. + increases  
 c. - decreases  
 d. - decreases



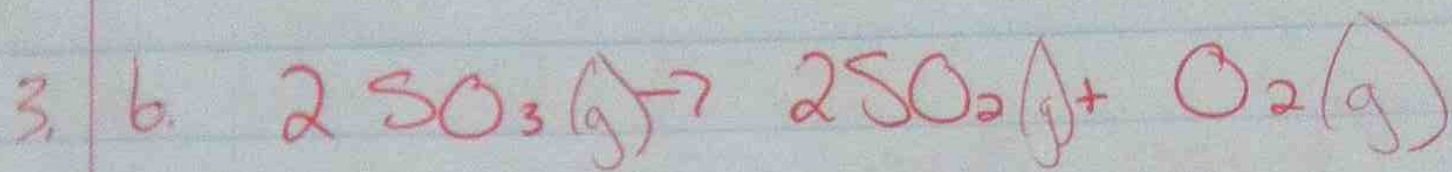
$$\Delta S = ((3 \text{ mol} \cdot 31.92 \frac{\text{J}}{\text{mol} \cdot \text{K}}) + (2 \cdot 188.72)) - ((2 \cdot 205.78) + (1 \cdot 248.11))$$

$$473.2 \frac{\text{J}}{\text{mol} \cdot \text{K}} - 659.67 \frac{\text{J}}{\text{mol} \cdot \text{K}} = \boxed{-186.47 \frac{\text{J}}{\text{mol} \cdot \text{K}}}$$

$$\Delta G = ((3 \text{ mol} \cdot 0) + (2 \text{ mol} \cdot -228.59)) - (2 \text{ mol} \cdot -33.05) + (1 \cdot -300.1)$$

$$= -457.18 \text{ kJ} - - 366.29 \text{ kJ}$$

$$= \boxed{-90.89 \text{ kJ}}$$

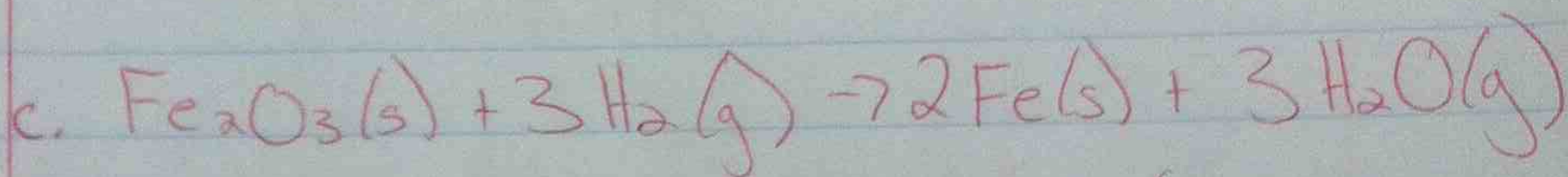


$$\Delta S = \left( (2 \text{ mol} \cdot 248.11 \frac{\text{J}}{\text{mol}\cdot\text{K}}) + (1 \text{ mol} \cdot 205.03) \right) - (2 \text{ mol} \cdot 256.65)$$

$$\Delta S = 701.25 - 513.3 = \boxed{187.95 \frac{\text{J}}{\text{K}}}$$

$$\Delta G = \left( (2 \text{ mol} \cdot -300.19 \frac{\text{kJ}}{\text{mol}}) + (1 \text{ mol} \cdot 0) \right) - (2 \text{ mol} \cdot -371.08)$$

$$= -600.38 \text{ kJ} + 742.16 = \boxed{+141.78 \text{ kJ}}$$



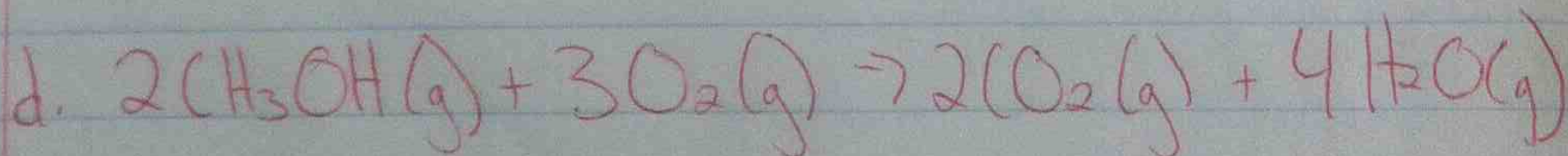
$$\Delta S = \left( (2 \text{ mol} \cdot 27.28 \frac{\text{J}}{\text{mol}\cdot\text{K}}) + (3 \cdot 188.72) \right) - \left( (1 \cdot 87.40) + (3 \cdot 130.57) \right)$$

$$620.72 \frac{\text{J}}{\text{K}} - 479.17 \frac{\text{J}}{\text{K}} = \boxed{141.55 \frac{\text{J}}{\text{K}}}$$

$$\Delta G = \left( (2 \text{ mol} \cdot 0) + (3 \cdot -228.59) \right) - \left( (1 \cdot -742.24) + (3 \cdot 0) \right)$$

$$= -685.77 \text{ kJ} + 742.24 \text{ kJ}$$

$$= \boxed{56.47 \text{ kJ}}$$



$$\Delta S = \left( (2 \text{ mol} \cdot 213.67 \frac{\text{J}}{\text{mol}\cdot\text{K}}) + (4 \cdot 188.72) \right) - \left( (2 \cdot 239.70) + (3 \cdot 205.03) \right)$$

$$= 1182.22 - 1094.49 = \boxed{87.73 \frac{\text{J}}{\text{K}}}$$

$$\Delta G = \left( (2 \text{ mol} \cdot -394.38 \frac{\text{kJ}}{\text{mol}}) + (4 \cdot -228.59) \right) - \left( (2 \cdot -162.42) + (3 \cdot 0) \right)$$

$$= -1703.4 \text{ kJ} + 324.84 \text{ kJ}$$

$$= \boxed{-1378.6 \text{ kJ}}$$

$$\begin{aligned} 4. a. \Delta G &= \Delta H - T\Delta S \\ &= -58.03 \text{ kJ} - (298 \text{ K} \cdot -.1766 \frac{\text{kJ}}{\text{K}}) \\ &= -58.03 \text{ kJ} + 52.63 \text{ kJ} \\ &= \boxed{-5.4 \text{ kJ}} \end{aligned}$$

$$\begin{aligned} b. 0 &= -58.03 \text{ kJ} - (T \cdot -.1766 \frac{\text{kJ}}{\text{K}}) \\ 58.03 \text{ kJ} &= + .1766 \frac{\text{kJ}}{\text{K}} \cdot T \\ T &= 328.5 \text{ K} \end{aligned}$$

c. Below this temperature the reaction is spontaneous

$$\begin{aligned} 5. \Delta G &= \Delta H - T\Delta S \\ -254.3 \text{ kJ} &= \Delta H - (298 \text{ K} \cdot .658 \frac{\text{kJ}}{\text{K}}) \\ -254.3 \text{ kJ} &= \Delta H - 196 \text{ kJ} \\ \Delta H &= \boxed{-58.3 \text{ kJ}} \end{aligned}$$