

If 50.3 kJ of heat was released when 5.48 g of formic acid are burned at constant pressure, then what is the enthalpy change of this reaction per mole of formic acid?

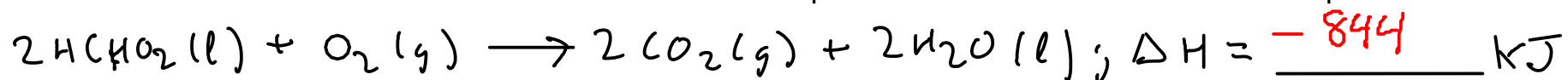
$$Q = -50.3 \text{ kJ} \quad \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2}$$

Find moles formic acid ...

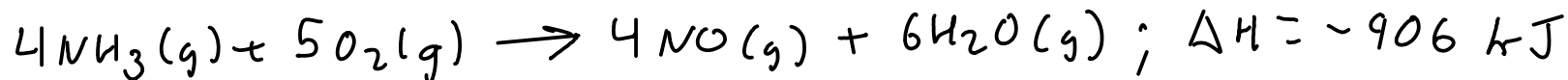
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

$$\Delta H = \frac{-50.3 \text{ kJ}}{0.1190631382 \text{ mol HCHO}_2} = \boxed{-422 \text{ kJ/mol HCHO}_2}$$

Based on the calculation above, can we complete this thermochemical equation?



We calculated the enthalpy change PER MOLE of formic acid (which is typical for calorimetry experiments), but the thermochemical equation written here is based on TWO moles of formic acid. We must double the enthalpy change.



What is the enthalpy change when 150. L of nitrogen monoxide are formed by this reaction at 25.0 C and 1.50 atm pressure?

- 1 - Convert 150. L of NO to moles using ideal gas equation.
- 2 - Convert moles NO to enthalpy change using thermochemical equation.

$$PV = nRT \quad \left| \quad P = 1.50 \text{ atm} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right.$$

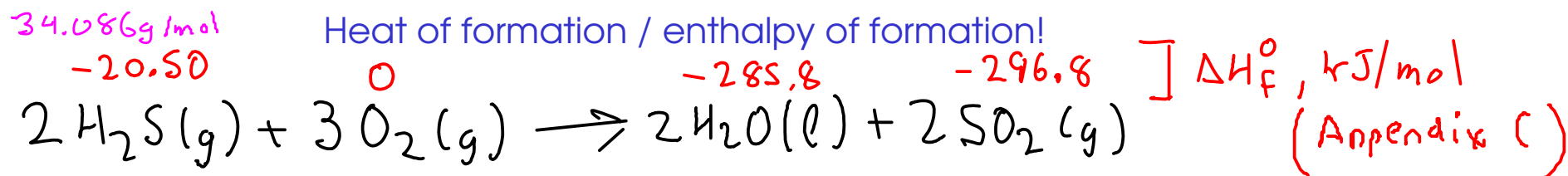
$$n = \frac{PV}{RT} \quad \left| \quad V = 150. \text{ L} \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \right.$$

$$\textcircled{1} \quad n_{\text{NO}} = \frac{(1.50 \text{ atm})(150. \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298.2 \text{ K})} =$$

$$= 9.194822849 \text{ mol NO}$$

$$4 \text{ mol NO} = -906 \text{ kJ}$$

$$\textcircled{2} \quad 9.194822849 \text{ mol NO} \times \frac{-906 \text{ kJ}}{4 \text{ mol NO}} = \boxed{-2080 \text{ kJ}}$$

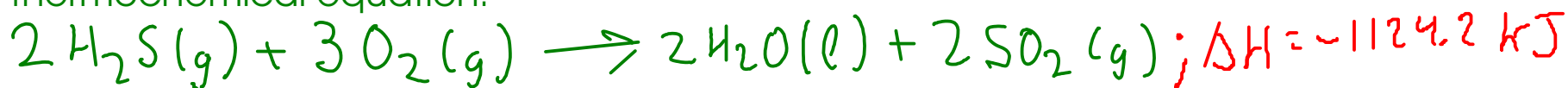


What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

- 1 - Calculate the enthalpy change for the reaction as written using Hess' Law.
- 2 - Convert 25.0 grams of hydrogen sulfide to moles using formula weight.
- 3 - Convert moles hydrogen sulfide to enthalpy change using thermochemical eqn.

$$\begin{aligned} \Delta H &= \sum \Delta_f^\circ H, \text{products} - \sum \Delta_f^\circ H, \text{reactants} \\ &= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)] \\ &= -1124.2 \text{ kJ} \end{aligned}$$

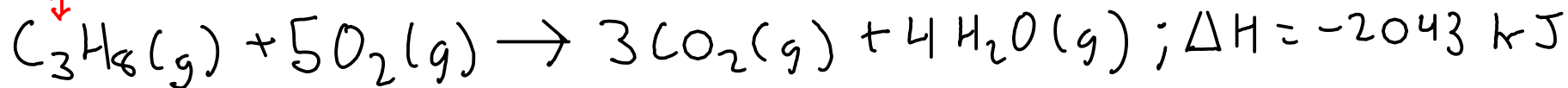
thermochemical equation:



$$34.086 \text{ g H}_2\text{S} = 1 \text{ mol H}_2\text{S} \quad | \quad 2 \text{ mol H}_2\text{S} = -1124.2 \text{ kJ}$$

$$25.0 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.086 \text{ g H}_2\text{S}} \times \frac{-1124.2 \text{ kJ}}{2 \text{ mol H}_2\text{S}} = \boxed{-412 \text{ kJ}}$$

propane



Calculate the volume of propane gas at 25.0 C and 1.08 atm required to provide 565 kJ of heat using the reaction above.

- 1 - Convert energy requirement to moles PROPANE using thermochemical equation.
- 2 - Convert moles propane to volume using ideal gas equation.

$$-2043 \text{ kJ} = \text{mol C}_3\text{H}_8$$

Since the reaction provides the energy, the energy requirement will be negative - from the reaction's point of view!

$$-565 \text{ kJ} \times \frac{\text{mol C}_3\text{H}_8}{-2043 \text{ kJ}} = 0.2765540871 \text{ mol C}_3\text{H}_8 \quad \textcircled{1}$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$n = 0.2765540871 \text{ mol C}_3\text{H}_8 \quad P = 1.08 \text{ atm}$$

$$R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

$$T = 25.0^\circ\text{C} = 298.2 \text{ K}$$

$$V = \frac{(0.2765540871 \text{ mol C}_3\text{H}_8)(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(298.2 \text{ K})}{(1.08 \text{ atm})}$$

$$= 6.27 \text{ L propane @ } 25.0^\circ\text{C} \text{ and } 1.08 \text{ atm}$$

END OF CHAPTER 6