

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?

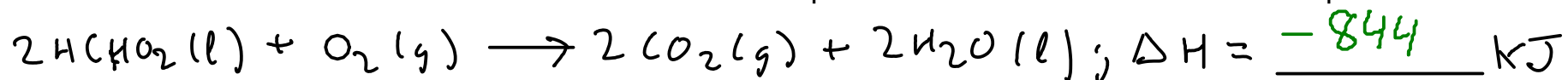
$$Q = -50.3 \text{ kJ} ; \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2}$$

Find moles formic acid:

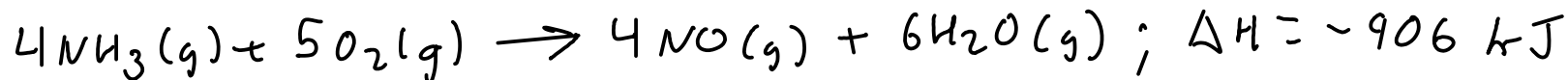
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

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

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



Since the equation is based on TWO MOLES of formic acid, and the enthalpy change above was calculated PER MOLE of formic acid, we need to multiply that enthalpy change by two to get the correct enthalpy change for the reaction.



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.

$$PV = nRT \quad \left| \quad \begin{array}{l} P = 1.50 \text{ atm} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ V = 150. \text{ L} \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \end{array} \right.$$

$$n = \frac{PV}{RT}$$

$$\textcircled{1} n_{\text{NO}} = \frac{(1.50 \text{ atm})(150. \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298.2 \text{ K})} = 9.194822849 \text{ mol NO}$$

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

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

34.086 g/mol
-20.50

Heat of formation / enthalpy of formation!

0

-285.8

-296.8

ΔH_f° , kJ/mol

(Appendix C)



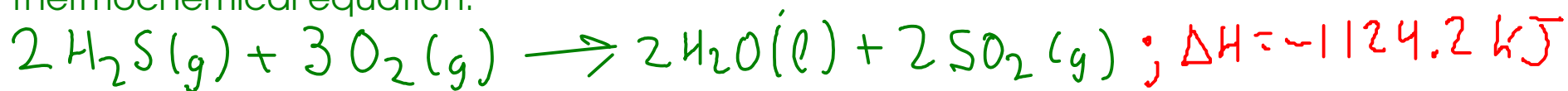
What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

- 1 - Calculate the enthalpy change of the reaction AS WRITTEN using Hess' Law
- 2 - Convert 25.0 g hydrogen sulfide to moles using formula weight.
- 3 - Convert moles hydrogen sulfide to enthalpy change using thermochemical equation.

$$\Delta H = \sum \Delta H_f^\circ, \text{products} - \sum \Delta H_f^\circ, \text{reactants}$$

$$\begin{aligned} \textcircled{1} &= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)] \\ &= -1124.2 \text{ kJ} \end{aligned}$$

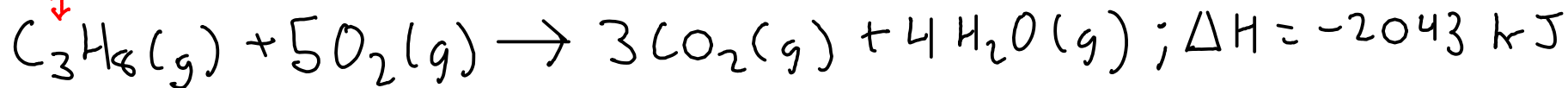
Thermochemical equation:



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

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

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 is providing the energy, the energy requirement has a NEGATIVE sign - from the point of view of the REACTION!

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

$$PV = nRT$$

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

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

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

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

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

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

END OF CHAPTER 6