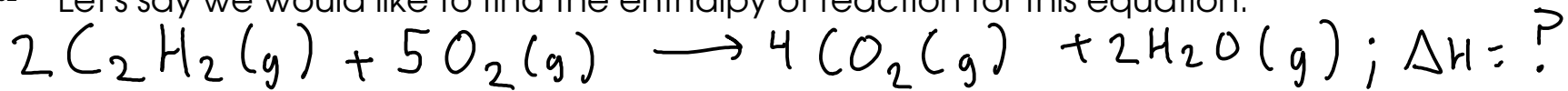
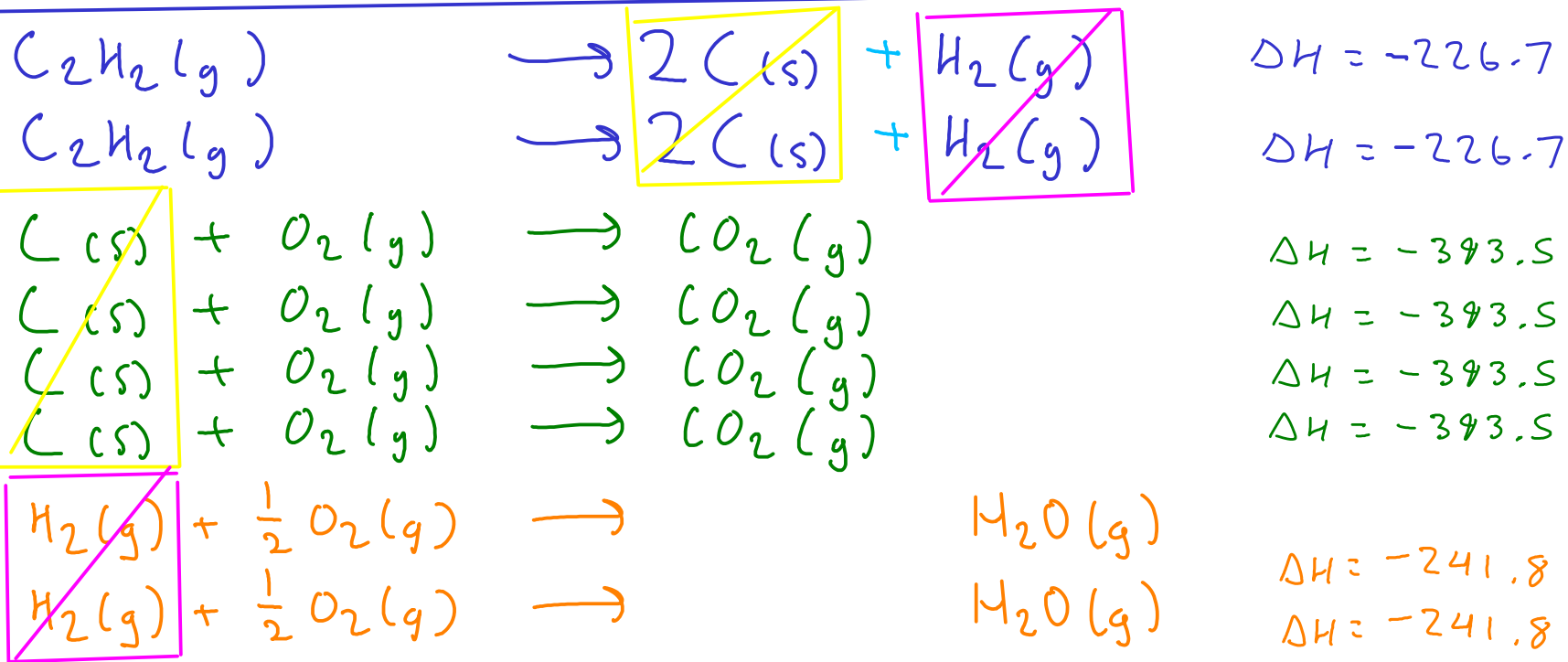
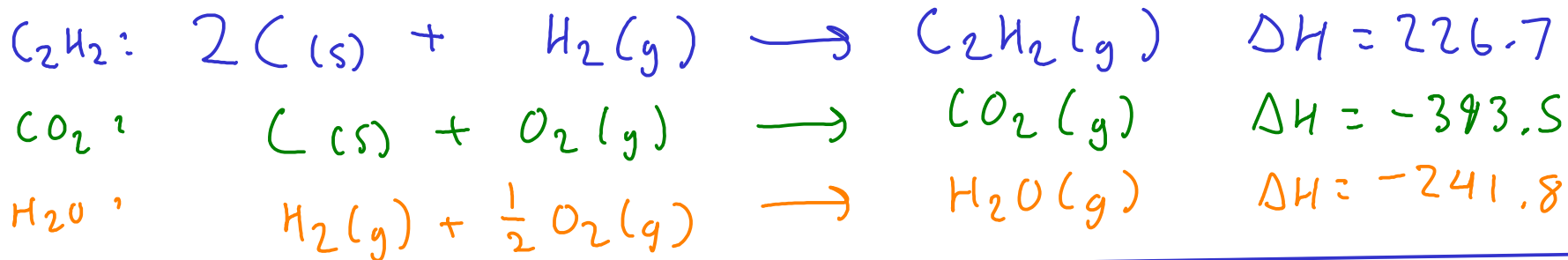


<sup>162</sup> Let's say we would like to find the enthalpy of reaction for this equation:



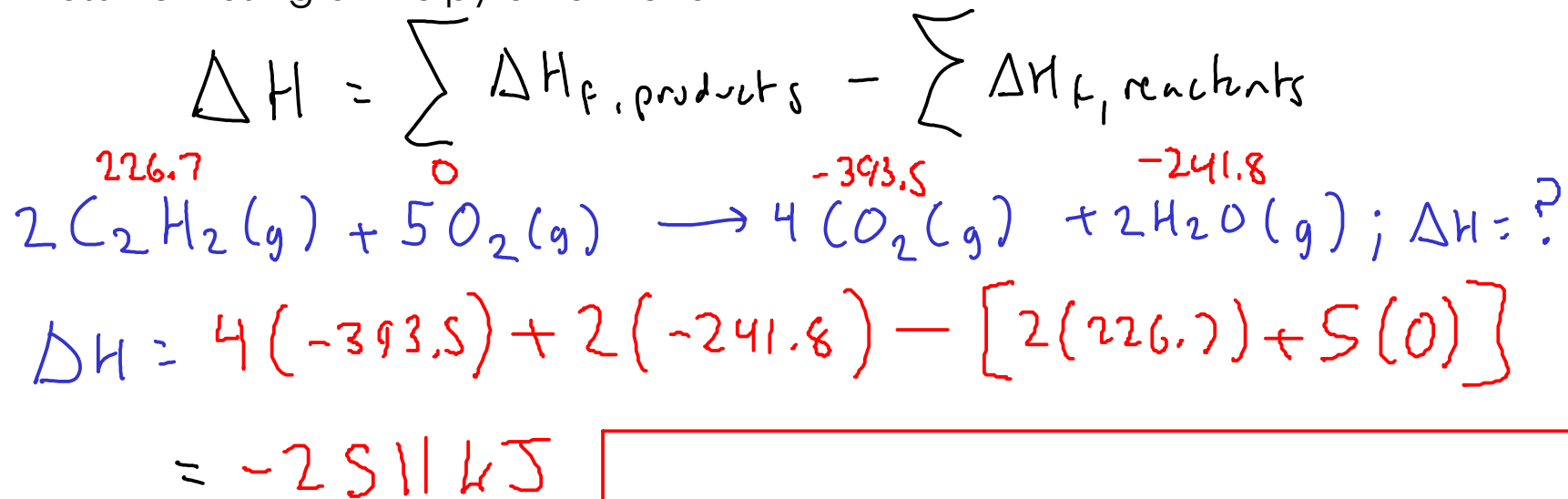
From  
A-8,  
text  
↓

Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.



$$\Delta H = 2(-226.7) + 4(-393.5) + 2(-241.8) = \boxed{-2511 \text{ kJ}}$$

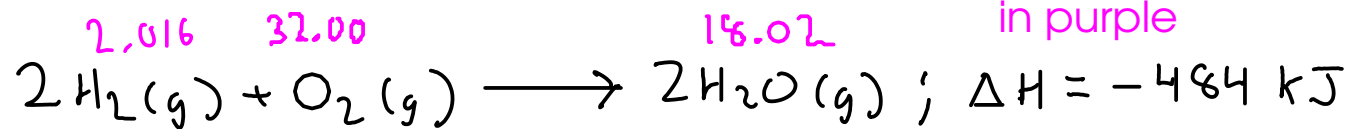
Hess' Law using enthalpy of formation:



See Appendix C in the textbook for enthalpy of formation data:  
p A-8 to A-11

\* Remember:

- Multiply each enthalpy by its stoichiometric coefficient from the reaction
- Enthalpy of formation of an element at its standard state is zero
- Watch phase labels. You will usually find SEVERAL enthalpies of formation for a given substance in different phases!
- For ionic substances in solution, remember that they exist as free ions, so look up the aqueous IONS!



Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

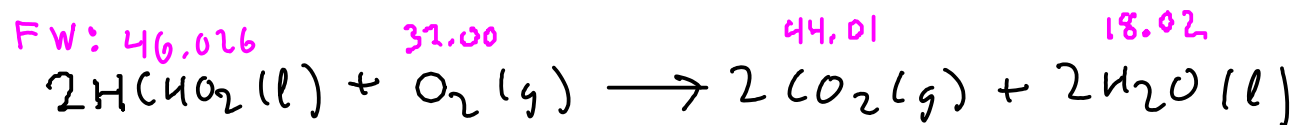
1 - Convert 1.00 kg hydrogen gas to moles using FORMULA WEIGHT

2 - Convert moles hydrogen gas to enthalpy change using THERMOCHEMICAL EQUATION

$$2.016 \text{ g H}_2 = 1 \text{ mol H}_2 \quad | \quad 2 \text{ mol H}_2 = -484 \text{ kJ} \quad | \quad \text{kg} = 10^3 \text{ g}$$

$$1.00 \text{ kg H}_2 \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{-484 \text{ kJ}}{2 \text{ mol H}_2} = \boxed{-120000 \text{ kJ per kg H}_2}$$

①                      ②



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?

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

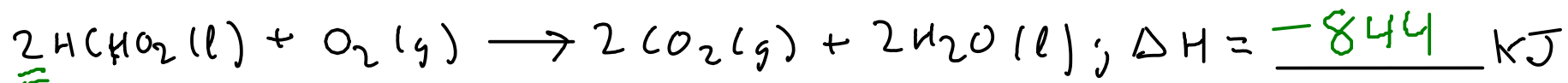
Find moles formic acid ... we know everything else.

$$46.026 \text{ g HCHO}_2 = \text{mol HCHO}_2$$

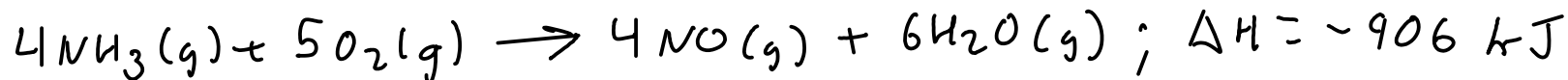
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

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

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



We calculated the enthalpy change PER MOLE of acid (which is typical in a calorimetry experiment), but the enthalpy term in the thermochemical equation is per TWO MOLES of acid. So we must multiply the calorimetry result by two for the thermochemical equation.



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 NO to moles using IDEAL GAS EQUATION

2 - Convert moles NO enthalpy change using THERMOCHEMICAL EQUATION

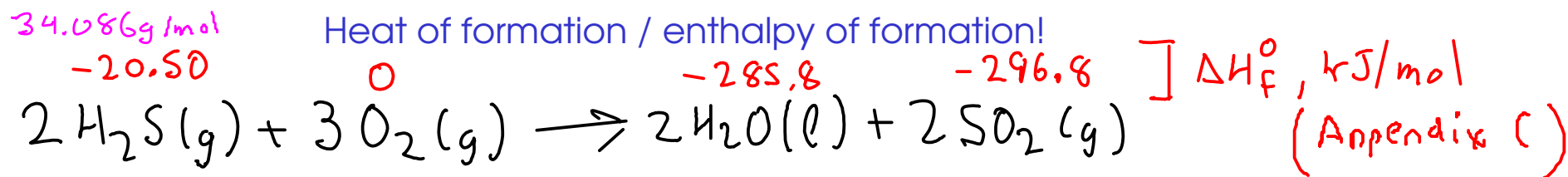
$$\textcircled{1} PV = nRT \quad \left| \quad P = 1.50 \text{ atm} \quad V = 150. \text{ L}$$

$$n = \frac{PV}{RT} \quad \left| \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \quad T = 25.0^\circ\text{C} = 298.2 \text{ K}$$

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

$$\textcircled{2} 4 \text{ mol NO} = -906 \text{ kJ}$$

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



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

- 1 - Calculate the enthalpy of the reaction as written using FORMATION HEATS
- 2 - Convert mass hydrogen sulfide to moles using formula weight.
- 3 - Convert moles hydrogen sulfide to enthalpy change using THERMOCHEMICAL EQUATION

$$\begin{aligned} \Delta H &= \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants} \\ &= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)] \\ &= -1124.2 \text{ kJ} \text{ (1)} \end{aligned}$$

So the THERMOCHEMICAL EQUATION is ...



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

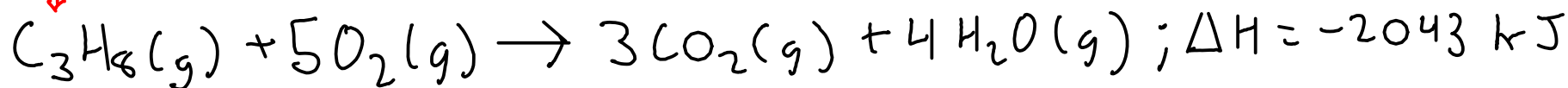
$$25.0 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.086 \text{ g H}_2\text{S}} \times \frac{-1124.2 \text{ kJ}}{2 \text{ mol H}_2\text{S}} = \boxed{-412 \text{ kJ}}$$

(2)                      (3)

The next page is an extra example problem for Chapter 6. This problem is here to show you that you can sometimes start with heat/energy/enthalpy requirements in a chemical calculation instead of a substance.

In the example problem, we start with an energy requirement (565 kJ). While you can't convert that to moles of energy (since that makes no sense!), you CAN use the thermochemical equation to convert directly from the energy requirement to moles of one of the substances - propane.

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 the energy requirement to moles propane using THERMOCHEMICAL EQUATION

2 - Convert moles propane to volume using IDEAL GAS EQUATION

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

Since the reaction is the SYSTEM here, the energy requirement has a NEGATIVE sign (it's LOSING, or providing) the energy...

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

$$PV = nRT$$

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

$$P = 1.08 \text{ atm} \quad R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

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

$$n = 0.2765540871 \text{ mol C}_3\text{H}_8$$

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

$$= 6.27 \text{ L C}_3\text{H}_8 @ \begin{matrix} 25.0^\circ\text{C} \\ 1.08 \text{ atm} \end{matrix}$$