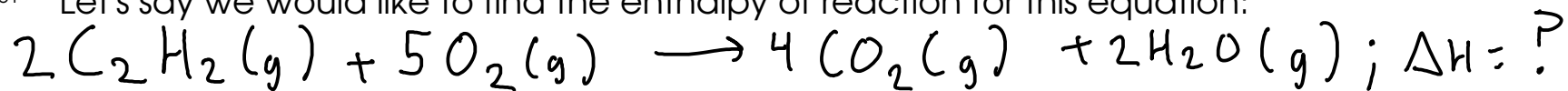
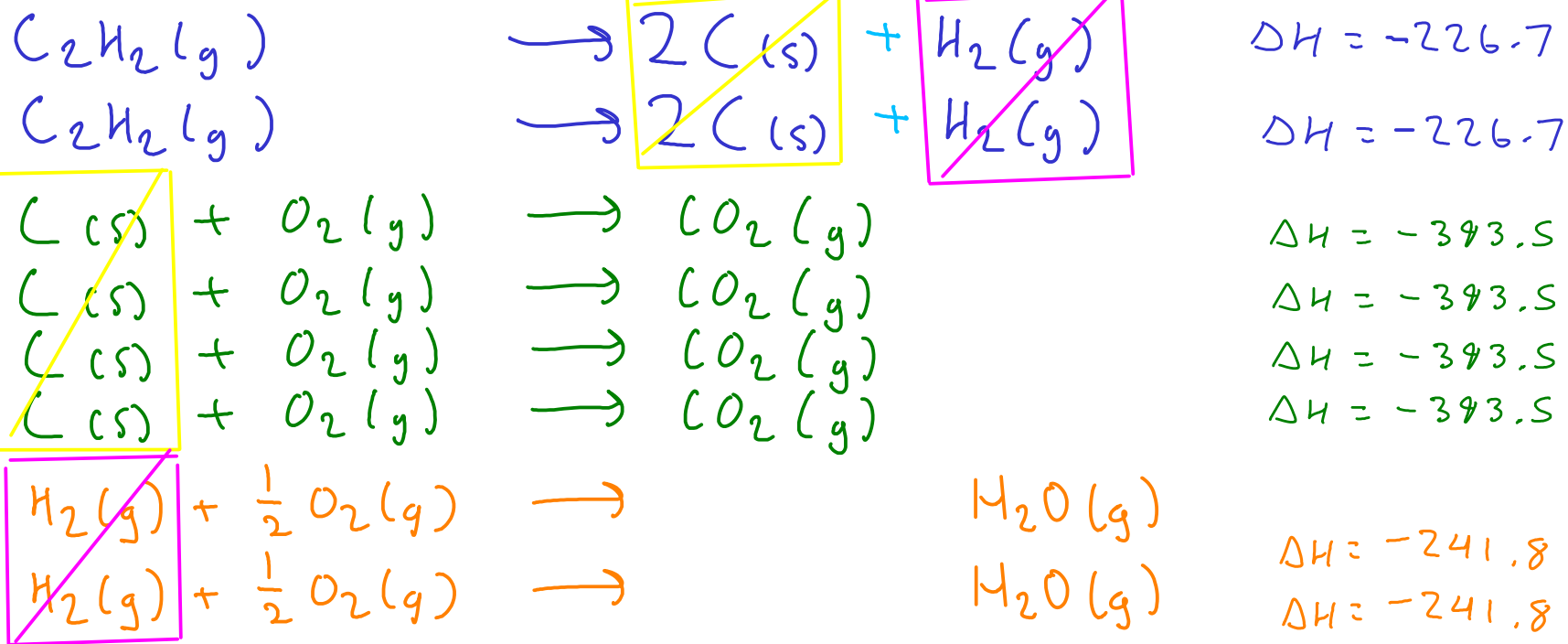
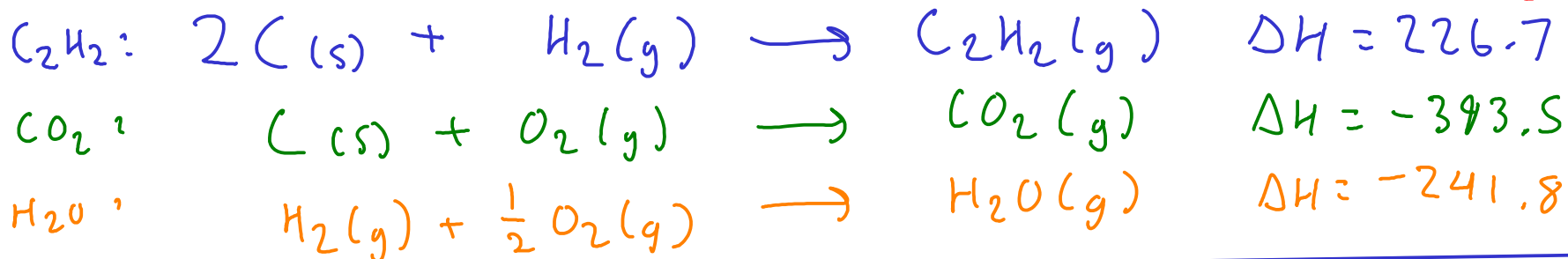


¹⁶¹ Let's say we would like to find the enthalpy of reaction for this equation:



From
A-8,
text
↓

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



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

$$\Delta H = \sum \Delta H_{f, \text{products}} - \sum \Delta H_{f, \text{reactants}}$$

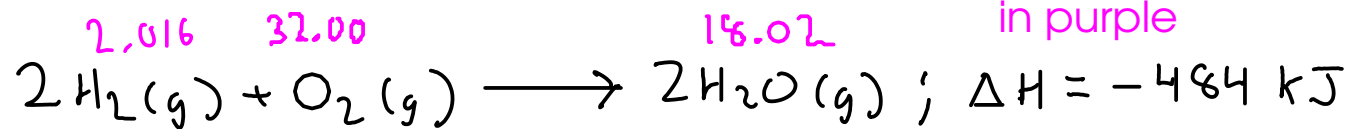
$$\Delta H = \left(\overset{\text{CO}_2}{4(-393.5)} + \overset{\text{H}_2\text{O}}{2(-241.8)} \right) - \left(2(226.7) + \text{S}(0) \right)$$

$$= \boxed{-2511 \text{ kJ}}$$

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!



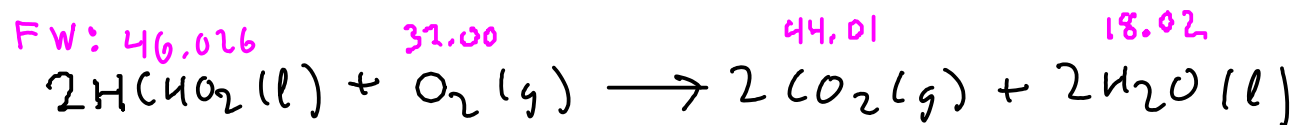
Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

1 - Convert 1.00 kg of hydrogen gas to moles using formula weight.

2 - Convert moles hydrogen gas to enthalpy change using thermochemical equation.

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

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



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?

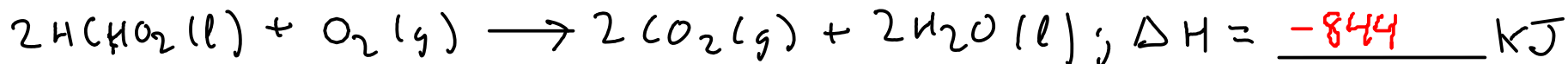
$$Q = -50.3 \text{ kJ} ; \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2}$$

Find moles formic acid:

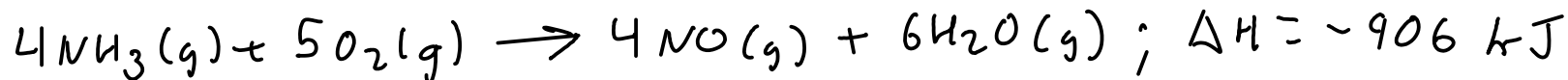
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

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

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



We calculated the enthalpy change PER MOLE of formic acid. This equation is based on TWO MOLES of formic acid. So, we double the enthalpy change above for two moles.



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 of NO to enthalpy change using thermochemical equation.

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

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

$$\Delta H_{\text{mol NO}} = -906 \text{ kJ}$$

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

34.086 g/mol
-20.50

Heat of formation / enthalpy of formation!

0

-285.8

-296.8

ΔH_f° , kJ/mol

(Appendix C)



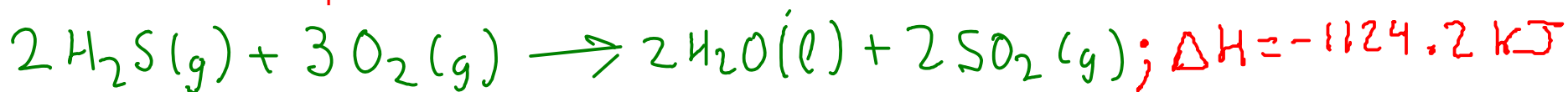
What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

- 1- Calculate the enthalpy change for the reaction as written using HEATS OF FORMATION.
- 2 - Convert 25.0 grams of hydrogen sulfide to moles using formula weight.
- 3 - Convert moles hydrogen sulfide to enthalpy change using thermochemical equation.

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

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

Thermochemical equation:



$$34.086 \text{ g H}_2\text{S} = 1 \text{ mol H}_2\text{S} \quad | \quad 2 \text{ mol H}_2\text{S} = -1124 \text{ kJ}$$

$$25.0 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.086 \text{ g H}_2\text{S}} \times \frac{-1124 \text{ kJ}}{2 \text{ mol H}_2\text{S}} = \boxed{-412 \text{ kJ}}$$

$\textcircled{2}$

$\textcircled{3}$