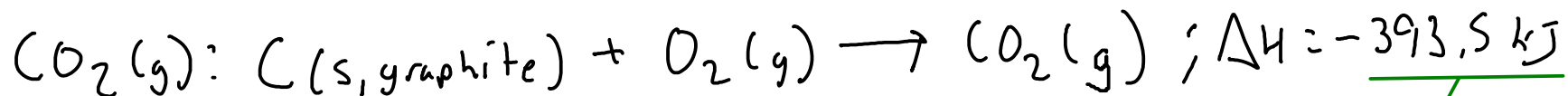
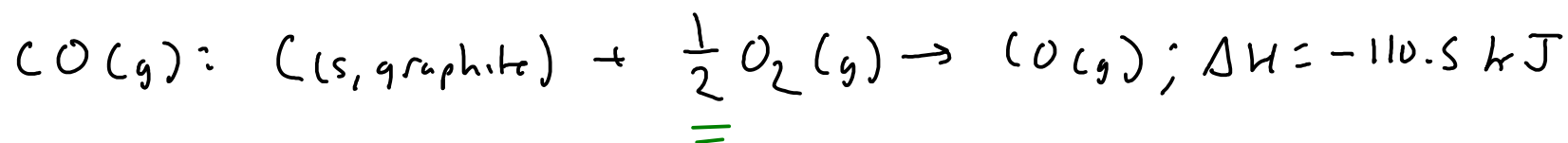


160 FORMATION REACTIONS

- A reaction that forms exactly one mole of the specified substance from its elements at their STANDARD STATE at 25C and 1 atm pressure.



heat of formation of carbon dioxide ΔH_f° or ΔH_f
"enthalpy of formation"



you may see fractional coefficients in these formation reactions, because you MUST form exactly one mole of the product!

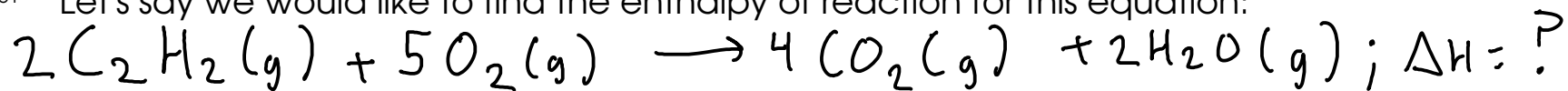
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

$$\Delta H_f^\circ, \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$$

- What are formation reactions good for?

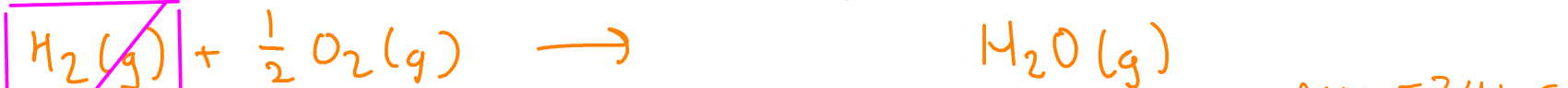
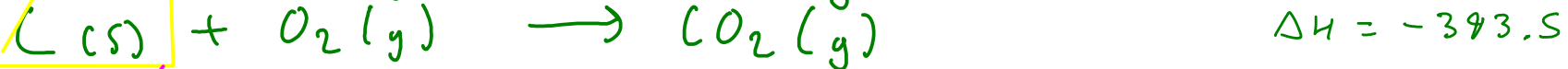
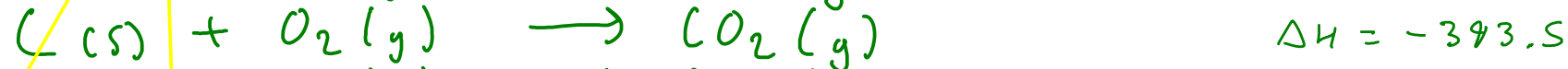
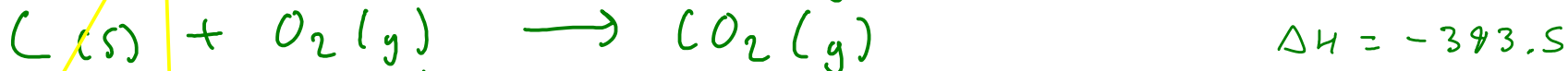
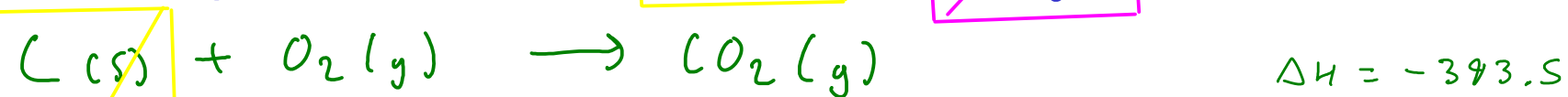
... finding enthalpies for more interesting reactions!

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.



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

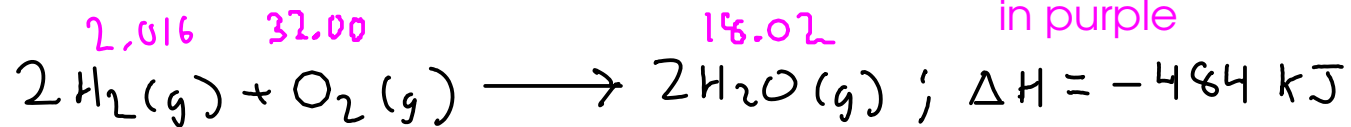
$$\Delta H = \sum \Delta H_{f, \text{products}} - \sum \Delta H_{f, \text{reactants}}$$

$$\Delta H = [4(-393.5) + 2(-241.8)] - [2(226.7) + 5(0)]$$
$$= \boxed{-2511 \text{ kJ}}$$

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!



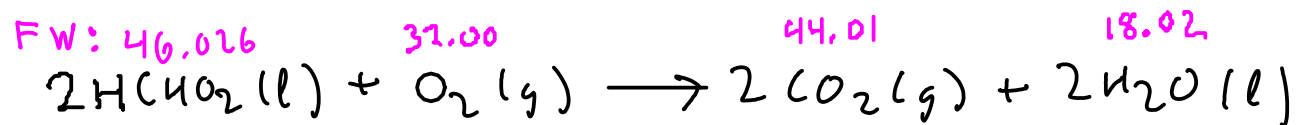
Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

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

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

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

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



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?

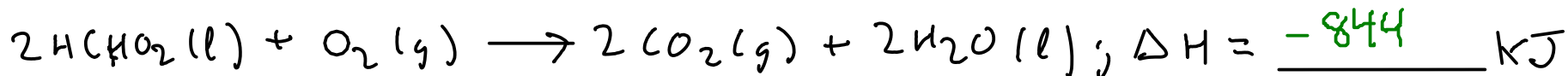
$$Q = -50.3 \text{ kJ} ; \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2}$$

Find moles formic acid...

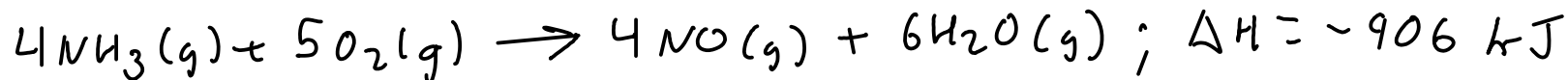
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

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

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



Since the equation is based on TWO moles of formic acid, we need to multiply the enthalpy change obtained in the calorimetry problem by two to find the enthalpy change of the reaction as written.



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.

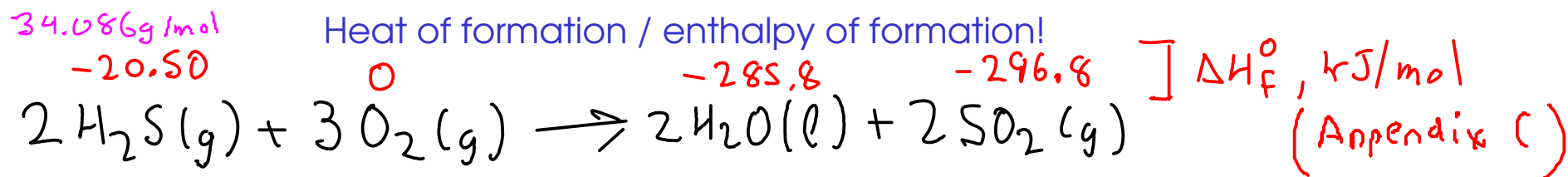
$$PV = nRT \quad \left| \quad \begin{array}{l} P = 1.50 \text{ atm} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ V = 150. \text{ L} \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \end{array} \right.$$

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

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

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

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



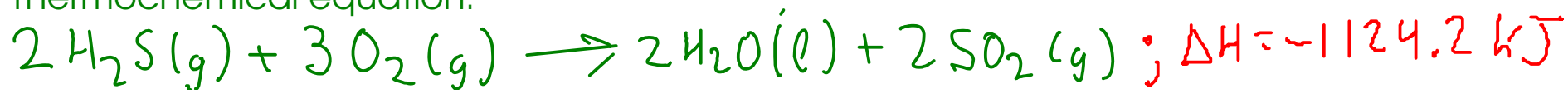
What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

- 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.

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

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

Thermochemical equation:



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

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