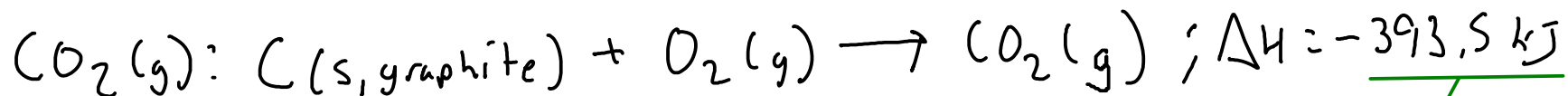
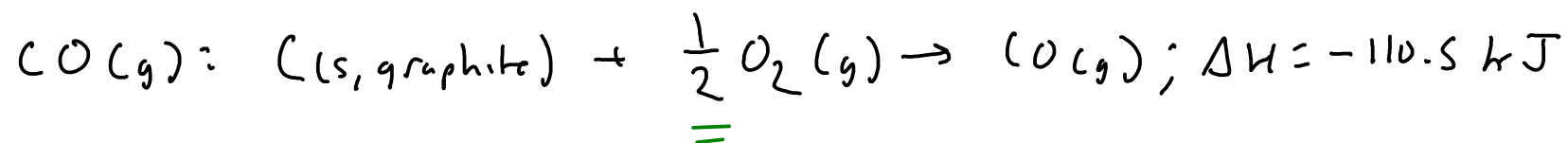


161 FORMATION REACTIONS

- 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!

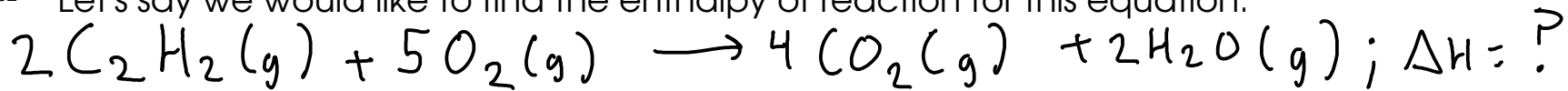
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

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

- What are formation reactions good for?

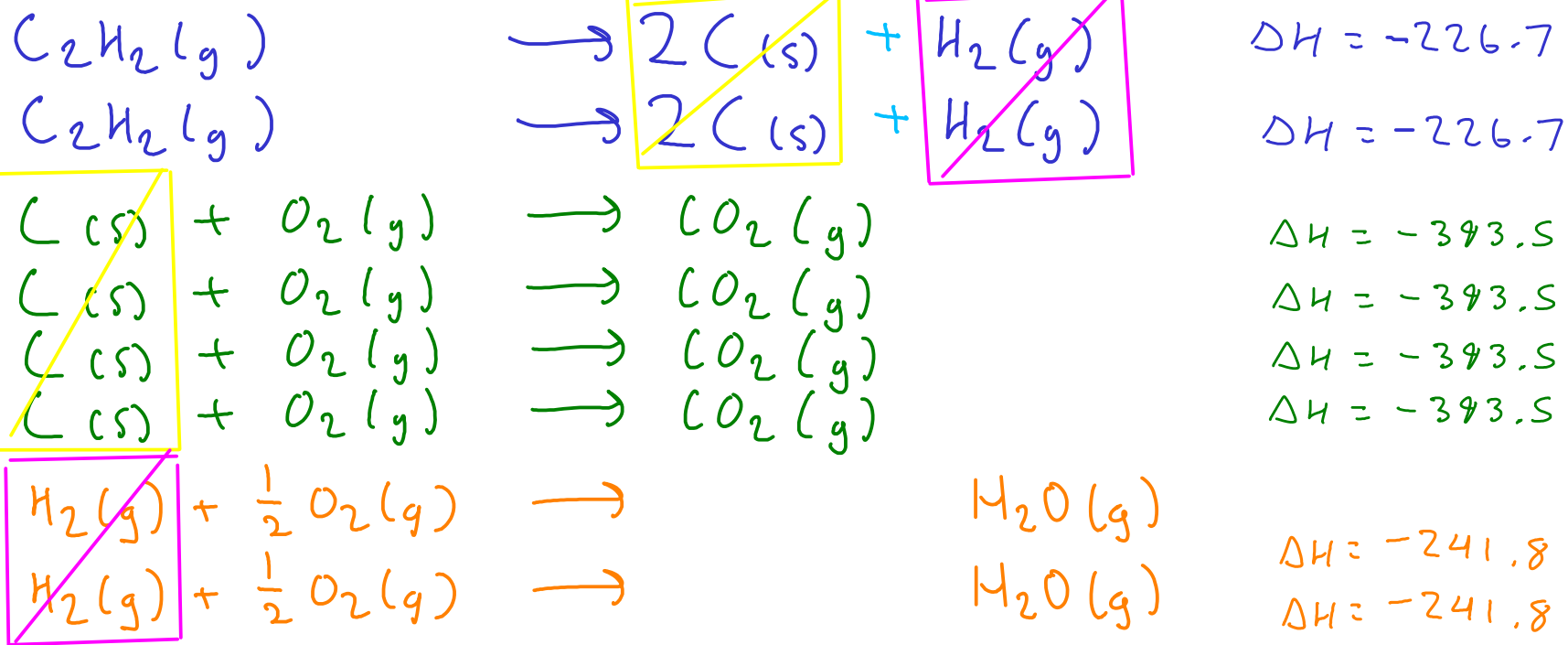
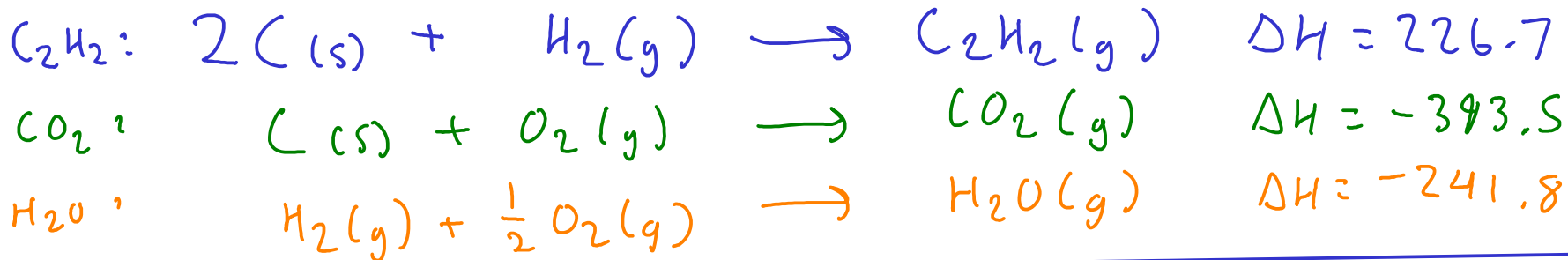
... finding enthalpies for more interesting reactions!

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



From
A-8,
text
↓

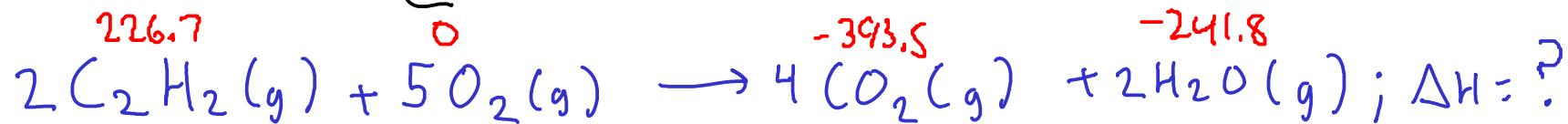
Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.



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

Hess' Law using enthalpy of formation:

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



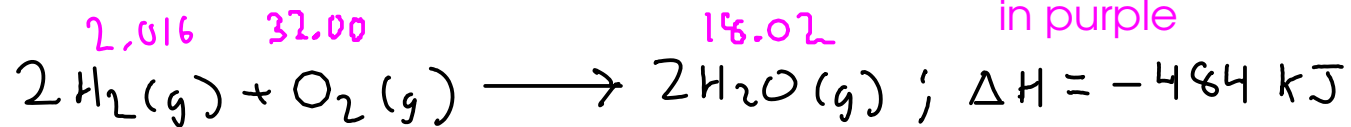
$$\Delta H = [4(-393.5) + 2(-241.8)] - [2(226.7) + 5(0)]$$

$$= -2811 \text{ kJ}$$

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.

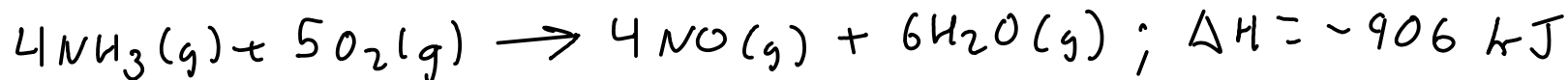
1 - Convert 1.00 kg hydrogen gas to moles. Use FORMULA WEIGHT.

2 - Convert moles hydrogen gas to enthalpy change. Use THERMOCHEMICAL EQUATION

$$\textcircled{1} \quad 2.016 \text{ g H}_2 = 1 \text{ mol H}_2 \quad \text{kg} = 10^3 \text{ g}$$

$$\textcircled{2} \quad 2 \text{ mol H}_2 = -484 \text{ kJ}$$

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



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 to enthalpy change using THERMOCHEMICAL EQUATION

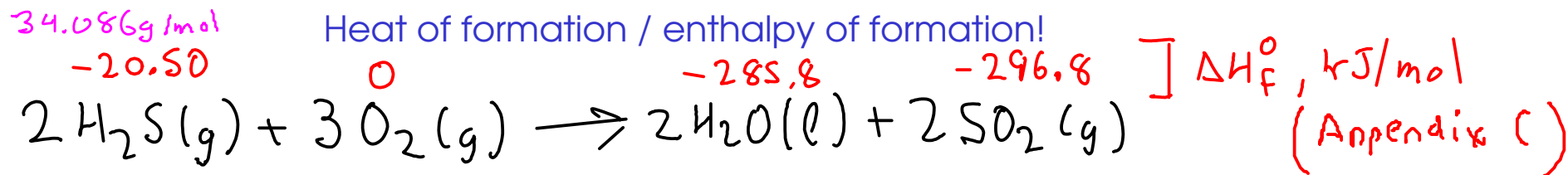
$$\textcircled{1} \quad PV = nRT \quad \left| \quad P = 1.50 \text{ atm} \quad V = 150. \text{ L} \right.$$

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

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

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

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



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

- 1 - Calculate the enthalpy change for the reaction as written. Use Hess's Law.
- 2 - Convert 25.0 grams hydrogen sulfide to moles using FORMULA WEIGHT.
- 3 - Convert moles hydrogen sulfide to enthalpy change using THERMOCHEMICAL EQUATION.

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

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

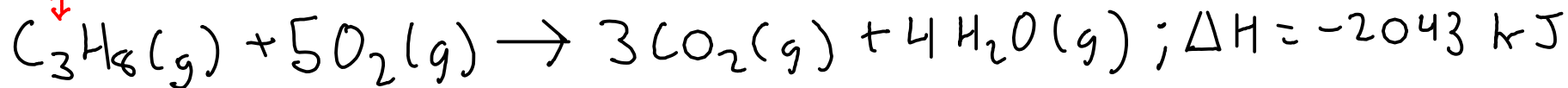
So the THERMOCHEMICAL EQUATION is ...



$$\textcircled{2} \quad 34.086 \text{ g H}_2\text{S} = 1 \text{ mol H}_2\text{S} \qquad \textcircled{3} \quad 2 \text{ mol H}_2\text{S} = -1124.2 \text{ kJ}$$

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

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 the 565 kJ energy requirement to moles propane using THERMOCHEMICAL EQUATION
- 2 - Convert moles propane to volume using IDEAL GAS EQUATION

$$\textcircled{1} \text{ mol C}_3\text{H}_8 = -2043 \text{ kJ}$$

Since the reaction is the system here, the energy requirement gets a negative sign. The reaction gives up this amount of energy!

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

$$\textcircled{2} \begin{array}{l} PV = nRT \\ V = \frac{nRT}{P} \end{array} \quad \begin{array}{l} P = 1.08 \text{ atm} \\ T = 25.0^\circ\text{C} = 298.2 \text{ K} \\ n = 0.2765540871 \text{ mol C}_3\text{H}_8 \end{array} \quad \begin{array}{l} R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \end{array}$$

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

$$= \boxed{6.27 \text{ L of C}_3\text{H}_8 \text{ @ } 25.0^\circ\text{C} + 1.08 \text{ atm}}$$

END OF CHAPTER 6