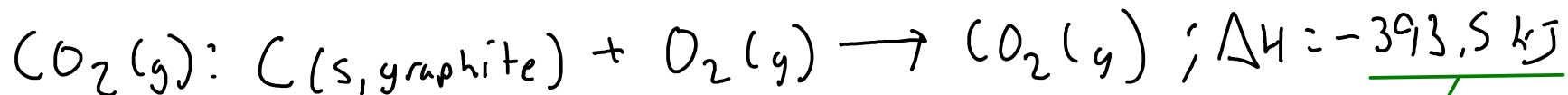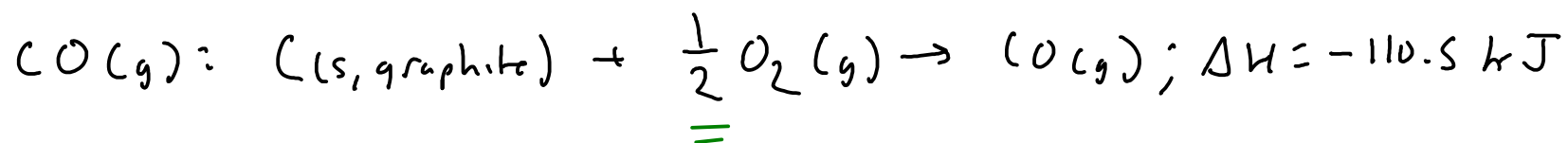FORMATION REACTIONS

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

$$CO_2(g): C(s, graphite) + O_2(g) \longrightarrow CO_2(g) \; ; \Delta H = -393.5 \, kJ$$

heat of formation of carbon dioxide ΔH_f^0 or ΔH_f
"enthalpy of formation"

$$CO(g): C(s, graphite) + \tfrac{1}{2}O_2(g) \rightarrow CO(g) \; ; \Delta H = -110.5 \, kJ$$

you may see fractional coefficients in these formation reactions, because you MUST form exactly one mole of the product!
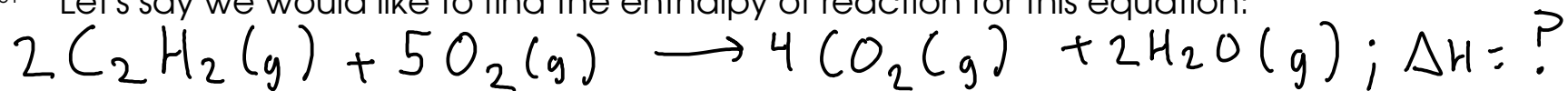
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.
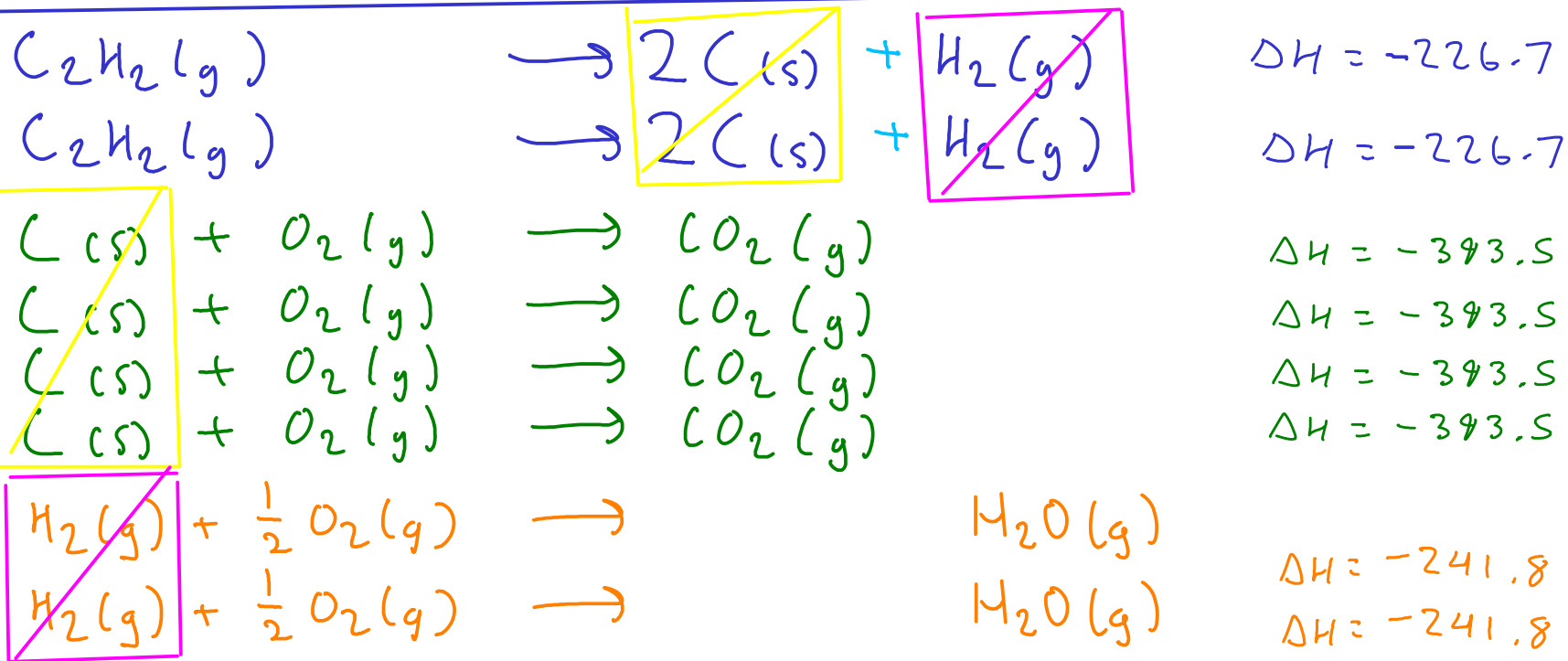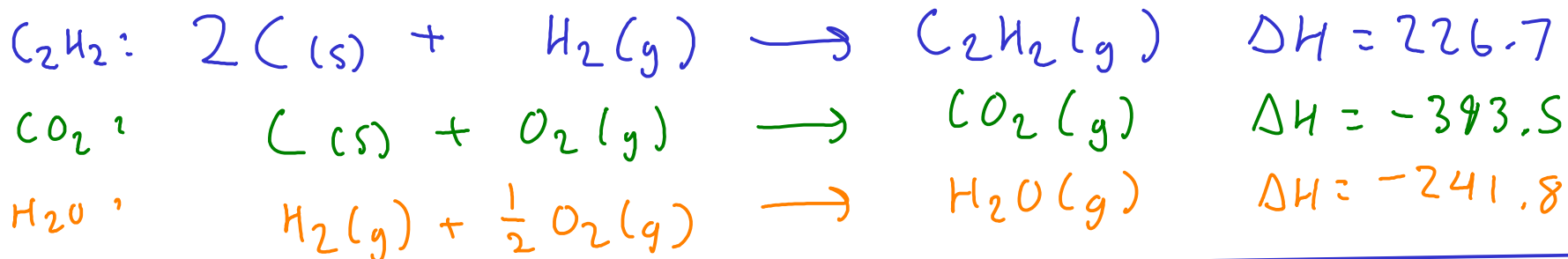
$$\Delta H_{f, O_2(g)}^0 = 0 \, kJ/mol$$

- What are formation reactions good for?

... finding enthalpies for more interesting reactions!

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

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g); \Delta H = ?$$

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

C_2H_2: $\quad 2C(s) + H_2(g) \longrightarrow C_2H_2(g) \quad \Delta H = 226.7$

CO_2: $\quad C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5$

H_2O: $\quad H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \quad \Delta H = -241.8$

$C_2H_2(g) \longrightarrow 2C(s) + H_2(g) \quad \Delta H = -226.7$

$C_2H_2(g) \longrightarrow 2C(s) + H_2(g) \quad \Delta H = -226.7$

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5$

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5$

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5$

$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -393.5$

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \quad \Delta H = -241.8$

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \quad \Delta H = -241.8$

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(\ell)$$

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

Hess' Law using enthalpy of formation:

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

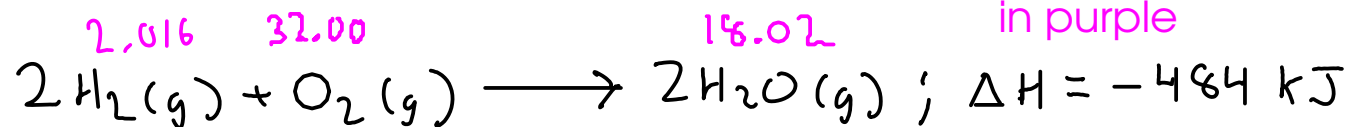$$\Delta H = \left(4(-393.5) + 2(-241.8) \right) - \left(2(226.7) + 5(0) \right)$$

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

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!
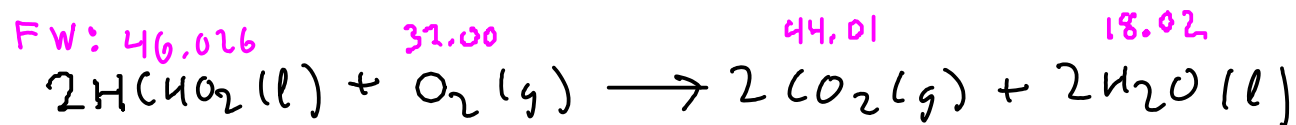
$$\Delta H_{f}, NaCl_{(aq)} = \Delta H_{f}, Na^{+}_{(aq)} + \Delta H_{f}, Cl^{-}_{(aq)}$$

Example problems:

$$2.016 \qquad 32.00 \qquad\qquad\qquad 18.02$$

$$2 H_2 (g) + O_2 (g) \longrightarrow 2 H_2O (g) \;; \Delta H = -484 \; kJ$$

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

1 - Convert 1.0 kg of hydrogen gas to moles using formula weight
2 - Convert moles hydrogen gas to enthalpy change using thermochemical equation

$$2.016 \, _g H_2 = mol \; H_2 \quad \Big| \quad 2 \, mol \; H_2 = -484 \; kJ \quad \Big\} \quad kg = 10^3 \, _g$$

$$1.0 \; kg \; H_2 \times \frac{10^3 \, _g}{kg} \times \frac{mol \; H_2}{2.016 \, _g \, H_2} \times \frac{-484 \; kJ}{2 \, mol \; H_2} = \boxed{-120000 \; kJ \; per \; kg \; H_2}$$

FW: 46.026   32.00   44.01   18.02

$$2 HCHO_2 (l) + O_2 (g) \longrightarrow 2 CO_2 (g) + 2 H_2O (l)$$

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?
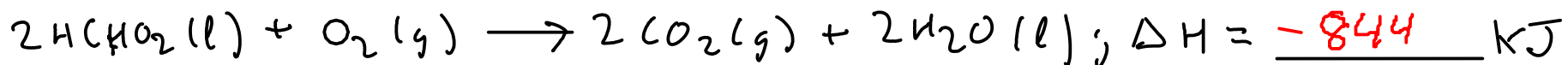
$$Q = -50.3 \text{ kJ} \qquad \Delta H = \frac{Q \text{ constant pressure}}{\text{mol } HCHO_2}$$

Find moles formic acid:

$$5.48 \text{ g } HCHO_2 \times \frac{\text{mol } HCHO_2}{46.026 \text{ g } HCHO_2} = 0.1190631382 \text{ mol } HCHO_2$$

$$\Delta H = \frac{-50.3 \text{ kJ}}{0.1190631382 \text{ mol } HCHO_2} = \boxed{-422 \text{ kJ} / \text{mol } HCHO_2}$$

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

$$2 HCHO_2 (l) + O_2 (g) \longrightarrow 2 CO_2 (g) + 2 H_2O (l) ; \Delta H = \underline{-844} \text{ kJ}$$

We calculated the enthalpy change per mole of formic acid, but this equation is written based on TWO MOLES of formic acid. So we need to double the experimentally-obetained value of enthalpy change for this equation.

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g) \ ; \ \Delta H = \sim 906 \ kJ$$
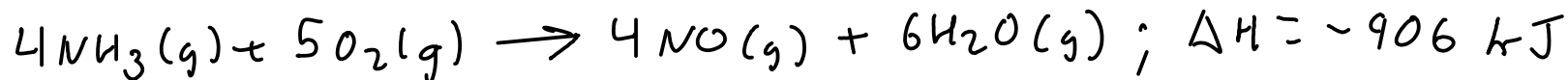
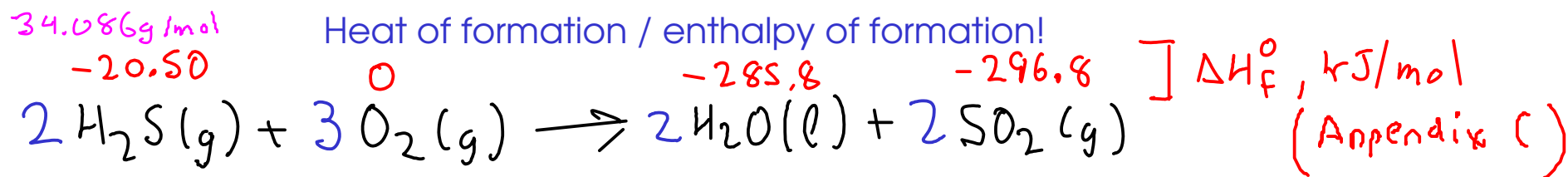What is the enthalpy change when 150. L of nitrogen monoxide are formed by this reaction at 25.0 C and 1.50 atm pressure?

1 - Convert 150. L of NO to moles NO using ideal gas equation.
2 - Convert moles NO to enthalpy change using thermochemical equation

$PV = nRT$ | $P = 1.50 \ atm$ $\qquad\qquad R = 0.08206 \ \dfrac{L \cdot atm}{mol \cdot K}$

$n = \dfrac{PV}{RT}$ | $V = 150. \ L$

$T = 25.0 \,°C = 298.2 \ K$

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

$4 \ mol \ NO = \sim 906 \ kJ$

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

34.086 g/mol

Heat of formation / enthalpy of formation!

$$-20.50 \qquad 0 \qquad\qquad -285.8 \qquad -296.8 \; \Big]\; \Delta H_f^0, \; kJ/mol$$

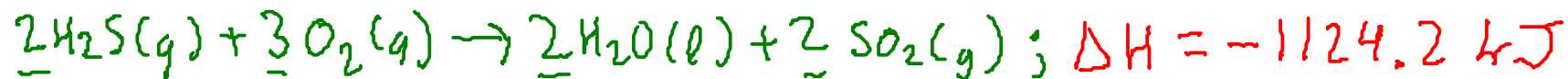$$2\,H_2S(g) + 3\,O_2(g) \longrightarrow 2\,H_2O(\ell) + 2\,SO_2(g) \qquad (Appendix\ C)$$

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

1 - Calculate the enthalpy of reaction from enthalpies of formation using Hess's Law.
2 - Convert 25.0 grams of hydrogen sulfide to moles using formula weight.
3 - Convert moles hydrogen sulfide to enthalpy change using the thermochemical equation that we found in #1

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① $\Delta H = \sum \Delta H_{f, \text{products}}^0 - \sum \Delta H_{f, \text{reactants}}^0$

$$= \Big[2(-285.8) + 2(-296.8) \Big] - \Big[2(-20.50) + 3(0) \Big]$$

$$= -1124.2 \; kJ$$

Thermochemical equation:

$$2\,H_2S(g) + 3\,O_2(g) \longrightarrow 2\,H_2O(\ell) + 2\,SO_2(g) \; ; \; \Delta H = -1124.2 \; kJ$$

$$34.086 \text{ g } H_2S = \text{mol } H_2S \; \Big| \; 2 \text{ mol } H_2S = -1124.2 \; kJ$$

$$25.0 \text{ g } H_2S \times \frac{\text{mol } H_2S}{34.086 \text{ g } H_2S} \times \frac{-1124.2 \; kJ}{2 \text{ mol } H_2S} = \boxed{-412 \; kJ}$$

propane

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g) \; ; \; \Delta H = -2043 \, kJ$$
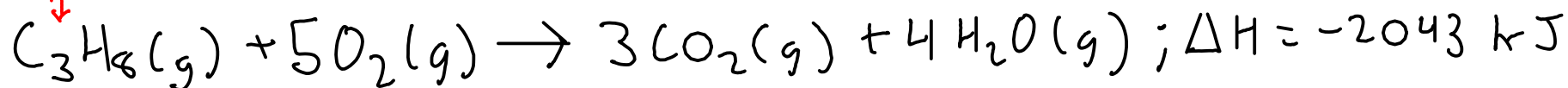
Calculate the volume of propane gas at 25.0 C and 1.08 atm required to provide 565 kJ of heat using the reaction above.

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1 - Convert energy requirement to moles propane using thermochemical equation.
2 - Convert moles propane to volume using ideal gas equation.

$$-2043 \, kJ = mol \; C_3H_8$$

① $-565 \, kJ \times \dfrac{mol \; C_3H_8}{-2043 \, kJ} = 0.2765540871 \, mol \; C_3H_8$

$$PV = nRT$$

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

$n = 0.2765540871 \, mol \; C_3H_8 \qquad T = 25.0°C = 298.2 \, K$

$R = 0.08206 \, \dfrac{L \cdot atm}{mol \cdot K} \qquad\qquad P = 1.08 \, atm$

$$V = \dfrac{(0.2765540871 \, mol \; C_3H_8)\left(0.08206 \, \frac{L \cdot atm}{mol \cdot K}\right)(298.2 \, K)}{(1.08 \, atm)}$$

$$= \boxed{6.27 \, L \; propane \; required}$$

END OF CHAPTER 6