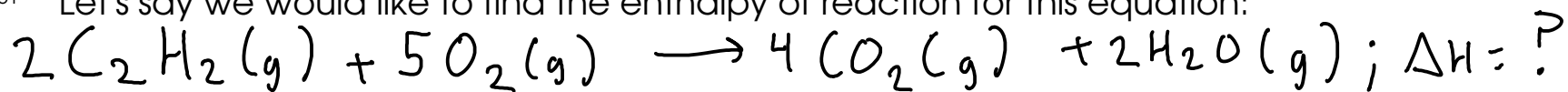Let's say we would like to find the enthalpy of reaction for this equation:

$$2 C_2 H_2 (g) + 5 O_2 (g) \longrightarrow 4 CO_2 (g) + 2 H_2 O (g); \Delta H = ?$$

2.

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

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

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

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

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

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

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$$2 C_2 H_2 (g) + 5 O_2 (g) \longrightarrow 4 CO_2 (g) + 2 H_2 O (\ell)$$
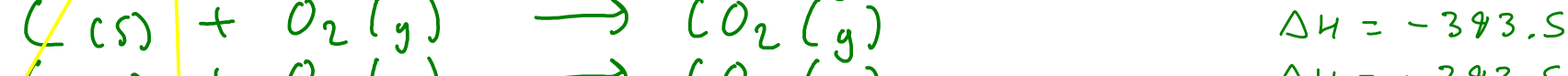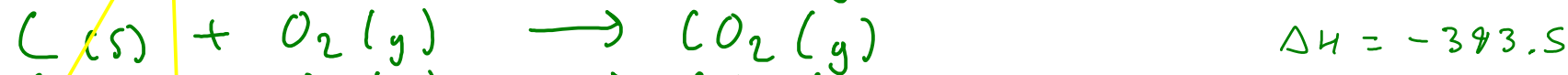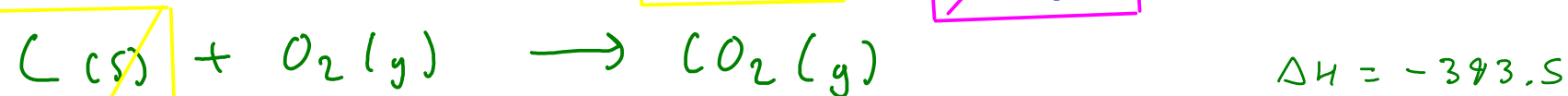
$$\Delta H = 2(-226.7) + 4(-393.5) + 2(-241.8) = \boxed{-2511 \text{ kJ}}$$

Hess' Law using enthalpy of formation:
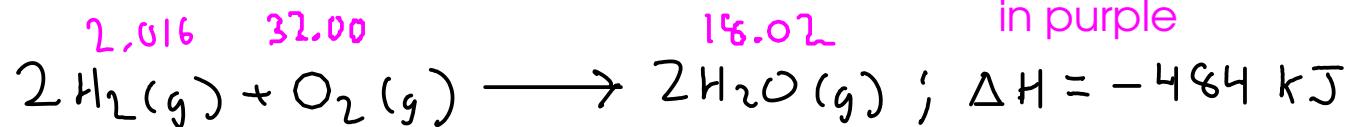
$$\Delta H = \sum \Delta H_{f, products} - \sum \Delta H_{f, reactants}$$

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

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

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!
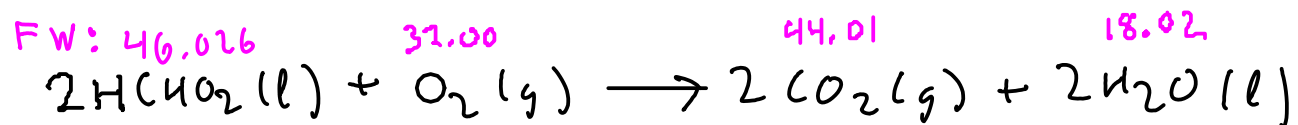
Example problems:

$$2.016 \qquad 32.00 \qquad\qquad\qquad 18.02$$
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \; ; \; \Delta H = -484 \; kJ$$

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 enthaply change using thermochemical equation

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$$2.016 \; g \; H_2 = mol \; H_2 \quad | \quad Kg = 10^3 g \quad | \quad 2 \, mol \; H_2 = -484 \; kJ$$

$$1.0 \; kg \; H_2 \times \frac{10^3 g}{Kg} \times \frac{mol \; H_2}{2.016 \; g \; H_2} \times \frac{-484 \; kJ}{2 \, mol \; H_2} = \boxed{\sim 120\,000 \; kJ \; per \; kg \; H_2}$$

FW: 46.026   32.00   44.01   18.02

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

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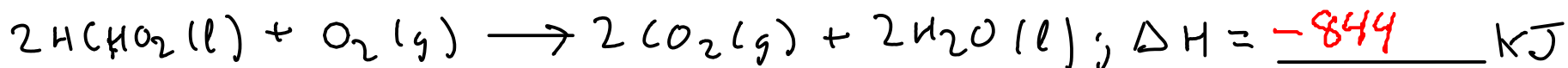
$$Q = -50.3 \text{ kJ} \quad ; \quad \Delta H = \frac{Q \text{ constant pressure}}{\text{mol } HCHO_2}$$

Find moles formic acid:

$$5.48 g \text{ } HCHO_2 \times \frac{\text{mol } HCHO_2}{46.026 g \text{ } HCHO_2} = 0.11906 \text{ mol } HCHO_2$$

$$\Delta H = \frac{-50.3 \text{ kJ}}{0.11906 \text{ mol } HCHO_2} = \boxed{-422 \text{ kJ}/\text{mol } HCHO_2}$$

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

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

We calculated the enthalpy change for a mole of formic acid being burned, and the equation has two moles being burned. Double the enthalpy change for two moles.

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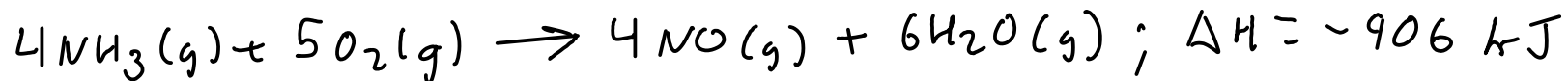
What is the enthalpy change when 150. L of nitrogen monoxide are formed by this reaction at 25.0 C and 1.50 atm pressure?

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

$$PV = nRT$$

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

$P = 1.50 \; atm$

$V = 150. \; L$

$T = 25.0°C = 298.2 \; K$

$R = 0.08206 \; \frac{L \cdot atm}{mol \cdot K}$

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

$$4 \; mol \; NO = -906 \; kJ$$

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

34.086g/mol

Heat of formation / enthalpy of formation!

$$\overset{-20.50}{2 H_2S(g)} + \overset{0}{3 O_2(g)} \longrightarrow \overset{-285.8}{2 H_2O(l)} + \overset{-296.8}{2 SO_2(g)} \quad] \Delta H_f^0, kJ/mol$$

(Appendix C)

What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?
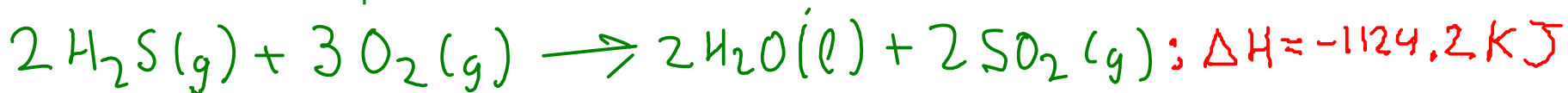
1 - Calculate the enthalpy change for the reaction as written using HEATS OF FORMATION and Hess's Law

2 - Convert 25.0 g of hydrogen sulfide to moles using formula weight

3 - Convert moles hydrogen sulfide to enthalpy change using thermochemical equation.

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

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

$$= -1124.2 \, kJ$$

Thermochemical equation:

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

34.086 g H_2S = mol H_2S | 2 mol H_2S = -1124.2 kJ

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

② ③