

One problem ...

PATH. The amount of energy required for a process depends on how the process is carried out.

Example: Driving from Florence to Columbia. How much energy is required? (gas)

2000 Jeep Cherokee vs 2008 Toyota Prius. The Jeep will use much more fuel than the Prius even though they start and end from exactly the same place. So the fuel usage is what we call a PATH FUNCTION, while the location is a STATE FUNCTION.

- so the heat of reaction depends on how the reaction is done.
- we need (for reporting) some kind of standard condition. At constant pressure, we can define a state function called ENTHALPY (H)

$$H = U + PV$$

$$\Delta H = Q_{\text{constant pressure}}$$

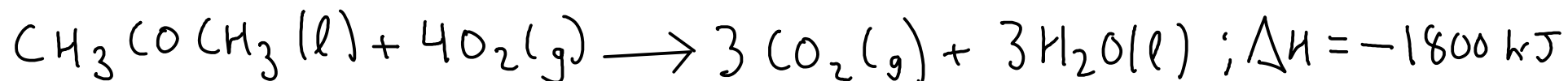
... we record the "enthalpy change of reaction" in our data books.

$$\Delta H_r$$

¹⁵⁸ SINCE the enthalpy change does NOT depend on path, this means that we can use standard values for enthalpy to predict the heat change in reactions that we have not tested in a calorimeter.

THERMOCHEMICAL EQUATIONS

- is like a regular chemical equation, except that phase labels are REQUIRED and the enthalpy for the reaction is given along with the equation.



- Why are phase labels required? Because phase changes either absorb or release energy.

$\Delta H = -1800 \text{ kJ}$... what does this mean?

$$1 \text{ mol CH}_3\text{COCH}_3 = -1800 \text{ kJ}$$

$$4 \text{ mol O}_2 = -1800 \text{ kJ}$$

$$3 \text{ mol CO}_2 = -1800 \text{ kJ}$$

$$3 \text{ mol H}_2\text{O} = -1800 \text{ kJ}$$

We treat the enthalpy change as if it's another product of the reaction!



What would be the enthalpy change when 25 g of water are produced by the reaction?

1 - Convert 25 g water to moles using formula weight.

2 - Convert moles water to enthalpy change using THERMOchemical equation.

$$18.016 \text{ g H}_2\text{O} = 1 \text{ mol H}_2\text{O} \quad | \quad 3 \text{ mol H}_2\text{O} = -1800 \text{ kJ}$$

$$25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{-1800 \text{ kJ}}{3 \text{ mol H}_2\text{O}} = -830 \text{ kJ}$$

This reaction is EXOTHERMIC: Energy is released from the reaction mixture to the surroundings.

This also equals Q, provided the pressure is constant.

(True for all combustions...)

A few more terms related to enthalpy:

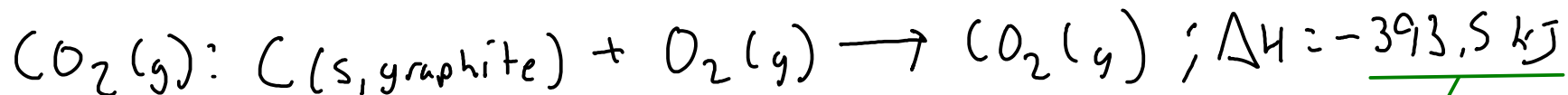
- Enthalpy of vaporization / heat of vaporization: The enthalpy change on vaporizing one mole of a substance. (from liquid to vapor)

- Enthalpy of fusion / heat of fusion: The enthalpy change when a mole of liquid changes to the solid state.

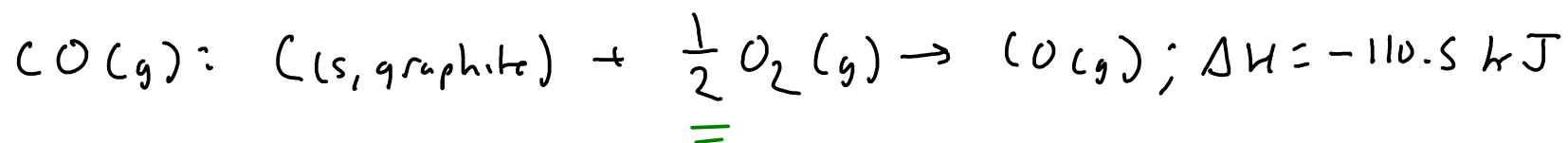
↑
Phase changes require energy, too!

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!

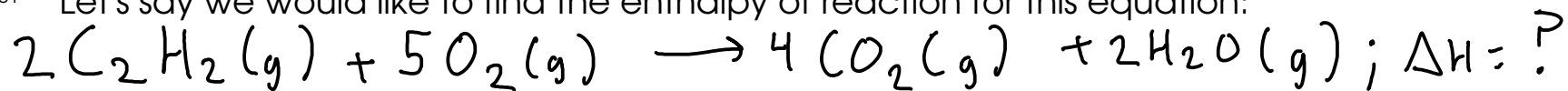
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

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

- What are formation reactions good for?

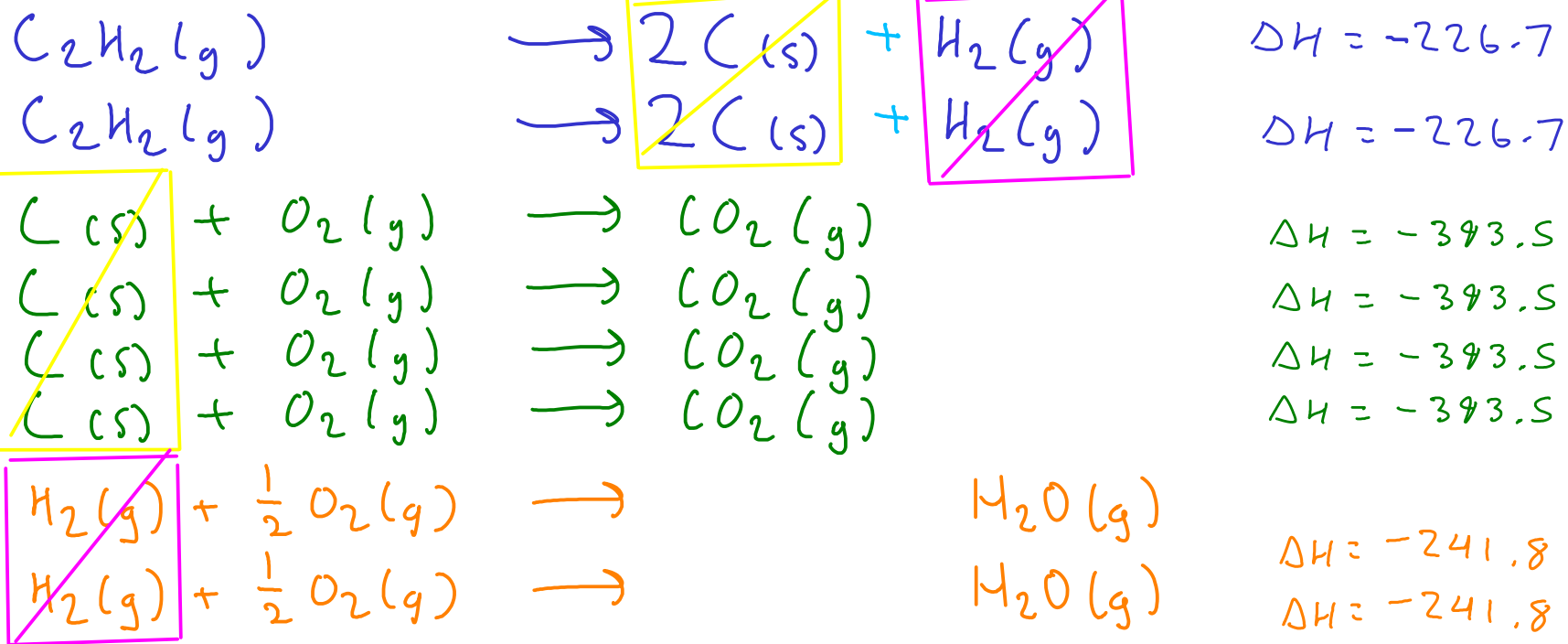
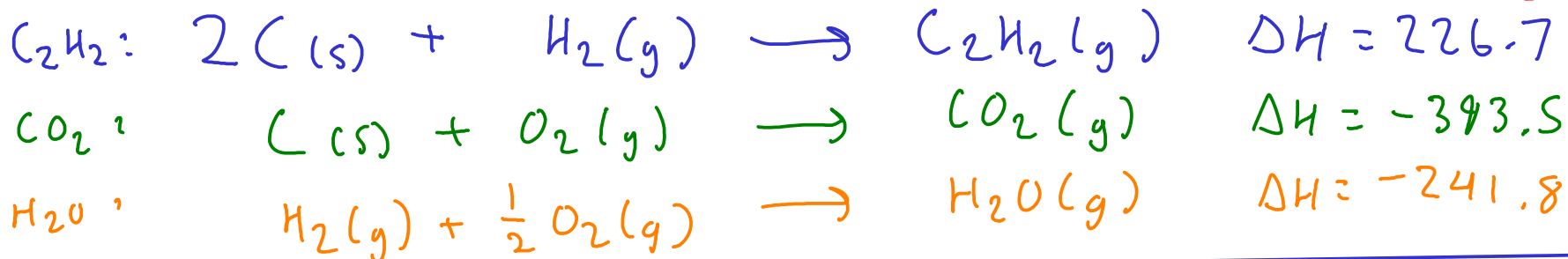
... finding enthalpies for more interesting reactions!

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.



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

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

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

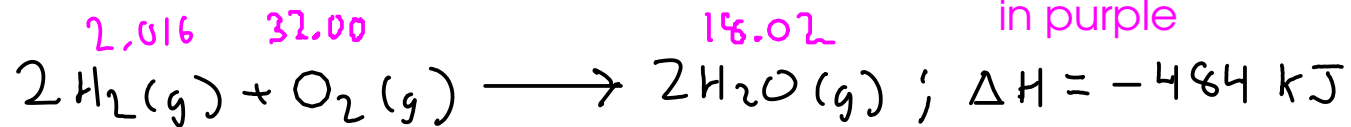
CO_2
 H_2O
 C_2H_2
 O_2

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

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.

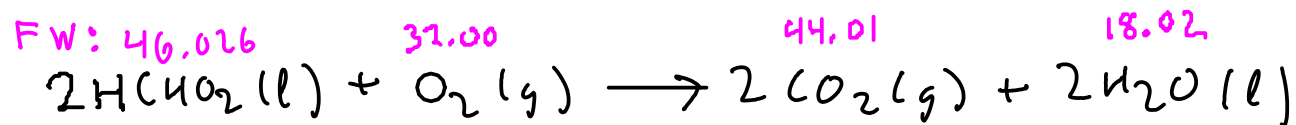
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- 1 - Convert 1.00 kg of hydrogen gas to moles using formula weight.
 - 2 - Convert moles hydrogen gas to enthalpy change using thermochemical equation.
-

$$2.016 \text{ g H}_2 = 1 \text{ mol H}_2$$

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

$$\text{kg} = 10^3 \text{ g}$$

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



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?

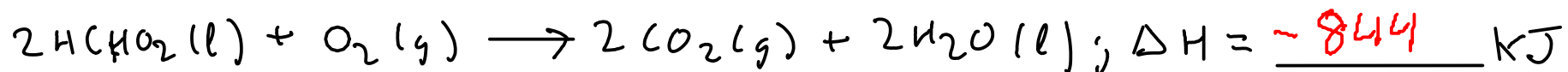
$$Q = -50.3 \text{ kJ} \quad ; \quad \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2}$$

Find moles formic acid: $46.026 \text{ g HCHO}_2 = 1 \text{ mol HCHO}_2$

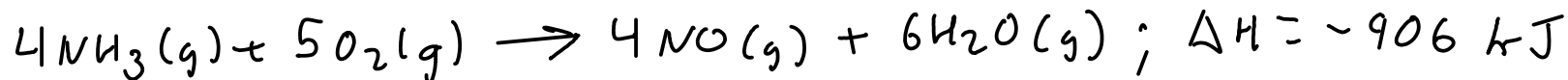
$$5.48 \text{ g HCHO}_2 \times \frac{1 \text{ mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

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

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



In the calorimetry experiment, we found the enthalpy change PER MOLE of formic acid, while the thermochemical equation is written per TWO moles of formic acid. So, we need to double the enthalpy change here ...



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.

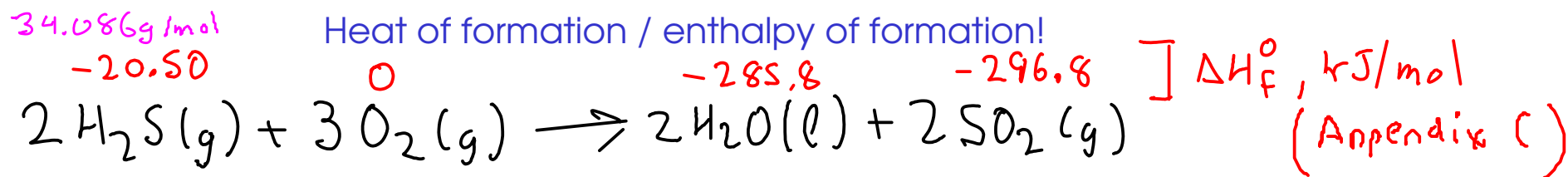
$$\textcircled{1} \quad PV = nRT \quad \left| \quad \begin{array}{l} P = 1.50 \text{ atm} \\ V = 150. \text{ L} \end{array} \right. \quad \begin{array}{l} R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \\ T = 25.0^\circ\text{C} = 298.2 \text{ K} \end{array}$$

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

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

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

$$\textcircled{2} \quad 9.194822849 \text{ mol NO} \times \frac{-906 \text{ kJ}}{4 \text{ mol NO}} = \boxed{-2080 \text{ kJ}}$$



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 HEATS OF FORMATION
- 2 - Convert 25.0 g of hydrogen sulfide to moles using formula weight.
- 3 - Convert moles hydrogen sulfide to enthalpy change using THERMOCHEMICAL equation.

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

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

Complete the thermochemical equation:



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

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

$\textcircled{2}$
 $\textcircled{3}$