

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)

Jeep Cherokee vs 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.

$$\Delta H_r$$

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.

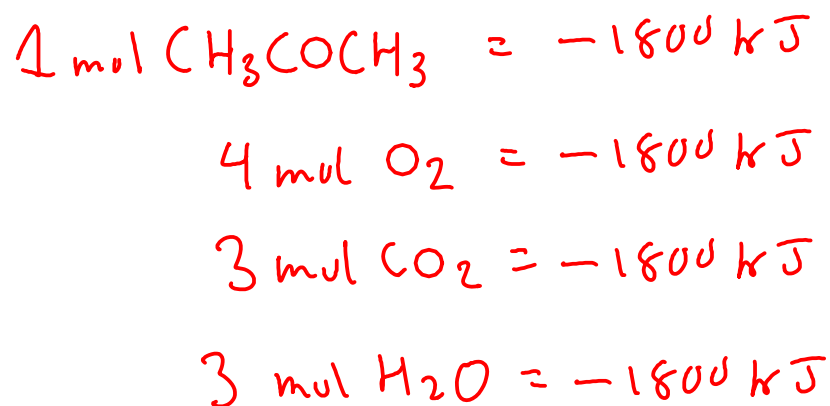
THERMOCHEMICAL EQUATIONS

- 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?



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

USING A THERMOCHEMICAL EQUATION

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

- 1) Convert 25 grams of water to moles. Use FORMULA WEIGHT.
- 2) Convert moles water to enthalpy change. Use THERMOCHEMICAL EQUATION.

$$\textcircled{1} \text{ H}_2\text{O} : \begin{array}{l} 1 - 2 \times 1.008 \\ 0 - 1 \times 16.00 \\ \hline 18.016 \text{ g H}_2\text{O} = \text{mol H}_2\text{O} \end{array}$$

$$\textcircled{2} 3 \text{ mol H}_2\text{O} = -1800 \text{ kJ}$$

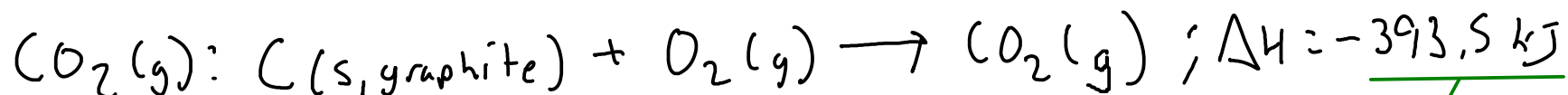
$$25 \text{ g H}_2\text{O} \times \frac{\text{mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{-1800 \text{ kJ}}{3 \text{ mol H}_2\text{O}} = \boxed{-830 \text{ kJ}}$$

$\textcircled{1}$
 $\textcircled{2}$

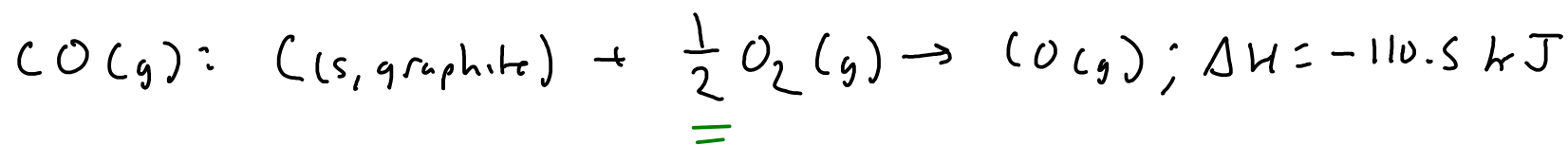
Notes: This is an EXOTHERMIC reaction ... energy is lost to the surroundings. (Negative value). Also, -830 kJ would be the heat observed in a calorimetry experiment per 25 grams of water produced, provided the calorimeter was run at constant pressure.

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
or "enthalpy of formation"



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

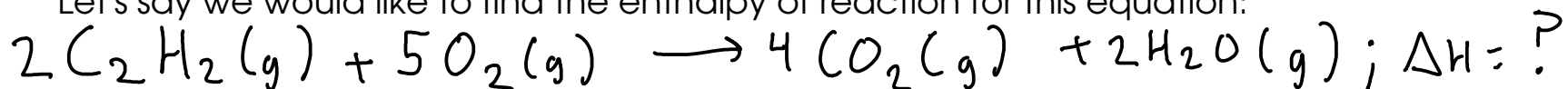
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

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

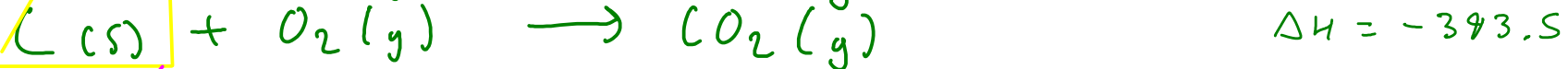
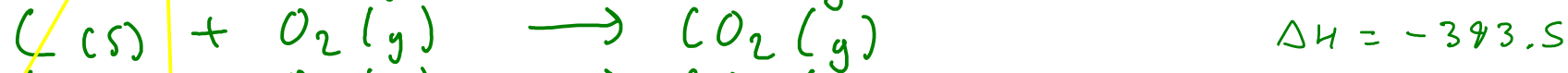
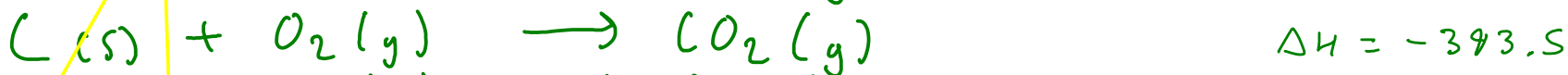
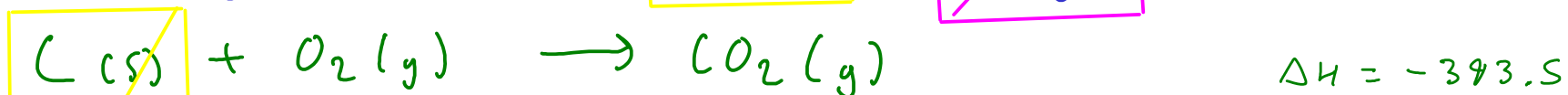
- What are formation reactions good for?

... finding enthalpies for more interesting reactions!

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



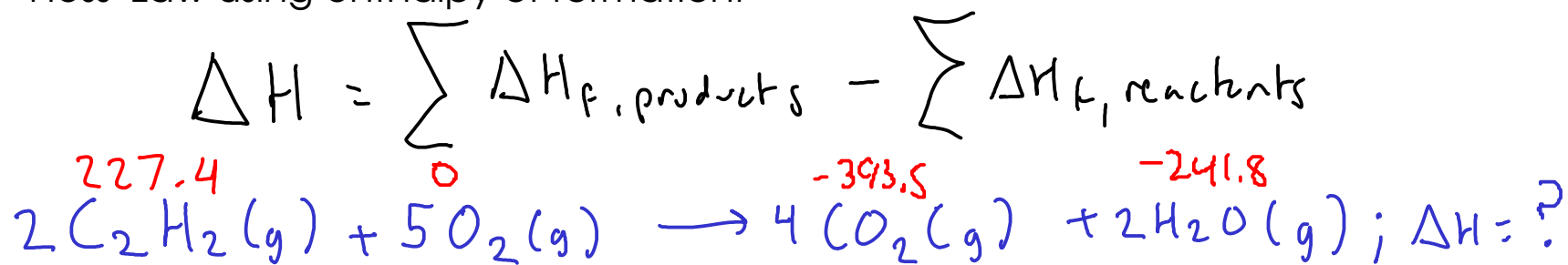
Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.



$$\Delta H = 2(-227.4) + 4(-393.5) + 2(-241.8) = \boxed{-2512.4 \text{ kJ}}$$

* Enthalpy of formation data taken from Openstax Chemistry Appendix G

Hess' Law using enthalpy of formation:



$$\Delta H = 4(-393.5) + 2(-241.8) - 2(227.4) - 5(0)$$

$$= -2812.4 \text{ kJ}$$

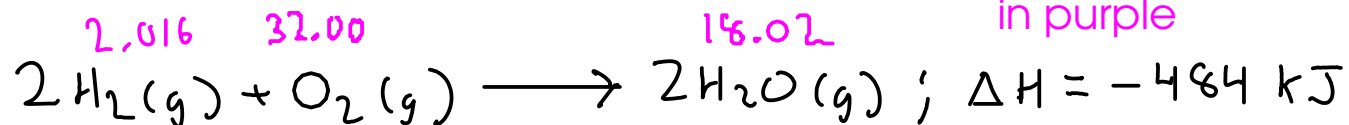
See Appendix G in the Openstax textbook for enthalpy of formation data!

* 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:

FORMULA WEIGHTS in g/mol
in purple



Calculate the enthalpy change for the combustion of 1000 g of hydrogen gas.

1) Convert 1000. grams of hydrogen gas to moles. Use FORMULA WEIGHT.

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

$$\textcircled{1} 2.016 \text{ g H}_2 = 1 \text{ mol H}_2 \quad \textcircled{2} 2 \text{ mol H}_2 = -484 \text{ kJ}$$

$$1000. \text{ g H}_2 \times \frac{\text{mol H}_2}{2.016 \text{ g H}_2} \times \frac{-484 \text{ kJ}}{2 \text{ mol H}_2} = \boxed{-120000 \text{ kJ}} \text{ per } 1000. \text{ g H}_2$$

①
②