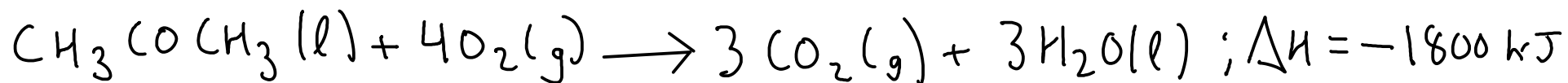


THERMOCHEMICAL EQUATIONS

- 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!

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

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

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

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

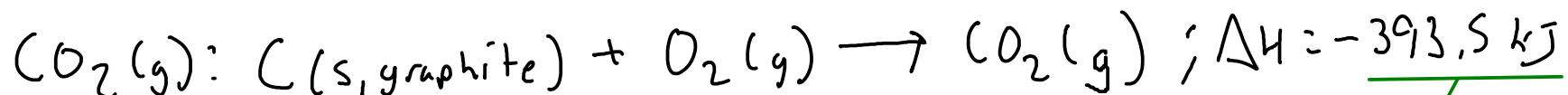
$\textcircled{1}$
 $\textcircled{2}$

↑ Exothermic process. (negative sign on Q or delta H)

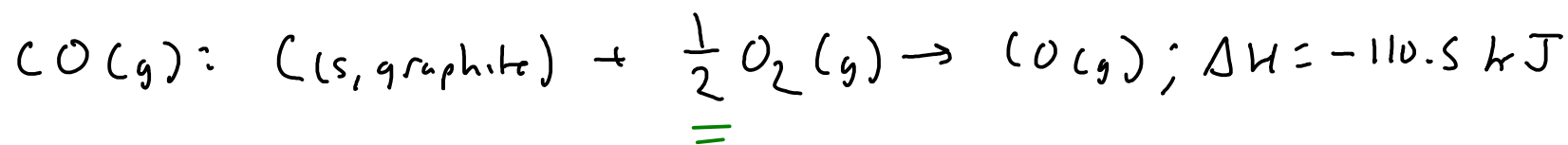
If the reaction was run at constant pressure, delta H would be equal to Q, so Q (heat) would also be -830 kJ.

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!

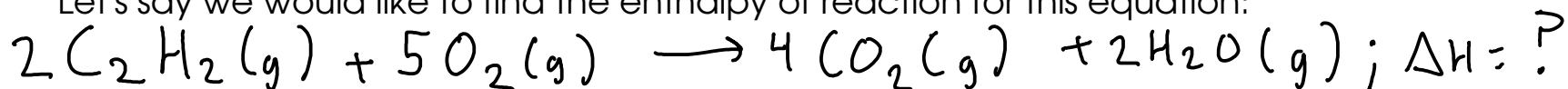
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

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

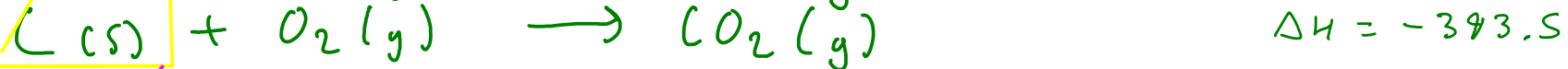
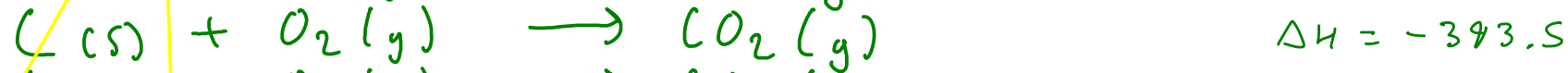
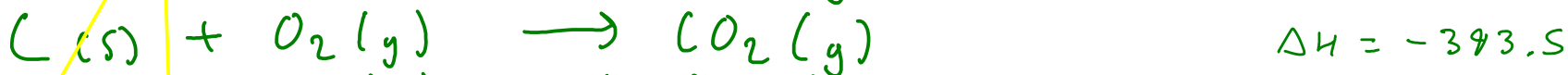
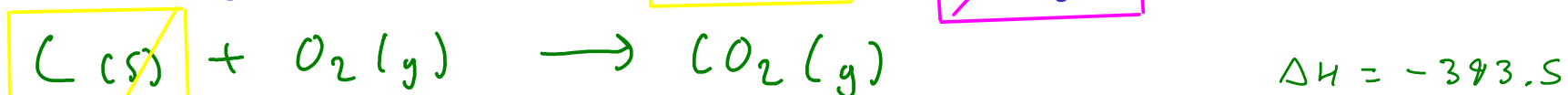
- What are formation reactions good for?

... finding enthalpies for more interesting reactions!

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



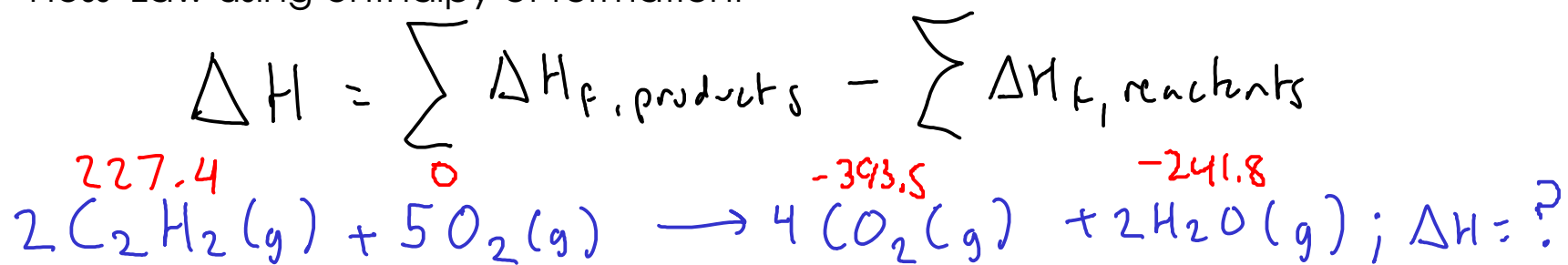
Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.



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

* Enthalpy of formation data taken from Openstax Chemistry Appendix G

Hess' Law using enthalpy of formation:



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

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

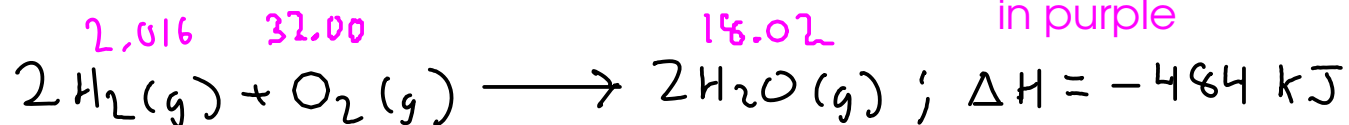
See Appendix G in the Openstax textbook for enthalpy of formation data!

* 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 g of hydrogen gas to moles. Use FORMULA WEIGHT.

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

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

$$1000 \text{ g H}_2 \times \frac{\overset{\text{mol H}_2}{2.016 \text{ g H}_2}}{\textcircled{1}} \times \frac{-484 \text{ kJ}}{\underset{\textcircled{2}}{2 \text{ mol H}_2}} = \boxed{-120000 \text{ kJ}} \text{ per kg H}_2$$