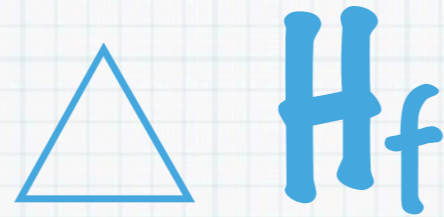


Standard Enthalpies of Formation



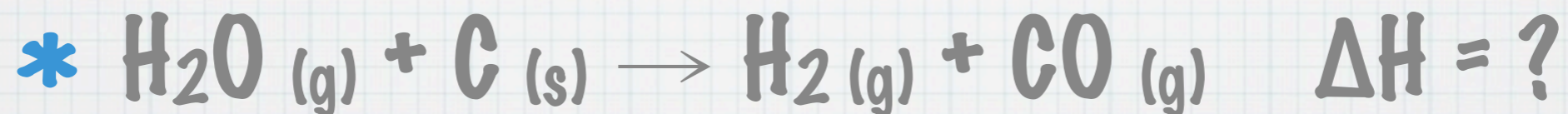
- * The standard enthalpy of formation is the quantity of energy associated with the formation of one mole of a substance from its elements in their standard states (at SATP)

Symbol: $\triangle H_f$

Unit: kJ/mol

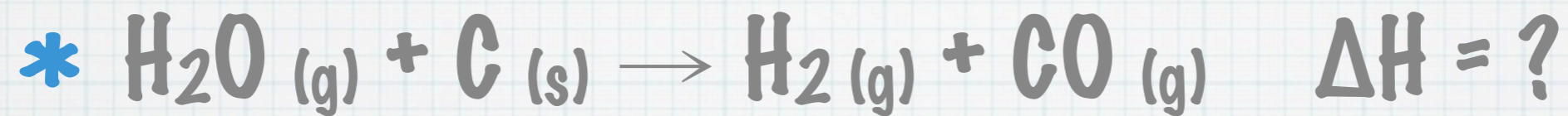
Calculating ΔH_f

* According to Hess's Law, the enthalpies of known equations may be used to calculate the enthalpy of an unknown reaction.

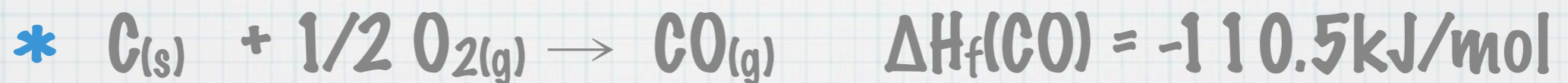
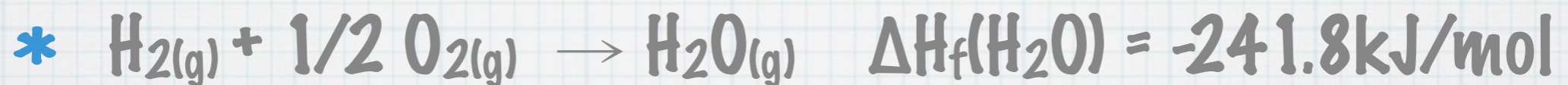


* Use the formation equations for each of the products and reactants to create the target equation.

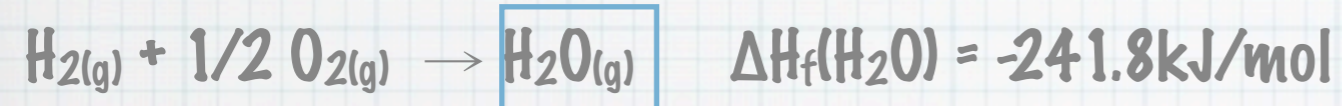
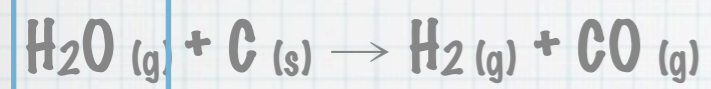
Example



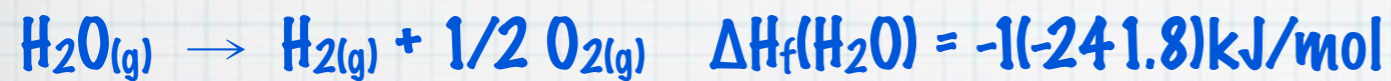
* Consider:



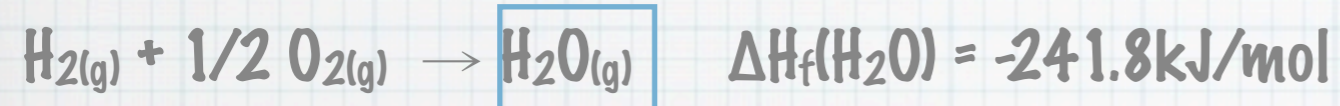
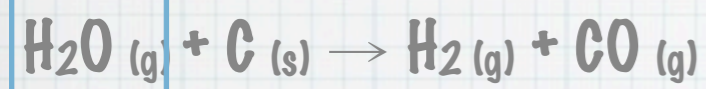
Equation 1



* H_2O on wrong side, reverse and change sign on ΔH



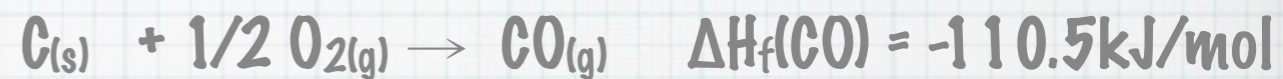
Equation 1



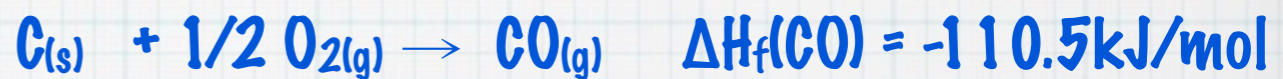
* H_2O on wrong side, reverse and change sign on ΔH



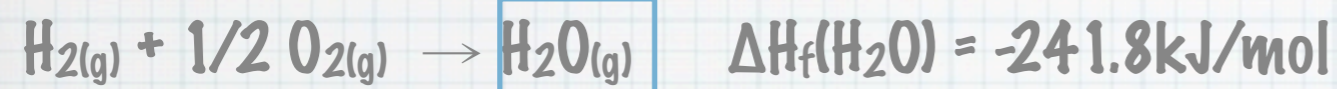
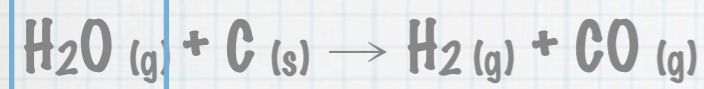
Equation 2



* Coefficients and sides match, ΔH stays the same



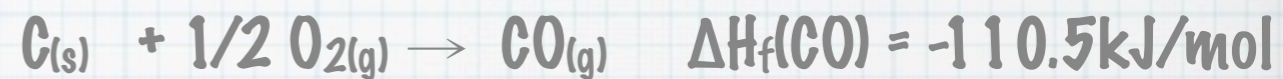
Equation 1



* H_2O on wrong side, reverse and change sign on ΔH



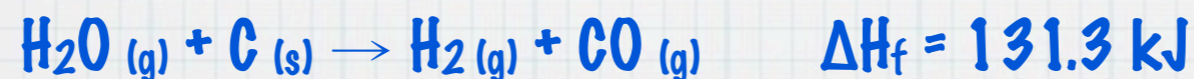
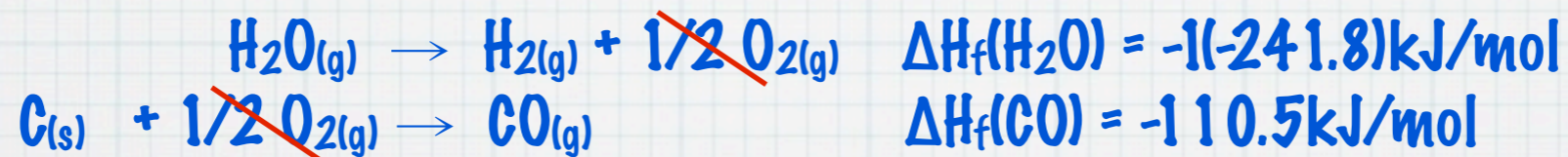
Equation 2



* Coefficients and sides match, ΔH stays the same



Combine Equations



Using Hess's Law

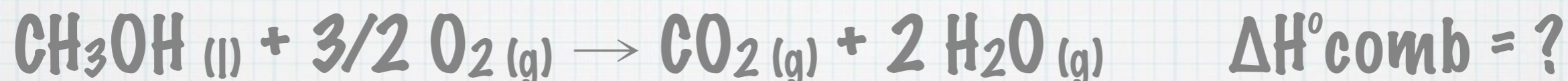
- * When all enthalpies of formation are known, the enthalpy of a reaction is equal to the sum of the enthalpies of formation of the products minus the sum of the enthalpies of formation of the reactants.

$$\Delta H_{\text{rxn}} = \sum n \Delta H_f (\text{products}) - \sum n \Delta H_f (\text{reactants})$$

n = the amount in moles of each product or reactant

Example

- * Calculate the heat of combustion of methanol.



$$\Delta H_{\text{CH}_3\text{OH}} = -239.9 \text{ kJ}$$

$$\Delta H_{\text{O}_2} = 0 \text{ kJ}$$

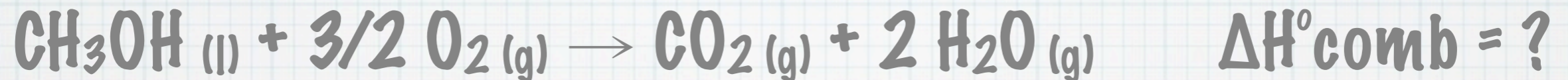
$$\Delta H_{\text{CO}_2} = -393.5 \text{ kJ}$$

$$\Delta H_{\text{H}_2\text{O}} = -241.8 \text{ kJ}$$

All atoms in their standard states have $H_f = 0$

Solution

$$* \Delta H_{rxn} = \sum n \Delta H_f (\text{products}) - \sum n \Delta H_f (\text{reactants})$$



$$\Delta H_{rxn} = [(1)(\text{CO}_2) + (2)(\text{H}_2\text{O})] - [(1)(\text{CH}_3\text{OH}) + (3/2)(\text{O}_2)]$$

$$\Delta H_{rxn} = [(1)(-393.5) + (2)(-241.8)] - [(1)(-239.2) + (3/2)(0)]$$

$$\Delta H_{rxn} = [-877.1] - [-239.2]$$

$$\Delta H_{rxn} = -637.9 \text{ kJ}$$

$$\Delta H_{\text{CH}_3\text{OH}} = -239.9 \text{ kJ}$$

$$\Delta H_{\text{O}_2} = 0 \text{ kJ}$$

$$\Delta H_{\text{CO}_2} = -393.5 \text{ kJ}$$

$$\Delta H_{\text{H}_2\text{O}} = -241.8 \text{ kJ}$$

The ΔH_{comb} is -637.9 kJ .

Homework

* Pg 323 #51-56

* Pg 324 # 1,4,6,11