



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

- 1 - Convert mass water to moles using formula weight.
- 2 - Convert moles water to enthalpy change using thermochemical equation.

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

$$25 \text{ g H}_2\text{O} \times \frac{\textcircled{1} 1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} \times \frac{\textcircled{2} -1800 \text{ kJ}}{3 \text{ mol H}_2\text{O}} = \boxed{-830 \text{ kJ}}$$

- This reaction is EXOTHERMIC, and it will release energy to the surroundings (making them warmer) - just like every other combustion reaction!
- If this reaction is done at constant pressure, the enthalpy change should equal the observed Q.

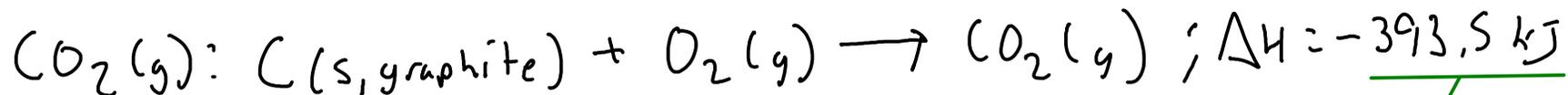
A few more terms related to enthalpy:

- 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.

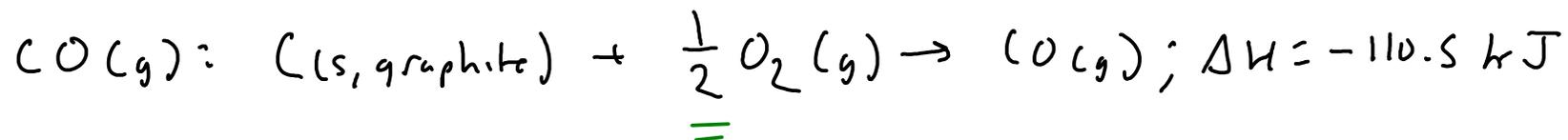
 Phase changes require energy, too!

## 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  $\Delta H_f^\circ$  or  $\Delta H_f$   
"enthalpy of formation"



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

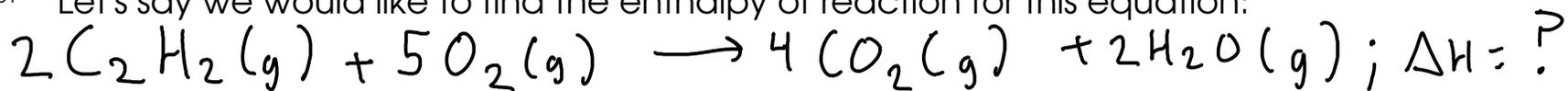
- The heat of formation for an element in its standard state at 25C and 1 atm is ZERO.

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

- What are formation reactions good for?

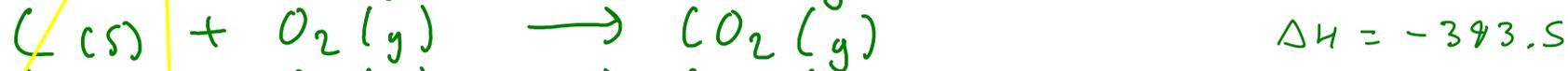
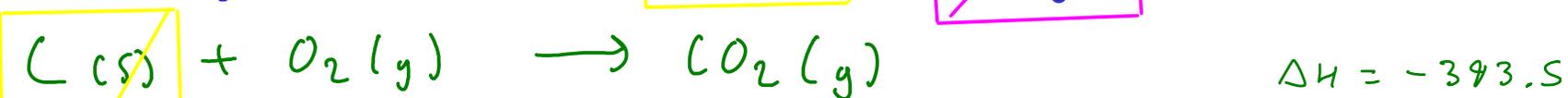
... finding enthalpies for more interesting reactions!

<sup>161</sup> 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.



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

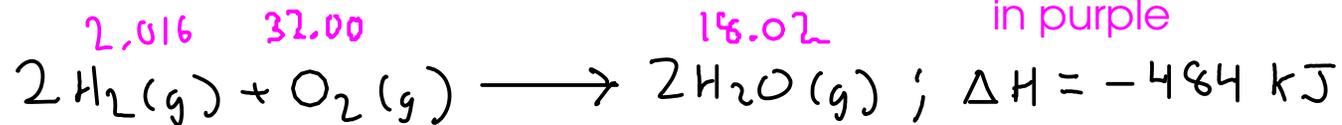
$$\Delta H = \sum \Delta H_{f, \text{products}} - \sum \Delta H_{f, \text{reactants}}$$

$$\Delta H = [4(-393.5) + 2(-241.8)] - [2(226.7) + 5(0)]$$
$$= -2511 \text{ kJ}$$

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!



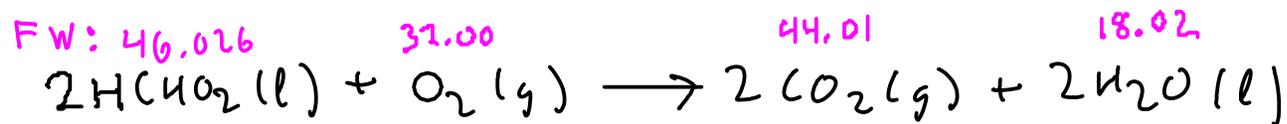
Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

1 - Convert 1.00 kg of hydrogen gas to moles using formula weight.

2 - Convert moles hydrogen gas to enthalpy change using thermochemical equation

$$2.016 \text{ g H}_2 = 1 \text{ mol H}_2 \quad | \quad \text{Kg} = 10^3 \text{ g} \quad | \quad 2 \text{ mol H}_2 = -484 \text{ kJ}$$

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



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?

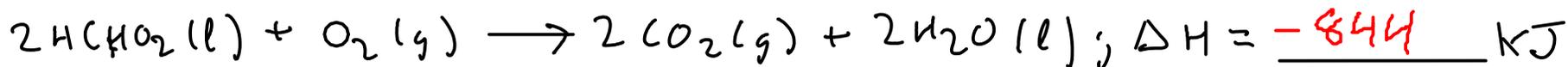
$$Q = -50.3 \text{ kJ} ; \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2}$$

Find moles formic acid:

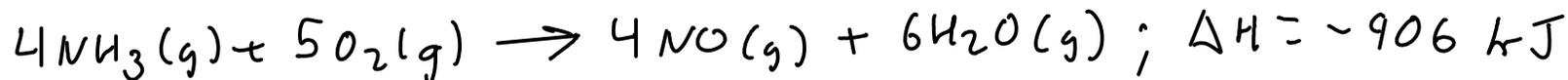
$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol HCHO}_2}{46.026 \text{ g HCHO}_2} = 0.1190631382 \text{ mol HCHO}_2$$

$$\Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol HCHO}_2} = \frac{-50.3 \text{ kJ}}{0.1190631382 \text{ mol HCHO}_2} = \boxed{-422 \frac{\text{kJ}}{\text{mol HCHO}_2}}$$

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



We calculated the enthalpy in the calorimetry problem per mole of formic acid, while the equation here is per TWO MOLES. So multiply the calculated enthalpy by two to get the enthalpy change for the reaction as written.



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 - Change 150. L of NO to moles using ideal gas equation.

2 - Change moles NO to enthalpy change using thermochemical equation.

$$PV = nRT \quad \left| \quad P = 1.50 \text{ atm} \quad V = 150. \text{ L} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right.$$

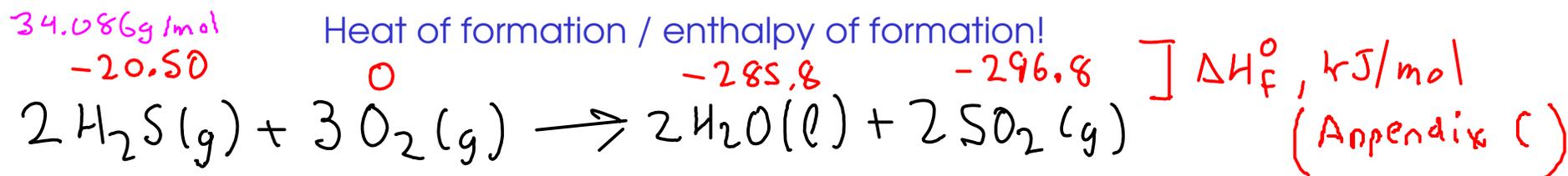
$$n = \frac{PV}{RT} \quad \left| \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \right.$$

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

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

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

$\textcircled{2}$



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 equation 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

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

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

Thermochemical equation:

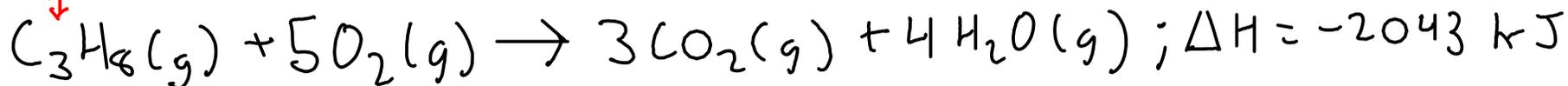


$$34.086 \text{ g H}_2\text{S} = 1 \text{ mol H}_2\text{S} \quad | \quad 2 \text{ mol H}_2\text{S} = -1124.2 \text{ kJ}$$

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

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

propane



Calculate the volume of propane gas at 25.0 C and 1.08 atm required to provide 565 kJ of heat using the reaction above.

- 1 - Convert energy requirement to MOLES PROPANE using THERMOCHEMICAL EQUATION
- 2 - Convert moles propane to volume using ideal gas equation

$$\text{mol C}_3\text{H}_8 = -2043 \text{ kJ}$$

Since the reaction provides the energy, the energy requirement will be negative from the reaction's point of view.

$$-565 \text{ kJ} \times \frac{\text{mol C}_3\text{H}_8}{-2043 \text{ kJ}} = 0.2765540871 \text{ mol C}_3\text{H}_8 \quad (1)$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$n = 0.2765540871 \text{ mol C}_3\text{H}_8$	$P = 1.08 \text{ atm}$
$R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$	
$T = 25.0^\circ\text{C} = 298.2 \text{ K}$	

$$V = \frac{(0.2765540871 \text{ mol C}_3\text{H}_8) (0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) (298.2 \text{ K})}{(1.08 \text{ atm})}$$

$$= 6.27 \text{ L propane @ } 25.0^\circ\text{C} \text{ } 1.08 \text{ atm}$$

## END OF CHAPTER 6