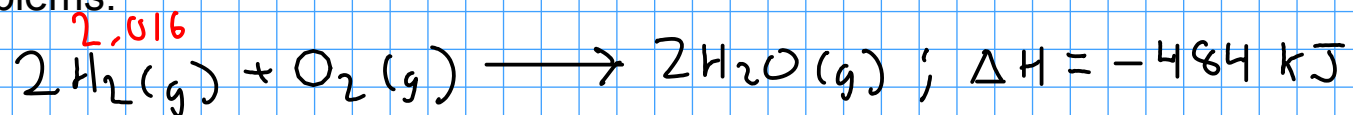


Example problems:



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.

$$2.016 \text{ g H}_2 = 1 \text{ mol H}_2, \quad \text{Kg} = 10^3 \text{ g}$$

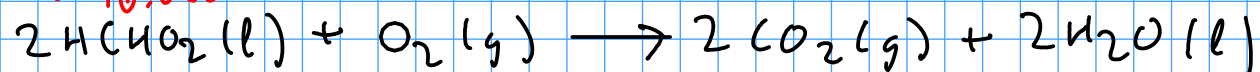
② Use the enthalpy of reaction (-484 kJ) to relate moles of hydrogen to enthalpy change.

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

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

$$= \boxed{-120000 \text{ kJ}}$$

FW: 46.026



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?

$\Delta H = ?$, $Q = -50.3 \text{ kJ}$, since we're at constant P,

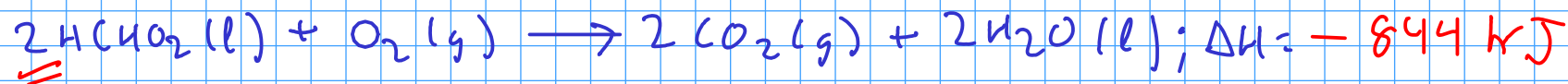
$$Q = \Delta H$$

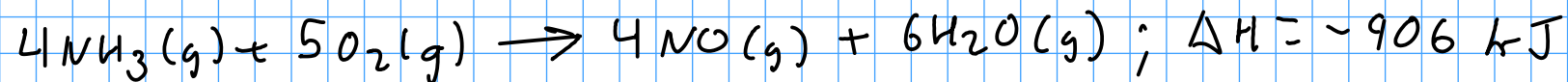
$$\Delta H = \frac{Q}{\text{moles HCHO}_2}$$

$$5.48 \text{ g HCHO}_2 \times \frac{\text{mol}}{46.026 \text{ g}} = 0.119 \text{ mol}$$

$$\Delta H = \frac{-50.3 \text{ kJ}}{0.119 \text{ mol}} = \boxed{-422 \frac{\text{kJ}}{\text{mol HCHO}_2}}$$

The thermochemical equation is written in terms of TWO moles of formic acid, so the enthalpy change of this reaction as written is -844 kJ





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 Calculate the number of moles of NO formed by using the ideal gas equation.

$$n = \frac{PV}{RT} \quad \begin{array}{l} P = 1.50 \text{ atm} \\ V = 150. \text{ L} \end{array} \quad \begin{array}{l} R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ T = 25.0^\circ = 298.2 \text{ K} \end{array}$$

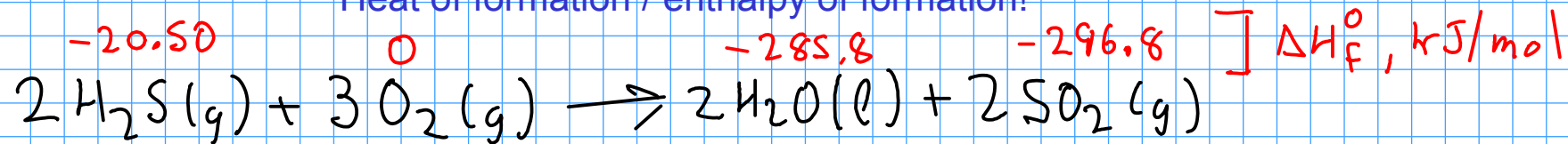
$$n = \frac{(1.50 \text{ atm})(150. \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298.2 \text{ K})} = 9.19 \text{ mol NO}$$

② Use the thermochemical equation to relate moles of NO and enthalpy!

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

$$9.19 \text{ mol NO} \times \frac{-906 \text{ kJ}}{4 \text{ mol NO}} = \boxed{-2080 \text{ kJ}}$$

Heat of formation / enthalpy of formation!



What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

To solve this problem, first calculate the enthalpy of reaction from standard enthalpies of formation. Use Hess's Law:

$$\begin{aligned}
 \Delta H &= \sum \Delta H_f^\circ, \text{ products} - \sum \Delta H_f^\circ, \text{ reactants} \\
 \Delta H &= 2(-285.8) + 2(-296.8) - [2(-20.50) + 3(0)] \\
 &\quad \text{H}_2\text{O} \quad \quad \quad \text{SO}_2 \quad \quad \quad \text{H}_2\text{S} \quad \quad \quad \text{O}_2 \\
 &= -1124.2 \dots \text{ for reaction as written}
 \end{aligned}$$

Next, we use the enthalpy of reaction to find the enthalpy change for 25.0 grams of hydrogen sulfide.

$$\text{H}_2\text{S}: 34.08 \text{ g H}_2\text{S} = 1 \text{ mol H}_2\text{S}$$

$$2 \text{ mol H}_2\text{S} = -1124.2 \text{ kJ}$$

$$25.0 \text{ g H}_2\text{S} \times \frac{1 \text{ mol H}_2\text{S}}{34.08 \text{ g H}_2\text{S}} \times \frac{-1124.2 \text{ kJ}}{2 \text{ mol H}_2\text{S}} = \boxed{-412 \text{ kJ}}$$