

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)

2000 Jeep Cherokee vs 2008 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)

$$H = U + PV$$

$$\Delta H = Q_{\text{constant pressure}}$$

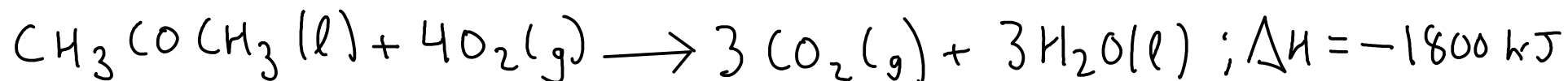
... we record the "enthalpy change of reaction" in our data books.

$$\Delta H_r$$

¹⁵⁸ 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?

$$1 \text{ mol CH}_3\text{COCH}_3 = -1800 \text{ kJ}$$

$$4 \text{ mol O}_2 = -1800 \text{ kJ}$$

$$3 \text{ mol CO}_2 = -1800 \text{ kJ}$$

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

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



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

1 - Convert 25 grams water to moles. Use FORMULA WEIGHT.

2 - Convert moles water to enthalpy change. Use THERMOCHEMICAL EQUATION.

$$\textcircled{1} \text{H}_2\text{O} - \text{H}: 2 \times 1.008$$

$$\quad \quad \quad \text{O}: 1 \times 16.00$$

$$\frac{\quad}{18.016 \text{ g H}_2\text{O}} = \text{mol H}_2\text{O}$$

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

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

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

This reaction is EXOTHERMIC!
(negative sign for enthalpy change)

A few more terms related to enthalpy:

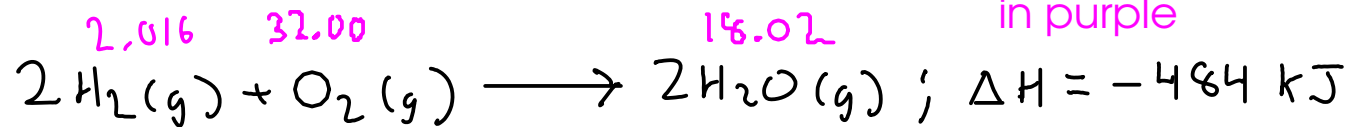
- Enthalpy of vaporization / heat of vaporization: The enthalpy change on vaporizing one mole of a substance. (from liquid to vapor)

- 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 Example problems:

FORMULA WEIGHTS in g/mol
in purple



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

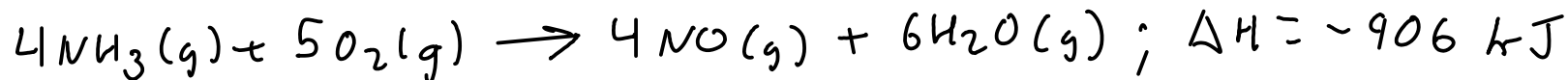
1 - Convert 1.00 kg hydrogen gas to moles. Use FORMULA WEIGHT.

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

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

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

① ②



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 - Convert 150. L of NO to moles. Use IDEAL GAS EQUATION.

2 - Convert moles NO to enthalpy change Use THERMOCHEMICAL EQUATION.

$$\textcircled{1} \quad PV = nRT \quad \left| \quad P = 1.50 \text{ atm} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right.$$

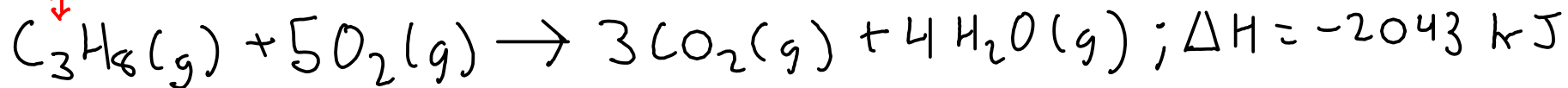
$$n = \frac{PV}{RT} \quad \left| \quad V = 150. \text{ L} \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \right.$$

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

$$\textcircled{2} \quad 4 \text{ mol NO} = -906 \text{ kJ}$$

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

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 the energy requirement to moles of propane. Use THERMOCHEMICAL EQUATION.
 2 - Convert moles propane to volume. Use IDEAL GAS EQUATION.

① $\text{mol C}_3\text{H}_8 = -2043 \text{ kJ}$ A bit about signs: 565 kJ is the amount of energy that needs to come out of the chemical reaction. So from the reaction's point of view, the energy requirement is -565 kJ.

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

② $PV = nRT$ | $n = 0.2765540871 \text{ mol C}_3\text{H}_8$ $P = 1.08 \text{ atm}$
 $V = \frac{nRT}{P}$ | $R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$ $T = 25.0^\circ\text{C} = 298.2 \text{ K}$

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

$$= \boxed{6.27 \text{ L of C}_3\text{H}_8 \text{ at } 25.0^\circ\text{C}, 1.08 \text{ atm}}$$

END OF CHAPTER 6