

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 g water to moles. Use FORMULA WEIGHT.

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

$$\textcircled{1} \text{H}_2\text{O} - \begin{array}{l} \text{H: } 2 \times 1.008 \\ \text{O: } 1 \times 16.00 \end{array}$$

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

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

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

Notes: (1) The reaction is EXOTHERMIC (releases energy to surroundings).

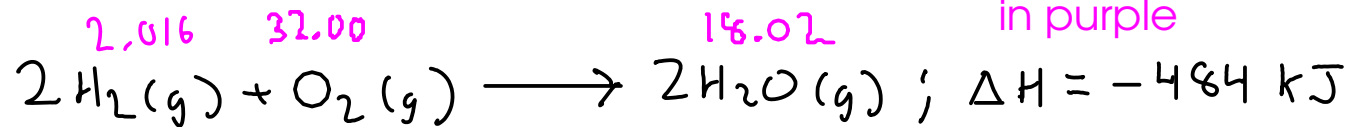
(2) Provided we are at constant pressure; $Q = -830 \text{ kJ}$

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!



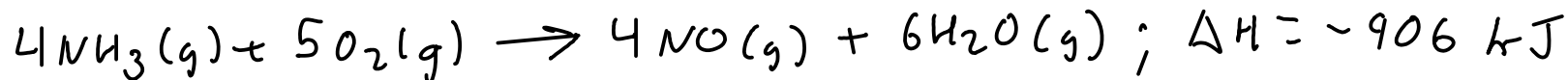
Calculate the enthalpy change for the combustion of 1.00 kg of hydrogen gas.

1 - Convert 1.00 kg hydrogen gas to moles. Use FORMULA WEIGHT (mass specified!)

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

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

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



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 NO to moles. Use IDEAL GAS EQUATION.

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

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

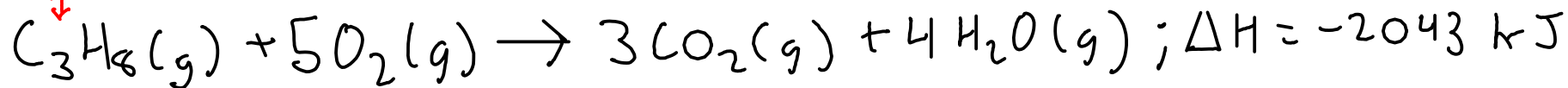
$$n = \frac{PV}{RT} \quad \left| \quad V = 150. \text{ L} \quad T = 25.0^\circ = 298.2 \text{ K} \right.$$

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

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

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

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 565 kJ to moles propane. Use THERMOCHEMICAL EQUATION.
- 2 - Convert moles propane to volume. Use IDEAL GAS EQUATION.

$$\textcircled{1} \text{ mol C}_3\text{H}_8 = -2043 \text{ kJ}$$

Note: The 565 kJ is from our (external observer's) point of view. The energy will have to come from the reaction. So, from the reaction's point of view, it's -565 kJ

$$-565 \text{ kJ} \times \frac{\text{mol C}_3\text{H}_8}{-2043 \text{ kJ}} = 0.276554087 \text{ (mol) C}_3\text{H}_8$$

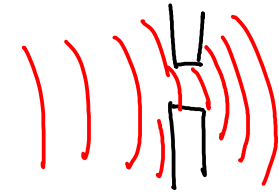
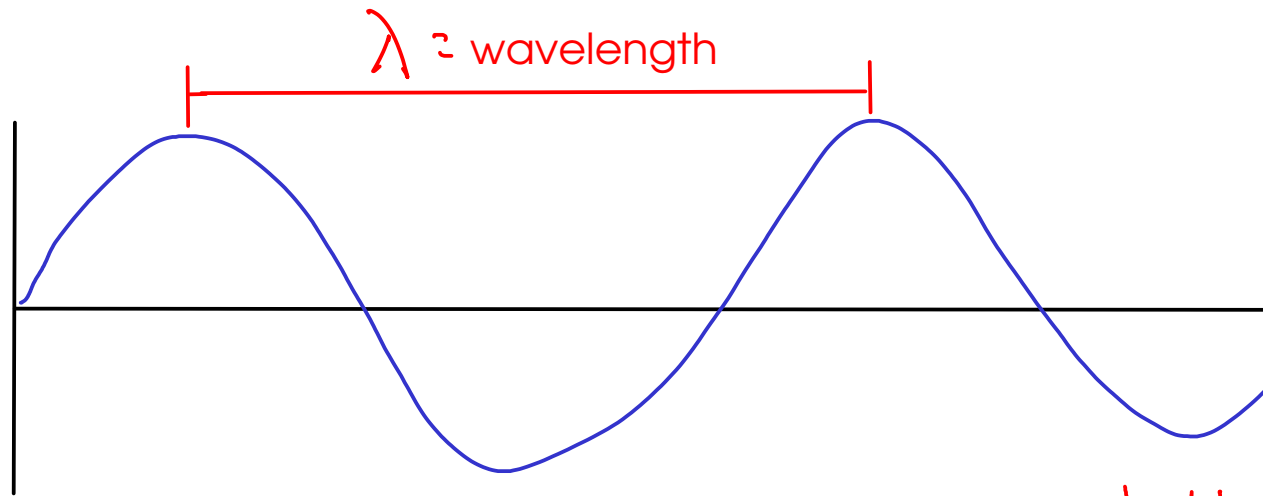
$$\textcircled{2} \begin{array}{l} PV = nRT \\ V = \frac{nRT}{P} \end{array} \left| \begin{array}{l} n = 0.276554087 \text{ (mol) C}_3\text{H}_8 \\ R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ T = 25.0^\circ\text{C} = 298.2 \text{ K} \\ P = 1.08 \text{ atm} \end{array} \right.$$

$$V = \frac{(0.276554087 \text{ (mol) C}_3\text{H}_8) (0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}) (298.2 \text{ K})}{(1.08 \text{ atm})}$$

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

END OF CHAPTER 6

LIGHT



Diffraction

$$\text{frequency} = \text{wavelengths} / \text{time} = \nu \quad \text{s}^{-1} : \text{Hertz } \nu, \text{ Hz}$$

- Light has properties of WAVES such as DIFFRACTION (it bends around small obstructions).
- Einstein noted that viewing light as a particle that carried an energy proportional to the FREQUENCY could explain the PHOTOELECTRIC EFFECT!

$$E_{\text{photon}} = h \nu$$

Planck's constant: $6.63 \times 10^{-34} \text{ J}\cdot\text{s}$

photon = particle or packet of light