$$CH_3COCH_3(l) + 4O_2(g) \longrightarrow 3CO_2(g) + 3H_2O(l)\ ;\ \Delta H = -1800\ kJ$$

What would be the enthapy 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.

H_2O: H: 2×1.008
O: 1×16.00

$18.016\ g\ H_2O = mol\ H_2O$

$3\ mol\ H_2O = -1800\ kJ$

$$25\ g\ H_2O \times \overset{\textcircled{1}}{\frac{mol\ H_2O}{18.016\ g\ H_2O}} \times \overset{\textcircled{2}}{\frac{-1800\ kJ}{3\ mol\ H_2O}} = \boxed{-830\ kJ}$$

This is an EXOTHERMIC reaction. 830 kJ are transferred to the SURROUNDINGS.
The enthalpy change here is the same as Q, as long as we do a constant pressure process.

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!

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.

$$CO_2(g):\ C(s,\text{graphite}) + O_2(g) \longrightarrow CO_2(g)\ ;\ \Delta H = \underline{-393.5\ \text{kJ}}$$

heat of formation of carbon dioxide ΔH_f° or ΔH_f

"enthalpy of formation"

$$CO(g):\ C(s,\text{graphite}) + \underline{\underline{\tfrac{1}{2}}}\,O_2(g) \rightarrow CO(g)\ ;\ \Delta H = -110.5\ \text{kJ}$$

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.

$$\Delta H^\circ_{f,\,O_2(g)} = 0\ \text{kJ/mol}$$

- What are formation reactions good for?

... finding enthalpies for more interesting reactions!

Let's say we would like to find the enthalpy of reaction for this equation:

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(g);\ \Delta H = ?$$

From A-8, text ↙

Hess' Law: If you add two reactions to get a new reaction, their enthalpies also add.

C_2H_2: $2C(s) + H_2(g) \longrightarrow C_2H_2(g)$ $\Delta H = 226.7$

CO_2: $C(s) + O_2(g) \longrightarrow CO_2(g)$ $\Delta H = -393.5$

H_2O: $H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)$ $\Delta H = -241.8$

$$C_2H_2(g) \longrightarrow \cancel{2C(s)} + \cancel{H_2(g)} \qquad \Delta H = -226.7$$

$$C_2H_2(g) \longrightarrow \cancel{2C(s)} + \cancel{H_2(g)} \qquad \Delta H = -226.7$$

$$\cancel{C(s)} + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5$$

$$\cancel{C(s)} + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5$$

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$$\cancel{H_2(g)} + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \qquad \Delta H = -241.8$$

$$\cancel{H_2(g)} + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \qquad \Delta H = -241.8$$

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(\ell)$$

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

Hess' Law using enthalpy of formation:

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

$$\overset{226.7}{2C_2H_2(g)} + \overset{0}{5O_2(g)} \rightarrow \overset{-393.5}{4CO_2(g)} + \overset{-241.8}{2H_2O(g)} ;\ \Delta H = ?$$

$$\Delta H = [4(-393.5) + 2(-241.8)] - [2(226.7) + 5(0)]$$

$$= -2511 \text{ kJ}$$

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!

Example problems:

FORMULA WEIGHTS in g/mol
in purple

$$2\overset{2.016}{H_2}(g) + \overset{32.00}{O_2}(g) \longrightarrow 2\overset{18.02}{H_2O}(g)\ ;\ \Delta H = -484\ \text{kJ}$$

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.

① $2.016\ g\ H_2 = mol\ H_2\ ,\ kg = 10^3\ g$

② $2\ mol\ H_2 = -484\ kJ$

$$1.00\ kg\ H_2 \times \frac{10^3\ g}{kg} \times \underset{\textcircled{1}}{\frac{mol\ H_2}{2.016\ g\ H_2}} \times \underset{\textcircled{2}}{\frac{-484\ kJ}{2\ mol\ H_2}} = \boxed{\begin{array}{c}-12\bar{0}000\ KJ \\ \text{per}\ kg\ H_2\end{array}}$$

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$; $\Delta H = -906$ 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 - Convert 150. L NO to moles. Use IDEAL GAS EQUATION.
2 - Convert moles NO to enthalpy change. Use THERMOCHEMICAL EQUATION.

① $PV = nRT$ | $P = 1.50\ \text{atm}$ $V = 150.\ \text{L}$

$n = \frac{PV}{RT}$ | $R = 0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$ $T = 25.0\ ^\circ\text{C} = 298.2\ \text{K}$

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

② 4 mol NO = −906 kJ

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

②