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.

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

heat of formation of carbon dioxide ΔH_f° or ΔH_f

"enthalpy of formation"

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

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.

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

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

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

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

Hess' Law using enthalpy of formation:

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

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

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

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!

163 Example problems:

FORMULA WEIGHTS in g/mol in purple

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

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

1 - Convert 1.00 kg hydrogen to moles using formula weight.
2 - Convert moles hydrogen gas to enthalpy change using thermochemical equation

2.016 g H_2 = mol H_2 | 2 mol H_2 = -484 kJ | kg = 10^3 g

$$1.00\ kg\ H_2 \times \frac{10^3\ g}{kg} \times \frac{mol\ H_2}{2.016\ g\ H_2} \times \frac{-484\ kJ}{2\ mol\ H_2} = -1\bar{2}0000\ kJ\ per\ kg\ H_2$$

FW: 46.026   32.00   44.01   18.02

$$2HCHO_2(l) + O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$$

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?

$$Q = -50.3 \text{ kJ} \; ; \quad \Delta H = \frac{Q_{\text{constant pressure}}}{\text{mol } HCHO_2}$$

Find moles formic acid...

$$5.48 \text{ g } HCHO_2 \times \frac{\text{mol } HCHO_2}{46.026 \text{ g } HCHO_2} = 0.119063182 \text{ mol } HCHO_2$$

$$\Delta H = \frac{-50.3 \text{ kJ}}{0.119063182 \text{ mol } HCHO_2} = -422 \text{ kJ/mol } HCHO_2$$

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

$$2HCHO_2(l) + O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l) \; ; \; \Delta H = \underline{-844} \text{ kJ}$$

Since the equation is based on TWO moles of formic acid, we need to multiply the enthalpy change obtained in the calorimetry problem by two to find the enthalpy change of the reaction as written.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)\ ;\ \Delta H = -906\ kJ$$

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 of NO to moles using ideal gas equation.
2 - Convert moles NO to enthalpy change using thermochemical equation.

$PV = nRT$

$n = \frac{PV}{RT}$

$P = 1.50\ atm$

$V = 150.\ L$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$

$T = 25.0°C = 298.2\ K$

① $$n_{NO} = \frac{(1.50\ atm)(150.\ L)}{\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(298.2\ K)} = 9.19482849\ mol\ NO$$

$4\ mol\ NO = -906\ kJ$

$$9.19482849\ mol\ NO \times \frac{-906\ kJ}{4\ mol\ NO} = \boxed{-2080\ kJ}$$

34.086g/mol

Heat of formation / enthalpy of formation!

$$\begin{array}{ccccccc} -20.50 & & 0 & & -285.8 & & -296.8 \\ 2H_2S(g) & + & 3O_2(g) & \rightarrow & 2H_2O(l) & + & 2SO_2(g) \end{array}$$

] ΔH_f^o, kJ/mol (Appendix C)

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 reaction AS WRITTEN using Hess' Law
2 - Convert 25.0 g hydrogen sulfide to moles using formula weight.
3- Covert moles hydrogen sulfide to enthalpy change using thermochemical equation.

$$\Delta H = \sum \Delta H^o_{f,products} - \sum \Delta H^o_{f,reactants}$$

(1) $= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)]$

$= -1124.2\ kJ$

Thermochemical equation:

$$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g)\ ;\ \Delta H = -1124.2\ kJ$$

$34.086\ g\ H_2S = mol\ H_2S$ | $2\ mol\ H_2S = -1124.2\ kJ$

$$25.0\ g\ H_2S \times \frac{mol\ H_2S}{34.086\ g\ H_2S} \times \frac{-1124.2\ kJ}{2\ mol\ H_2S} = \boxed{-412\ kJ}$$