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.

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

- 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:

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

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

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

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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:

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

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

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

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

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

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

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

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

$$1.00\ kg\ H_2 \times \frac{10^3\ g}{kg} \times \frac{mol\ H_2}{2.016\ g\ H_2} \times \frac{-484\ kJ}{2\ mol\ H_2} = -120\,000\ kJ\ per\ kg\ H_2$$

(1) (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?

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

Find moles formic acid:

$$5.48\ g\ HCHO_2 \times \frac{mol\ HCHO_2}{46.026\ g\ HCHO_2} = 0.119063138\ mol\ HCHO_2$$

$$\Delta H = \frac{-50.3\ kJ}{0.119063138\ mol\ HCHO_2} = -422\ kJ/mol\ HCHO_2$$

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

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

We calculated the enthalpy change for a mole of formic acid being burned, and the equation has two moles being burned. Double the enthalpy change for two moles.

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

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 150L of NO to moles using ideal gas equation.
2 - Convert moles NO to enthalpy change using thermochemical equation.

$PV = nRT$

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

$P = 1.50\ atm$

$V = 150.\ L$

$T = 25.0^\circ C = 298.2\ K$

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

① $$n_{NO} = \frac{(1.50\ atm)(150.\ L)}{\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(298.2\ K)} = 9.194822849\ mol\ NO$$

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

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

②

34.086g/mol

Heat of formation / enthalpy of formation!

$$\overset{-20.50}{2 H_2S(g)} + \overset{0}{3 O_2(g)} \rightarrow \overset{-285.8}{2 H_2O(l)} + \overset{-296.8}{2 SO_2(g)} \quad \Big] \Delta H_f^o, \text{kJ/mol}$$

(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 for the reaction as written using HESS'S LAW and ENTHALPIES OF FORMATION.

2 - Convert 25.0 grams of hydrogen sulfide to moles using formula weight.

3 - Convert moles hydrogen sulfide to enthalpy change using THERMOCHEMICAL EQUATION.

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(1) $\Delta H = \sum \Delta H^o_{f,\text{products}} - \sum \Delta H^o_{f,\text{reactants}}$

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

$$= -1124.2 \text{ kJ}$$

So the thermochemical equation is ...

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

34.086 g H_2S = mol H_2S	2 mol H_2S = -1124.2 kJ

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

(2) (3)