Announcements

This is the week of November 19th - 23rd.
Group I is going the lab on Molecular Volume; Group II is doing the Gases tutorial.

More on the X-mas exam date December 17th 9 am – 12:00 pm.
Rooms:

Section 006:

Room	**Student Number**
HSB 240	01860 - 13710
AH Stage	13715 - 98264

Hess's Law **applies to all reactions**, and has two important consequences

i) ΔH for $A \rightarrow B$; $= -\Delta H$ for $B \rightarrow A$

ii) $n(\Delta H \text{ for } A \rightarrow B) = \Delta H$ for $nA \rightarrow nB$

Example: Calculate ΔH^0 for

$$C(gr) + 2S(s) \rightarrow CS_2(\ell)$$

Given that:

i) $C(gr) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_1^0 = -394$ kJ mol^{-1}

ii) $S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H_2^0 = -297$ kJmol^{-1}

iii) $CS_2(\ell) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$ $\Delta H^0_3 = -1077$ kJ mol^{-1}

To obtain the **LHS** of the desired reaction, we add i) + 2x ii) together.

iv) $C(gr) + O_2(g) + 2S(s) + 2O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$

$= C(gr) + 2S(s) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$

$$\Delta H_4^0 = -394 + 2x(-297) = -988 \text{ kJ mol}^{-1}$$

To get the **RHS** of the desired reaction, take the reverse of iii)

v) $CO_2(g) + 2SO_2\,(g) \rightarrow CS_2(\ell) + 3O_2(g)$ $\Delta H_5^0 = +1077$ kJ mol^{-1}

Add iv) + v)

$$C(gr) + \cancel{3O_2(g)} + 2S(s) + \cancel{CO_2(g)} + \cancel{2SO_2(g)} \rightarrow CS_2(\ell) + \cancel{3O_2(g)} + \cancel{CO_2(g)} + \cancel{2SO_2(g)}$$

= desired equation $\Delta H^0 = \Delta H_4^0 + \Delta H_5^0$

$= -988 + 1077 = 89$ kJ mol^{-1}

Enthalpies of Formation

Most wisely used enthalpy values are standard enthalpies of formation, ΔH_f^0 which are defined as:

"the **enthalpy change** for the formation of 1 mole of a substance at 1 bar from its **constituent elements** In their standard state at 25^0 C (298 K)

The standard state of an element is its most stable physical form at 25^0 C and 1 atmosphere (in Si units 1 bar)

For all elements in their standard states $\Delta H_f^0 = 0$

For example: C(gr) : 0 but C(s, diamond) : 1.895 kJ mol^{-1}
and C(g, atom) : 716.7 kJ mol^{-1}

Similarly, $O_2(g)$: 0.0 but $O_3(g)$: 142.7 kJ mol^{-1} and O(g, atom) : 249.2 kJ mol^{-1}

Note: ΔH_f^0 of $Cl_2(g)$, $Br_2(\ell)$ and $I_2(s)$ = 0.0 but for $Br_2(g)$: 30.91 kJ mol^{-1}

Thus energy is required to convert any element from its standard state to a different allotropic form

See M&H Appendix 1, p. 608-609 for tables of ΔH_f^0 or a very limited list in Table 8.3 p. 210

Heats of formation for compounds are also given

For example: ΔH_f^0 for $CH_4(g)$ =-74.8 kJ mol^{-1}

This represents the ΔH for the reaction of formation:

$$C(gr) + 2H_2(g) \rightarrow CH_4(g)$$

Make sure you can write formation equations

The ΔH^0 for any reaction can be calculated if the ΔH_f^0 of all reactants and products are known

Example:
Calculate ΔH^0 for the production of ethane (C_2H_6) from acetylene (C_2H_2) and hydrogen (H_2)

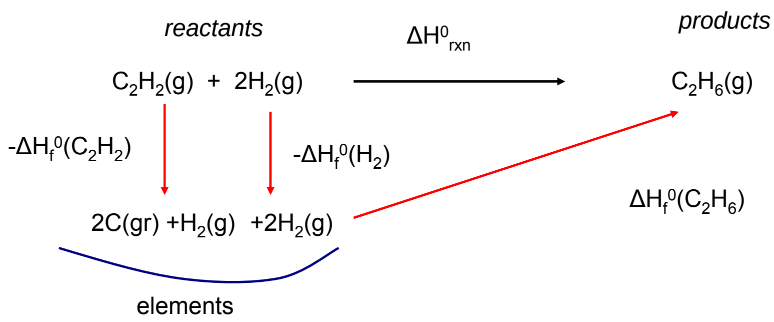
Begin with the balanced equation:

$$C_2H_2(g) + 2H_2(g) \rightarrow C_2H_6(g)$$

Look up the appropriate ΔH_f^0's:

$C_2H_2(g)$; that is, $2C(gr) + H_2(g) \rightarrow C_2H_2(g)$ ΔH_f^0 = 226.7 kJ mol^{-1}
$H_2(g)$ ΔH_f^0 = 0
$C_2H_6(g)$; that is, $2C(gr) + 3H_2(g) \rightarrow C_2H_6(g)$ ΔH_f^0 = -84.7 kJ mol^{-1}

The formation of products can be viewed as proceeding through the sequence:



By Hess's Law:

$\Delta H^0_{rxn} = -\Delta H_f^0(C_2H_2) - 2\Delta H_f^0(H_2) + \Delta H_f^0(C_2H_6)$

$= -226.7 - 2x(0.0) + (-84.7)$ kJ mol^{-1} = -311.4 kJ mol^{-1}

For this example, note that $-\Delta H_f^0(C_2H_2) - 2\Delta H_f^0(H_2) + \Delta H_f^0(C_2H_6)$
$= -\Delta H_f^0(C_2H_2 + 2H_2) + \Delta H_f^0(C_2H_6)$
$= -\Delta H_f^0(\text{reactants}) + \Delta H_f^0(\text{products})$

In general for any reaction:

$$\Delta H^0_{rxn} = \sum\Delta H_f^0(\text{products}) - \sum\Delta H_f^0(\text{reactants})$$

Thus, the **enthalpy of any reaction** can be calculated if the ΔH_f^0 of **all products and reactants** are known

Example:
Calculate ΔH^o for the combustion of heptane, $C_7H_{16}(\ell)$

Start by writing a balanced equation with all substances in their standard state; that is, 25^o C, 1 bar pressure

$C_7H_{16}(\ell) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(\ell)$

Look up required ΔH_f^o / kJ mol^{-1}

$C_7H_{16}(\ell)$: -224.0

$O_2(g)$: 0.0

$CO_2(g)$: -393.5

$H_2O(\ell)$: -285.8

$\therefore$ $\sum\Delta H_f^o$(products) - $\sum\Delta H_f^o$(reactants)

= 7x(-383.5) + 8x(-285.8) -1x(-224.0)-11x(0.0)

= -4816.9 kJ mol^{-1}