

Announcements

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

Rooms:

Section 006:

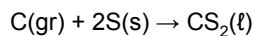
Room	Student Number
HSB 240	01860 - 13710
AH Stage	13715 - 98264

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

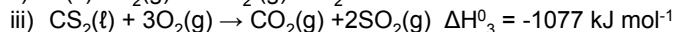
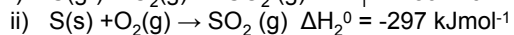
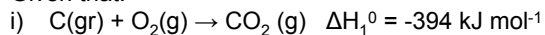
i) ΔH for $A \rightarrow B$; = $-\Delta H$ for $B \rightarrow A$

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

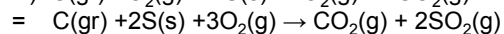
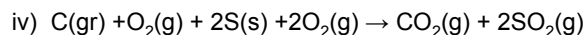
Example: Calculate ΔH° for



Given that:

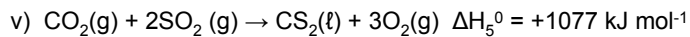


To obtain the **LHS** of the desired reaction, we add i) + 2x ii) together.

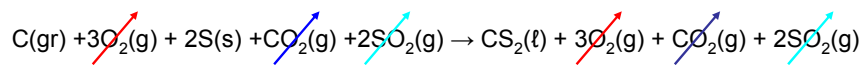


$$\Delta H_4^\circ = -394 + 2x(-297) = -988 \text{ kJ mol}^{-1}$$

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



Add iv) + v)



= desired equation

$$\begin{aligned} \Delta H^0 &= \Delta H_4^0 + \Delta H_5^0 \\ &= -988 + 1077 = 89 \text{ kJ mol}^{-1} \end{aligned}$$

Enthalpies of Formation

Most widely 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° C (298 K)

The standard state of an element is its most stable physical form at 25° 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 $O(g, \text{atom})$: 249.2 kJ mol⁻¹

Note: ΔH_f° of $Cl_2(g)$, $Br_2(l)$ and $I_2(s)$ = 0.0 but for $Br_2(g)$: 30.91 kJ mol⁻¹

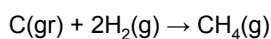
Thus energy is required to convert any element from its standard state to a different allotropic form

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

Heats of formation for compounds are also given

For example: ΔH_f° for $CH_4(g)$ = -74.8 kJ mol⁻¹

This represents the ΔH for the reaction of formation:



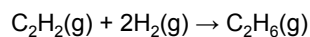
Make sure you can write formation equations

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

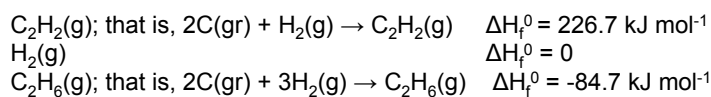
Example:

Calculate ΔH° for the production of ethane (C_2H_6) from acetylene (C_2H_2) and hydrogen (H_2)

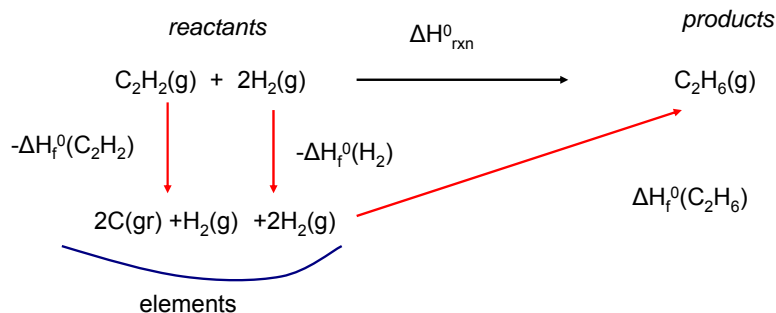
Begin with the balanced equation:



Look up the appropriate ΔH_f° 's:



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



By Hess's Law:

$$\begin{aligned} \Delta H_{\text{rxn}}^{\circ} &= -\Delta H_{\text{f}}^{\circ}(\text{C}_2\text{H}_2) - 2\Delta H_{\text{f}}^{\circ}(\text{H}_2) + \Delta H_{\text{f}}^{\circ}(\text{C}_2\text{H}_6) \\ &= -226.7 - 2 \times (0.0) + (-84.7) \text{ kJ mol}^{-1} = -311.4 \text{ kJ mol}^{-1} \end{aligned}$$

For this example, note that $-\Delta H_{\text{f}}^{\circ}(\text{C}_2\text{H}_2) - 2\Delta H_{\text{f}}^{\circ}(\text{H}_2) + \Delta H_{\text{f}}^{\circ}(\text{C}_2\text{H}_6)$
 $= -\Delta H_{\text{f}}^{\circ}(\text{C}_2\text{H}_2 + 2\text{H}_2) + \Delta H_{\text{f}}^{\circ}(\text{C}_2\text{H}_6)$
 $= -\Delta H_{\text{f}}^{\circ}(\text{reactants}) + \Delta H_{\text{f}}^{\circ}(\text{products})$

In general for any reaction:

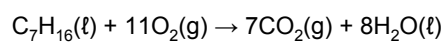
$$\Delta H_{\text{rxn}}^{\circ} = \sum \Delta H_{\text{f}}^{\circ}(\text{products}) - \sum \Delta H_{\text{f}}^{\circ}(\text{reactants})$$

Thus, the **enthalpy of any reaction** can be calculated if the $\Delta H_{\text{f}}^{\circ}$ of **all products and reactants** are known

Example:

Calculate ΔH° for the combustion of heptane, $\text{C}_7\text{H}_{16}(\ell)$

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



Look up required ΔH_f° / kJ mol⁻¹



$$\begin{aligned} \therefore \quad & \Sigma \Delta H_f^\circ(\text{products}) - \Sigma \Delta H_f^\circ(\text{reactants}) \\ & = 7 \times (-393.5) + 8 \times (-285.8) - 1 \times (-224.0) - 11 \times (0.0) \\ & = -4816.9 \text{ kJ mol}^{-1} \end{aligned}$$