

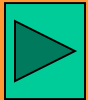
Industrial Chemistry

Hess's law

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Hess's Law



Hess's Law and its experimental verification



Hess's Law calculations, 4 examples.



Hess's Law and calculations

Hess's law states that

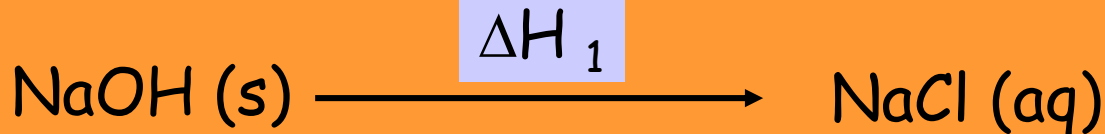
"enthalpy change is independent
of the route taken"



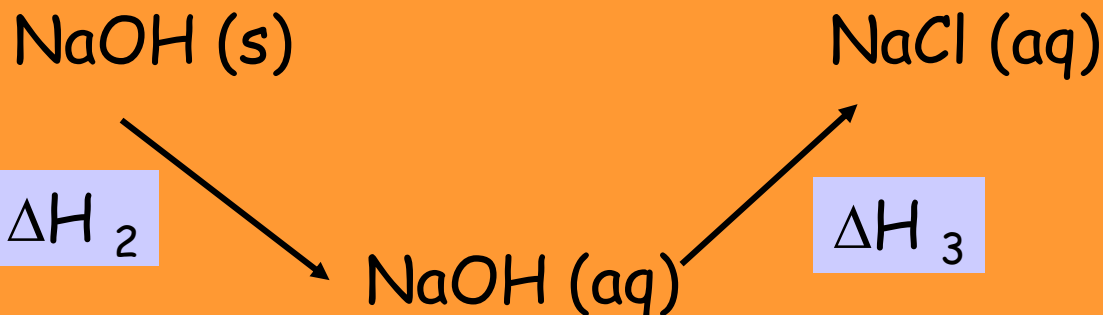
Verification of Hess's Law

ΔH = enthalpy change

Route 1



Route 2



The conversion of solid NaOH to NaCl solution can be achieved by two possible routes. **Route 1** is a single-step process, (adding HCl (aq) directly to the solid NaOH) and **Route 2** is a two-step process (dissolve the solid NaOH in water, then adding the HCl(aq)) All steps are exothermic.

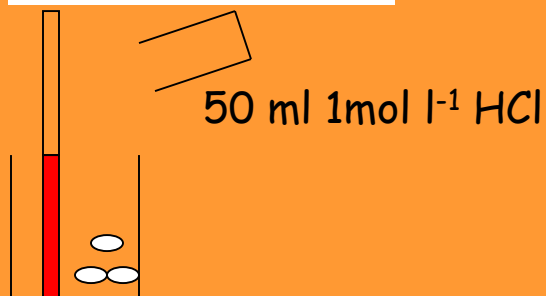
If Hess's Law applies, the enthalpy change for **route 1** must be the same as the overall change for **route 2**.

$$\Delta H_1 = \Delta H_2 + \Delta H_3$$



Verification of Hess's Law

Route 1 ΔH_1



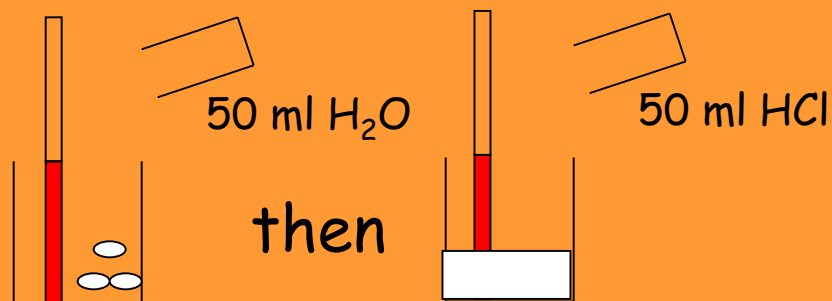
2.50g of NaOH added to a dry, insulated beaker.

Before adding the acid, its temperature is recorded. The final temperature after adding the acid is also recorded.

Knowing the specific heat capacity for water, it is then possible to calculate the Enthalpy change for this reaction.

$$\Delta H_1 = c m \Delta T$$

Route 2 $\Delta H_2 + \Delta H_3$



1. 2.50g of NaOH added to a dry, insulated beaker.

2. Before adding the water, its temperature is recorded. The final temperature rise after adding the water is also recorded. ΔH_2

3. Now add the acid, again, recording the final temperature. ΔH_3

Use the equation below to calculate ΔH_2 and ΔH_3

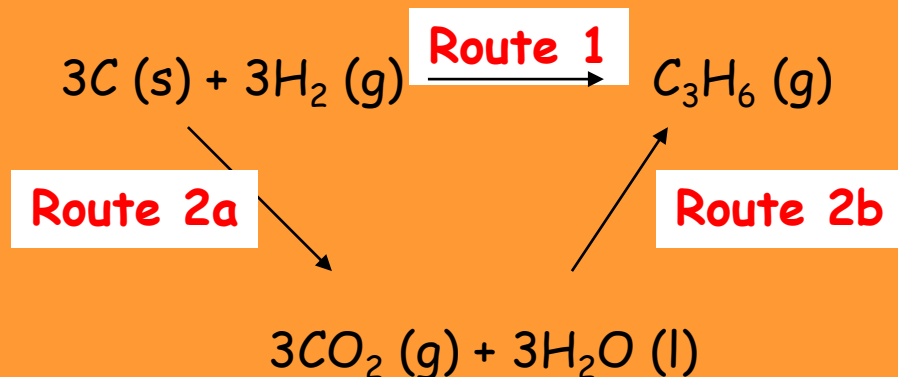
$$\Delta H = c m \Delta T$$

$$\Delta H_1 = \Delta H_2 + \Delta H_3 \text{ will verify Hess's Law}$$



Hess's Law Calculations

Hess's Law can be used to calculate enthalpy changes that cannot be directly measured by experiment.



Route 1 cannot be carried out in a lab, as carbon and hydrogen will not combine directly.

The enthalpy of combustion reactions can act as a stepping stone which enables a link with carbon and hydrogen (the reactants) with propene (the product)

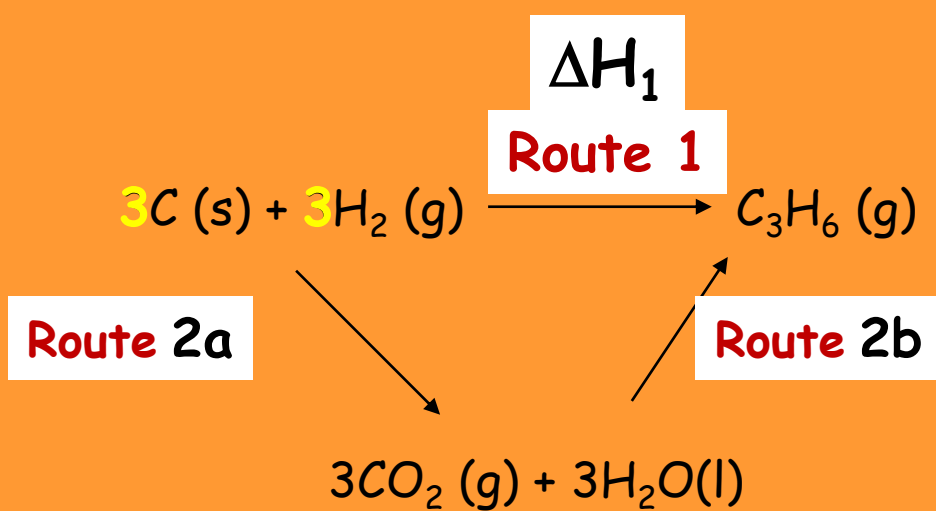
Route 2a involves the combustion of **both carbon** and **hydrogen**



Route 2b involves the **reverse** combustion of propane



Example 1



$$\Delta H_1 = \Delta H_{2a} + \Delta H_{2b}$$

Route 2a $\Delta H_c \text{C} = -394 \text{ kJ mol}^{-1}$

$\Delta H_c \text{H} = -286 \text{ kJ mol}^{-1}$

$$\Delta H_{2a} = -(3 \times 394) = -1182 \text{ kJ mol}^{-1} \quad + \quad -(3 \times 286) = -858 \text{ kJ mol}^{-1}$$

$$\Delta H_{2a} = -2040 \text{ kJ}$$

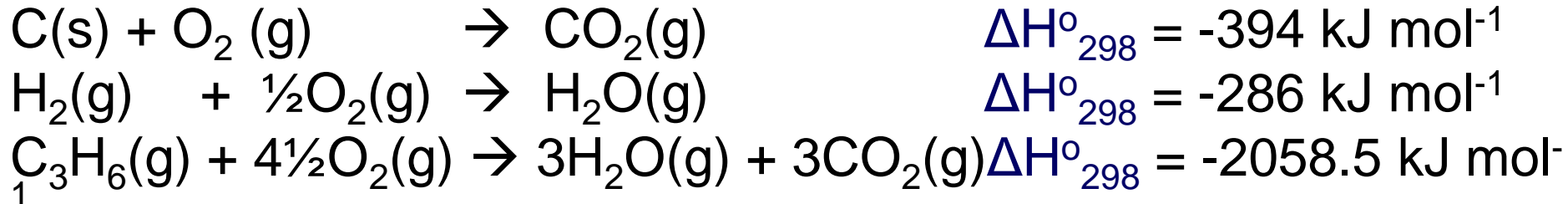
Route 2b $\Delta H_c \text{Propene} = -2058.5 \text{ kJ mol}^{-1}$

$$\Delta H_{2b} = +2058.5 \text{ kJ (note the reverse sign)}$$

$$\Delta H_1 = -2040 \text{ kJ} \quad (+2058.5) \quad = +18.5 \text{ kJ mol}^{-1}$$



Alternative approach for example 1



Re-write the equations so that the reactants and products are on the same side of the "arrow" as the equation you are interested in. Multiply each equation so that there are the same number of moles of each constituent also.



$$\Delta H_c = 3 \times -394 \text{ kJ}$$



$$\Delta H_c = 3 \times -286 \text{ kJ}$$



$$\Delta H_c = +2058.5 \text{ kJ}$$

Equation has been reversed; (enthalpy now has opposite sign)





$$\Delta H_c = 3 \times -394$$

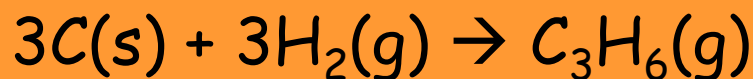


$$\Delta H_c = 3 \times -286$$



$$\Delta H_c = +2058.5$$

Now add the equations and also the corresponding enthalpy values



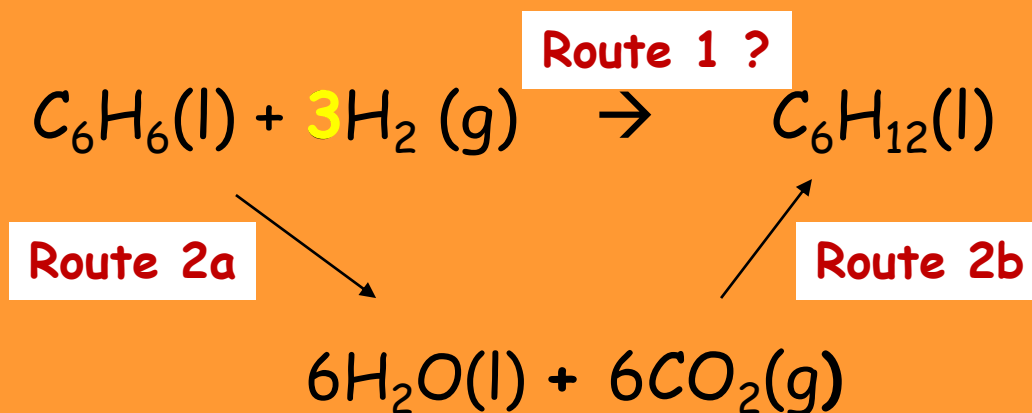
$$\Delta H_f = (3 \times -394) + (3 \times -286) + (+2058.5)$$

$$\Delta H_f = \underline{+18.5 \text{ kJ mol}^{-1}}$$



Calculate the enthalpy change for the reaction:

Example 2



The products of combustion act as a stepping stone which enables a link to be made with benzene and hydrogen (the reactants) with cyclohexane (the product).

Route 2a involves the combustion of **both benzene and hydrogen**



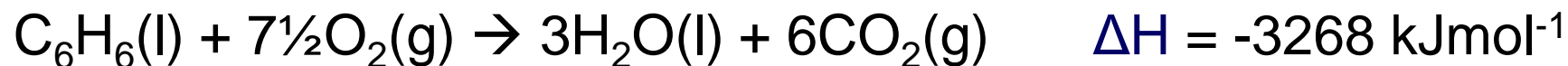
Route 2b involves the **reverse** combustion of cyclohexane



$$\Delta H_1 = \Delta H_{2a} + \Delta H_{2b} = (-3268 + (3 \times -286)) + 3924 = 202 \text{ kJ mol}^{-1}$$



Alternative approach for example 2



Re-write the equations so that the reactants and products are on the same side of the "arrow" as the equation you are interested in. Multiply each equation so that there are the same number of moles of each constituent also.



Equation has been reversed; (enthalpy now has opposite sign)





$$\Delta\text{H}_c = -3268$$



$$\Delta\text{H}_c = 3 \times -286$$



$$\Delta\text{H}_c = +3924$$

Now add the equations and also the corresponding enthalpy values



$$\Delta\text{H}_f = -3268 + (3 \times -286) + 3924$$

$$\Delta\text{H}_f = -202 \text{ kJ mol}^{-1}$$



3. Use the enthalpy changes of combustion shown in the table to work out the enthalpy change of formation of ethyne, C₂H₂.

Substance	C(graphite)	H ₂ (g)	C ₂ H ₂ (g)
ΔH°(combustion)	-394 kJmol ⁻¹	-286 kJmol ⁻¹	-1299 kJmol ⁻¹

“Using the Second method”

“Required” equation, 2C(s) + H₂(g) → C₂H₂(g) ΔH_f = ?

- $2\text{C(s)} + 2\cancel{\text{O}_2\text{(g)}} \rightarrow 2\cancel{\text{CO}_2\text{(g)}} \quad \Delta H_c = 2 \times -394 \text{ kJ mol}^{-1}$
- $\text{H}_2\text{(g)} + \frac{1}{2}\cancel{\text{O}_2\text{(g)}} \rightarrow \cancel{\text{H}_2\text{O(g)}} \quad \Delta H_c = -286 \text{ kJ mol}^{-1}$
- $\text{C}_2\text{H}_2\text{(g)} + 2\frac{1}{2}\text{O}_2\text{(g)} \rightarrow 2\text{CO}_2\text{(g)} + \text{H}_2\text{O(g)} \quad \Delta H_c = -1299 \text{ kJ mol}^{-1}$
- $2\cancel{\text{CO}_2\text{(g)}} + \cancel{\text{H}_2\text{O(g)}} \rightarrow \text{C}_2\text{H}_2\text{(g)} + 2\frac{1}{2}\cancel{\text{O}_2\text{(g)}} \quad \Delta H_c = +1299 \text{ kJ mol}^{-1}$

Adding “bulleted” equations gives us $2\text{C(graphite)} + \text{H}_2\text{(g)} \rightarrow \text{C}_2\text{H}_2\text{(g)}$

$$\Delta H_f = (2 \times -394) + (-286) + 1299$$

$$\Delta H_f = +225 \text{ kJ mol}^{-1}$$



4. Using the following standard enthalpy changes of formation, $\Delta H^\circ_f / \text{kJmol}^{-1}$: $\text{CO}_2(\text{g})$, -394; $\text{H}_2\text{O}(\text{g})$, -286; $\text{C}_2\text{H}_5\text{OH}(\text{l})$, -278 calculate the standard enthalpy of combustion of ethanol i.e. the enthalpy change for the reaction $\text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$

“Using the Second method”



Add bulleted equations



Solve equation for ΔH_c

$$\Delta H_c = +278 + (2 \times -394) + (3 \times -286)$$

$$\Delta H_c = -1368 \text{ kJ mol}^{-1}$$

