Enthalpy of Combustion and Hess's Law

Enthalpies of Combustion, $\Delta_c H$ can be obtained relatively easily <u>**by experiment**</u> using apparatus like a Thiemann calorimeter or a Bomb calorimeter.

Write a balanced chemical equation for reaction that would represent the ΔH_c for **ethane**.

→		$\Delta_c H$ (**C_2H_6**)= ....kJ mol^{-1}

Write a balanced chemical equation for reaction that would represent the ΔH_c for **ethene**.

→		$\Delta_c H$(**C_2H_4**)= ....kJ mol^{-1}

Write a balanced chemical equation for reaction that would represent the ΔH_c for **hydrogen**.

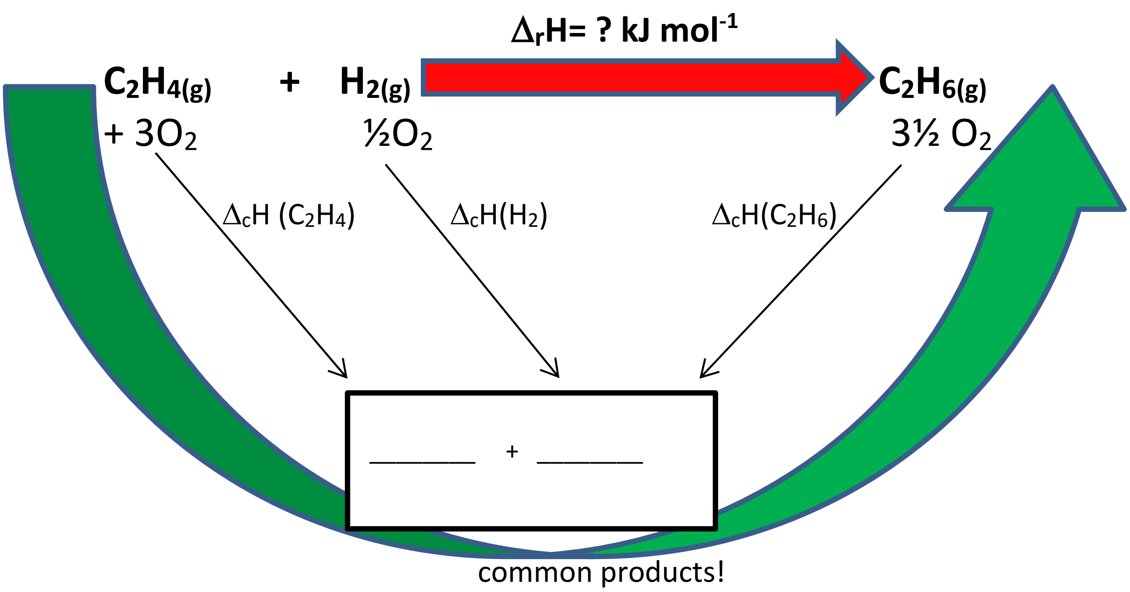
→		$\Delta_c H$(**H_2**)= ....kJ mol^{-1}

Now consider the following chemical reaction. It represents the conversion of ethene into ethane.

ethene	+ hydrogen	→	ethane		**$\Delta_r H$= ? kJ mol^{-1}**
C_2H_4	+ H_2	→	C_2H_6		

The enthalpy change ΔH_r for this reaction is impossible to obtain by a simple experiment.

However, there is a way of working it out by applying Hess's Law. Complete the box below.



A couple of important things to notice:

- The combustion of both the reactants and products gives the **same products** in the same ratios, i.e. carbon dioxide and water. These common products form a kind of 'steppingstone' on the journey along the alternative route!
- The oxygen required for each of the combustion reaction is the same on both sides of the reaction and so it 'cancels out' in the main reaction which is the conversion of ethene and hydrogen into ethane. Put another way, it appears in the main reaction but isn't involved in the main reaction!

So, if we can't go via the red arrow, i.e. directly, then we can get from reactants to products by travelling along the green arrow. This involves adding the enthalpy changes that form part of the route. Each enthalpy change is in ( ).

You will notice that two of the arrows are in the direction that you are travelling. These are the enthalpies of combustion for both reacting molecules. Notice how one of the legs of the journey is heading the 'wrong way up the arrow'. You simply need to remember to reverse the sign of this enthalpy change before adding it to the others.

So, writing it out **carefully in full**:

$$\Delta_r H = (\Delta_c H(C_2H_4)) + (\Delta_c H(H_2)) + (-\Delta_c H(C_2H_6))$$

ROUTE 1 = [……………………………………………ROUTE 2 ……………………………………………]

Putting in the values:

$$\Delta_r H = (\ldots\ldots\ldots) + (\ldots\ldots\ldots) + (-\ldots\ldots\ldots) = \ldots\ldots\ldots \text{kJmol}^{-1}$$

Now have a go at drawing a Hess Cycle for the very similar reaction which is:

$$C_3H_{6(g)} + H_{2(g)} \xrightarrow{\Delta_r H = \ ?\ \text{kJ mol}^{-1}}$$