

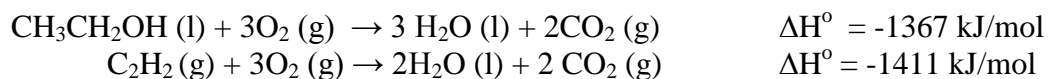
Hess's Laws

Part 1

Hess's Law of Heat Summation states: the enthalpy change of a physical or chemical process depends upon the beginning and end conditions (reactants versus products). The enthalpy change is independent of the chemical pathways (number of reaction steps). **It is the sum of enthalpy changes of all the individual steps in the process**

- ▶ Each ΔH for the steps in a reaction mechanism can be added to calculate the ΔH for the overall reaction
- ▶ In this type of question, the reaction steps will be provided, as well as their ΔH . The overall reaction will also be provided ** You have to manipulate the reaction steps to get them to add up to the overall reaction. FLIP, MULTIPLY, CANCEL

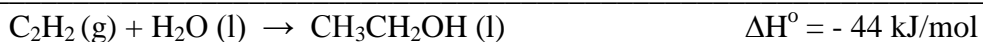
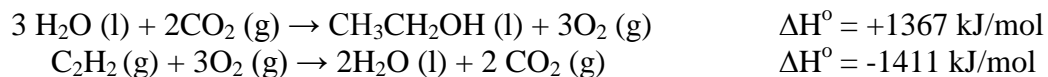
Example #1: Ethene, C_2H_2 (g) reacts with water to form ethanol, CH_3CH_2OH (l)
Determine the enthalpy change of this reaction, given the following thermochemical equations:



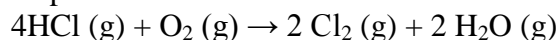
So our challenge is to manipulate the two given equations to get our overall reaction of :
 $C_2H_2 (g) + H_2O (l) \rightarrow CH_3CH_2OH (l) \quad \Delta H^\circ = ?$

Study what you have : we need the $CH_3CH_2OH (l)$ on the product side and we need to get rid of the 2 $CO_2 (g)$ and the 3 $O_2 (g)$ by canceling out . REMEMBER IF WE FLIP A REACTION AROUND - THE ΔH° WILL BE THE INVERSE.

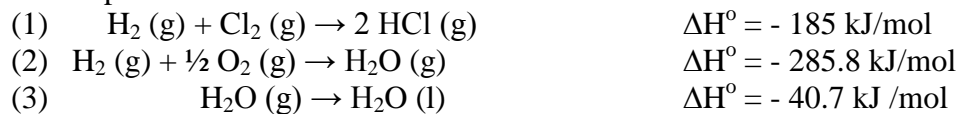
Flip the first reaction:



Example #2 : Mixing household cleaners can result in the production of hydrogen chloride gas, HCl (g). Not only is this gas dangerous in its own right, but it also reacts with oxygen to form chlorine gas and water vapour. The overall reaction is :



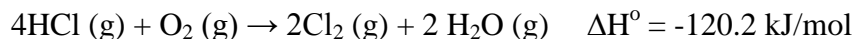
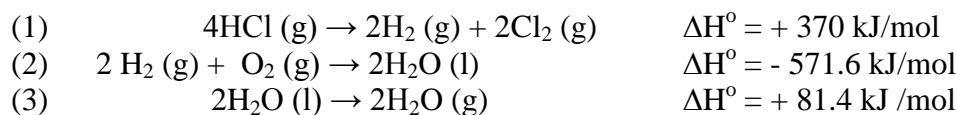
The steps are :



Step (1) flip and multiply by 2 (including the ΔH° !)

Step (2) double to get rid of the H_2

Step (3) double and flip to make sure there are 2 $\text{H}_2\text{O (g)}$ and the $\text{H}_2\text{O (l)}$ are gone



Hess's Law Part 2

Hess's Law can also be done using standard molar enthalpies of formation (ΔH°_f)

- ▶ The energy change required to form a compound can be calculated based on the energy required to make its elements
- ▶ UNDER STANDARD CONDITIONS THE STANDARD MOLAR ENTHALPY OF FORMATION OF ANY ELEMENT IS ZERO - elements are free!
- ▶ These concepts can be applied when calculating the enthalpy change of a chemical reaction
- ▶ **Add the standard molar heats of formation of the products, add the standard molar heats of the reactants, and then subtract** The values would have to be provided for you, or all values except one with the total.
- ▶
$$\Delta\text{H}^\circ_{\text{rxn}} = \sum \Delta\text{H}^\circ_f \text{ products} - \sum \Delta\text{H}^\circ_f \text{ reactants}$$

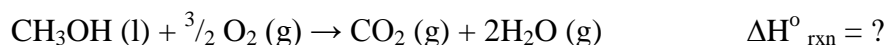
(this formula is on back of periodic table)

Example #1

Calculate the standard enthalpy change ($\Delta H^\circ_{\text{rxn}}$) for the combustion of methanol (CH_3OH), given the following values :

Compound	ΔH°_f (kJ/mol)
CH_3OH	-238.6
CO_2	-393.5
$\text{H}_2\text{O (g)}$	-241.8

Balance in terms of one mole of the methanol.



Apply your formula : $\Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_f \text{ products} - \sum \Delta H^\circ_f \text{ reactants}$
Remember to use the coefficients in your balanced equation

$$\begin{aligned} \Delta H^\circ_{\text{rxn}} &= \sum [(n \Delta H^\circ_f \text{CO}_2) + (n \Delta H^\circ_f \text{H}_2\text{O})] - \sum [(n \Delta H^\circ_f \text{CH}_3\text{OH}) + (n \Delta H^\circ_f \text{O}_2)] \\ &= [(1 \text{ mol} \times -393.5 \text{ kJ/mol}) + (2 \text{ mol} \times -241.8 \text{ kJ/mol})] - [(1 \text{ mol} \times -238.6 \text{ kJ/mol}) + (\frac{3}{2} \text{ mol} \times 0 \text{ kJ/mol})] \\ &= (-877.1 \text{ kJ}) - (-238.6 \text{ kJ}) \\ &= -638.5 \text{ kJ} \end{aligned}$$

Calculation of ΔH Using Bond Energies

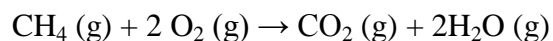
(Another way to calculate $\Delta H^\circ_{\text{rxn}}$)

- ◆ Every bond contains energy
- ◆ The energy trapped in reactant bonds can be compared to the energy trapped in product bonds - the difference is $\Delta H^\circ_{\text{rxn}}$
- ◆ The formula : $\Delta H^\circ_{\text{rxn}} = \Delta H_{\text{break bonds}} - \Delta H_{\text{form bonds}}$

Which is the same as $\Delta H^\circ_{\text{rxn}} = \sum \text{bond energies of reactants} - \sum \text{bond energies of products}$
(This equation is also on the back of your periodic table)

- ◆ Bond energies are always provided (or total value and all except one)
- ◆ Count all bonds and multiply by the appropriate bond energy value

Example : Calculate using bond energies, the ΔH for the combustion of methane.



Bond	Bond Energies (kJ/mol)
C-H	413
O=O	498
C=O	745
O-H	467

$$\Delta H^\circ_{\text{rxn}} = \sum \text{bond energies of reactants} - \sum \text{bond energies of products}$$

$$= \sum [(4 \times 413 \text{ kJ/mol}) + (2 \text{ mol} \times 498 \text{ kJ/mol})] - \sum [(2 \times 745 \text{ kJ/mol}) + (2 \times 2 \text{ mol} \times 467 \text{ kJ/mol})]$$

$$= 2648 \text{ kJ} - 3358 \text{ kJ}$$

$$= -710 \text{ kJ}$$

This type of question would provide the bond energies of all of them and calculate the total, or they would provide the total and all energies except one for you to solve.

AND FINALLY ; a hierarchal order of increasing magnitude of energy changes:

Physical Changes = tens of kJ/mol

Chemical Changes = hundreds to thousands of kJ/mol

Nuclear Changes = millions to billions of kJ/mol