

ENTHALPY OF FORMATION, ΔH_f

Enthalpy of formation is the molar enthalpy change for the formation of a compound from the elements in their standard states.

Example Question 1

Write the formation equation and give the enthalpy of formation for liquid methanol, $\text{CH}_3\text{OH}(\text{l})$

SOLUTION

Step 1: Write the equation to produce one mole of product.	$\rightarrow \text{CH}_3\text{OH}(\text{l})$
Step 2: Write the reactants as the elements in their standard states.	$\text{C}(\text{s}) + \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$
Step 3: Balance the equation keeping the product at one mole.	$\text{C}(\text{s}) + 2 \text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$
Step 4: Add the enthalpy of formation. Enthalpies of formation are found in the information package.	$\text{C}(\text{s}) + 2 \text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l}) \quad \Delta H_f = -239.1 \text{ kJ/mol}$

⇒ Try this one:

Write the formation equation and give the enthalpy of formation for ammonium nitrate, $\text{NH}_4\text{NO}_3(\text{s})$.

Go to the end of this lesson for the answer.

NOTES

- The standard state of an element is the element's most stable state at SATP (25 °C and 100 kPa). For example, oxygen's standard state is diatomic and gaseous, $\text{O}_2(\text{g})$, because this is oxygen's most stable state at SATP. Oxygen can also exist as a triatomic molecule (ozone, O_3) at SATP but this is not oxygen's most stable state.
- The enthalpy of formation for an element in its standard state is zero.
- Enthalpies of formation are found on page 18 of the course manual.
- The degree symbol (°) written with standard enthalpy of formation, ΔH_f° , just means that the value is for standard conditions. We will always assume standard conditions, and not include the symbol.

The enthalpy of formation for a substance can be thought of as a measure of the energy of the substance relative to the elements that make up the substance. Example: As a compound, one mole of methanol has 239.1 kJ less energy than the carbon, hydrogen, and oxygen.

In other words, we can use enthalpies of formation as a relative measure of the energy of a substance.

Therefore, we can calculate the enthalpy change (energy change) for a reaction by calculating the change in the enthalpies of formation between the reactants and the products. The formula is . . .

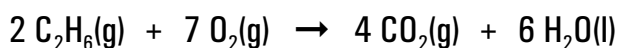
$$\Delta H = \sum n \cdot \Delta H_{f(\text{products})} - \sum n \cdot \Delta H_{f(\text{reactants})} \quad \text{where } \Sigma \text{ means summation (add all the values together).}$$

Example Question 2

Use enthalpies of formation to calculate the molar enthalpy of combustion for ethane?

SOLUTION

Step 1: Write the balanced equation for the reaction.



Step 2: Look up the enthalpy of formation for each reactant and product.

Substance	$\text{C}_2\text{H}_6(\text{g})$	$\text{O}_2(\text{g})$	$\text{CO}_2(\text{g})$	$\text{H}_2\text{O}(\text{l})$
Enthalpy of Formation	-83.8 kJ/mol	0 kJ/mol (element in standard state)	-393.5 kJ/mol	-285.8 kJ/mol

Step 3: Calculate the enthalpy change as the change in the enthalpies of formation (take into account the amount of each substance in the balanced equation).

$$\begin{aligned} \Delta H &= \sum n \cdot \Delta H_{f(\text{products})} - \sum n \cdot \Delta H_{f(\text{reactants})} \\ &= [4 \text{ mol}(-393.5 \text{ kJ/mol}) + 6 \text{ mol}(-285.8 \text{ kJ/mol})] - [2 \text{ mol}(-83.8 \text{ kJ/mol}) + 7 \text{ mol}(0 \text{ kJ/mol})] \\ &\quad \text{carbon dioxide} \qquad \qquad \text{water} \qquad \qquad \text{ethane} \qquad \qquad \text{oxygen} \\ &= [-3288.8 \text{ kJ}] - [-167.6 \text{ kJ}] \\ &= -3121.2 \text{ kJ} \end{aligned}$$

Step 4: The question asks for molar enthalpy of combustion.

$$\Delta H_{\text{comb}} = \frac{\Delta H}{n} = \frac{-3121.2 \text{ kJ}}{2 \text{ mol C}_2\text{H}_6} = -1560.6 \text{ kJ/mol C}_2\text{H}_6$$

Therefore, the molar enthalpy of combustion for ethane is $-1560.6 \text{ kJ/mol C}_2\text{H}_6$

Answer to **Try this one:**

