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, $CH_3OH(l)$

SOLUTION

Step 1: Write the equation to produce one mole of product.	$\rightarrow$ **$CH_3OH(l)$**
Step 2: Write the reactants as the elements in their standard states.	**$C(s)$** + **$H_2(g)$** + **$O_2(g)$** $\rightarrow$ $CH_3OH(l)$
Step 3: Balance the equation keeping the product at one mole.	$C(s)$ + **2** $H_2(g)$ + **½** $O_2(g)$ $\rightarrow$ $CH_3OH(l)$
Step 4: Add the enthalpy of formation. Enthalpies of formation are found in the information package.	$C(s) + 2\ H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow CH_3OH(l)$ **$\Delta H_f = -239.1$ kJ/mol**

✏ **Try this one:**

Write the formation equation and give the enthalpy of formation for ammonium nitrate, $NH_4NO_3(s)$.

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, $O_2(g)$, 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, $\Delta H_f°$, just means that the value is for standard conditions. We will alway 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 calculated the change in the enthalpies of formation between the reactants and the products. The formula is . . .

$\mathbf{\Delta H = \sum n \cdot \Delta H_{f\,(products)} - \sum n \cdot \Delta H_{f\,(reactants)}}$ where $\sum$ 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.

$$2\,C_2H_6(g) + 7\,O_2(g) \rightarrow 4\,CO_2(g) + 6\,H_2O(l)$$

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

Substance	$C_2H_6(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(l)$
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).

$$\Delta H = \sum n \cdot \Delta H_{f\,(products)} - \sum n \cdot \Delta H_{f\,(reactants)}$$

$$= [\underbrace{4\,\cancel{mol}(-393.5\,kJ/\cancel{mol})}_{\text{carbon dioxide}} + \underbrace{6\,\cancel{mol}(-285.8\,kJ/\cancel{mol})}_{\text{water}}] - [\underbrace{2\,\cancel{mol}(-83.8\,kJ/\cancel{mol})}_{\text{ethane}} + \underbrace{7\,\cancel{mol}(0\,kJ/\cancel{mol})}_{\text{oxygen}}]$$

$$= [-3288.8\,kJ] - [-167.6\,kJ]$$

$$= -3121.2\,kJ$$

Step 4: The question asks for <u>molar</u> enthalpy of combustion.

$$\Delta H_{comb} = \frac{\Delta H}{n} = \frac{-3121.2\,kJ}{2\,mol\ C_2H_6} = -1560.6\,kJ/mol\ C_2H_6$$

Therefore, the molar enthalpy of combustion for ethane is -1560.6 kJ/mol C_2H_6

Answer to **Try this one:**

$$N_2(g) + 2\,H_2(g) + \frac{3}{2}\,O_2(g) \rightarrow NH_4NO_3(s) \qquad \Delta H_f = -365.6\,kJ/mol$$