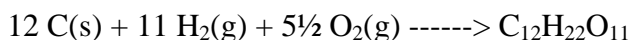


## Values of $H_f^\circ$ from Standard Enthalpies of Combustion

To measure directly the heat of formation of sucrose,  $C_{12}H_{22}O_{11}$ , you would have to carry out the following reaction:



But no one has ever been able to figure out how to make this reaction occur directly under any conditions, so there is no direct way to measure the  $H_f^\circ$  of sucrose. How, then, can we obtain values of  $H_f^\circ$  for compounds such as sucrose?

If the compound in question can be burned - which is usually far easier to do than make it from its elements - then we have a source of energy data that we can use to calculate its  $H_f^\circ$ . This is because the products of combustion are nearly always compounds whose values of  $H_f^\circ$  are known or can be measured by direct means. The combustion of sucrose in an atmosphere of pure oxygen proceeds by the following equation:



If the standard enthalpy change for this reaction can be measured, and if we can look up the values of  $H_f^\circ$  for three of the four chemicals in the equation, then we can use the enthalpies of formation equation to find the  $H_f^\circ$  of the remaining substance, sucrose.

The above equation is a combustion equation and data tables would have  $H_c^\circ$  values (standard enthalpy of combustion). The combustion of sucrose has a value of -5639.7 kJ/mol :

$$H_c^\circ = -5639.7 \text{ kJ/mol} \qquad H_c^\circ = [(12)\text{CO}_2\text{(g)} + (11)\text{H}_2\text{O(l)}] - [\text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(s)} + (12)\text{O}_2\text{(g)}]$$

All values are in kJ/moles.

$$-5639.7 = [(12)(-393.5) + (11)(-285.8)] - [\text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(s)} + (12)(0)]$$

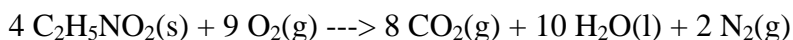
$$-5639.7 = -4722 - 3143.8 - [\text{C}_{12}\text{H}_{22}\text{O}_{11}]$$

$$2226.1 = -[\text{C}_{12}\text{H}_{22}\text{O}_{11}]$$

Therefore the  $H_f^\circ$  of sucrose,  $C_{12}H_{22}O_{11}$ , is -2226.1 kJ/mole

## Sample problem

One of the "building blocks" for proteins such as those in muscles and sinews is an amino acid called glycine,  $\text{C}_2\text{H}_5\text{NO}_2$ . The equation for its combustion is



The value of  $H_c^\circ$  for glycine is  $-973.49 \text{ kJ/mole}$ . Using this information and the values of  $H_f^\circ$  calculate the  $H_f^\circ$  for glycine.

## Solution

For this problem, enthalpy of formation equation becomes

$$\Delta H^\circ = [(8)\text{CO}_2(\text{g}) + (10)\text{H}_2\text{O}(\text{l}) + (2)\text{N}_2(\text{g})] - [(4)\text{C}_2\text{H}_5\text{NO}_2(\text{s}) + (9)\text{O}_2(\text{g})]$$

No, I didn't forget the  $H_c^\circ$ . The first term  $\Delta H^\circ$ , is obtained from the standard heat of combustion of glycine. Since the chemical equation for this reaction is for the combustion of *four* moles of glycine, we have to multiple  $H_{\text{combustion}}^\circ$  by four.

$$H^\circ = 4 \text{ mol} \times -973.49 \text{ kJ/mol} = -3894.0 \text{ kJ}$$

Now we can substitute into enthalpy of formation equation the correct values.

$$\begin{aligned} -3894.0 \text{ kJ} &= [(8)-393.5 + (10)-285.8 + (2)0] - [(4)\text{C}_2\text{H}_5\text{NO}_2 + (9)0] \\ -3894.0 &= -3148.0 - 2858 - [(4)\text{C}_2\text{H}_5\text{NO}_2] \\ 2112.0 &= -[(4)\text{C}_2\text{H}_5\text{NO}_2] \end{aligned}$$

Therefore by rearranging we get

$$H_f^\circ \text{ for } \text{C}_2\text{H}_5\text{NO}_2 = \frac{-2112.0 \text{ kJ/mole}}{4 \text{ moles}} = -528.0 \text{ kJ/mole}$$

Thus, the standard heat of formation of glycine is  $-528.0 \text{ kJ/mol}$ , and we have seen how we can determine this quantity without making glycine directly from its elements.