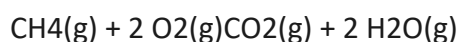


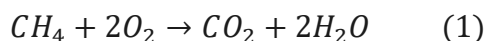
Question #83146

scientist measures the standard enthalpy change for the following reaction to be -816.9 kJ:



Based on this value and the standard enthalpies of formation for the other substances, the standard enthalpy of formation of $\text{H}_2\text{O}(\text{g})$ is kJ/mol.

Solution:



According to the chemical equation (1) and Hess's law [1], the standard enthalpy of formation of H_2O is equal to:

$$\Delta H_f^\circ(\text{H}_2\text{O}) = \frac{\Delta H_{\text{reaction}}^\circ - \Delta H_f^\circ(\text{CO}_2) + \Delta H_f^\circ(\text{CH}_4)}{2}$$
$$\Delta H_f^\circ(\text{H}_2\text{O}) = \frac{-816.9 - (-393.509) + (-74.9)}{2} = -249.1455 \approx -249.15 \text{ kJ/mol}$$

Answer:

The standard enthalpy of formation of H_2O is equal to -249.15 (based on the standard enthalpy change for the following reaction (measured) and the standard enthalpies of formation for the other substances (CH_4 , CO_2) [2]).

References:

[1] https://en.wikipedia.org/wiki/Hess%27s_law

[2] https://en.wikipedia.org/wiki/Standard_enthalpy_of_formation

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