A 0.424 g sample of liquid C_5H_{12} was combusted completely using excess oxygen inside a bomb (constant volume) calorimeter, with the products being carbon dioxide and liquid water. The calorimeter's heat capacity is 4.782 kJ $°C^{-1}$. If the temperature inside the calorimeter increased from 25.0 °C to 33.4 °C, determine ΔH for this reaction with respect to the system in kJ mol^{-1} at 298 K. Do not worry about how realistic the final answer is.

Solution:

Balanced equation for task:

$C_5H_{12} + 8O_2 = 5CO_2 + 6H_2O$;

As the reaction proceeds, heat, produced by combustion of the hydrocarbon is transferred to calorimeter and heats him up from 25 to 33.4 degrees Celsius. Consequently, amount of energy transferred to calorimeter is equal to amount, produced by combusustion of hydrocarbon:

$Q_c = Q_1$;

$C*(t_2 – t_1) = Q_1$;

$4.782 \frac{kJ}{°C}*(33.4\ °C – 25°) = Q_1$;

Q_1=40.1688 kJ;

(where Q_c – amount of heat received by calorimeter, Q_1 – amount of heat produced by combustion, C – heat capacity of calorimeter).

This amount of heat is produced by 0.424 g of pentane. We need to determine its amount of moles:

$n(C_5H_{12}) = \frac{m(C_5H_{12})}{M\ (C_5H_{12})} = \frac{0.424\ g}{72\ g/mol}$ = 0.005889 moles.

As the enthalpy sign is opposite to sign of heat effect, and enthalpy change for reaction is counted for one mole of reactant, we get following:

$\Delta H_{reaction} = \frac{-40.1688\ kJ}{0.005889\ mol}$ = -6820.9883 kJ/mol.

Answer:

Enthalpy change in this reaction is -6820.9883 kJ/mol.