Answer on Question #56951 - Chemistry - Other

Question:

A 2.56 g sample of anthracene, C14H10, was burned to heat an aluminum calorimeter (mass=948 g). The calorimeter contained 1.50 L of water with an initial temperature of 20.5 degrees Celsius and a final temperature of 34.3 degrees Celsius. a) Calculate the molar heat of combustion of anthracene b) Write the thermochemical equation, two ways, for the complete combustion of anthracene c) If the actual value of delta H = -7150 kJ/mol, what is the percentage error?

Solution

First we need to calculate the heat of combustion.

$q=cm\Delta T$

The amount of heat required to change water temperature:

$q_1 = 4200\times1.5\times(34.3-20.5) = 86940$ J

The amount of heat required to change aluminium bomb temperature:

$q_2 = 897\times0.948\times(34.3-20.5) = 11730$ J

Total amount of heat absorbed is

$q = q_1 + q_2 = 98670$ J

The amount of substance of antracene is

n = m/M = 2.56/178.23 = 0.014 mol

The molar heat of combustion is

$q_m = 98670/0.014 =$ $\mathbf{7.048\times10^6}$ **J/mol**

A balanced equation is

$2\ C_{14}H_{10} + 33\ O_2 = 28\ CO_2 + 10\ H_2O$

In this reaction two moles of anthracene reacts, so the amount of heat is

$qrxn = 2\times7.048\times10^6 = 1.4096\times10^7$ J or 14096 kJ

The thermochemical equation is

$\mathbf{2\ C_{14}H_{10} + 33\ O_2 = 28\ CO_2 + 10\ H_2O \quad + 14096\ kJ}$

or

$\mathbf{2\ C_{14}H_{10} + 33\ O_2 = 28\ CO_2 + 10\ H_2O \quad \Delta H = -14096\ kJ}$

The percentage error is

$$\varepsilon = \frac{7150 - 7048}{7150} \times 100\% = \mathbf{1.4\%}$$