

## Answer on Question #56951 - Chemistry - Other

### Question:

A 2.56 g sample of anthracene, C<sub>14</sub>H<sub>10</sub>, 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:

$$q_1 = 4200 \times 1.5 \times (34.3 - 20.5) = 86940 \text{ J}$$

The amount of heat required to change aluminium bomb temperature:

$$q_2 = 897 \times 0.948 \times (34.3 - 20.5) = 11730 \text{ J}$$

Total amount of heat absorbed is

$$q = q_1 + q_2 = 98670 \text{ J}$$

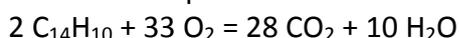
The amount of substance of anthracene is

$$n = m/M = 2.56/178.23 = 0.014 \text{ mol}$$

The molar heat of combustion is

$$q_m = 98670/0.014 = \mathbf{7.048 \times 10^6 \text{ J/mol}}$$

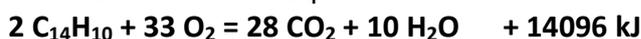
A balanced equation is



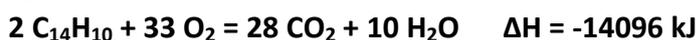
In this reaction two moles of anthracene reacts, so the amount of heat is

$$q_{rxn} = 2 \times 7.048 \times 10^6 = 1.4096 \times 10^7 \text{ J or } 14096 \text{ kJ}$$

The thermochemical equation is



or



The percentage error is

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