Answer on Question #81849 – Chemistry – General Chemistry

We have a calorimeter bomb with a heat capacity of 555 J/K. In this bomb, 1000.0 mL of water is placed. 2.476 g of a solid are burned in the calorimetric bomb. The temperature of the bomb and water increases by 2.05°C. The molar mass of this solid is 573.3 g/mol. How much heat would be released (in kJ, and note that we want the amount of heat that comes out) if we burned 0.155 mol of this solid in the bomb calorimetric? The specific heat of the water is 4.184 J/K/g. Make an approximation that the density of the water is 1.00 g/mL.

Solution:

The heat released by the reaction will be absorbed by the water in the calorimeter and the calorimeter itself:

$q_{\text{lost by solid}} = -(q_{\text{gained by water}} + q_{\text{gained by calorimeter}})$

$q_{water} = m \times C \times \Delta T$

$q_{water} = (1000.0\ g) \times (4.184\ J/g/°C) \times (2.05°C)$

$q_{water} = 8577.2\ J = 8.58\ kJ$

$q_{cal} = C_{cal} \times \Delta T$

$q_{cal} = (555\ J/°C) \times (2.05°C) = 1137.75\ J = 1.14\ kJ$

$q_{reaction} = -(8.58\ kJ + 1.14\ kJ) = 9.72\ kJ$

$Q_{reaction}$ (1 mol) = (9.72 kJ) / (2.476 g) $\times$ (573.3 g/mol) = 2250 kJ/mol

$Q_{reaction}$ (0.155 mol) = 2250 kJ/mol $\times$ 0.155 mol = 348.8 kJ