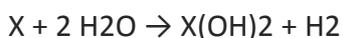


Answer on Question #81918, Chemistry / Organic Chemistry

A coffee-cup (constant pressure) calorimeter is used to carry out the following reaction in an unknown volume of water (where X is a hypothetical metal):

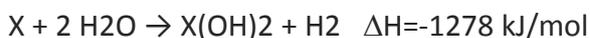


In this process, the water temperature rose from 25.0 °C to 37.1 °C. If 0.00579 mol of "X" was consumed during the reaction, and the ΔH of this reaction with respect to the system is -1278 kJ mol⁻¹, what volume of water (in mL) was present in the calorimeter?

The specific heat of water is 4.184 J g⁻¹ °C⁻¹

Solution

We should find the heat that was released by the reaction:



1278 kJ is released when 1 mole of "X" reacts with water

x kJ is released when 0.00579 mol of "X" react with water

Solve the proportion:

$$\frac{1278}{x} = \frac{1}{0.00579}$$
$$x = 7.39962$$

$$Q_{\text{released}} = 7.39962 \text{ kJ}$$

$$Q_{\text{released}} = Q_{\text{absorbed by water}}$$

$$Q_{\text{absorbed by water}} = cm\Delta T$$

$$7.39962 \times 10^3 \text{ J} = 4.184 \text{ J/g} \times ^\circ\text{C} \times m_{\text{water}} \times (37.1^\circ\text{C} - 25.0^\circ\text{C})$$

$$m_{\text{water}} = 146.16 \text{ g}$$

$$d_{\text{water}} = 1 \text{ g/cm}^3$$

$$V = m/d$$

$$V_{\text{water}} = 146.16 \text{ g} / 1 \text{ g/cm}^3 = 146.16 \text{ cm}^3 = 146.16 \text{ mL}$$

Answer: 146.16 mL