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):

X + 2 H2O → X(OH)2 + H2

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:

$X + 2\ H_2O \rightarrow X(OH)_2 + H_2 \quad \Delta H=-1278$ kJ/mol

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:

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

$Q_{released}$ = 7.39962 kJ

$Q_{relased}=Q_{absorbed\ by\ water}$

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

7.39962×10^3 J = 4.184 J/g$\times^{o}$C $\times m_{water}\times(37.1^{o}$ C-25.0^{0}C)

m_{water} = 146.16 g

d_{water} = 1g/cm^3

V=m/d

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

Answer: 146.16 mL

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