Task #82213

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 H_2O → $X(OH)_2$ + H_2

In this process, the water temperature rose from 25.0 °C to 32.2 °C. If 0.00803 mol of "X" was consumed during the reaction, and the ΔH of this reaction with respect to the system is -1798 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.

Firstly, we should find heat that was released by the reaction, using ΔH.

1 mole(“X”) – 1798 kJ

0.00803 mole(“X”) – x kJ

x = 0.00803 mole * 1798 kJ/1 mole = 14.44 kJ

x – Q, which was released.

$Q = Q_{water}$

Q = c*m*Δt, Δt = 32.2°C – 25.0°C = 7.2°C

Q = c*ρ*V* Δt

V = Q/ c*ρ*Δt

V = 14.44kJ/4.184 (J/g*°C) * 1 g/cm^3 * 7.2 °C = 479.44 cm^3 = 479.44 ml

Answer:

V = 14.44kJ/4.184 (J/g*°C) * 1 g/cm^3 * 7.2 °C = 479.44 cm^3 = 479.44 ml