Task #82145

A coffee-cup (constant pressure) calorimeter is used to carry out the following reaction in 500 mL of water (where X is a hypothetical metal):

$X + 2 H_2O \rightarrow X(OH)_2 + H_2$

In this process, the water temperature started at 25.0 °C and increased as the reaction progressed. If 0.00857 mol of "X" was consumed during the reaction, and the ΔH of this reaction with respect to the system is -1031 kJ mol^{-1} , what was the final temperature of the water in the calorimeter?

The specific heat of water is 4.184 J g^{-1} $°C^{-1}$

Solution.

Find the enthalpy change of water in this reaction ΔH (H_2O):

$\Delta H (H_2O) = n(X) * \Delta H$

$H (H_2O) = 8835.67$ J

$\Delta H (H_2O) = C_{water} * m(H_2O) * \Delta T$

$\Delta T = \Delta H (H_2O) / C_{water} * m(H_2O)$

$\Delta T = 8835.67/4.184*500 = 4.2$ °C

$\Delta T = T_2 - T_1$

$T_2 = \Delta T + T_1$

$T_2 = 4.2 + 25 = 29.2$ °C

Answer:

$T_2 = 4.2 + 25 = 29.2$ °C