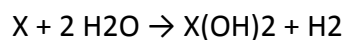


Task #82145

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):



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 $-1031 \text{ kJ mol}^{-1}$, what was the final temperature of the water in the calorimeter?

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

Solution.

Find the enthalpy change of water in this reaction $\Delta H (\text{H}_2\text{O})$:

$$\Delta H (\text{H}_2\text{O}) = n(\text{X}) * \Delta H$$

$$H (\text{H}_2\text{O}) = 8835.67 \text{ J}$$

$$\Delta H (\text{H}_2\text{O}) = C_{\text{water}} * m(\text{H}_2\text{O}) * \Delta T$$

$$\Delta T = \Delta H (\text{H}_2\text{O}) / C_{\text{water}} * m(\text{H}_2\text{O})$$

$$\Delta T = 8835.67 / 4.184 * 500 = 4.2 \text{ }^\circ\text{C}$$

$$\Delta T = T_2 - T_1$$

$$T_2 = \Delta T + T_1$$

$$T_2 = 4.2 + 25 = 29.2 \text{ }^\circ\text{C}$$

Answer:

$$T_2 = 4.2 + 25 = 29.2 \text{ }^\circ\text{C}$$