

Question

During electrolysis of aqueous solution containing 500 ml of NaBr, the cathode was obtained 0, 56 l hydrogen under normal conditions. Find the mass in grams of the substance was obtained in the anode and the pH of the solution.

Given

$$V_{\text{solution}} = 500 \text{ ml}$$

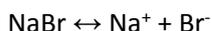
$$V_{\text{H}_2} = 0.56 \text{ l (normal condition)}$$

$$m = ?$$

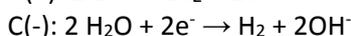
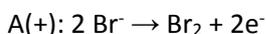
$$\text{pH} = ?$$

Solution

NaBr dissociates in the solution as follows

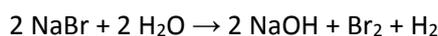


The electrode reactions are



The substance obtained in the anode is bromine (Br_2)

Overall reaction is



Hydrogen and bromine are formed with equimolar ratio. $V_m = 22.4 \text{ l/mol}$ (under normal conditions),

$M_{\text{Br}_2} = 159.8 \text{ g/mol}$. So, we may write a proportion

$$22.4, \text{ l/mol (H}_2) - 159.8, \text{ g/mol (Br}_2)$$

$$0.56, \text{ l (H}_2) - m, \text{ g (Br}_2)$$

Whence

$$m = \frac{159.8 \cdot 0.56}{22.4} = 3.995 \text{ g}$$

Since NaOH is formed, solution has basic pH that is calculated by the following equation

$$\text{pH} = 14 - \text{pOH} = 14 + \log [\text{OH}^-]$$

Molarity of hydroxyl ions is

$$[\text{OH}^-] = n(\text{OH}^-) / V_{\text{solution}}$$

By the reaction equation we know that 1 mole of H_2 corresponds to 2 moles of NaOH. So, we have the proportion

$$22.4, \text{ l/mol (H}_2) - 2, \text{ moles (NaOH)}$$

$$0.56, \text{ l (H}_2) - n, \text{ moles (NaOH)}$$

Whence

$$n = \frac{0.56 \cdot 2}{22.4} = 0.05 \text{ moles}$$

$$n(\text{OH}^-) = n(\text{NaOH}) = 0.05 \text{ moles}$$

Thus, the solution molarity

$$[\text{OH}^-] = 0.05 / 0.5 = 0.1 \text{ mol/l}$$

$$\text{pH} = 14 + \log 0.1 = 14 - 1 = 13$$

Answer: $m(\text{Br}_2) = 3.995 \text{ g}$, $\text{pH} = 13$