Question

During electrolysis of aqueous solution containing 500 ml of NaBr, the cathode was obtained o, 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

$V_{solution}$= 500 ml
V_{H2} = 0.56 l (normal condition)

m = ?
pH = ?

Solution

NaBr dissociates in the solution as follows
$NaBr \leftrightarrow Na^+ + Br^-$

The electrode reactions are
A(+): $2\ Br^- \rightarrow Br_2 + 2e^-$
C(-): $2\ H_2O + 2e^- \rightarrow H_2 + 2OH^-$

The substance obtained in the anode is bromine (Br_2)

Overal reaction is

$$2\ NaBr + 2\ H_2O \rightarrow 2\ NaOH + Br_2 + H_2$$

Hydrogen and bromine are formed with equimolar ratio. V_m = 22.4 l/mol (under normal conditions), M_{Br2} = 159.8 g/mol. So, we may write a proportion
22.4, l/mol (H_2) – 159.8, g/mol (Br_2)
0.56, l (H_2) – m, g (Br_2)
Whence

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

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

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

Molarity of hydrixyl ions is

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

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, l/mol (H_2) – 2, moles (NaOH)
0.56, l (H_2) – n, moles (NaOH)
Whence

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

$n(OH^-)$ = n(NaOH) = 0.05 moles

Thus, the solution molarity

$$[OH^-] = 0.05 / 0.5 = 0.1\ mol/l$$
$$pH = 14 + log\ 0.1 = 14 - 1 = 13$$

Answer: m(Br_2) = 3.995 g, pH = 13