

Answer on Question# 56944 - Chemistry - General chemistry

Question:

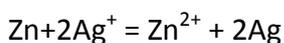
A battery is constructed at 25°C using the voltaic cell with initial concentrations $\text{Zn}|\text{Zn}^{2+}$ (0.100 M) $||$ Ag^+ (1.500 M) $|\text{Ag}$

How much does the cell voltage drop when 95% of the capacity (i.e., the concentration of Ag drops to 5% of its starting value) is consumed? The standard reduction potentials for the two half-cells are



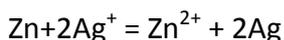
Solution

The balanced equation is



The final concentration of Ag^+ is $0.05 \times 1.5 = 0.075 \text{ M}$

The ICE table of this reaction is as follows



	Zn	2Ag ⁺	Zn ²⁺	2Ag
Initial	-	1.5	0.1	-
Change	-	-2x	+x	-
Equilibrium	-	0.075	0.1+0.713 = 0.813	-

$$1.5 - 2x = 0.075;$$

$$x = 0.713$$

The Nernst equation for this reaction is (for final concentrations)

$$E = E^\circ - \frac{RT}{zF} \ln Q$$

$$E^\circ = 0.80 - (-0.76) = 1.56 \text{ V}$$

$$E = 1.56 - \frac{0.059}{2} \lg \frac{0.813}{0.075^2} = 1.50 \text{ V}$$

The Nernst equation for this reaction is (for initial concentrations)

$$E = 1.56 - \frac{0.059}{2} \lg \frac{0.1}{1.5^2} = 1.6 \text{ V}$$

The cell voltage drop is $1.600 - 1.50 = 0.10 \text{ V}$

Answer: 0.10 V.