Corrections to previous calculations

Question: How much oxygen will be consumed in a sealed chamber containing an electrochemical oxygen sensor (R-22 or equal) that is in air at 25°C?

Given:

Faraday's Law: 96,500 coulombs = 1 equivalent

Reduction half reaction for oxygen:

$4OH^- = O_2 + 2H_2O + 4e^-$

therefore: 4 equivalents/mole

1 ampere = 1 coulomb/second

therefore: $127 \times 10^{-6} = 127 \times 10^{-6}$ coulombs/second

(worst case current in air at 25°C_

Gas Law:

$V = nRT/P$ (assume T = 25°C = 295°K)

Where:

n = number of moles

R = 0.08205 liter-atm-mole^{-1}-deg^{-1} (Gas Law Constant)

T = temperature (°K)

P = pressure (atmospheres)

Calculations:

127×10^{-6} coulombs / 96,500 coulombs/equivalent = 1.32×10^{-9} equivalents (each second)

\# of moles (each second) = 1.32×10^{-9} equivalents / 4 eq/mole = 3.29×10^{-10} moles

The volume of O_2 consumed (each second):

$$V = nRT/P$$

$$= \frac{(3.29 \times 10^{-10} \text{ moles})(0.08205 \text{ liter-atm-mole}^{-1}\text{-deg}^{-1})(295°K)}{(1 \text{ atm})}$$

$= 8.0 \times 10^{-9}$ liters of O_2 (each second)