

ELECTROCHEMISTRY

ELECTROLYTIC CELLS AND ELECTROLYSIS

CHEMISTRY 165 // SPRING 2020

PRACTICE PROBLEM 1

A current of 10.23 A is passed through a solution of silver nitrate for 1800. seconds. What mass of silver can be plated onto the cathode electrode?

— *answer* —

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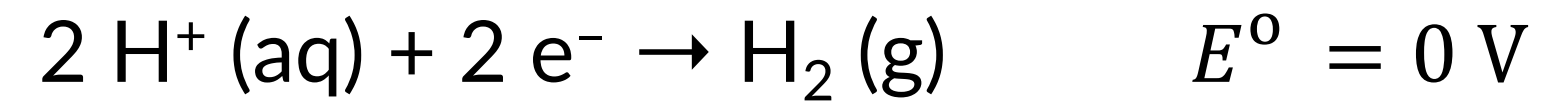
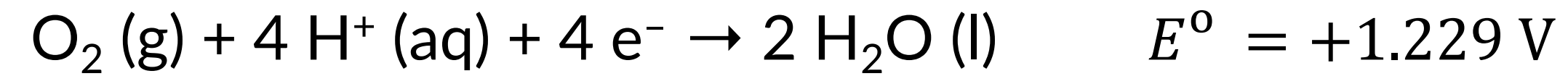
— *answer* —

The reaction we are interested in is the cathodic reaction: $\text{Ag}^+ (\text{aq}) + \text{e}^- \rightarrow \text{Ag} (\text{s})$.

We can determine the amount of Ag using the stoichiometry of the equation above where $n = 1$, $F = 96500 \frac{\text{C}}{\text{mol e}^-}$, and $1 \text{ A} = 1 \frac{\text{C}}{\text{s}}$.

$$m_{\text{Ag}} = 1800. \text{ s} \times \frac{10.23 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mol e}^-}{96500 \text{ C}} \times \frac{1 \text{ mol Ag}}{1 \text{ mol e}^-} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 20.59 \text{ g Ag}$$

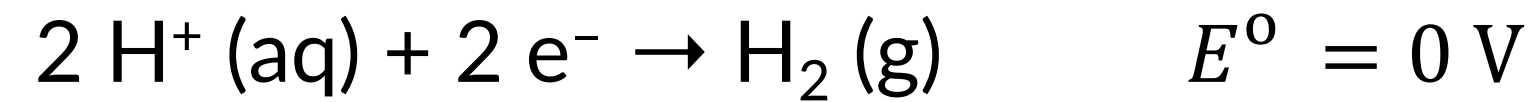
PRACTICE PROBLEM 2



An electrolytic cell is constructed from the reactions above. How many moles of $\text{O}_2 (\text{g})$, at 25°C and 1.00 atm , are produced if the electrolytic cell is operated at a current of 0.025 A for 1.0 hour ?

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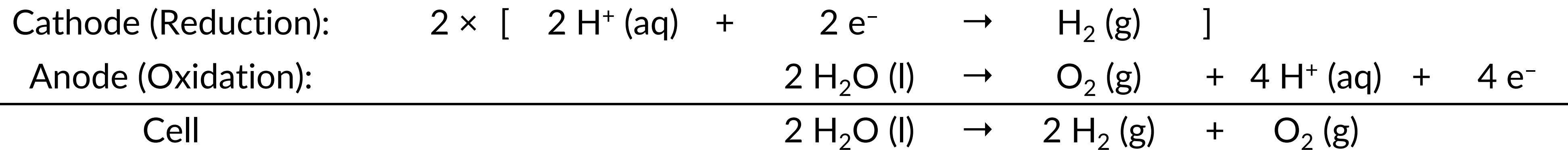
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First, write out the balanced net ionic equation for the **nonspontaneous** electrolytic cell reaction where $E_{\text{cell}}^\circ < 0 \text{ V}$ and $\Delta G^\circ > 0$.



We can determine the amount of O_2 using the stoichiometry of the equation above where $n = 4$, $F = 96500 \frac{\text{C}}{\text{mol e}^-}$, and $1 \text{ A} = 1 \frac{\text{C}}{\text{s}}$.

$$n_{\text{O}_2} = 1.0 \text{ hr} \times \frac{3600 \text{ s}}{1 \text{ hr}} \times \frac{0.025 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mol e}^-}{96500 \text{ C}} \times \frac{1 \text{ mol O}_2}{4 \text{ mol e}^-} = 2.3 \times 10^{-4} \text{ mol O}_2$$