An electrochemical cell is constructed with an open switch, as shown in the diagram above. A strip of Sn and a strip of unknown metal, X are used as electrodes. When the switch is closed, the mass of the Sn electrode increases. The half-reactions are shown below.

\[
\begin{align*}
\text{Sn}^{2+} (aq) + 2 \text{e}^- & \rightarrow \text{Sn}(s) & E^\circ &= -0.14 \text{ V} \\
\text{X}^{3+} (aq) + 3 \text{e}^- & \rightarrow \text{X}(s) & E^\circ &= ?
\end{align*}
\]

(a) In the diagram above, label the electrode that is the cathode. Justify your answer.

(b) In the diagram above, draw an arrow indicating the direction of electron flow in the external circuit when the switch is closed.

(c) If the standard cell potential \(E^\circ_{\text{cell}}\) is +0.60 V, what is the standard potential, in volts for the \(X^{3+}/X\) electrode?

(d) Identify metal X.

(e) Write balanced net-ionic equation for the overall chemical reaction occurring in the cell.

(f) In the cell, the concentration of \(\text{Sn}^{2+}\) is changed from 1.0 \(M\) to 0.50 \(M\), and the concentration of \(\text{X}^{3+}\) is changed from 1.0 \(M\) to 0.10 \(M\).
   (i) Substitute all appropriate values for determining the cell potential, \(E_{\text{cell}}\), into the Nernst equation. (Do not do any calculations.)
   (ii) On the basis of your response in (f) (i), will the cell potential be greater than, less than, or equal to \(E^\circ_{\text{cell}}\)? Justify your answer.
AP REVIEW QUESTIONS – Electrochemistry - Answers

Answer:
(a) tin electrode is the cathode; cathode is the site of reduction (gain in electrons) and will convert metal ions into a metal.

(b) (see diagram)

(c) red: \( \text{Sn}^{2+} (aq) + 2 \text{e}^- \rightarrow \text{Sn(s)} \) \( E^\circ = -0.14 \text{ V} \)

oxd: \( \text{X(s)} - 3 \text{e}^- \rightarrow \text{X}^{3+} (aq) \) \( E^\circ = +0.74 \text{ V} \)

\[ E^\circ_{\text{cell}} = +0.60 \text{ V} \]

red: \( \text{X}^{3+} (aq) + 3 \text{e}^- \rightarrow \text{X(s)} \) \( E^\circ = -0.74 \text{ V} \)

(d) Cr

(e) \( 3 \text{Sn}^{2+} + 2 \text{Cr} \rightarrow 3 \text{Sn} + 2 \text{Cr}^{3+} \)

(f) (i) \( E^\circ_{\text{cell}} = 0.60 - \frac{0.0592}{6} \log \left( \frac{(0.10)^2}{(0.50)^3} \right) \)

(ii) greater; •
An external direct-current power supply is connected to two platinum electrodes immersed in a beaker containing 1.0 M CuSO$_4$($aq$) at 25°C, as shown in the diagram above. As the cell operates, copper metal is deposited onto one electrode and O$_2(g)$ is produced at the other electrode. The two reduction half-reactions for the overall reaction that occurs in the cell are shown in the table below.

<table>
<thead>
<tr>
<th>Half-Reaction</th>
<th>$E^0$(V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>O$_2(g)$ + 4 H$^+$($aq$) + 4 e$^-$ $\rightarrow$ 2 H$_2$O($l$)</td>
<td>+1.23</td>
</tr>
<tr>
<td>Cu$^{2+}(aq)$ + 2 e$^-$ $\rightarrow$ Cu($s$)</td>
<td>+0.34</td>
</tr>
</tbody>
</table>

(a) On the diagram, indicate the direction of electron flow in the wire.

(b) Write a balanced net ionic equation for the electrolysis reaction that occurs in the cell.

(c) Predict the algebraic sign of $\Delta G^\circ$ for the reaction. Justify your prediction.

(d) Calculate the value of $\Delta G^\circ$ for the reaction.

An electric current of 1.50 amps passes through the cell for 40.0 minutes.

(e) Calculate the mass, in grams, of the Cu($s$) that is deposited on the electrode.

(f) Calculate the dry volume, in liters measured at 25°C and 1.16 atm, of the O$_2(g)$ that is produced.
AP REVIEW QUESTIONS – Electrochemistry - Answers

Answer:
(a) from the right to the left

(b) \[ 2 \text{Cu}^{2+}(aq) + 2 \text{H}_2\text{O}(l) \rightarrow 2 \text{Cu}(s) + \text{O}_2(g) + 4 \text{H}^+(aq) \]

(c) \(+\), a non-spontaneous reaction that requires the input of energy to take place

(d) \[ E^\circ = +0.34\text{v} + (-1.23\text{v}) = -0.89\text{v}; \quad \Delta G^\circ = -nF \Delta E^\circ = -(4)(96500)(-0.89) = 343540 \text{ J} = 340 \text{ kJ} \]

(e) \[ (1.50 \text{amps})(2400 \text{sec}) = 3600 \text{coul.}; \quad 3600 \text{coul.} \times \frac{1 \text{mol e}^-}{96500 \text{coul}} \times \frac{1 \text{mol Cu}}{2 \text{mol e}^-} \times \frac{63.55 \text{g}}{1 \text{mol Cu}} = 1.19 \text{g Cu} \]

(f) \[ 1.19 \text{g Cu} = 0.187 \text{mol Cu}; \quad \text{using a 2:1 ratio from equation in part (b), this gives } 0.00933 \text{ mol O}_2 \]

\[ V = \frac{nRT}{P} = \frac{(0.00933 \text{ mol})(0.0821 \frac{\text{L atm}}{\text{mol K}})(298 \text{K})}{1.16 \text{ atm}} = 0.197 \text{L O}_2 \]
In a hydrogen-oxygen fuel cell, energy is produced by the overall reaction represented above.

(a) When the fuel cell operates at 25°C and 1.00 atm for 78.0 minutes, 0.0746 mol of \( O_2(g) \) is consumed. Calculate the volume of \( H_2(g) \) consumed during the same time period. Express your answer in liters measured at 25°C and 1.00 atm.

(b) Given that the fuel cell reaction takes place in an acidic medium,

(i) write the two half reactions that occur as the cell operates,
(ii) identify the half reaction that takes place at the cathode, and
(iii) determine the value of the standard potential, \( E^\circ \), of the cell.

(c) Calculate the charge, in coulombs, that passes through the cell during the 78.0 minutes of operation as described in part (a).

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**Answer:**

(a) volume of \( H_2 = (2)(\text{mol. } O_2)(\text{molar volume @ 25°C}) = (2)(0.0746 \text{ mol})(24.45 \text{ L mol}^{-1}) = 3.65 \text{ L} \)

(b) (i) \( O_2(g) + 4 \text{ H}^+(aq) + 4 \text{ e}^- \rightarrow 2 \text{ H}_2\text{O(l)} \quad E^\circ = +1.23 \text{ V} \) (ii) cathode reaction

2 \( H_2(g) \rightarrow 4 \text{ H}^+(aq) + 4 \text{ e}^- \quad E^\circ = 0.00 \text{ V} \)

(iii) cell potential = \( 1.23 \text{ V} \)

(c) \( 0.0746 \text{ mol } O_2 \times \frac{4 \text{ mol } \text{e}^-}{1 \text{ mol } O_2} \times \frac{96500 \text{ coul}}{1 \text{ mol } \text{e}^-} = 28,800 \text{ coul.} \)