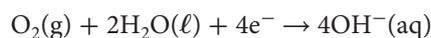


(Student textbook page 658)

$$E^{\circ}_{\text{cathode}} = +0.40 \text{ V}$$

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \\ &= +0.40 \text{ V} - (-0.45 \text{ V}) \\ &= +0.85 \text{ V} \end{aligned}$$

14. The standard cell potentials are based on 1.0 mol/L concentrations and SATP conditions. The concentrations and conditions in the environment where corrosion occurs will be different. The cell potentials will be different to those calculated.
15. Aluminum is a stronger reducing agent than iron is and will act as the anode, protecting the iron cathode.
16. The proximity to the Atlantic ocean results in a climate where salt-laden fog and generally humid conditions are common. This provides an electrolyte to connect dissimilar metals, making corrosion more common.
17. The tin covers the steel so that water and oxygen from the environment cannot reach the steel. This keeps the iron in the steel from corroding.
18. If the can is scratched, water and oxygen from the environment can now reach the steel under the tin. The iron in the steel now corrodes faster in contact with the tin than it would on its own. This happens because tin is a weaker reducing agent and less reactive than iron and acts as a cathode in each miniature galvanic cell on the surface of the can. The tin provides a large area of available cathodes for the small galvanic cells, and iron acts as the anode of each cell, and thus rusting is facilitated.

(Student textbook page 664)

19. A galvanic cell converts chemical energy into electrical energy, whereas an electrolytic cell converts electrical energy into chemical energy.
20. a. Reaction at anode: $2\text{Cl}^{-}(\ell) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^{-}$
 b. Reaction at cathode: $\text{Ca}^{2+}(\ell) + 2\text{e}^{-} \rightarrow \text{Ca}(\ell)$
 c. $\text{CaCl}_2(\ell) \rightarrow \text{Ca}(\ell) + \text{Cl}_2(\text{g})$
21. a. Reaction at negative electrode:
 $\text{Li}^{+}(\ell) + 1\text{e}^{-} \rightarrow \text{Li}(\ell)$
 b. Reaction at positive electrode:
 $2\text{Br}^{-}(\ell) \rightarrow \text{Br}_2(\ell) + 2\text{e}^{-}$
 c. $2\text{Li}^{+}(\ell) + 2\text{Br}^{-}(\ell) \rightarrow 2\text{Li}(\ell) + \text{Br}_2(\ell)$
22. The negative sign means that the reaction is not spontaneous. This negative value represents the minimum potential difference that you would have to

apply from an external power source to drive the cell reaction for the electrolytic cell.

23. The current in the electrolytic cell flows in one direction. Because alternating current constantly changes direction, it would not be able to drive the redox reaction in the electrolytic cell.
24. Sodium is solid at standard temperatures and will neither conduct an electric current nor be purified by electrolysis in an aqueous solution. Thus, Davy heated the salts until they melted and applied electrolysis to the molten solids to isolate sodium.

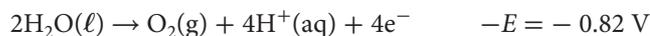
Answers to Caption Questions

Figure 10.5 (Student textbook page 637): $\text{Cu}^{2+}(\text{aq})$ ions gain electrons and deposit on the copper electrode. This causes the concentration of $\text{Cu}^{2+}(\text{aq})$ to decrease. The $\text{Zn}(\text{s})$ atoms lose electrons, and $\text{Zn}^{2+}(\text{aq})$ ions go into solution, causing the concentration of $\text{Zn}^{2+}(\text{aq})$ to increase.

Figure 10.7 (Student textbook page 638): Solid platinum is a very weak reducing agent, so it is less likely to participate in the reaction.

Figure 10.25 (Student textbook page 662): The non-standard reduction potentials for the reactions that occur at each electrode are as follows:

Oxidation:



Reduction:



The number of electrons gained during reduction must be the same as the number of electrons lost during oxidation. The equations show that the reduction reaction occurs twice for every oxidation reaction. Therefore, 2 mol $\text{H}_2(\text{g})$ is produced for every 1 mol $\text{O}_2(\text{g})$.

Answers to Section 10.1 Review Questions (Student textbook page 648)

- It is important to keep each half-cell separate so that an instantaneous reaction does not occur, "short-circuiting" the voltmeter.
- The salt bridge is needed to complete the circuit so that the cell can operate. The electrolyte in the salt bridge will have negative ions flow towards the anode and positive ions flow towards the cathode. This movement of ions is responsible for completing the flow of electricity in the cell. The ends of the salt bridge are plugged with cotton or glass wool to avoid the mixing of materials from the half-cell and to stop the flow of electrolytes into the half-cells.