21.2 Half-Cells and Cell Potentials

Connecting to Your World
Batteries provide current to power lights and many kinds of electronic devices—such as the personal digital assistant shown here. One type of battery that you're probably familiar with is a car battery. The electrical potential between the negative and positive terminals of a car battery is 12 V. Other common batteries are the ones used to power flashlights, radios, and CD players. They are commonly known as AAA, AA, C, and D batteries. The electrical potential in these batteries is 1.5 V. In this section, you will learn how to calculate electrical potential.

Electrical Potential
The electrical potential of a voltaic cell is a measure of the cell's ability to produce an electric current. Electrical potential is usually measured in volts (V). The potential of an isolated half-cell cannot be measured. For example, you cannot measure the electrical potential of a zinc half-cell or of a copper half-cell separately. When these two half-cells are connected to form a voltaic cell, however, the difference in potential can be measured. The electrical potential of a 1 M zinc–copper cell is +1.10 V.

The electrical potential of a cell results from a competition for electrons between two half-cells. The half-cell that has a greater tendency to acquire electrons will be the one in which reduction occurs. Oxidation occurs in the other half-cell as there is a loss of electrons. The tendency of a given half-reaction to occur as a reduction is called the reduction potential. The half-cell in which reduction occurs has a greater reduction potential than the half-cell in which oxidation occurs. The difference between the reduction potentials of the two half-cells is called the cell potential.

\[
\text{cell potential} = \left( \frac{\text{reduction potential}}{\text{of half-cell in which reduction occurs}} \right) - \left( \frac{\text{reduction potential}}{\text{of half-cell in which oxidation occurs}} \right)
\]

Or:

\[
E_{\text{cell}} = E_{\text{red}} - E_{\text{oxid}}
\]

Figure 21.8 A working voltaic cell can be constructed using a lemon and strips of copper and zinc.

Guide for Reading

Key Concepts
- What causes the electrical potential of an electrochemical cell?
- What value is assigned to the electrical potential of a standard hydrogen electrode?
- How can you determine the standard reduction potential of a half-cell?
- How can you interpret the cell potential of a redox reaction in terms of the spontaneity of the reaction?

Vocabulary
- electrical potential
- reduction potential
- cell potential
- standard cell potential
- standard hydrogen electrode

Reading Strategy
Monitoring Your Understanding
After you read the section, identify what you don't yet understand about half-cells and cell potentials. Reread the section or ask for help. Then, write your clarified understanding.

Connect to Your World

FOCUS

Objectives
21.2.1 Identify what causes the electrical potential of an electrochemical cell.
21.2.2 Define the standard hydrogen electrode.
21.2.3 Describe how the standard reduction potential of a half-cell is determined.
21.2.4 Interpret the meaning of the sign of the standard cell potential.

Guide for Reading

Build Vocabulary

Word Parts
Potential comes from the Latin potens meaning “power.” Electrical potential is the ability of a cell to produce an electric current. Ask students, What does a weather forecast mean when it says that there is a potential for rain? (It could possibly rain.)

Reading Strategy

Summarize
Read the first page of the text and, with the participation of the class, summarize the contents in a sentence or two. Ask students to write the summary in their notebooks. Encourage them to continue to summarize the principal ideas under each heading.

INSTRUCT

Ask, Which of the batteries mentioned is rechargeable? (the car battery) Do you use other rechargeable batteries? (Students may have experienced charging batteries in items such as cell phones, electric toothbrushes, and small vacuum cleaners.)
Section 21.2 (continued)

Electrical Potential

Use Visuals

Figure 21.8 Check student understanding of electrochemical, or voltaic, cells by ask students to identify the anode and cathode in the lemon battery. (Zinc is the anode and copper is the cathode.) What process goes on at the anode? (oxidation) What process goes on the cathode? (reduction) In which process are electrons lost? (oxidation) What role does the lemon play in the battery? (It is the salt bridge.

Standard Cell Potential

Use Visuals

Figure 21.9 Have students study the diagram. Ask, What is the significance of the standard hydrogen electrode? How is it different from other electrodes? (It has been assigned a standard reduction potential of 0.00 V at 25°C.) What does the standard hydrogen half-cell consist of? (A platinum electrode with hydrogen gas at atmospheric pressure bubbling over its surface; the electrode is immersed in 1.00M acid solution at 25°C.)

Discuss

Students may be confused by the arbitrary assignment of 0.00 V to the hydrogen half-cell. Use the Celsius temperature scale as an analogy. When the Celsius temperature scale was invented, 0°C was arbitrarily assigned to the temperature at which water freezes. All measured temperatures are relative to that assigned value. Temperatures higher than the freezing point of water are positive temperatures. Temperatures lower than 0°C are negative.

Standard Cell Potential

The standard cell potential \( E^{\circ}_{\text{cell}} \) is the measured cell potential when the ion concentrations in the half-cells are 1M, any gases are at a pressure of 101 kPa, and the temperature is 25°C. The symbols \( E^{\circ}_{\text{red}} \) and \( E^{\circ}_{\text{oxid}} \) represent the standard reduction potentials for the reduction and oxidation half-cells, respectively. The relationship between these values follows the general relationship for cell potential given on the previous page.

\[
E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red}} - E^{\circ}_{\text{oxid}}
\]

Because half-cell potentials cannot be measured directly, scientists have chosen an arbitrary electrode to serve as a reference. The standard hydrogen electrode is used with other electrodes so the reduction potentials of the other cells can be measured. The standard reduction potential of the hydrogen electrode has been assigned a value of 0.00 V. The standard hydrogen electrode, which is illustrated in Figure 21.9, consists of a platinum electrode immersed in a solution with a hydrogen-ion concentration of 1M. The solution is at 25°C. The electrode itself is a small square of platinum foil coated with finely divided platinum, known as platinum black. Hydrogen gas at a pressure of 101 kPa is bubbled around the platinum electrode. The half-cell reaction that occurs at the platinum black surface is as follows.

\[
2H^+ (aq, 1M) + 2e^- \rightarrow H_2(g, 101 kPa) \quad E^{\circ}_{\text{H}_2} = 0.00 \text{ V}
\]

The double arrows in the equation indicate that the reaction is reversible. The symbol \( E^{\circ}_{\text{red}} \) represents the standard reduction potential of H⁺. The standard reduction potential of H⁺ is the tendency of H⁺ ions to acquire electrons and be reduced to \( H_2(g) \). Whether this half-cell reaction occurs as a reduction or as an oxidation is determined by the reduction potential of the half-cell to which the standard hydrogen electrode is connected.

Checkpoint What components and conditions are required for a standard hydrogen electrode?

Standard Reduction Potentials

A voltaic cell can be made by connecting a standard hydrogen half-cell to a standard zinc half-cell, as shown in Figure 21.10. To determine the overall reaction for this cell, first identify the half-cell in which reduction takes place. In all electrochemical cells, reduction takes place at the cathode and oxidation takes place at the anode. A voltmeter gives a reading of +0.76 V when the zinc electrode is connected to the negative terminal and the hydrogen electrode is connected to the positive terminal. The zinc is oxidized, which means that it is the anode. Hydrogen ions are reduced, which means that the hydrogen electrode is the cathode. You can now write the half-reactions and the overall cell reaction.

Oxidation: \( \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^- \) (at anode)

Reduction: \( 2H^+ (aq) + 2e^- \rightarrow H_2(g) \) (at cathode)

Cell reaction: \( \text{Zn}(s) + 2H^+(aq) \rightarrow \text{Zn}^{2+}(aq) + H_2(g) \)

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Differentiated Instruction

Special Needs

Give students small cards with \( e^- \) printed on them in large letters. Divide students into pairs. Let one student role play being an anode and the other a cathode. The anode should hand one or more cards to the cathode.
You can determine the standard reduction potential of a half-cell by using a standard hydrogen electrode and the equation for standard cell potential. In the zinc–hydrogen cell, zinc was oxidized and hydrogen ions were reduced. Let $E_{\text{red}}^{\circ} = E_{\text{Zn}^{2+}}^{\circ}$ and $E_{\text{oxid}}^{\circ} = E_{\text{H}^+}^{\circ}$ in the standard cell potential equation.

$$E_{\text{cell}}^{\circ} = E_{\text{red}}^{\circ} - E_{\text{oxid}}^{\circ}$$

The cell potential ($E_{\text{cell}}^{\circ}$) was measured at $+0.76 \text{ V}$. The reduction potential of the hydrogen half-cell is a defined standard: $E_{\text{H}^+}^{\circ}$ always equals $0.00 \text{ V}$. Substituting these values into the preceding equation will give the standard reduction potential for the zinc half-cell.

$$+0.76 \text{ V} = 0.00 \text{ V} - E_{\text{Zn}^{2+}}^{\circ}$$

$$E_{\text{Zn}^{2+}}^{\circ} = -0.76 \text{ V}$$

The standard reduction potential for the zinc half-cell is $-0.76 \text{ V}$. The value is negative because the tendency of zinc ions to be reduced to zinc metal in this cell is less than the tendency of hydrogen ions to be reduced to hydrogen gas ($H_2$). Consequently, the zinc ions are not reduced. Instead, the opposite occurs: Zinc metal is oxidized to zinc ions.

Many different half-cells can be paired with the hydrogen half-cell in a similar manner. Using this method, the standard reduction potential for each half-cell can be obtained. For a standard copper half-cell, for example, the measured standard cell potential is $+0.34 \text{ V}$. Copper is the cathode, and Cu$^{2+}$ ions are reduced to Cu metal when the cell operates. The hydrogen half-cell is the anode, and $H_2$ gas is oxidized to $H^+$ ions. You can calculate the standard reduction potential for copper as follows.

$$E_{\text{Cu}^{2+}}^{\circ} = E_{\text{red}}^{\circ} - E_{\text{oxid}}^{\circ}$$

$$E_{\text{cell}}^{\circ} = E_{\text{Cu}^{2+}}^{\circ} - E_{\text{H}^+}^{\circ}$$

$$+0.34 \text{ V} = E_{\text{Cu}^{2+}}^{\circ} - 0.00$$

$$E_{\text{Cu}^{2+}}^{\circ} = +0.34 \text{ V}$$

This value is positive because the tendency for copper ions to be reduced in the cell is greater than the tendency of hydrogen ions to be reduced.

Differentiated Instruction

Less Proficient Readers

Have pairs of students look at the diagram of the voltaic cell in Figure 21.10. Have them trace the movement of the electrons from one electrode to the other. They should be able to explain that the electrons travel up from the zinc anode into the external wire, through the voltmeter, and down through the wire into the hydrogen/platinum cathode. Have them identify where oxidation is taking place and where reduction is taking place. Refer them to Table 21.2 to find the appropriate half-cell reactions and have them write the overall reaction for the cell.

Answers to...

Figure 21.10 Reduction occurs at the cathode. In this cell hydrogen ions are reduced to hydrogen gas.

Checkpoint

The standard hydrogen electrode consists of a platinum electrode immersed in a 1M acid solution at 25°C.
Discuss

Explain that the two half-cells of a voltaic cell are competing for electrons and that oxidation or reduction could occur in either cell. The half-cell with the more positive reduction potential will win the competition and undergo reduction; oxidation will occur in the other half-cell. The potential produced by the electrochemical cell is the difference in the reduction potentials of the two half-cell reactions. Explain that the potential value of any half-cell is obtained by measuring it against the standard hydrogen electrode. Point out that Table 21.2 contains tabulated values for some half-cells. These can be used to calculate cell potentials.

### Table 21.2

<table>
<thead>
<tr>
<th>Electrode</th>
<th>Half-reaction</th>
<th>$E^\circ$ (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li$^+$/Li</td>
<td>Li$^+$ + e$^-$ $\rightarrow$ Li</td>
<td>-3.05</td>
</tr>
<tr>
<td>K$^+$/K</td>
<td>K$^+$ + e$^-$ $\rightarrow$ K</td>
<td>-2.93</td>
</tr>
<tr>
<td>Ba$^{2+}$/Ba</td>
<td>Ba$^{2+}$ + 2e$^-$ $\rightarrow$ Ba</td>
<td>-2.90</td>
</tr>
<tr>
<td>Ca$^{2+}$/Ca</td>
<td>Ca$^{2+}$ + 2e$^-$ $\rightarrow$ Ca</td>
<td>-2.87</td>
</tr>
<tr>
<td>Na$^+$/Na</td>
<td>Na$^+$ + e$^-$ $\rightarrow$ Na</td>
<td>-2.71</td>
</tr>
<tr>
<td>Mg$^{2+}$/Mg</td>
<td>Mg$^{2+}$ + 2e$^-$ $\rightarrow$ Mg</td>
<td>-2.37</td>
</tr>
<tr>
<td>Al$^{3+}$/Al</td>
<td>Al$^{3+}$ + 3e$^-$ $\rightarrow$ Al</td>
<td>-1.66</td>
</tr>
<tr>
<td>H$^+$/H$_2$</td>
<td>2H$^+$ + 2e$^-$ $\rightarrow$ H$_2$</td>
<td>-0.83</td>
</tr>
<tr>
<td>Zn$^{2+}$/Zn</td>
<td>Zn$^{2+}$ + 2e$^-$ $\rightarrow$ Zn</td>
<td>-0.76</td>
</tr>
<tr>
<td>Cr$^{3+}$/Cr</td>
<td>Cr$^{3+}$ + 3e$^-$ $\rightarrow$ Cr</td>
<td>-0.74</td>
</tr>
<tr>
<td>Fe$^{2+}$/Fe</td>
<td>Fe$^{2+}$ + 2e$^-$ $\rightarrow$ Fe</td>
<td>-0.44</td>
</tr>
<tr>
<td>Hg$^{2+}$/Hg</td>
<td>2Hg + 2e$^-$ $\rightarrow$ 2Hg</td>
<td>-0.42</td>
</tr>
<tr>
<td>Cd$^{2+}$/Cd</td>
<td>Cd$^{2+}$ + 2e$^-$ $\rightarrow$ Cd</td>
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</tr>
<tr>
<td>PbSO$_4$/Pb</td>
<td>PbSO$_4$ + 2e$^-$ $\rightarrow$ Pb + SO$_4^{2-}$</td>
<td>-0.36</td>
</tr>
<tr>
<td>Co$^{2+}$/Co</td>
<td>Co$^{2+}$ + 2e$^-$ $\rightarrow$ Co</td>
<td>-0.28</td>
</tr>
<tr>
<td>Ni$^{2+}$/Ni</td>
<td>Ni$^{2+}$ + 2e$^-$ $\rightarrow$ Ni</td>
<td>-0.25</td>
</tr>
<tr>
<td>Sn$^{2+}$/Sn</td>
<td>Sn$^{2+}$ + 2e$^-$ $\rightarrow$ Sn</td>
<td>-0.14</td>
</tr>
<tr>
<td>Pb$^{2+}$/Pb</td>
<td>Pb$^{2+}$ + 2e$^-$ $\rightarrow$ Pb</td>
<td>-0.13</td>
</tr>
<tr>
<td>Fe$^{3+}$/Fe</td>
<td>Fe$^{3+}$ + 3e$^-$ $\rightarrow$ Fe</td>
<td>-0.036</td>
</tr>
<tr>
<td>H$^+$/H$_2$</td>
<td>2H$^+$ + 2e$^-$ $\rightarrow$ H$_2$</td>
<td>0.000</td>
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<tr>
<td>AgCl/Ag</td>
<td>AgCl + e$^-$ $\rightarrow$ Ag + Cl$^-$</td>
<td>+0.22</td>
</tr>
<tr>
<td>Hg$_2$Cl$_2$/Hg</td>
<td>Hg$_2$Cl$_2$ + 2e$^-$ $\rightarrow$ 2Hg + 2Cl$^-$</td>
<td>+0.27</td>
</tr>
<tr>
<td>Cu$^{2+}$/Cu</td>
<td>Cu$^{2+}$ + 2e$^-$ $\rightarrow$ Cu</td>
<td>+0.34</td>
</tr>
<tr>
<td>O$_2$/OH$^-$</td>
<td>O$_2$ + 2H$^+$ + 4e$^-$ $\rightarrow$ 4OH$^-$</td>
<td>+0.40</td>
</tr>
<tr>
<td>Cu$^+/Cu$</td>
<td>Cu$^+$ + e$^-$ $\rightarrow$ Cu$^0$</td>
<td>+0.82</td>
</tr>
<tr>
<td>I$_2$/I$^-$</td>
<td>I$_2$ + 2e$^-$ $\rightarrow$ 2I$^-$</td>
<td>+0.54</td>
</tr>
<tr>
<td>Fe$^{3+}$/Fe$^{2+}$</td>
<td>Fe$^{3+}$ + e$^-$ $\rightarrow$ Fe$^{2+}$</td>
<td>+0.77</td>
</tr>
<tr>
<td>Hg$_2$Cl$_2$/Hg</td>
<td>Hg$_2$Cl$_2$ + 2e$^-$ $\rightarrow$ 2Hg + 2Cl$^-$</td>
<td>+0.79</td>
</tr>
<tr>
<td>Ag$^+/Ag$</td>
<td>Ag$^+$ + e$^-$ $\rightarrow$ Ag</td>
<td>+0.80</td>
</tr>
<tr>
<td>O$_2$/H$_2$O (pH 7)</td>
<td>O$_2$ + 4H$^+$ + 4e$^-$ $\rightarrow$ 2H$_2$O</td>
<td>+0.82</td>
</tr>
<tr>
<td>Hg$^{2+}$/Hg</td>
<td>Hg$^{2+}$ + 2e$^-$ $\rightarrow$ Hg</td>
<td>+0.85</td>
</tr>
<tr>
<td>Br$_2$/Br$^-$</td>
<td>Br$_2$ + 2e$^-$ $\rightarrow$ 2Br$^-$</td>
<td>+1.07</td>
</tr>
<tr>
<td>O$_2$/H$_2$O</td>
<td>O$_2$ + 4H$^+$ + 4e$^-$ $\rightarrow$ 2H$_2$O</td>
<td>+1.23</td>
</tr>
<tr>
<td>MnO$_2$/Mn$^{2+}$</td>
<td>MnO$_2$ + 4H$^+$ + 4e$^-$ $\rightarrow$ Mn$^{2+}$ + 2H$_2$O</td>
<td>+1.28</td>
</tr>
<tr>
<td>Cr$_2$O$_7^{2-}$/Cr$^{3+}$</td>
<td>Cr$_2$O$_7^{2-}$ + 14H$^+$ + 6e$^-$ $\rightarrow$ 2Cr$^{3+}$ + 7H$_2$O</td>
<td>+1.33</td>
</tr>
<tr>
<td>Cl$_2$/Cl$^-$</td>
<td>Cl$_2$ + 2e$^-$ $\rightarrow$ 2Cl$^-$</td>
<td>+1.36</td>
</tr>
<tr>
<td>PbO$_2$/Pb$^{2+}$</td>
<td>PbO$_2$ + 4H$^+$ + 4e$^-$ $\rightarrow$ Pb$^{2+}$ + 2H$_2$O</td>
<td>+1.46</td>
</tr>
<tr>
<td>MnO$_4^-$/Mn$^{2+}$</td>
<td>MnO$_4^-$ + 8H$^+$ + 5e$^-$ $\rightarrow$ Mn$^{2+}$ + 4H$_2$O</td>
<td>+1.51</td>
</tr>
<tr>
<td>PbO$_2$/PbSO$_4$</td>
<td>PbO$_2$ + 4H$^+$ + SO$_4^{2-}$ + 2e$^-$ $\rightarrow$ PbSO$_4$ + 2H$_2$O</td>
<td>+1.69</td>
</tr>
<tr>
<td>F$_2$/F$^-$</td>
<td>F$_2$ + 2e$^-$ $\rightarrow$ 2F$^-$</td>
<td>+2.87</td>
</tr>
</tbody>
</table>

---

**Differentiated Instruction**

**Special Needs**

Some students may have trouble remembering which reactions in Table 21.2 have the greatest tendency to occur as reductions. Have them compare the table with the activity series of metals in Table 21.1. Point out that the most active metals are at the top of Table 21.1; ions of these same metals are at the top of Table 21.2. Because active metals lose electrons easily, they are most easily oxidized. By the same reasoning, their ions are least likely to be reduced.
Calculating Standard Cell Potentials

To function, a cell must be constructed of two half-cells. The half-cell reactions having the more positive (or less negative) reduction potential occurs as a reduction in the cell. With this in mind, it is possible to write cell reactions and calculate cell potentials for cells without actually assembling them. You can simply use the known standard reduction potentials for the half-reactions at the bottom of the table in which reduction will occur. This information can then be used to find the resulting cell potential. If the cell potential for a given redox reaction is positive, then the reaction is spontaneous as written. If the cell potential is negative, then the reaction is nonspontaneous. This latter reaction will be spontaneous in the reverse direction, however, and the cell potential will then have a numerically equal but positive value.

### CONCEPTUAL PROBLEM 21.1

**Determining Reaction Spontaneity**

Many hearing aid batteries contain zinc and silver oxide. Show that the following redox reaction between Zn and Ag⁺ is spontaneous.

\[
Zn(s) + 2Ag⁺(aq) \rightarrow Zn²⁺(aq) + 2Ag(s)
\]

1. **Analyze** Identify the relevant concepts. Identify the half-reactions, and calculate the standard cell potential \(E_{cell}^°\). If \(E_{cell}^°\) is positive, the reaction is spontaneous.

2. **Solve** Apply concepts to this situation. The half-reactions from the equation are:

   - Oxidation: \(Zn(s) \rightarrow Zn²⁺(aq) + 2e⁻\)
   - Reduction: \(Ag⁺(aq) + e⁻ \rightarrow Ag(s)\)

   Use Table 21.2 to look up the standard reduction potentials for the half-cells.

\[
\begin{align*}
Zn²⁺(aq) + 2e⁻ & \rightarrow Zn(s) & E_{Zn²⁺}^° & = -0.76 V \\
Ag⁺(aq) + e⁻ & \rightarrow Ag(s) & E_{Ag⁺}^° & = +0.80 V
\end{align*}
\]

Calculate the standard cell potential.

\[
E_{cell}^° = E_{Zn²⁺}^° - E_{Ag⁺}^° = -0.76 V - (+0.80 V) = -1.56 V
\]

\(E_{cell}^° < 0\), so the reaction is spontaneous.

### Practice Problems

9. Determine whether the following redox reaction will occur spontaneously.

\[3Zn²⁺(aq) + 2Cr(s) \rightarrow 3Zn(s) + 2Cr³⁺(aq)\]

10. Is this redox reaction spontaneous as written?

\[Co²⁺(aq) + Fe(s) \rightarrow Fe²⁺(aq) + Co(s)\]

### Expected Outcome

Several cells could be created in each category. Typical cells might be Na/Na⁺ with Sn²⁺/Sn, +2.57 V; Mg²⁺/Mg with Pb²⁺/Pb, +2.24 V; Ca²⁺/Ca with Al³⁺/Al, 1.21 V; PbSO₄/Pb with Ni²⁺/Ni, +0.11 V.

### Answers

9. nonspontaneous; \(E_{cell}^° = -0.02 V\)

10. yes; \(E_{cell}^° = +0.16 V\)

### Practice Problems Plus

Chapter Assessment Problems Plus 41 is related to Conceptual problem 21.1.
Discuss
As students work Sample Problems 21.1 and 21.2, they may think that when a half-cell reaction must be multiplied by a factor such as 2 or 3, the \( E^0 \) value must also be multiplied by that factor. Point out that even though two or three times as many electrons are present, the tendency for the electrons to flow is not two or three times greater. The tendency, which is measured by the \( E^0 \) value, remains the same.

Discuss
Point out that until now it has been assumed that reactions in standard half-cells always go in the spontaneous direction. For example, a Ag/Cu cell can be constructed in such a way that the only substance available for oxidation is silver metal and the only substance available for reduction is copper ion. In this case, the reaction is nonspontaneous and the cell potential is negative. Thus, the sign of the cell potential is an indicator of reaction spontaneity. Students will learn in the next section that nonspontaneous reactions can be driven by an outside source of energy.

Sample Problem 21.1

1. \( \text{Ni}(s) \rightarrow \text{Ni}^2+(aq) + 2e^- \) (at anode)
2. \( \text{Fe}_3^+(aq) + e^- \rightarrow \text{Fe}^{2+}(aq) \) (at cathode)

The half-cell with the more positive reduction potential is the one in which reduction occurs (the cathode). That means that the oxidation reaction occurs at the anode. Add the half-reactions, making certain that the number of electrons lost equals the number of electrons gained.

Practice Problems Plus

A voltaic cell is constructed using the following half-reactions:
1. \( \text{Li}^+(aq) + e^- \rightarrow \text{Li}(s) \) \( E^0_{\text{Li}^+/\text{Li}(s)} = -3.05 \text{ V} \)
2. \( \text{Mg}^{2+}(aq) + 2e^- \rightarrow \text{Mg}(s) \) \( E^0_{\text{Mg}^{2+}/\text{Mg}(s)} = -2.37 \text{ V} \)

Determine the standard cell reaction and the standard cell potential.

\[ 2\text{Li}(s) + \text{Mg}^{2+}(aq) \rightarrow \text{Mg}(s) + 2\text{Li}^+(aq); \quad 0.68 \text{ V} \]
### Sample Problem 21.2

**Calculating the Standard Cell Potential**

Calculate the standard cell potential for the voltaic cell described in Sample Problem 21.1. The half-reactions are as follows.

- \( \text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq) \) \( E_{\text{Fe}^{3+/2+}}^0 = +0.77 \text{ V} \)
- \( \text{Ni}^2+(aq) + 2e^- \rightarrow \text{Ni}(s) \) \( E_{\text{Ni}^{2+/0}}^0 = -0.25 \text{ V} \)

1. **Analyze** List the knowns and the unknown.
   - Knowns: \( E_{\text{Fe}^{3+/2+}}^0 = +0.77 \text{ V} \)
   - Knowns: \( E_{\text{Ni}^{2+/0}}^0 = -0.25 \text{ V} \)
   - Cathode: \( \text{Fe}^{3+} \) half-cell

2. **Calculate** Solve for the unknown.

   \[
   E_{\text{cell}} = E_{\text{red}} - E_{\text{oxi}} = E_{\text{Fe}^{3+/2+}}^0 - E_{\text{Ni}^{2+/0}}^0
   \]

   \[
   = +0.77 \text{ V} - (-0.25 \text{ V})
   \]

   \[
   = 1.02 \text{ V}
   \]

3. **Evaluate** Does the result make sense?

   If the reduction potential of the reduction is positive, and the reduction potential of the oxidation is negative, \( E_{\text{cell}} \) must be positive.

### Practice Problems

13. A voltaic cell is constructed using the following half-reactions.
   - \( \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) \) \( E_{\text{Cu}^{2+/0}}^0 = +0.34 \text{ V} \)
   - \( \text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s) \) \( E_{\text{Ag}^{+/0}}^0 = +0.80 \text{ V} \)

14. A voltaic cell is constructed using the following half-reactions.
   - \( \text{Al}^{3+}(aq) + 3e^- \rightarrow \text{Al}(s) \) \( E_{\text{Al}^{3+/0}}^0 = -1.66 \text{ V} \)
   - \( \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) \) \( E_{\text{Cu}^{2+/0}}^0 = +0.34 \text{ V} \)

Calculate the standard cell potential.

### 21.2 Section Assessment

15. **Key Concept** What causes the electrical potential of a cell?
16. **Key Concept** What is the electrical potential of a standard hydrogen electrode?
17. **Key Concept** How can you find the standard reduction potential of a half-cell?
18. **Key Concept** What cell potential values indicate a spontaneous reaction? A nonspontaneous reaction?
19. The standard reduction potential for a cadmium half-cell is \( -0.40 \text{ V} \). What does this mean?

### Practice Problems Plus

**Answers**

13. \( E_{\text{cell}}^0 = +2.00 \text{ V} \)
14. \( E_{\text{cell}}^0 = +0.46 \text{ V} \)

**Evaluate Understanding**

Ask students to use Table 21.2 to create an electrochemical cell that will operate spontaneously. Have them calculate the cell potential and write the equations for the two half-reactions and the overall cell reaction. Finally, have them use the shorthand method to represent the cell.

**Reteach**

Write a nonspontaneous redox reaction on the board, for example:

\( [\text{Cd}^{2+}(aq) + e^- \rightarrow \text{Cd}(s)] \) Ask, What element is being oxidized? (cadmium) Have students write the half-reaction with student input. \( [\text{Ag}(s) \rightarrow \text{Ag}^+(aq) + e^-] \) Guide students through the process of calculating the cell potential \( (-1.20 \text{ V}) \) and ask whether the cell operates spontaneously. (No, it has a negative potential.)

**Connecting Concepts**

**Batteries** Reread the discussion on batteries in Section 21.1. Use Figure 21.6 and Table 21.2 to calculate the standard cell potential of a voltaic cell in a lead–acid battery.

**Connecting Concepts**

**Interactive Textbook** Test yourself on the concepts in Section 21.2.

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### Section 21.2 Assessment

15. Electrical potential of cell results from competition for electrons between two half-cells.
16. The standard reduction potential of the hydrogen electrode has been assigned a value of 0.00 V at 25°C.
17. by connecting it to a standard hydrogen electrode and measuring the cell potential
18. If the cell potential for a redox reaction is positive, the reaction is spontaneous; if it is negative, the reaction is nonspontaneous.
19. The cadmium half-cell has a tendency to undergo oxidation.

**Connecting Concepts**

**Interactive Textbook** Test yourself on the concepts in Section 21.2.

- \( \text{Pb} + \text{PbO}_2 + 2\text{H}_2\text{SO}_4^{2-} \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O}; E_{\text{cell}}^0 = +2.05 \text{ V} \)

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If your class subscribes to the Interactive Textbook, use it to review key concepts in Section 21.2.