

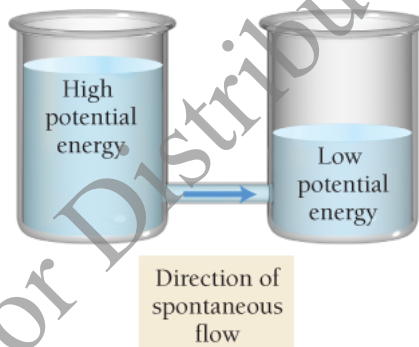
19.4: Standard Electrode Potentials

Key Concept Video Standard Electrode Potentials

As we have just seen, the standard cell potential (E_{cell}°) for an electrochemical cell depends on the specific half-reactions occurring in the half-cells and is a measure of the potential energy difference (per unit charge) between the two electrodes. We can think of the electrode in each half-cell as having its own individual potential, called the **standard electrode potential** ϕ . The overall standard cell potential (E_{cell}°) is the difference between the two standard electrode potentials.

We can better understand this idea with an analogy. Consider two water tanks with different water levels connected by a common pipe, as shown in Figure 19.5. The water in each tank has its own level and corresponding potential energy. When the tanks are connected, water flows from the tank with the higher water level (higher potential energy) to the tank with a lower water level (lower potential energy). Similarly, each half-cell in an electrochemical cell has its own charge and corresponding electrode potential. *When the cells are connected, electrons flow from the electrode with greater potential energy (more negatively charged) to the electrode with less potential energy (more positively charged).*

Figure 19.5 An Analogy for Electrode Potential



One limitation to this analogy is that, unlike the water level in a tank, we cannot measure the electrode potential in a half-cell directly—we can only measure the overall potential that occurs when two half-cells are combined in a whole cell. However, we can arbitrarily assign a potential of zero to the electrode in a *particular* type of half-cell and then measure all other electrode potentials relative to that zero.

The half-cell electrode that is normally chosen to have a potential of zero is the **standard hydrogen electrode (SHE)** ϕ . This half-cell consists of an inert platinum electrode immersed in 1 M HCl with hydrogen gas at 1 atm bubbling through the solution, as shown in Figure 19.6. When the standard hydrogen electrode acts as the cathode, the following half-reaction occurs:

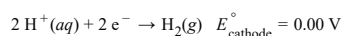
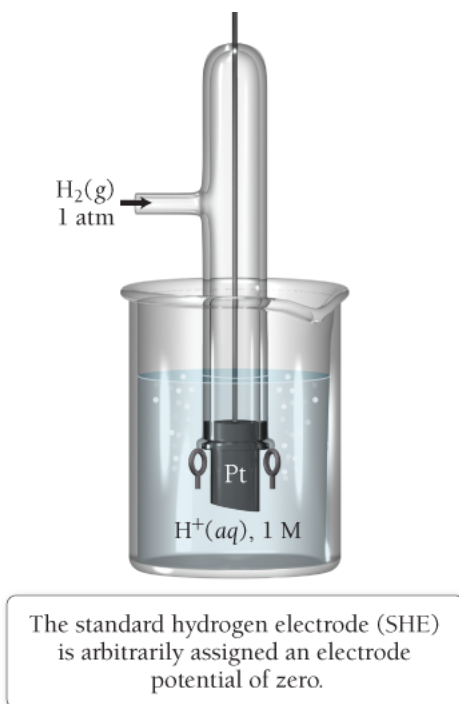


Figure 19.6 The Standard Hydrogen Electrode

All other electrode potentials are then measured relative to the SHE.

Standard Hydrogen Electrode (SHE)



If we connect the standard hydrogen electrode to an electrode in another half-cell of interest, we can measure the potential difference (or voltage) between the two electrodes. Since we assigned the standard hydrogen electrode zero voltage, we can determine the electrode potential of the other half-cell.

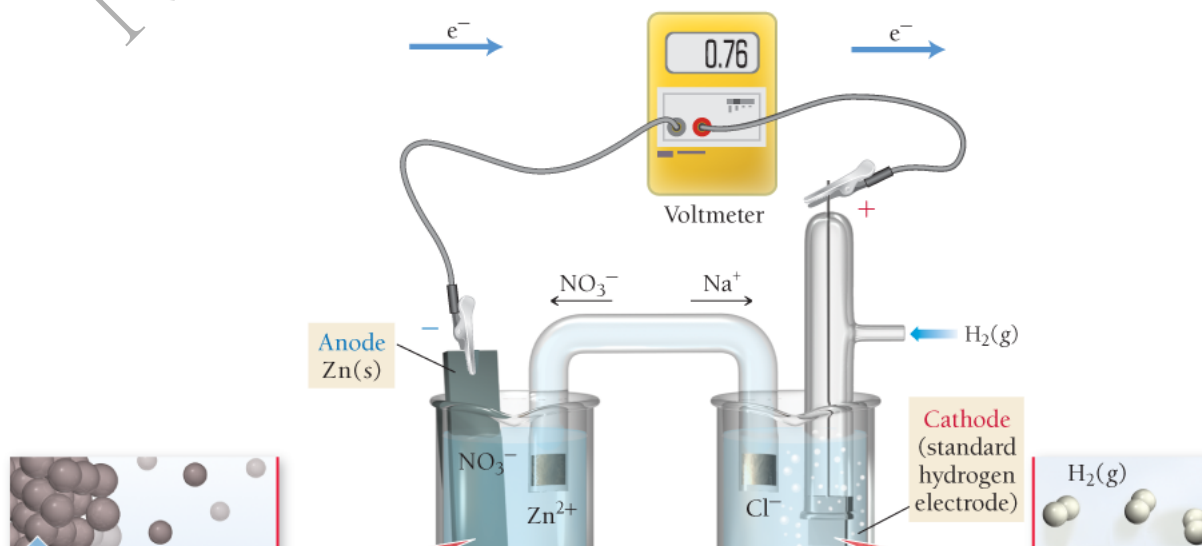
Consider the electrochemical cell shown in Figure 19.7. In this electrochemical cell, Zn is oxidized to Zn^{2+} and H^+ is reduced to H_2 under standard conditions (all solutions are 1 M in concentration and all gases are 1 atm in pressure) and at 25 °C. Electrons travel from the anode (where oxidation occurs) to the cathode (where reduction occurs), so we define E_{cell}° as the difference in voltage between the cathode (final state) and the anode (initial state).

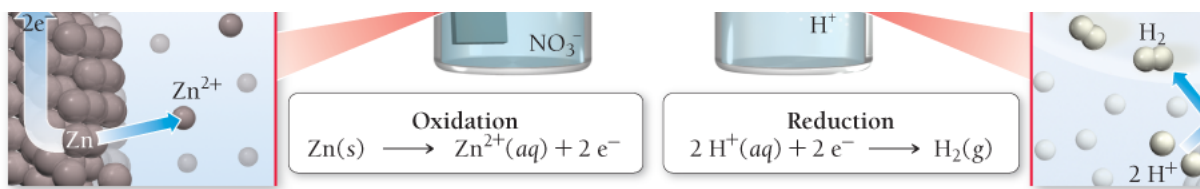
$$\begin{aligned} E_{\text{cell}}^\circ &= E_{\text{final}}^\circ - E_{\text{initial}}^\circ \\ &= E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ \end{aligned}$$

Figure 19.7 Measuring Electrode Potential

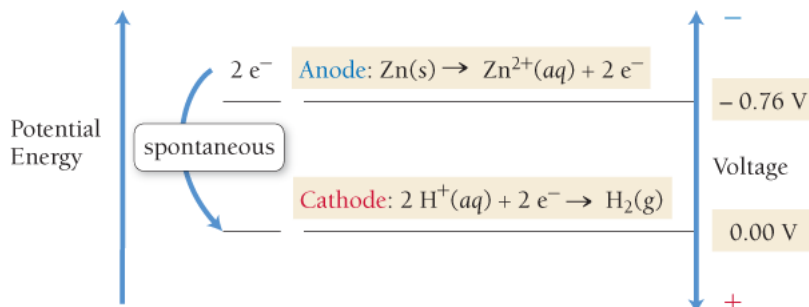
Because the electrode potential of the SHE is zero, the electrode potential for the oxidation of Zn is equal to the cell potential.

Measuring Half-Cell Potential with the SHE





The measured cell potential for this cell is +0.76 V. The anode (in this case, Zn/Zn^{2+}) is at a more negative voltage (higher potential energy) than the cathode (in this case, the SHE). Therefore, electrons spontaneously flow from the anode to the cathode. We can diagram the potential energy and the voltage as follows:



Referring back to our water tank analogy, the zinc half-cell is like the water tank with the higher water level. Electrons therefore flow from the zinc electrode to the standard hydrogen electrode.

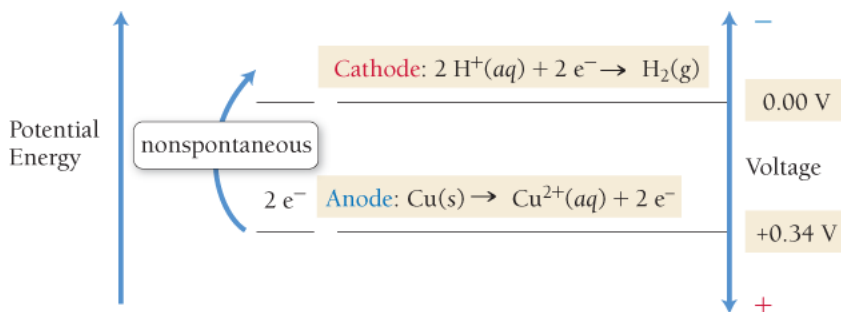
Because we assigned the SHE a potential of zero (0.00 V), we can determine the electrode potential for the Zn/Zn^{2+} half-cell (the anode) from the measured cell potential (E_{cell}°).

$$\begin{aligned}
 E_{\text{cell}}^{\circ} &= E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\
 0.76\text{ V} &= 0.00\text{ V} - E_{\text{anode}}^{\circ} \\
 E_{\text{anode}}^{\circ} &= -0.76\text{ V}
 \end{aligned}$$

The potential for the Zn/Zn^{2+} electrode is *negative*. The negative potential indicates that an electron at the Zn/Zn^{2+} electrode has greater potential energy than it does at the SHE. *Remember that the more negative the electrode potential is, the greater the potential energy of an electron at that electrode.*

What would happen if we connected an electrode in which the electron has *more positive* potential (than the standard hydrogen electrode) to the standard hydrogen electrode? That electrode would then have a more *positive* voltage (lower potential energy for an electron).

For example, suppose we connect the standard hydrogen electrode to a Cu electrode immersed in a 1 M Cu^{2+} solution. The measured cell potential for this cell is -0.34 V. The anode (Cu/Cu^{2+}) is at a more positive voltage (lower potential energy) than the cathode (the SHE). Therefore, electrons do *not* spontaneously flow from the anode to the cathode. We diagram the potential energy and the voltage of this cell as follows:



The copper half-cell is like the water tank with the lower water level, and electrons *do not* spontaneously flow

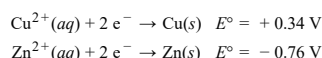
from the copper electrode to the standard hydrogen electrode.

We can determine the electrode potential for the Cu/Cu^{2+} half-cell (the anode) from the measured cell potential as we did for the Zn/Zn^{2+} half-cell.

$$\begin{aligned} E_{\text{cell}}^{\circ} &= E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ} \\ -0.34 \text{ V} &= 0.00 \text{ V} - E_{\text{anode}}^{\circ} \\ E_{\text{anode}}^{\circ} &= +0.34 \text{ V} \end{aligned}$$


The potential for the Cu/Cu^{2+} electrode is *positive*. The positive potential indicates that an electron at the Cu/Cu^{2+} electrode has *lower* potential energy than it does at the SHE. The more positive the electrode potential, the lower the potential energy of an electron at that electrode.

By convention, standard electrode potentials are written for *reduction* half-reactions. We write the standard electrode potentials for the two half-reactions just discussed as:



We can see that the Cu/Cu^{2+} electrode is positive relative to the SHE (and will therefore tend to draw electrons *away* from the SHE) and that the Zn/Zn^{2+} electrode is negative relative to the SHE (and will therefore tend to repel electrons toward the SHE). The standard electrode potentials for a number of common half-reactions are listed in [Table 19.1](#).

Table 19.1 Standard Electrode Potentials at 25 °C

| Reduction Half-Reaction | | |
|---|---|--|
| <p>Stronger oxidizing agent</p>  | $\text{F}_2(\text{g}) + 2 \text{e}^{-}$ | $\rightarrow 2 \text{F}^{-}(\text{aq})$ |
| | $\text{H}_2\text{O}_2(\text{aq}) + 2 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{PbO}_2(\text{s}) + 4 \text{H}^{+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow \text{PbSO}_4(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{MnO}_4^{-}(\text{aq}) + 4 \text{H}^{+}(\text{aq}) + 3 \text{e}^{-}$ | $\rightarrow \text{MnO}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{MnO}_4^{-}(\text{aq}) + 8 \text{H}^{+}(\text{aq}) + 5 \text{e}^{-}$ | $\rightarrow \text{Mn}^{2+}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$ |
| | $\text{Au}^{3+}(\text{aq}) + 3 \text{e}^{-}$ | $\rightarrow \text{Au}(\text{s})$ |
| | $\text{PbO}_2(\text{s}) + 4 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow \text{Pb}^{2+}(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{Cl}_2(\text{g}) + 2 \text{e}^{-}$ | $\rightarrow 2 \text{Cl}^{-}(\text{aq})$ |
| | $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^{+}(\text{aq}) + 6 \text{e}^{-}$ | $\rightarrow 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\text{l})$ |
| | $\text{O}_2(\text{g}) + 4 \text{H}^{+}(\text{aq}) + 4 \text{e}^{-}$ | $\rightarrow 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{MnO}_2(\text{s}) + 4 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow \text{Mn}^{2+}(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{IO}_3^{-}(\text{aq}) + 6 \text{H}^{+}(\text{aq}) + 5 \text{e}^{-}$ | $\rightarrow \frac{1}{2} \text{I}_2(\text{aq}) + 3 \text{H}_2\text{O}(\text{l})$ |
| | $\text{Br}_2(\text{l}) + 2 \text{e}^{-}$ | $\rightarrow 2 \text{Br}^{-}(\text{aq})$ |
| | $\text{VO}_2^{+}(\text{aq}) + 2 \text{H}^{+}(\text{aq}) + \text{e}^{-}$ | $\rightarrow \text{VO}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ |
| | $\text{NO}_3^{-}(\text{aq}) + 4 \text{H}^{+}(\text{aq}) + 3 \text{e}^{-}$ | $\rightarrow \text{NO}(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$ |
| | $\text{ClO}_2(\text{g}) + \text{e}^{-}$ | $\rightarrow \text{ClO}_2^{-}(\text{aq})$ |
| | $\text{Ag}^{+}(\text{aq}) + \text{e}^{-}$ | $\rightarrow \text{Ag}(\text{s})$ |
| | $\text{Fe}^{3+}(\text{aq}) + \text{e}^{-}$ | $\rightarrow \text{Fe}^{2+}(\text{aq})$ |
| | $\text{O}_2(\text{g}) + 2 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow \text{H}_2\text{O}_2(\text{aq})$ |
| | $\text{MnO}_4^{-}(\text{aq}) + \text{e}^{-}$ | $\rightarrow \text{MnO}_4^{2-}(\text{aq})$ |
| | $\text{I}_2(\text{s}) + 2 \text{e}^{-}$ | $\rightarrow 2 \text{I}^{-}(\text{aq})$ |
| | $\text{Cu}^{+}(\text{aq}) + \text{e}^{-}$ | $\rightarrow \text{Cu}(\text{s})$ |
| | $\text{O}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l}) + 4 \text{e}^{-}$ | $\rightarrow 4 \text{OH}^{-}(\text{aq})$ |
| | $\text{Cu}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow \text{Cu}(\text{s})$ |
| | $\text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-}$ | $\rightarrow \text{H}_2\text{SO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$ |
| | $\text{Cu}^{2+}(\text{aq}) + \text{e}^{-}$ | $\rightarrow \text{Cu}^{+}(\text{aq})$ |



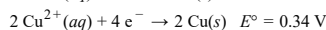
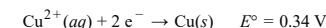
 Weaker
 oxidizing agent

| | |
|---|---|
| $\text{Sn}^{4+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Sn}^{2+}(\text{aq})$ |
| $2 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{H}_2(\text{g})$ |
| $\text{Fe}^{3+}(\text{aq}) + 3 \text{e}^{-}$ | $\longrightarrow \text{Fe}(\text{s})$ |
| $\text{Pb}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Pb}(\text{s})$ |
| $\text{Sn}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Sn}(\text{s})$ |
| $\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Ni}(\text{s})$ |
| $\text{Cd}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Cd}(\text{s})$ |
| $\text{Fe}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Fe}(\text{s})$ |
| $\text{Cr}^{3+}(\text{aq}) + \text{e}^{-}$ | $\longrightarrow \text{Cr}^{2+}(\text{aq})$ |
| $\text{Cr}^{3+}(\text{aq}) + 3 \text{e}^{-}$ | $\longrightarrow \text{Cr}(\text{s})$ |
| $\text{Zn}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Zn}(\text{s})$ |
| $2 \text{H}_2\text{O}(\text{l}) + 2 \text{e}^{-}$ | $\longrightarrow \text{H}_2(\text{g}) + 2 \text{OH}^{-}(\text{aq})$ |
| $\text{Mn}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Mn}(\text{s})$ |
| $\text{Al}^{3+}(\text{aq}) + 3 \text{e}^{-}$ | $\longrightarrow \text{Al}(\text{s})$ |
| $\text{Mg}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Mg}(\text{s})$ |
| $\text{Na}^{+}(\text{aq}) + \text{e}^{-}$ | $\longrightarrow \text{Na}(\text{s})$ |
| $\text{Ca}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Ca}(\text{s})$ |
| $\text{Ba}^{2+}(\text{aq}) + 2 \text{e}^{-}$ | $\longrightarrow \text{Ba}(\text{s})$ |
| $\text{K}^{+}(\text{aq}) + \text{e}^{-}$ | $\longrightarrow \text{K}(\text{s})$ |
| $\text{Li}^{+}(\text{aq}) + \text{e}^{-}$ | $\longrightarrow \text{Li}(\text{s})$ |

Summarizing Standard Electrode Potentials:

- The electrode potential of the standard hydrogen electrode (SHE) is exactly zero.
- The electrode in any half-cell with a greater tendency to undergo reduction is positively charged relative to the SHE and therefore has a positive E° .
- The electrode in any half-cell with a lesser tendency to undergo reduction (or greater tendency to undergo oxidation) is negatively charged relative to the SHE and therefore has a negative E° .
- The cell potential of any electrochemical cell (E_{cell}°) is the difference between the electrode potentials of the cathode and the anode ($E_{\text{cell}}^{\circ} = E_{\text{cat}}^{\circ} - E_{\text{an}}^{\circ}$).
- E_{cell}° is positive for spontaneous reactions and negative for nonspontaneous reactions.

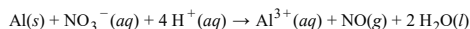
Multiplying a half-reaction by some constant does *not* affect the value of E° for the half-reaction.



Example 19.4 illustrates how to calculate the standard potential of an electrochemical cell from the standard electrode potentials of the half-reactions.

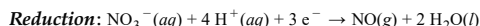
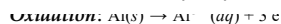
Example 19.4 Calculating Standard Potentials for Electrochemical Cells from Standard Electrode Potentials of the Half-Reactions

Use tabulated standard electrode potentials to calculate the standard cell potential for this reaction occurring in an electrochemical cell at 25 °C (The equation is balanced.)



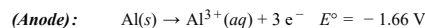
SOLUTION Begin by separating the reaction into oxidation and reduction half-reactions. (In this case, you can readily see that Al(s) is oxidized. In cases where it is not so apparent, you may want to assign oxidation states to determine the correct half-reactions.)

Oxidation: $\text{Al}(\text{s}) \rightarrow \text{Al}^{3+}(\text{aq}) + 3 \text{e}^{-}$

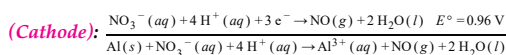


Look up the standard electrode potentials for each half-reaction in [Table 19.1](#). Add the half-cell reactions together to obtain the overall redox equation. Calculate the standard cell potential by subtracting the electrode potential of the anode from the electrode potential of the cathode.

Oxidation

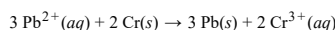


Reduction



$$\begin{aligned} E_{\text{cell}}^{\circ} &= E_{\text{cat}}^{\circ} - E_{\text{an}}^{\circ} \\ &= 0.96 \text{ V} - (-1.66 \text{ V}) \\ &= 2.62 \text{ V} \end{aligned}$$

FOR PRACTICE 19.4 Use tabulated standard electrode potentials to calculate the standard cell potential for this reaction occurring in an electrochemical cell at 25 °C. (The equation is balanced.)



Interactive Worked Example 19.4 Calculating Standard Potentials for Electrochemical Cells from Standard Electrode Potentials of the Half-Reactions

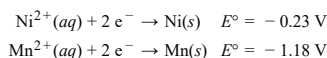
Conceptual Connection 19.2 Standard Electrode Potentials

Predicting the Spontaneous Direction of an Oxidation–Reduction Reaction

The following mnemonics (NIO and PIR) can help you predict the spontaneous direction of redox reactions: N.I.O.–More Negative Is Oxidation P.I.R.–More Positive Is Reduction

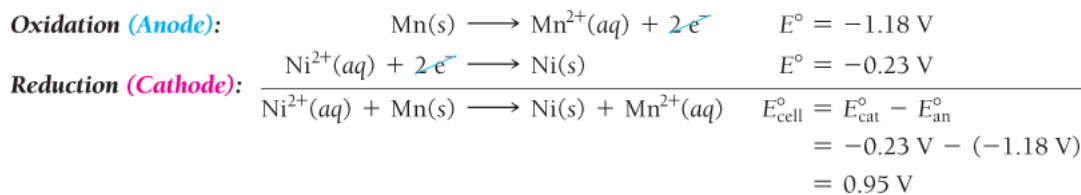
To determine the spontaneous direction of an oxidation–reduction reaction, we examine the electrode potentials of the two relevant half-reactions in [Table 19.1](#). The half-reaction with the more *negative* electrode potential tends to lose electrons and therefore undergo oxidation. (Remember that negative charge repels electrons.) The half-reaction having the more *positive* electrode potential tends to gain electrons and therefore undergo reduction. (Remember that positive charge attracts electrons.)

Consider the two reduction half-reactions:



Because the manganese half-reaction has a more negative electrode potential, it repels electrons and proceeds in the reverse direction (oxidation). Because the nickel half-reaction has the more positive (or less negative) electrode potential, it attracts electrons and proceeds in the forward direction.

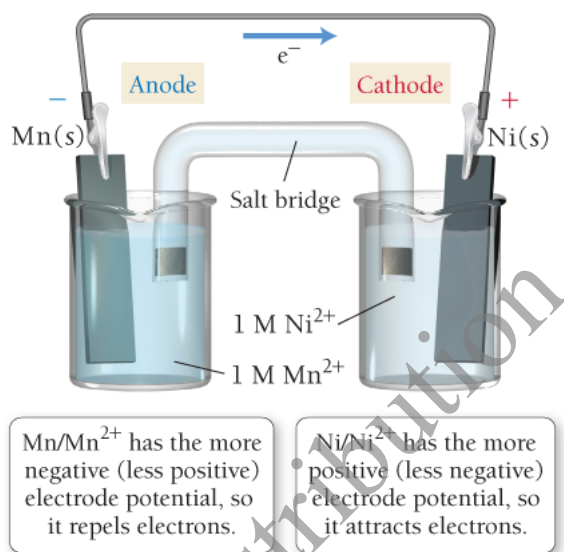
We can confirm this by calculating the standard electrode potential for manganese acting as the anode (oxidation) and nickel acting as the cathode (reduction).



The overall cell potential is positive, indicating a spontaneous reaction. Consider the electrochemical cell corresponding to this spontaneous redox reaction in [Figure 19.8](#). We draw the manganese half-cell on the left as the anode and the nickel half-cell on the right as the cathode. Electrons flow from the anode to the cathode.

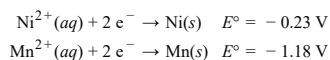
Figure 19.8 Mn/Ni²⁺ Electrochemical Cell

Since the reduction of Mn^{2+} is listed below the reduction of Ni^{2+} in Table 19.1, the reduction of Ni^{2+} is spontaneous when paired with the oxidation of Mn.



Another way to predict the spontaneity of a redox reaction is to note the relative positions of the two half-reactions in [Table 19.1](#). The table lists half-reactions in order of *decreasing* electrode potential, so the half-reactions near the top of the table—those having large *positive* electrode potentials—attract electrons and therefore tend to occur in the forward direction. Half-reactions near the bottom of the table—those having large *negative* electrode potentials—repel electrons and therefore tend to occur in the reverse direction. In other words, as we move down [Table 19.1](#), the half-reactions become less likely to occur in the forward direction and more likely to occur in the reverse direction. As a result, *any reduction half-reaction listed is spontaneous when paired with the reverse of a half-reaction that appears below it in Table 19.1*.

For example, if we return to our two previous half-reactions involving manganese and nickel, we see that the manganese half-reaction is listed below the nickel half-reaction in [Table 19.1](#).



Therefore, the nickel reaction occurs in the forward direction (reduction) and the manganese reaction occurs in the reverse direction (oxidation).

Recall from [Section 8.9](#) that an *oxidizing agent* causes the *oxidation* of another substance (and is itself reduced) and that a *reducing agent* causes the *reduction* of another substance (and is itself oxidized).

Summarizing the Prediction of Spontaneous Direction for Redox Reactions:

- The half-reaction with the more *positive* electrode potential attracts electrons more strongly and undergoes reduction. (Substances listed at the top of [Table 19.1](#) tend to undergo reduction; they are good oxidizing agents.)

reduction. (Substances listed at the top of Table 19.1 tend to undergo reduction; they are good oxidizing agents.)

- The half-reaction with the more *negative* electrode potential repels electrons more strongly and undergoes oxidation. (Substances listed near the bottom of Table 19.1 tend to undergo oxidation; they are good reducing agents.)
- Any reduction reaction in Table 19.1 is spontaneous when paired with the *reverse* of any of the reactions listed below it on the table.

Example 19.5 Predicting Spontaneous Redox Reactions and Sketching Electrochemical Cells

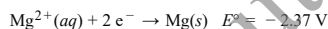
Without calculating E°_{cell} , predict whether each of these redox reactions is spontaneous (when the reactants and products are in their standard states). If the reaction is spontaneous as written, make a sketch of the electrochemical cell in which the reaction could occur. If the reaction is not spontaneous as written, write an equation for the direction in which the spontaneous reaction occurs and sketch the corresponding electrochemical cell. In your sketches, make sure to label the anode (which should be drawn on the left), the cathode, and the direction of electron flow.

- a. $\text{Fe(s)} + \text{Mg}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Mg(s)}$
 b. $\text{Fe(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Pb(s)}$

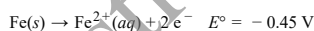
SOLUTION

- a. $\text{Fe(s)} + \text{Mg}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Mg(s)}$

This reaction involves the reduction of Mg^{2+} :

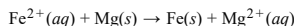


and the oxidation of Fe:



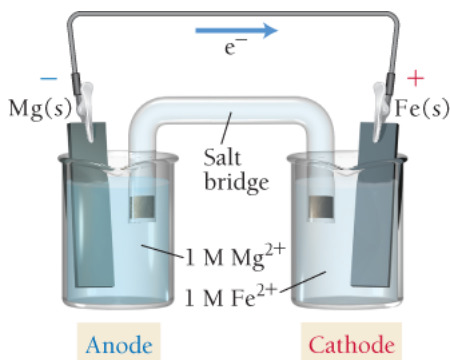
The magnesium half-reaction has the more negative electrode potential and therefore repels electrons more strongly and undergoes oxidation. The iron half-reaction has the more positive electrode potential and therefore attracts electrons more strongly and undergoes reduction. So the reaction as written is *not* spontaneous. (The reaction pairs the reduction of Mg^{2+} with the reverse of a half-reaction *above* it in Table 19.1—such pairings are not spontaneous.)

However, the reverse reaction is spontaneous:



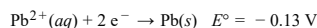
The corresponding electrochemical cell is shown in Figure 19.9.

Figure 19.9 Mg/Fe^{2+} Electrochemical Cell

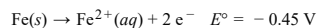


- b. $\text{Fe(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Pb(s)}$

This reaction involves the reduction of Pb^{2+} :

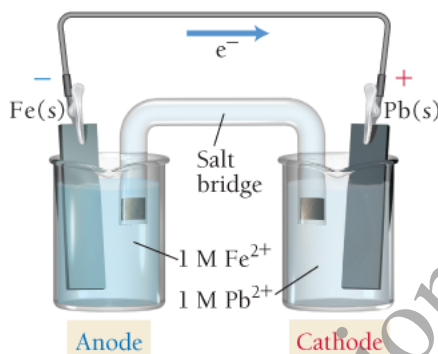


and the oxidation of iron:



The iron half-reaction has the more negative electrode potential and therefore repels electrons and undergoes oxidation. The lead half-reaction has the more positive electrode potential and therefore attracts electrons and undergoes reduction. The reaction is spontaneous as written. (The reaction pairs the reduction of Pb^{2+} with the reverse of a half-reaction *below* it in Table 19.1—such pairings are always spontaneous.) The corresponding electrochemical cell is shown in Figure 19.10.

Figure 19.10 Fe/Pb^{2+} Electrochemical Cell



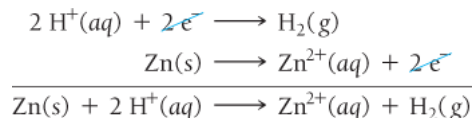
FOR PRACTICE 19.5 Are these redox reactions spontaneous under standard conditions?

- $\text{Zn}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Ni}(\text{s})$
- $\text{Zn}(\text{s}) + \text{Ca}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Ca}(\text{s})$

Conceptual Connection 19.3 Selective Oxidation


Predicting Whether a Metal Will Dissolve in Acid

Recall from Chapter 16 that acids dissolve some metals. Most acids dissolve metals by the reduction of H^{+} ions to hydrogen gas and the corresponding oxidation of the metal to its ion. For example, if solid Zn is submersed into hydrochloric acid, the following reaction occurs:

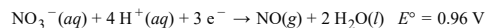



We observe the reaction as the dissolving of the zinc and the bubbling of hydrogen gas. The zinc is oxidized and the H^{+} ions are reduced. Notice that this reaction involves the pairing of a reduction half-reaction (the reduction of H^{+}) with the reverse of a half-reaction that is listed below it in Table 19.1. Therefore, this reaction is spontaneous.

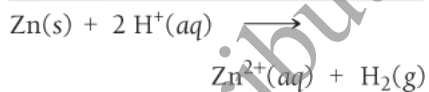
What happens, in contrast, if we pair the reduction of H^{+} with the oxidation of Cu? The reaction is not spontaneous because it involves pairing the reduction of H^{+} with the reverse of a half-reaction that is listed *above* it in the table. Consequently, copper does not react with H^{+} and does not dissolve in acids such as HCl. In

general, metals with reduction half-reactions listed below the reduction of H^+ to H_2 in Table 19.1  dissolve in acids, while metals listed above it do not.

An important exception to this rule is nitric acid (HNO_3), which can oxidize metals through the reduction half-reaction:



Since this half-reaction is above the reduction of H^+ in Table 19.1 , HNO_3 can oxidize metals (such as copper) that can't be oxidized by HCl.



When zinc is immersed in hydrochloric acid, the zinc is oxidized, forming ions that become solvated in the solution. Hydrogen ions are reduced, forming bubbles of hydrogen gas.

Conceptual Connection 19.4 Metals Dissolving in Acids

Interactive

Not for Distribution

Not for Distribution

Not for Distribution

Not for Distribution