Changes in reaction conditions can have a tremendous effect on the course of a redox reaction. For example, under standard conditions, the reaction of Co(s) with Ni^{2+}(aq) to form Ni(s) and Co^{2+}(aq) occurs spontaneously, but if we reduce the concentration of Ni^{2+} by a factor of 100, so that [Ni^{2+}] is 0.01 M, then the reverse reaction occurs spontaneously instead. The relationship between voltage and concentration is one of the factors that must be understood to predict whether a reaction will be spontaneous.

The Relationship between Cell Potential and Free Energy

Electrochemical cells convert chemical energy to electrical energy and vice versa. The total amount of energy produced by an electrochemical cell, and thus the amount of energy available to do electrical work, depends on both the cell potential and the total number of electrons that are transferred from the reductant to the oxidant during the course of a reaction. The resulting electric current is measured in coulombs (C). The SI unit of measure for the number of electrons that pass a given point in 1 second; it is defined as $6.25 \times 10^{18}$ $e^-$/s and relates electron potential (in volts) to energy (in joules): 1 J/1 V = 1 C., an SI unit that measures the number of electrons passing a given point in 1 s. A coulomb relates energy (in joules) to electrical potential (in volts). Electric current is measured in amperes (A); the fundamental SI unit of electric current; it is defined as the flow of 1 C/s past a given point: 1 A = 1 C/s.; 1 A is defined as the flow of 1 C/s past a given point (1 C = 1 A·s):

$\left( \dfrac{1 \text{ J}}{1 \text{ V}} = 1 \text{ C} = 1 \text{ A} \cdot \text{s} \right)$

In chemical reactions, however, we need to relate the coulomb to the charge on a mole of electrons. Multiplying the
charge on the electron by Avogadro’s number gives us the charge on 1 mol of electrons, which is called the faraday (F). The charge on 1 mol of electrons; it is obtained by multiplying the charge on the electron by Avogadro’s number, named after the English physicist and chemist Michael Faraday (1791–1867):

\[
F = \left( 1.60218 \times 10^{-19} \text{C} \right) \frac{6.02214 \times 10^{23}}{1 \text{ mole}} \tag{17.4.2}
\]

\[
F = 9.64855 \times 10^{4} \text{ C/mole} \simeq 96,486 \text{ C/(mol e^{-})}
\]

The total charge transferred from the reductant to the oxidant is therefore \( nF \), where \( n \) is the number of moles of electrons.

**Michael Faraday (1791–1867)**

Faraday was a British physicist and chemist who was arguably one of the greatest experimental scientists in history. The son of a blacksmith, Faraday was self-educated and became an apprentice bookbinder at age 14 before turning to science. His experiments in electricity and magnetism made electricity a routine tool in science and led to both the electric motor and the electric generator. He discovered the phenomenon of electrolysis and laid the foundations of electrochemistry. In fact, most of the specialized terms introduced in this chapter (electrode, anode, cathode, and so forth) are due to Faraday. In addition, he discovered benzene and invented the system of oxidation state numbers that we use today. Faraday is probably best known for “The Chemical History of a Candle,” a series of public lectures on the chemistry and physics of flames.

The maximum amount of work that can be produced by an electrochemical cell (\( w_{\text{max}} \)) is equal to the product of the cell potential (\( E_{\text{cell}} \)) and the total charge transferred during the reaction (\( nF \)):

\[
w_{\text{(max)}} = nFE_{\text{cell}} \tag{17.4.3}
\]

Work is expressed as a negative number because work is being done by a system (an electrochemical cell with a positive potential) on its surroundings.

As you learned in Chapter 16, the change in free energy (\( \Delta G \)) is also a measure of the maximum amount of work that can be performed during a chemical process (\( \Delta G = w_{\text{max}} \)). Consequently, there must be a relationship between the potential of an electrochemical cell and \( \Delta G \), the most important thermodynamic quantity discussed in Chapter 16. This relationship is as follows:

\[
\Delta G = -nFE_{\text{cell}} \tag{17.4.4}
\]

A spontaneous redox reaction is therefore characterized by a *negative* value of \( \Delta G \) and a *positive* value of \( E_{\text{cell}} \), consistent with our earlier discussions. When both reactants and products are in their standard states, the relationship between \( \Delta G^{\circ} \) and \( E^{\circ}_{\text{cell}} \) is as follows:

\[
\Delta G^{\circ}(\circ) = -nFE^{\circ}_{\text{cell}} \tag{17.4.5}
\]

2
Note the Pattern

A spontaneous redox reaction is characterized by a negative value of $\Delta G^\circ$, which corresponds to a positive value of $E^\circ_{\text{cell}}$.

Example 17.4.1

Suppose you want to prepare elemental bromine from bromide using the dichromate ion as an oxidant. Using the data in Table 17.3.1, calculate the free-energy change ($\Delta G^\circ$) for this redox reaction under standard conditions. Is the reaction spontaneous?

Given: redox reaction

Asked for: $\Delta G^\circ$ for the reaction and spontaneity

Strategy:

A From the relevant half-reactions and the corresponding values of $E^\circ$, write the overall reaction and calculate $E^\circ_{\text{cell}}$ using Equation 17.2.2.

B Determine the number of electrons transferred in the overall reaction. Then use Equation 17.4.5 to calculate $\Delta G^\circ$. If $\Delta G^\circ$ is negative, then the reaction is spontaneous.

Solution:

A As always, the first step is to write the relevant half-reactions and use them to obtain the overall reaction and the magnitude of $E^\circ$. From the Standard Reduction Potentials Table, we can find the reduction and oxidation half-reactions and corresponding $E^\circ$ values:

\[
\begin{align*}
\text{cathode:} & \quad \text{Cr}_2\text{O}_7^{2-} (aq) + 14 H^+(aq) + 6e^- \rightarrow 2\text{Cr}^{3+}(aq) + 7H_2O(l) \quad E_{\text{cathode}}^{\circ} = 1.23\;V \\
\text{anode:} & \quad 2\text{Br}^-(aq) \rightarrow \text{Br}_2(aq) + 2e^- \quad E_{\text{anode}}^{\circ} = 1.09\;V
\end{align*}
\]

To obtain the overall balanced chemical equation, we must multiply both sides of the oxidation half-reaction by 3 to obtain the same number of electrons as in the reduction half-reaction, remembering that the magnitude of $E^\circ$ is not affected:

\[
\begin{align*}
\text{cathode:} & \quad \text{Cr}_2\text{O}_7^{2-} (aq) + 14 H^+(aq) + 6e^- \rightarrow 2\text{Cr}^{3+}(aq) + 7H_2O(l) \quad E_{\text{cathode}}^{\circ} = 1.23\;V \\
\text{anode:} & \quad 6\text{Br}^-(aq) \rightarrow 3\text{Br}_2(aq) + 6e^- \quad E_{\text{anode}}^{\circ} = 1.09\;V \\
\text{overall:} & \quad \text{Cr}_2\text{O}_7^{2-} (aq) + 14 H^+(aq) + 6\text{Br}^-(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 3\text{Br}_2(aq) + 7H_2O(l) \quad E_{\text{cell}}^{\circ} = 0.14\;V
\end{align*}
\]
We can now calculate $\Delta G^\circ$ using Equation 17.4.5 Because six electrons are transferred in the overall reaction, the value of $n$ is 6:

$$\Delta G^{o} = -nFE_{cell}^{o} = \left ( 6 \; \text{mol} \right ) \left ( 96,468 \; \text{J/} \; \text{(V} \cdot \text{mol}) \right ) \left ( 0.14 \; \text{V} \right )$$
$$= -8.1 \times 10^{4} \; \text{J}$$
$$= -81 \; \text{kJ/mol Cr}_2\text{O}_7$$

Thus $\Delta G^\circ$ is $-81 \; \text{kJ}$ for the reaction as written, and the reaction is spontaneous.

Exercise

Use the data in Table 17.3.1 to calculate $\Delta G^\circ$ for the reduction of ferric ion by iodide:

$$2\text{Fe}^{3+}(aq) + 2\text{I}^-(aq) \rightarrow 2\text{Fe}^{2+}(aq) + \text{I}_2(s)$$

Is the reaction spontaneous?

Answer: $-44 \; \text{kJ/mol I}_2$; yes

Potentials for the Sums of Half-Reactions

Although Table 17.3.1 and Standard Reduction Potentials at 25°C list several half-reactions, many more are known. When the standard potential for a half-reaction is not available, we can use relationships between standard potentials and free energy to obtain the potential of any other half-reaction that can be written as the sum of two or more half-reactions whose standard potentials are available. For example, the potential for the reduction of $\text{Fe}^{3+}(aq)$ to $\text{Fe}(s)$ is not listed in the table, but two related reductions are given:

$$\text{Fe}^{3+}(aq) + e^{-} \rightarrow \text{Fe}^{2+}(aq) \; \; \; E^o=0.77 \; \text{V}$$
$$\text{Fe}^{2+}(aq) + 2e^{-} \rightarrow \text{Fe}(aq) \; \; \; E^o=-0.45 \; \text{V}$$

Although the sum of these two half-reactions gives the desired half-reaction, we cannot simply add the potentials of two reductive half-reactions to obtain the potential of a third reductive half-reaction because $E^o$ is not a state function. However, because $\Delta G^\circ$ is a state function, the sum of the $\Delta G^\circ$ values for the individual reactions gives us $\Delta G^\circ$ for the overall reaction, which is proportional to both the potential and the number of electrons ($n$) transferred. To obtain the value of $E^o$ for the overall half-reaction, we first must add the values of $\Delta G^\circ$ ($= -nFE^o$) for each individual half-reaction to obtain $\Delta G^\circ$ for the overall half-reaction:

$$\text{Fe}^{3+}(aq) + e^{-} \rightarrow \text{Fe}^{2+}(aq) \; \; \; \Delta G^\circ = -0.77 \; \text{V}$$
$$\text{Fe}^{2+}(aq) + 2e^{-} \rightarrow \text{Fe}(aq) \; \; \; \Delta G^\circ = -0.45 \; \text{V}$$

$$\Delta G^\circ = -0.77 + (-0.45) = -1.22 \; \text{V}$$

$$\text{Fe}^{3+}(aq) + 2e^{-} \rightarrow \text{Fe}(aq) \; \; \; \Delta G^\circ = -1.22 \; \text{V}$$
\[
\text{Fe}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Fe}(\text{s}) \quad \Delta G^{o} = -2F(-0.457\text{V})
\]
\[
\text{Fe}^{3+}(\text{aq}) + e^{-} \rightarrow \text{Fe}(\text{s}) \quad \Delta G^{o} = -1F(0.77\text{V}) + (-2F(-0.457\text{V})
\]

Solving the last expression for \(\Delta G^{o}\) for the overall half-reaction,
\[
\Delta G^{o} = F\left[\Delta G^{o}(\text{Fe}^{2+}(\text{aq}) + 2e^{-} \rightarrow \text{Fe}(\text{s})) + \Delta G^{o}(\text{Fe}^{3+}(\text{aq}) + e^{-} \rightarrow \text{Fe}(\text{s}))\right] = F\left(0.13\text{V}\right)
\]

(17.4.9)

Three electrons \((n = 3)\) are transferred in the overall reaction (Equation 17.4.8), so substituting into Equation 17.4.5 and solving for \(E^{o}\) gives the following:
\[
\Delta G^{o} = -nFE_{cell}^{o}
\]
\[
F\left(0.13\text{V}\right) = -3F E_{cell}^{o}
\]
\[
E_{cell}^{o} = -\dfrac{0.13\text{V}}{3} = -0.043\text{V}
\]

This value of \(E^{o}\) is very different from the value that is obtained by simply adding the potentials for the two half-reactions (0.32 V) and even has the opposite sign.

**Note the Pattern**

Values of \(E^{o}\) for half-reactions cannot be added to give \(E^{o}\) for the sum of the half-reactions; only values of \(\Delta G^{o} = -nFE_{cell}^{o}\) for half-reactions can be added.

**The Relationship between Cell Potential and the Equilibrium Constant**

We can use the relationship between \(\Delta G^{o}\) and the equilibrium constant \(K\), defined in Chapter 16, to obtain a relationship between \(E^{o}\) and \(K\). Recall that for a general reaction of the type \(aA + bB \rightarrow cC + dD\), the standard free-energy change and the equilibrium constant are related by the following equation:
\[
\Delta G^{o} = -RT\ln K
\]

(Equation 17.4.10)

Given the relationship between the standard free-energy change and the standard cell potential (Equation 17.4.5), we can write
\[
-nFE_{cell}^{o} = -RT\ln K
\]

(Equation 17.4.11)

Rearranging this equation,
\[
E_{cell}^{o} = \left(\dfrac{RT}{nF}\right)\ln K
\]

(Equation 17.4.12)

For \(T = 298\text{K}\), Equation 17.4.12 can be simplified as follows:
\[ E_{\text{cell}}^{o} = \left( \dfrac{RT}{nF} \right) \ln K = \left[ \dfrac{8.314 \text{ J}/(\text{mol} \cdot \text{K})}{96,486 \text{ J}/(\text{V} \cdot \text{mol})} \right] 2.303 \log K = \left( \dfrac{0.0591}{n} \right) \log K \tag{17.4.13} \]

Thus \( E_{\text{cell}}^{o} \) is directly proportional to the logarithm of the equilibrium constant. This means that large equilibrium constants correspond to large positive values of \( E_{\text{cell}}^{o} \) and vice versa.

---

**Example 17.4.2**

Use the data in Table 17.3.1 to calculate the equilibrium constant for the reaction of metallic lead with \( \text{PbO}_2 \) in the presence of sulfate ions to give \( \text{PbSO}_4 \) under standard conditions. (This reaction occurs when a car battery is discharged.) Report your answer to two significant figures.

**Given:** redox reaction

**Asked for:** \( K \)

**Strategy:**

A Write the relevant half-reactions and potentials. From these, obtain the overall reaction and \( E_{\text{cell}}^{o} \).

B Determine the number of electrons transferred in the overall reaction. Use Equation 17.4.13 to solve for \( \log K \) and then \( K \).

**Solution:**

A The relevant half-reactions and potentials from Table 17.3.1 are as follows:

\[
\begin{align*}
\text{cathode:} & \quad \text{PbO}_2(\text{s}) + \text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2e^- \rightarrow \text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l}) & E_{\text{cathode}}^{o} = 1.69 \text{ V} \\
\text{anode:} & \quad \text{Pb}(\text{s}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + 2e^- & E_{\text{anode}}^{o} = 0.36 \text{ V} \\
\text{overall:} & \quad \text{Pb}(\text{s}) + \text{PbO}_2(\text{s}) + 2\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) \rightarrow 2\text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l}) & E_{\text{cell}}^{o} = 2.05 \text{ V} \\
\end{align*}
\]

B Two electrons are transferred in the overall reaction, so \( n = 2 \). Solving Equation 17.4.13 for \( \log K \) and inserting the values of \( n \) and \( E_{\text{cell}}^{o} \),

\[
\ln K = \dfrac{nE^{o}}{0.0591 \text{ V}} = \dfrac{2(2.05 \text{ V})}{0.0591 \text{ V}} = 69.37
\]

\[
K = 2.3 \times 10^{69}
\]

Thus the equilibrium lies far to the right, favoring a discharged battery (as anyone who has ever tried unsuccessfully to start a car after letting it sit for a long time will know).
Exercise

Use the data in Table 17.3.1 to calculate the equilibrium constant for the reaction of Sn$^{2+}$(aq) with oxygen to produce Sn$^{4+}$(aq) and water under standard conditions. Report your answer to two significant figures. The reaction is as follows:

$$2\text{Sn}^{2+}(aq) + \text{O}_2(g) + 4\text{H}^+(aq) \rightleftharpoons 2\text{Sn}^{4+}(aq) + 2\text{H}_2\text{O}(l)$$

Answer: $1.2 \times 10^{73}$

Figure 17.4.1 summarizes the relationships that we have developed based on properties of the system—that is, based on the equilibrium constant, standard free-energy change, and standard cell potential—and the criteria for spontaneity ($\Delta G^o < 0$). Unfortunately, these criteria apply only to systems in which all reactants and products are present in their standard states, a situation that is seldom encountered in the real world. A more generally useful relationship between cell potential and reactant and product concentrations, as we are about to see, uses the relationship between $\Delta G$ and the reaction quotient $Q$ developed in Chapter 16.

Figure 17.4.1 The Relationships among Criteria for Thermodynamic Spontaneity The three properties of a system that can be used to predict the spontaneity of a redox reaction under standard conditions are $K$, $\Delta G^o$, and $E^{\circ}_{\text{cell}}$. If we know the value of one of these quantities, then these relationships enable us to calculate the value of the other two. The signs of $\Delta G^o$ and $E^{\circ}_{\text{cell}}$ and the magnitude of $K$ determine the direction of spontaneous reaction under standard conditions.

The Effect of Concentration on Cell Potential: The Nernst Equation

Recall from Chapter 16 that the actual free-energy change for a reaction under nonstandard conditions, $\Delta G$, is given as follows:

$$\Delta G = \Delta G^o + RT \ln Q \tag{17.4.14}$$

We also know that $\Delta G = -nF E_{\text{cell}}$ and $\Delta G^o = -nF E^{\circ}_{\text{cell}}$. Substituting these expressions into Equation 17.4.14, we obtain

$$-nF E_{\text{cell}} = -nF E^{\circ}_{\text{cell}} + RT \ln Q \tag{17.4.15}$$

Dividing both sides of this equation by $-nF$,

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln Q \tag{17.4.16}$$
Equation 17.4.16 is called the Nernst equation. An equation for calculating cell potentials $E_{\text{cell}}$ under nonstandard conditions; it can be used to determine the direction of spontaneous reaction for any redox reaction under any conditions:

$$E_{\text{cell}} = E^{\circ} - \frac{(RT/nF)\ln Q}{T}$$

after the German physicist and chemist Walter Nernst (1864–1941), who first derived it. The Nernst equation is arguably the most important relationship in electrochemistry. When a redox reaction is at equilibrium ($\Delta G = 0$), Equation 17.4.16 reduces to Equation 17.4.12 because $Q = K$, and there is no net transfer of electrons (i.e., $E_{\text{cell}} = 0$).

Substituting the values of the constants into Equation 17.4.16 with $T = 298$ K and converting to base-10 logarithms give the relationship of the actual cell potential ($E_{\text{cell}}$), the standard cell potential ($E^{\circ}_{\text{cell}}$), and the reactant and product concentrations at room temperature (contained in $Q$):

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591 \ V}{F} \ln Q$$

Equation 17.4.17 allows us to calculate the potential associated with any electrochemical cell at 298 K for any combination of reactant and product concentrations under any conditions. We can therefore determine the spontaneous direction of any redox reaction under any conditions, as long as we have tabulated values for the relevant standard electrode potentials. Notice in Equation 17.4.17 that the cell potential changes by $0.0591/n$ V for each 10-fold change in the value of $Q$ because $\log 10 = 1$.

Example 17.4.3

In the exercise in Example 6, you determined that the following reaction proceeds spontaneously under standard conditions because $E^{\circ}_{\text{cell}} > 0$ (which you now know means that $\Delta G^{\circ} < 0$):

$$2\text{Ce}^{4+}(aq) + 2\text{Cl}^{-}(aq) \rightarrow 2\text{Ce}^{3+}(aq) + \text{Cl}_2(g)$$

Calculate $E$ for this reaction under the following nonstandard conditions and determine whether it will occur spontaneously: $[\text{Ce}^{4+}] = 0.013 \ \text{M}$, $[\text{Ce}^{3+}] = 0.60 \ \text{M}$, $[\text{Cl}^{-}] = 0.0030 \ \text{M}$, $\text{Cl}_2 = 1.0 \ \text{atm}$, and $T = 25^\circ \text{C}$.

Given: balanced redox reaction, standard cell potential, and nonstandard conditions

Asked for: cell potential

Strategy:

Determine the number of electrons transferred during the redox process. Then use the Nernst equation to find the cell potential under the nonstandard conditions.

Solution:
We can use the information given and the Nernst equation to calculate \( E_{\text{cell}} \). Moreover, because the temperature is 25°C (298 K), we can use Equation 17.4.17 instead of 19.46. The overall reaction involves the net transfer of two electrons:

\[
2\text{Ce}^{4+}(aq) + 2e^- \rightarrow 2\text{Ce}^{3+}(aq) \\
2\text{Cl}^- (aq) \rightarrow \text{Cl}_2(g) + 2e^-
\]

so \( n = 2 \). Substituting the concentrations given in the problem, the partial pressure of \( \text{Cl}_2 \), and the value of \( E_{\text{cell}}^o \) into Equation 17.4.17,

\[
E_{\text{cell}} = E_{\text{cell}}^o - \frac{0.0591\;V}{F} \log Q
\]

\[
= 0.25\;V - \frac{0.0591\;V}{2} \log n \left( \frac{[\text{Ce}^{3+}]^2 P_{\text{Cl}_2}}{[\text{Ce}^{4+}]^2 [\text{Cl}^-]^2} \right)
\]

\[
= 0.25\;V - \left( 0.0296\;V \times 8.37 \right) = 0.00\;V
\]

Thus the reaction will not occur spontaneously under these conditions (because \( E = 0 \) V and \( \Delta G = 0 \)). The composition specified is that of an equilibrium mixture.

Exercise

In the exercise in Example 6, you determined that molecular oxygen will not oxidize MnO\(_2\) to permanganate via the reaction

\[
4\text{MnO}_2(s) + 3\text{O}_2(g) + 4\text{OH}^-(aq) \rightarrow 4\text{MnO}_4^-(aq) + 2\text{H}_2\text{O}(l)
\]

\( E_{\text{cell}}^o = -0.20\;V \)

Calculate \( E_{\text{cell}} \) for the reaction under the following nonstandard conditions and decide whether the reaction will occur spontaneously: pH 10, \( P(\text{O}_2) = 0.20 \) atm, \( [\text{MnO}_4^-] = 1.0 \times 10^{-4} \) M, and \( T = 25^\circ C \).

Answer: \( E_{\text{cell}} = -0.22 \) V; the reaction will not occur spontaneously.

Applying the Nernst equation to a simple electrochemical cell such as the Zn/Cu cell discussed in Section 17.1.4 allows us to see how the cell voltage varies as the reaction progresses and the concentrations of the dissolved ions change. Recall that the overall reaction for this cell is as follows:

\[
\text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s)
\]

\( E_{\text{cell}}^o = 1.10\;V \)

The reaction quotient is therefore \( Q = [\text{Zn}^2+/][\text{Cu}^{2+}] \). Suppose that the cell initially contains 1.0 M Cu\(^{2+}\) and 1.0 \times 10^{-6} M Zn\(^{2+}\). The initial voltage measured when the cell is connected can then be calculated from Equation 17.4.17:

\[
E_{\text{cell}} = E_{\text{cell}}^o - \frac{0.0591\;V}{n} \log \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}
\]
Thus the initial voltage is greater than $E^\circ$ because $Q < 1$. As the reaction proceeds, $[Zn^{2+}]$ in the anode compartment increases as the zinc electrode dissolves, while $[Cu^{2+}]$ in the cathode compartment decreases as metallic copper is deposited on the electrode. During this process, the ratio $Q = [Zn^{2+}]/[Cu^{2+}]$ steadily increases, and the cell voltage therefore steadily decreases. Eventually, $[Zn^{2+}] = [Cu^{2+}]$, so $Q = 1$ and $E_{cell} = E^\circ_{cell}$. Beyond this point, $[Zn^{2+}]$ will continue to increase in the anode compartment, and $[Cu^{2+}]$ will continue to decrease in the cathode compartment. Thus the value of $Q$ will increase further, leading to a further decrease in $E_{cell}$. When the concentrations in the two compartments are the opposite of the initial concentrations (i.e., $1.0 \ M \ Zn^{2+}$ and $1.0 \times 10^{-6} \ M \ Cu^{2+}$), $Q = 1.0 \times 10^6$, and the cell potential will be reduced to $0.92 \ V$.

The variation of $E_{cell}$ with log $Q$ over this range is linear with a slope of $-0.0591/n$, as illustrated in Figure 17.4.2. As the reaction proceeds still further, $Q$ continues to increase, and $E_{cell}$ continues to decrease. If neither of the electrodes dissolves completely, thereby breaking the electrical circuit, the cell voltage will eventually reach zero. This is the situation that occurs when a battery is “dead.” The value of $Q$ when $E_{cell} = 0$ is calculated as follows:

$$E_{\text{-cell}} = E_{\text{-cell}}^\circ - \frac{0.0591\ V}{n} \log(Q)$$

$$E^\circ = \frac{0.0591\ V}{n} \log(Q)$$

$$\log(Q) = \frac{E^\circ \times n}{0.0591\ V} = \frac{1.10\ \cancel{V} \times 2}{0.0591\ \cancel{V}} = 37.23$$

$$Q = 10^{37.23} = 1.7 \times 10^{37}$$

Figure 17.4.2 The Variation of $E_{cell}$ with Log $Q$ for a Zn/Cu Cell Initially, log $Q < 0$, and the voltage of the cell is greater than $E^\circ_{cell}$. As the reaction progresses, log $Q$ increases, and $E_{cell}$ decreases. When $[Zn^{2+}] = [Cu^{2+}]$, log $Q = 0$ and $E_{cell} = E^\circ_{cell} = 1.10 \ V$. As long as the electrical circuit remains intact, the reaction will continue, and log $Q$ will increase until $Q = K$ and the cell voltage reaches zero. At this point, the system will have reached equilibrium.

Recall that at equilibrium, $Q = K$. Thus the equilibrium constant for the reaction of Zn metal with $Cu^{2+}$ to give Cu metal and $Zn^{2+}$ is $1.7 \times 10^{37}$ at 25°C.
Concentration Cells

A voltage can also be generated by constructing an electrochemical cell in which each compartment contains the same redox active solution but at different concentrations. The voltage is produced as the concentrations equilibrate. Suppose, for example, we have a cell with 0.010 M AgNO$_3$ in one compartment and 1.0 M AgNO$_3$ in the other. The cell diagram and corresponding half-reactions are as follows:

\[
\begin{align*}
\text{cathode: } & \quad \text{Ag}^{+}(aq, \ 1.0 \ M) + e^{-} \rightarrow \text{Ag}(s) \quad \text{(17.4.22)} \\
\text{anode: } & \quad \text{Ag}(s) \rightarrow \text{Ag}^{+}(aq, \ 0.010 \ M) + e^{-} \quad \text{(17.4.23)} \\
\text{overall: } & \quad \text{Ag}^{+}(aq, \ 1.0 \ M) \rightarrow \text{Ag}^{+}(aq, \ 0.010 \ M) \quad \text{(17.4.24)}
\end{align*}
\]

As the reaction progresses, the concentration of Ag$^+$ will increase in the left (oxidation) compartment as the silver electrode dissolves, while the Ag$^+$ concentration in the right (reduction) compartment decreases as the electrode in that compartment gains mass. The total mass of Ag(s) in the cell will remain constant, however. We can calculate the potential of the cell using the Nernst equation, inserting 0 for $E^{\circ}_{\text{cell}}$ because $E^{\circ}_{\text{cathode}} = -E^{\circ}_{\text{anode}}$:

\[
E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{0.0591 \ V}{n} \log Q = 0 - \left( \frac{0.0591 \ V}{1} \right) \log \frac{0.010}{1.0} = 0.12 \quad \text{(17.4.25)}
\]

An electrochemical cell of this type, in which the anode and cathode compartments are identical except for the concentration of a reactant, is called a concentration cell. An electrochemical cell in which the anode and the cathode compartments are identical except for the concentration of a reactant.. As the reaction proceeds, the difference between the concentrations of Ag$^+$ in the two compartments will decrease, as will $E_{\text{cell}}$. Finally, when the concentration of Ag$^+$ is the same in both compartments, equilibrium will have been reached, and the measured potential difference between the two compartments will be zero ($E_{\text{cell}} = 0$).

Example 17.4.4

Calculate the voltage in a galvanic cell that contains a manganese electrode immersed in a 2.0 M solution of MnCl$_2$ as the cathode, and a manganese electrode immersed in a $5.2 \times 10^{-2}$ M solution of MnSO$_4$ as the anode ($T = 25^\circ C$).

**Given:** galvanic cell, identities of the electrodes, and solution concentrations

**Asked for:** voltage

**Strategy:**

A Write the overall reaction that occurs in the cell.

B Determine the number of electrons transferr
Substitute this value into the Nernst equation to calculate the voltage.

\[
E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0591 \text{ V}}{n} \log Q = 0 - \frac{0.0591 \text{ V}}{2} \log \frac{5.2 \times 10^{-2}}{2.0} = 0.047
\]

**Solution:**

A. This is a concentration cell, in which the electrode compartments contain the same redox active substance but at different concentrations. The anions (Cl\(^-\) and SO\(_4^{2-}\)) do not participate in the reaction, so their identity is not important. The overall reaction is as follows:

\[
\text{Mn}^{2+}(\text{aq, 2.0 M}) \rightarrow \text{Mn}^{2+}(\text{aq, 5.2} \times 10^{-2} \text{ M})
\]

B. For the reduction of Mn\(^{2+}\)(aq) to Mn(s), \(n = 2\). We substitute this value and the given Mn\(^{2+}\) concentrations into Equation 17.4.17:

Thus manganese will dissolve from the electrode in the compartment that contains the more dilute solution and will be deposited on the electrode in the compartment that contains the more concentrated solution.

**Exercise**

Suppose we construct a galvanic cell by placing two identical platinum electrodes in two beakers that are connected by a salt bridge. One beaker contains 1.0 M HCl, and the other a 0.010 M solution of Na\(_2\)SO\(_4\) at pH 7.00. Both cells are in contact with the atmosphere, with P(O\(_2\)) = 0.20 atm. If the relevant electrochemical reaction in both compartments is the four-electron reduction of oxygen to water, \(\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(l)\), what will be the potential when the circuit is closed?

**Answer:** 0.41 V

**Using Cell Potentials to Measure Solubility Products**

Because voltages are relatively easy to measure accurately using a voltmeter, electrochemical methods provide a convenient way to determine the concentrations of very dilute solutions and the solubility products (\(K_{sp}\)) of sparingly soluble substances. As you learned in Chapter 16, solubility products can be very small, with values of less than or equal to \(10^{-30}\). Equilibrium constants of this magnitude are virtually impossible to measure accurately by direct methods, so we must use alternative methods that are more sensitive, such as electrochemical methods.

To understand how an electrochemical cell is used to measure a solubility product, consider the cell shown in Figure 17.4.3, which is designed to measure the solubility product of silver chloride: \(K_{sp} = [\text{Ag}^+] [\text{Cl}^-]\). In one compartment, the cell contains a silver wire inserted into a 1.0 M solution of Ag\(^+\); the other compartment contains a silver wire inserted into a 1.0 M Cl\(^-\) solution saturated with AgCl. In this system, the Ag\(^+\) ion concentration in the first compartment equals \(K_{sp}\).

We can see this by dividing both sides of the equation for \(K_{sp}\) by [Cl\(^-\)] and substituting: 

\[
[\text{Ag}^+] = K_{sp} / [\text{Cl}^-] = K_{sp} / 1.0 = K_{sp}
\]

The overall cell reaction is as follows:

\[
\text{Ag}^+(\text{aq, concentrated}) \rightarrow \text{Ag}^+(\text{aq, dilute})
\]
Thus the voltage of the concentration cell due to the difference in \([Ag^+]\) between the two cells is as follows:

\[
E_{\text{cell}} = 0 - \left( \frac{0.0591 \; V}{1} \right) \log \left( \frac{[Ag^{+}]_{\text{dilute}}}{[Ag^{+}]_{\text{concentrated}}} \right) = -0.0591 \; V \tag{17.4.26}
\]

![Galvanic Cell for Measuring the Solubility Product of AgCl](image)

**Figure 17.4.3. A Galvanic Cell for Measuring the Solubility Product of AgCl** One compartment contains a silver wire inserted into a 1.0 M solution of \(Ag^+\), and the other compartment contains a silver wire inserted into a 1.0 M \(Cl^-\) solution saturated with AgCl. The potential due to the difference in \([Ag^+]\) between the two cells can be used to determine \(K_{sp}\).

By closing the circuit, we can measure the potential caused by the difference in \([Ag^+]\) in the two cells. In this case, the experimentally measured voltage of the concentration cell at 25°C is 0.580 V. Solving Equation 17.4.26 for \(K_{sp}\),

\[
\log K_{sp} = \frac{-E_{\text{cell}}}{0.0591 \; V} = \frac{-0.580 \; \cancel{V}}{0.0591 \; \cancel{V}} = -9.81 \tag{17.4.27}
\]

\(K_{sp} = 1.5 \times 10^{-10}\)

Thus a single potential measurement can provide the information we need to determine the value of the solubility product of a sparingly soluble salt.

**Example 17.4.5**

To measure the solubility product of lead(II) sulfate (PbSO₄) at 25°C, you construct a galvanic cell like the one shown in Figure 17.4.3, which contains a 1.0 M solution of a very soluble Pb²⁺ salt [lead(II) acetate trihydrate] in one compartment that is connected by a salt bridge to a 1.0 M solution of Na₂SO₄ saturated with PbSO₄ in the other. You then insert a Pb electrode into each compartment and close the circuit. Your voltmeter shows a voltage of 230 mV. What is \(K_{sp}\) for PbSO₄? Report your answer to two significant figures.

**Given:** galvanic cell, solution concentrations, electrodes, and voltage

**Asked for:** \(K_{sp}\)
Strategy:

A From the information given, write the equation for $K_{sp}$. Express this equation in terms of the concentration of $\text{Pb}^{2+}$.

B Determine the number of electrons transferred in the electrochemical reaction. Substitute the appropriate values into Equation 17.4.26 and solve for $K_{sp}$.

Solution:

A You have constructed a concentration cell, with one compartment containing a 1.0 M solution of $\text{Pb}^{2+}$ and the other containing a dilute solution of $\text{Pb}^{2+}$ in 1.0 M $\text{Na}_2\text{SO}_4$. As for any concentration cell, the voltage between the two compartments can be calculated using the Nernst equation. The first step is to relate the concentration of $\text{Pb}^{2+}$ in the dilute solution to $K_{sp}$:

\[
\left[ \text{Pb}^{2+} \right] \left[ \text{SO}_4^{2-} \right] = K_{sp}
\]

\[
\left[ \text{Pb}^{2+} \right] = \frac{K_{sp}}{1.0 \text{ M}} = K_{sp}
\]

B The reduction of $\text{Pb}^{2+}$ to $\text{Pb}$ is a two-electron process and proceeds according to the following reaction:

\[
\text{Pb}^{2+} (\text{aq, concentrated}) \rightarrow \text{Pb}^{2+} (\text{aq, dilute})
\]

so

\[
E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.0591 \text{ V}}{n} \log Q
\]

\[
0.230 \text{ V} = 0 - \frac{0.0591 \text{ V}}{2} \log \left( \frac{\left[ \text{Pb}^{2+} \right]_{\text{dilute}}}{\left[ \text{Pb}^{2+} \right]_{\text{concentrated}}} \right) = -0.0296 \text{ V} \log \left( \frac{K_{sp}}{1.0} \right)
\]

\[
-7.77 = \log \left( \frac{K_{sp}}{1.0} \right)
\]

\[
1.7 \times 10^{-8} = K_{sp}
\]

Exercise

A concentration cell similar to the one described in Example 11 contains a 1.0 M solution of lanthanum nitrate $[\text{La(NO}_3)_3]$ in one compartment and a 1.0 M solution of sodium fluoride saturated with $\text{LaF}_3$ in the other. A metallic La strip is inserted into each compartment, and the circuit is closed. The measured potential is 0.32 V. What is the $K_{sp}$ for $\text{LaF}_3$? Report your answer to two significant figures.

Answer: $5.7 \times 10^{-17}$

Using Cell Potentials to Measure Concentrations

Another use for the Nernst equation is to calculate the concentration of a species given a measured potential and the concentrations of all the other species. We saw an example of this in Example 11, in which the experimental conditions
were defined in such a way that the concentration of the metal ion was equal to $K_{sp}$. Potential measurements can be used to obtain the concentrations of dissolved species under other conditions as well, which explains the widespread use of electrochemical cells in many analytical devices. Perhaps the most common application is in the determination of $[H^+]$ using a pH meter, as illustrated in Example 12.

**Example 17.4.6**

Suppose a galvanic cell is constructed with a standard Zn/Zn$^{2+}$ couple in one compartment and a modified hydrogen electrode in the second compartment (Figure 17.1.3). The pressure of hydrogen gas is 1.0 atm, but $[H^+]$ in the second compartment is unknown. The cell diagram is as follows:

\[
\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq, 1.0 M}) \parallel \text{H}^+(\text{aq, ? M}) \mid \text{H}_2(\text{g, 1.0 atm}) \mid \text{Pt(s)}
\]

What is the pH of the solution in the second compartment if the measured potential in the cell is 0.26 V at 25°C?

**Given:** galvanic cell, cell diagram, and cell potential

**Asked for:** pH of the solution

**Strategy:**

A Write the overall cell reaction.

B Substitute appropriate values into the Nernst equation and solve for $-\log[H^+]$ to obtain the pH.

**Solution:**

A Under standard conditions, the overall reaction that occurs is the reduction of protons by zinc to give H$_2$ (note that Zn lies below H$_2$ in Table 17.3.1):

\[
\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g}) \quad E_{o} = 0.76 \text{ V}
\]

B By substituting the given values into the simplified Nernst equation (Equation 17.4.17), we can calculate $[H^+]$ under nonstandard conditions:

\[
E_{\text{cell}} = E_{\text{cell}}^{o} - \frac{0.0591 \text{ V}}{n} \log \left( \frac{[\text{Zn}^{2+}]P_{\text{H}_2}}{[\text{H}^+]^2} \right)
\]

\[
0.26 \text{ V} = 0.76 \text{ V} - \left( \frac{0.0591 \text{ V}}{2} \right) \log \left( \frac{(1.0)(1.0)}{[\text{H}^+]^2} \right) = -0.0296 \text{ V} \log \left( \frac{1}{[\text{H}^+]^2} \right) = -2 \log [\text{H}^+] = 16.9
\]

\[
8.46 = -\log[\text{H}^+] \quad \Rightarrow \quad 8.5 = \text{pH}
\]
Thus the potential of a galvanic cell can be used to measure the pH of a solution.

Exercise

Suppose you work for an environmental laboratory and you want to use an electrochemical method to measure the concentration of Pb$^{2+}$ in groundwater. You construct a galvanic cell using a standard oxygen electrode in one compartment ($E^\circ_{\text{cathode}} = 1.23$ V). The other compartment contains a strip of lead in a sample of groundwater to which you have added sufficient acetic acid, a weak organic acid, to ensure electrical conductivity. The cell diagram is as follows:

$$\text{Pb(s) | Pb}^{2+}(\text{aq, } ? \text{ M}) || \text{H}^+(\text{aq}, 1.0 \text{ M})| \text{O}_2(\text{g, 1.0 atm})|\text{Pt(s)}$$

When the circuit is closed, the cell has a measured potential of 1.62 V. Use Table 17.7.1 and Standard Reduction Potentials at 25°C to determine the concentration of Pb$^{2+}$ in the groundwater.

Answer: 1.2 x 10$^{-9}$ M

Summary

A coulomb (C) relates electrical potential, expressed in volts, and energy, expressed in joules. The current generated from a redox reaction is measured in amperes (A), where 1 A is defined as the flow of 1 C/s past a given point. The faraday (F) is Avogadro’s number multiplied by the charge on an electron and corresponds to the charge on 1 mol of electrons. The product of the cell potential and the total charge is the maximum amount of energy available to do work, which is related to the change in free energy that occurs during the chemical process. Adding together the $\Delta G$ values for the half-reactions gives $\Delta G$ for the overall reaction, which is proportional to both the potential and the number of electrons ($n$) transferred. Spontaneous redox reactions have a negative $\Delta G$ and therefore a positive $E_{\text{cell}}$. Because the equilibrium constant $K$ is related to $\Delta G$, $E^\circ_{\text{cell}}$ and $K$ are also related. Large equilibrium constants correspond to large positive values of $E^\circ$. The Nernst equation allows us to determine the spontaneous direction of any redox reaction under any reaction conditions from values of the relevant standard electrode potentials. Concentration cells consist of anode and cathode compartments that are identical except for the concentrations of the reactant. Because $\Delta G = 0$ at equilibrium, the measured potential of a concentration cell is zero at equilibrium (the concentrations are equal). A galvanic cell can also be used to measure the solubility product of a sparingly soluble substance and calculate the concentration of a species given a measured potential and the concentrations of all the other species.

Key Takeaway

- The Nernst equation can be used to determine the direction of spontaneous reaction for any redox reaction in aqueous solution.

Key Equations

Charge on a mole of electrons (faraday)
Equation 17.4.2: \( F \approx 96,486 \text{ J/(V\cdot mol)} \)

Maximum work from an electrochemical cell

Equation 17.4.3: \( w_{\text{max}} = -nF\varepsilon_{\text{cell}} \)

Relationship between \( \Delta G^\circ \) and \( \Delta E^\circ \)

Equation 17.4.5: \( \Delta G^\circ = -nF\varepsilon_{\text{cell}} \)

Relationship between \( \Delta G^\circ \) and \( K \) for a redox reaction

Equation 17.4.10: \( \Delta G^\circ = -RT \ln K \)

Relationship between \( \Delta E^\circ \) and \( K \) for a redox reaction at 25°C

Equation 17.4.12: \( \left( \frac{RT}{nF} \right) \ln K \)

Equation 17.4.13: \( \left( \frac{0.0591}{n} \right) \log K \)

Relationship between \( \Delta G^\circ \) and \( Q \)

Equation 17.4.14: \( \Delta G = \Delta G^\circ + RT \ln Q \)

Relationship between \( E_{\text{cell}} \) and \( Q \) at 25°C

Equation 17.4.17: \( E_{\text{cell}} = E_{\text{cell}}^\circ - \left( \frac{0.0591}{F} \right) \ln Q \)

Conceptual Problems

1. State whether you agree or disagree with this reasoning and explain your answer: Standard electrode potentials arise from the number of electrons transferred. The greater the number of electrons transferred, the greater the measured potential difference. If 1 mol of a substance produces 0.76 V when 2 mol of electrons are transferred—as in \( \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^- \)—then 0.5 mol of the substance will produce 0.76/2 V because only 1 mol of electrons is transferred.

2. What is the relationship between the measured cell potential and the total charge that passes through a cell? Which of these is dependent on concentration? Which is dependent on the identity of the oxidant or the reductant? Which is dependent on the number of electrons transferred?

3. In the equation \( w_{\text{max}} = -nF\varepsilon_{\text{cell}} \), which quantities are extensive properties and which are intensive properties?

4. For any spontaneous redox reaction, \( E \) is positive. Use thermodynamic arguments to explain why this is true.

5. State whether you agree or disagree with this statement and explain your answer: Electrochemical methods are especially useful in determining the reversibility or irreversibility of reactions that take place in a cell.
6. Although the sum of two half-reactions gives another half-reaction, the sum of the potentials of the two half-reactions cannot be used to obtain the potential of the net half-reaction. Why? When does the sum of two half-reactions correspond to the overall reaction? Why?

7. Occasionally, you will find high-quality electronic equipment that has its electronic components plated in gold. What is the advantage of this?

8. Blood analyzers, which measure pH, P(CO₂) and P(O₂) are frequently used in clinical emergencies. For example, blood P(CO₂) is measured with a pH electrode covered with a plastic membrane that is permeable to CO₂. Based on your knowledge of how electrodes function, explain how such an electrode might work. Hint: CO₂(g) + H₂O(l) → HCO₃⁻(aq) + H⁺(aq).

9. Concentration cells contain the same species in solution in two different compartments. Explain what produces a voltage in a concentration cell. When does V = 0 in such a cell?

10. Describe how an electrochemical cell can be used to measure the solubility of a sparingly soluble salt.

Answers

1. extensive: \( w_{\text{max}} \) and \( n \); intensive: \( E^{\circ}_{\text{cell}} \)

2. Gold is highly resistant to corrosion because of its very positive reduction potential.

Numerical Problems

1. The chemical equation for the combustion of butane is as follows:

\[
( C_{4}H{10} (g) + \frac{13}{2}O_{2} (g) \rightarrow 2CO_{2} (g) + 5H_{2}O (g) )
\]

This reaction has \( \Delta H^{\circ} = -2877 \text{ kJ/mol} \). Calculate \( E^{\circ}_{\text{cell}} \) and then determine \( \Delta G^{\circ} \). Is this a spontaneous process? What is the change in entropy that accompanies this process at 298 K?

2. How many electrons are transferred during the reaction Pb(s) + Hg₂Cl₂(s) → PbCl₂(aq) + 2Hg(l)? What is the standard cell potential? Is the oxidation of Pb by Hg₂Cl₂ spontaneous? Calculate \( \Delta G^{\circ} \) for this reaction.

3. For the cell represented as Al(s)||Al³⁺(aq)||Sn²⁺(aq), Sn⁴⁺(aq)||Pt(s), how many electrons are transferred in the redox reaction? What is the standard cell potential? Is this a spontaneous process? What is \( \Delta G^{\circ} \)?

4. Explain why the sum of the potentials for the half-reactions Sn²⁺(aq) + 2e⁻ → Sn(s) and Sn⁴⁺(aq) + 2e⁻ → Sn²⁺(aq) does not equal the potential for the reaction Sn⁴⁺(aq) + 4e⁻ → Sn(s). What is the net cell potential? Compare the values of \( \Delta G^{\circ} \) for the sum of the potentials and the actual net cell potential.
Based on the Standard Reduction Potentials at 25°C, do you agree with the proposed potentials for the following half-reactions? Why or why not?

1. \( \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s), \ E^\circ = 0.68 \text{ V} \)
2. \( \text{Ce}^{4+}(aq) + 4e^- \rightarrow \text{Ce}(s), \ E^\circ = -0.62 \text{ V} \)

For each reaction, calculate \( E^\circ_{\text{cell}} \) and then determine \( \Delta G^\circ \). Indicate whether each reaction is spontaneous.

1. \( 2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g) \)
2. \( \text{K}_2\text{S}_2\text{O}_6(aq) + \text{I}_2(s) \rightarrow 2\text{KI}(aq) + 2\text{K}_2\text{SO}_4(aq) \)
3. \( \text{Sn}(s) + \text{CuSO}_4(aq) \rightarrow \text{Cu}(s) + \text{SnSO}_4(aq) \)

What is the standard change in free energy for the reaction between \( \text{Ca}^{2+} \) and \( \text{Na}(s) \) to give \( \text{Ca}(s) \) and \( \text{Na}^+ \)? Do the sign and magnitude of \( \Delta G^\circ \) agree with what you would expect based on the positions of these elements in the periodic table? Why or why not?

In acidic solution, permanganate (\( \text{MnO}_4^- \)) oxidizes \( \text{Cl}^- \) to chlorine gas, and \( \text{MnO}_4^- \) is reduced to \( \text{Mn}^{2+}(aq) \).

1. Write the balanced chemical equation for this reaction.
2. Determine \( E^\circ_{\text{cell}} \).
3. Calculate the equilibrium constant.

Potentiometric titrations are an efficient method for determining the endpoint of a redox titration. In such a titration, the potential of the solution is monitored as measured volumes of an oxidant or a reductant are added. Data for a typical titration, the potentiometric titration of Fe(II) with a 0.1 M solution of Ce(IV), are given in the following table. The starting potential has been arbitrarily set equal to zero because it is the change in potential with the addition of the oxidant that is important.

<table>
<thead>
<tr>
<th>Titrant (mL)</th>
<th>( E ) (mV)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.00</td>
<td>50</td>
</tr>
<tr>
<td>6.00</td>
<td>100</td>
</tr>
<tr>
<td>9.00</td>
<td>255</td>
</tr>
<tr>
<td>10.00</td>
<td>960</td>
</tr>
<tr>
<td>11.00</td>
<td>1325</td>
</tr>
<tr>
<td>12.00</td>
<td>1625</td>
</tr>
<tr>
<td>14.00</td>
<td>1875</td>
</tr>
</tbody>
</table>

1. Write the balanced chemical equation for the oxidation of Fe\(^{2+}\) by Ce\(^{4+}\).
2. Plot the data and then locate the endpoint.
3. How many millimoles of Fe\(^{2+}\) did the solution being titrated originally contain?
10. The standard electrode potential \( (E^\circ) \) for the half-reaction \( \text{Ni}^{2+} (aq) + 2e^- \rightarrow \text{Ni}(s) \) is \(-0.257 \, \text{V}\). What pH is needed for this reaction to take place in the presence of 1.00 atm \( \text{H}_2(g) \) as the reductant if \( [\text{Ni}^{2+}] \) is 1.00 M?

11. The reduction of Mn(VII) to Mn(s) by \( \text{H}_2(g) \) proceeds in five steps that can be readily followed by changes in the color of the solution. Here is the redox chemistry:

   1. \( \text{MnO}_4^- (aq) + e^- \rightarrow \text{MnO}_4^{2-} (aq); \quad E^\circ = +0.56 \, \text{V} \) (purple \( \rightarrow \) dark green)
   2. \( \text{MnO}_4^{2-} (aq) + 2e^- + 4\text{H}^+ (aq) \rightarrow \text{MnO}_2(s); \quad E^\circ = +2.26 \, \text{V} \) (dark green \( \rightarrow \) dark brown solid)
   3. \( \text{MnO}_2(s) + e^- + 4\text{H}^+ (aq) \rightarrow \text{Mn}^{3+} (aq); \quad E^\circ = +0.95 \, \text{V} \) (dark brown solid \( \rightarrow \) red-violet)
   4. \( \text{Mn}^{3+} (aq) + e^- \rightarrow \text{Mn}^{2+} (aq); \quad E^\circ = +1.51 \, \text{V} \) (red-violet \( \rightarrow \) pale pink)
   5. \( \text{Mn}^{2+} (aq) + 2e^- \rightarrow \text{Mn}(s); \quad E^\circ = -1.18 \, \text{V} \) (pale pink \( \rightarrow \) colorless)

   1. Is the reduction of \( \text{MnO}_4^- \) to \( \text{Mn}^{3+} \) (aq) by \( \text{H}_2(g) \) spontaneous under standard conditions? What is \( E^\circ_{\text{cell}} \)?
   2. Is the reduction of \( \text{Mn}^{3+} \) (aq) to Mn(s) by \( \text{H}_2(g) \) spontaneous under standard conditions? What is \( E^\circ_{\text{cell}} \)?

12. Mn(III) can disproportionate (both oxidize and reduce itself) by means of the following half-reactions:

   \[
   \begin{align*}
   \text{Mn}^{3+} (aq) + e^- & \rightarrow \text{Mn}^{2+} (aq); \quad E^\circ = +1.51 \, \text{V} \\
   \text{Mn}^{3+} (aq) + 2\text{H}_2\text{O}(l) & \rightarrow \text{MnO}_2(s) + 4\text{H}^+ (aq) + e^-; \quad E^\circ = +0.95 \, \text{V}
   \end{align*}
   \]

   1. What is \( E^\circ \) for the disproportionation reaction?
   2. Is disproportionation more or less thermodynamically favored at low pH than at pH 7.0? Explain your answer.
   3. How could you prevent the disproportionation reaction from occurring?

13. For the reduction of oxygen to water, \( E^\circ = 1.23 \, \text{V} \). What is the potential for this half-reaction at pH 7.00? What is the potential in a 0.85 M solution of NaOH?

14. The biological molecule abbreviated as NADH (reduced nicotinamide adenine dinucleotide) can be formed by reduction of NAD\(^+\) (nicotinamide adenine dinucleotide) via the half-reaction \( \text{NAD}^+ + \text{H}^+ + 2e^- \rightarrow \text{NADH}; \quad E^\circ = -0.32 \, \text{V} \).

   1. Would NADH be able to reduce acetate to pyruvate?
   2. Would NADH be able to reduce pyruvate to lactate?
   3. What potential is needed to convert acetate to lactate?

   \[
   \begin{align*}
   \text{acetate} + \text{CO}_2 + 2\text{H}^+ + 2e^- & \rightarrow \text{pyruvate} + \text{H}_2\text{O}; \quad E^\circ = +0.70 \, \text{V} \\
   \text{pyruvate} + 2\text{H}^+ + 2e^- & \rightarrow \text{lactate}; \quad E^\circ = -0.185 \, \text{V}
   \end{align*}
   \]

15. Given the following biologically relevant half-reactions, will FAD (flavin adenine dinucleotide), a molecule used to transfer electrons whose reduced form is FADH\(_2\), be an effective oxidant for the conversion of acetaldehyde to acetate at pH 4.00?
acetate + 2H⁺ + 2e⁻ → acetaldehyde + H₂O; \( E^\circ = -0.58 \) V

FAD + 2H⁺ + 2e⁻ → FADH₂; \( E^\circ = -0.18 \) V

16. Ideally, any half-reaction with \( E^\circ > 1.23 \) V will oxidize water as a result of the half-reaction \( \text{O}_2(g) + 4\text{H}^+(aq) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(l) \).

   1. Will FeO₄²⁻ oxidize water if the half-reaction for the reduction of Fe(VI) → Fe(III) is FeO₄²⁻(aq) + 8H⁺(aq) + 3e⁻ → Fe³⁺(aq) + 4H₂O; \( E^\circ = 1.9 \) V?

   2. What is the highest pH at which this reaction will proceed spontaneously if [Fe³⁺] = [FeO₄²⁻] = 1.0 M and \( P_{\text{O}_2} = 1.0 \) atm?

17. Under acidic conditions, ideally any half-reaction with \( E^\circ > 1.23 \) V will oxidize water via the reaction

\[ \text{O}_2(g) + 4\text{H}^+(aq) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(l). \]

   1. Will aqueous acidic KMnO₄ evolve oxygen with the formation of MnO₂?

   2. At pH 14.00, what is \( E^\circ \) for the oxidation of water by aqueous KMnO₄ (1 M) with the formation of MnO₂?

   3. At pH 14.00, will water be oxidized if you are trying to form MnO₂ from MnO₄²⁻ via the reaction

\[ 2\text{MnO}_4^{2-}(aq) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{MnO}_2(s) + \text{O}_2(g) + 4\text{OH}^-(aq) \]

18. Complexing agents can bind to metals and result in the net stabilization of the complexed species. What is the net thermodynamic stabilization energy that results from using CN⁻ as a complexing agent for Mn³⁺/Mn²⁺?

\[ \text{Mn}^{3+}(aq) + \text{e}^- \rightarrow \text{Mn}^{2+}(aq); \ E^\circ = +1.51 \) V

\[ \text{Mn(CN)}_6^{3-}(aq) + \text{e}^- \rightarrow \text{Mn(CN)}_6^{4-}(aq); \ E^\circ = -0.24 \) V

19. You have constructed a cell with zinc and lead amalgam electrodes described by the cell diagram

\[ \text{Zn(Hg)}(s) | \text{Zn(NO}_3)_2(aq) || \text{Pb(NO}_3)_2(aq) | \text{Pb(Hg)}(s). \]

If you vary the concentration of Zn(NO₃)₂ and measure the potential at different concentrations, you obtain the following data:

<table>
<thead>
<tr>
<th>Zn(NO₃)₂ (M)</th>
<th>( E_{\text{cell}} ) (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0005</td>
<td>0.7398</td>
</tr>
<tr>
<td>0.002</td>
<td>0.7221</td>
</tr>
<tr>
<td>0.01</td>
<td>0.7014</td>
</tr>
</tbody>
</table>

   1. Write the half-reactions that occur in this cell.

   2. What is the overall redox reaction?

   3. What is \( E^\circ_{\text{cell}} \)? What is \( \Delta G^\circ \) for the overall reaction?

   4. What is the equilibrium constant for this redox reaction?
20. Hydrogen gas reduces Ni$^{2+}$ according to the following reaction: Ni$^{2+}$(aq) + H$_2$(g) → Ni(s) + 2H$^+$(aq); $E^\circ_{\text{cell}} = -0.25$ V; $\Delta H = 54$ kJ/mol.

1. What is $K$ for this redox reaction?
2. Is this reaction likely to occur?
3. What conditions can be changed to increase the likelihood that the reaction will occur as written?
4. Is the reaction more likely to occur at higher or lower pH?

21. The silver–silver bromide electrode has a standard potential of 0.07133 V. What is $K_{sp}$ of AgBr?

Answers

18. 6e$^-$; $E^\circ_{\text{cell}} = 1.813$ V; the reaction is spontaneous; $\Delta G^\circ = -525$ kJ/mol Al.

16. yes; $E^\circ = 0.40$ V

17. 1. yes; $E^\circ = 0.45$ V
   2. 0.194 V

20. 3. yes; $E^\circ = 0.20$ V

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