Standard Reduction Potentials

In Section 10.1, you learned how a galvanic cell, similar to the one in **Figure 1(a)**, converts the chemical energy from a spontaneous redox reaction into electrical energy. Although the design of a galvanic cell makes the cell impractical as an energy source for a portable electronic device, it is useful for showing the chemical processes that are involved. It is also useful for showing how commercial galvanic cells, such as 1.5 V alkaline dry cells (**Figure 1(b)**), work. An important design feature of any galvanic cell is its cell potential.

Figure 1 Both (a) an experimental zinc–copper galvanic cell and (b) commercial 1.5 V alkaline dry cells operate on the same principle.

Cell Potential

The electrons in each half-cell of a galvanic cell have different amounts of electric potential energy. This is analogous to two apples hanging from a tree. The apple hanging from an upper branch of the tree has more gravitational potential energy than the apple hanging from a lower branch. **Cell potential** is a measure of the electric potential difference or voltage between two half-cells. Cell potential is analogous to the difference in gravitational potential energy between the two apples hanging at different heights on the tree.

The SI unit for cell potential is the volt (symbol: V). One volt is defined as one joule of energy per coulomb of electric charge, or 1 V = 1 J/C. The coulomb is the SI unit for electric charge. (1 C is equivalent to the charge of 6.25×10^{18} electrons.)

Since cell potential is a ratio of energy to charge, it is unaffected by the number of electrons transferred as a galvanic cell operates. For example, an alkaline D cell transfers electrons, and therefore electrical energy, for a longer time than a smaller AAA cell because it contains a greater quantity of reactants. However, the quantity of the energy per charge (in joules per coulomb) transferred by both cells is identical: 1.5 J/C or 1.5 V. Factors that influence cell potential include the design of the cell, the redox reaction involved, and the concentration of the reactants in the cell. For example, the lithium-ion cell and the alkaline cell have different cell potentials because they use different redox reactions.

The cell potential for a galvanic cell is measured using a voltmeter. The magnitude of the reading indicates the difference in the ability of the two half-cells to attract electrons. The sign of the voltmeter reading identifies the anode, the cathode, and the direction of electron flow in the cell. A voltmeter typically has two terminals: a red (positive) terminal and a black (negative) terminal. A positive reading on the voltmeter means that the positive terminal is connected to the cathode of the cell and the negative terminal is connected to the anode. A negative reading means that the connections are reversed. Once the anode and the cathode are determined, the direction of electron flow—from the anode to the cathode, or vice versa—is known. ● CAREER LINK

cell potential the electric potential difference (voltage) between the two half-cells in a galvanic cell; the SI unit is the volt, and the unit symbol is V (1 V = 1 J/C)

LEARNING TIP

Electrode Mnemonic
These mnemonics may help you remember the process that occurs at each electrode:
- RED CAT—REDuction always occurs at the CAThode.
- AN OX—OXidation always occurs at the ANode.
Can you think of your own mnemonics?

Standard Cells and Standard Cell Potentials

Cell potential varies depending on the concentration of the chemicals in a cell. To study cells more easily, chemists have defined standard conditions under which cells operate. A **standard cell** is a galvanic cell in which all the entities are at SATP, with a concentration of 1.0 mol/L for solutions. For example, the line notation for a standard zinc–copper cell (**Figure 1(a)**) is

$$Zn(s)\,|\,Zn^{2+}(aq)\,||\,Cu^{2+}(aq)\,|\,Cu(s)$$

$$\qquad\quad 1\ mol/L \qquad 1\ mol/L$$
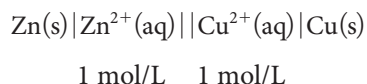
The **standard cell potential** ($\Delta E^{\circ}_{r\,(cell)}$) is the electric potential difference of a cell that is operating under standard conditions. It is a measure of the potential energy difference, per unit of charge, between the cathode and the anode. The degree symbol is used to indicate standard conditions. The **standard reduction potential** (E°_r) indicates the ability of a half-cell to attract electrons, and therefore undergo reduction. As a cell operates, the half-cell with the greater (more positive) reduction potential gains electrons from the half-cell with the lower reduction potential. Therefore, the standard cell potential is the difference between the reduction potentials of the two half-cells:

$$\Delta E^{\circ}_{r\,(cell)} = E^{\circ}_{r\,(cathode)} - E^{\circ}_{r\,(anode)}$$

standard cell a galvanic cell in which all the entities involved in the half-cell reactions are at SATP and the solutions have a concentration of 1.0 mol/L

standard cell potential ($\Delta E^{\circ}_{r\,(cell)}$) the electric potential difference of a galvanic cell that is operating under standard conditions

standard reduction potential (E°_r) the ability of a half-cell to attract electrons in a cell that is operating under standard conditions

Developing a Reduction Potentials Table

Cell potentials are measured with a voltmeter. The electric potential of a half-cell cannot be measured because a half-cell reaction cannot occur on its own. Chemists have overcome the difficulty in predicting cell potentials for combinations of half-cells by assigning the **standard hydrogen half-cell** an electric potential of exactly 0 V (**Figure 2**). The standard hydrogen half-cell consists of an inert platinum electrode in contact with an acidic solution containing 1.0 mol/L hydrogen ions through which hydrogen gas is bubbled at a pressure of 100 kPa. The electric potentials of other half-cells can be determined relative to the standard hydrogen half-cell.

standard hydrogen half-cell the galvanic cell from which all the half-cell potentials are determined; $E^{\circ}_r = 0$ V (by definition)
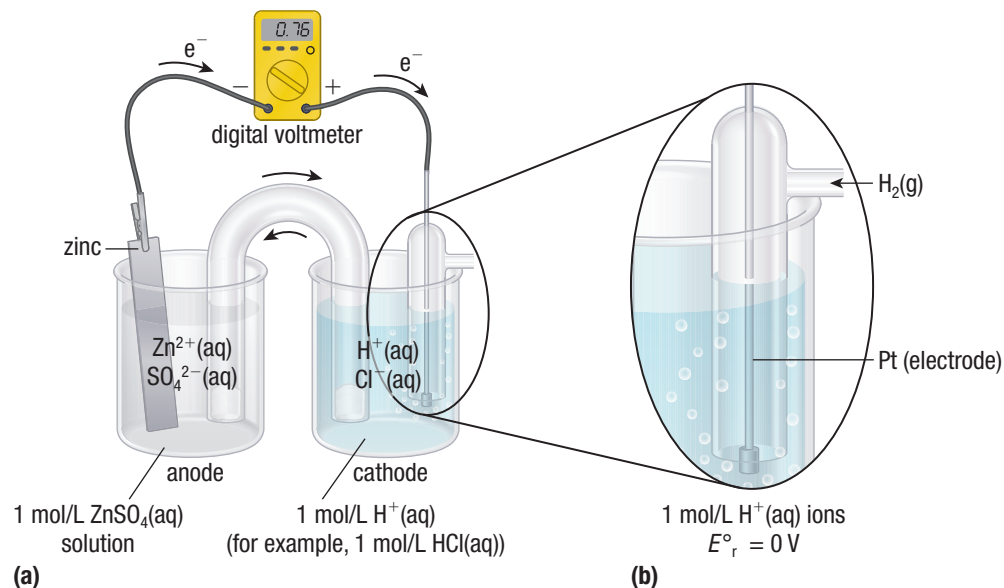
Figure 2 (a) In this galvanic cell, hydrogen is reduced and zinc is oxidized. (b) In the hydrogen half-cell, hydrogen gas at 100 kPa passes over a platinum electrode in a 1 mol/L solution of hydrogen ions. The reduction of hydrogen at this electrode is arbitrarily assigned a value of exactly 0 V.

The cell shown in Figure 2(a) contains zinc ions in one half-cell and hydrogen ions in the other half-cell. Due to their positive charge, both zinc ions and hydrogen ions compete for the electrons—both ions can be oxidizing agents. However, the positive

cell potential indicates that electrons flow from the zinc anode to the platinum cathode. Therefore, the hydrogen ion is a stronger oxidizing agent than the zinc ion.

The cell potential for the zinc–hydrogen cell is 0.76 V. Therefore, the reduction potential of the hydrogen ions is greater than the reduction potential of the zinc ions by 0.76 V. Since the reduction potential of the hydrogen ions is defined as exactly 0 V, the reduction potential of the zinc ions, measured relative to the hydrogen ions, must be -0.76 V. This value is a measure of how much stronger hydrogen ions are as oxidizing agents than zinc ions (**Figure 3**). The two reduction half-reactions can be listed in order of decreasing reduction potential as follows:
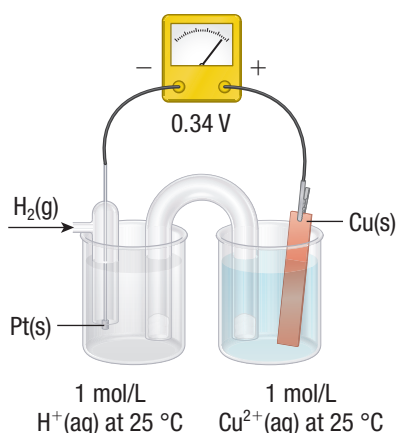
$$2\,H^+(aq) + 2e^- \rightleftharpoons H_2(g) \quad E^\circ_r = 0\text{ V}$$

$$Zn^{2+}(aq) + 2\,e^- \rightleftharpoons Zn(s)\ E^\circ_r = -0.76\text{ V}$$
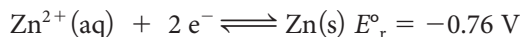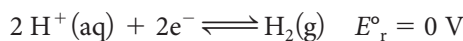
If the zinc half-cell in Figure 2 is replaced with a copper half-cell, the reading on the meter changes to –0.34 V. The change in sign means that electrons are now flowing in the opposite direction—from the hydrogen half-cell to the copper half-cell. Therefore, copper is the cathode and platinum is the anode. If the connections to the electrodes are reversed, the reading changes to a positive value of 0.34 V (**Figures 4** and **5**). This means that the copper(II) ion is a better competitor for electrons (a better oxidizing agent) than the hydrogen ion by 0.34 V.
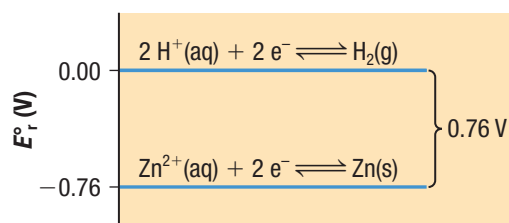
Figure 3 $H^+(aq)$ is a stronger oxidizing agent than $Zn^{2+}(aq)$ by 0.76 V.
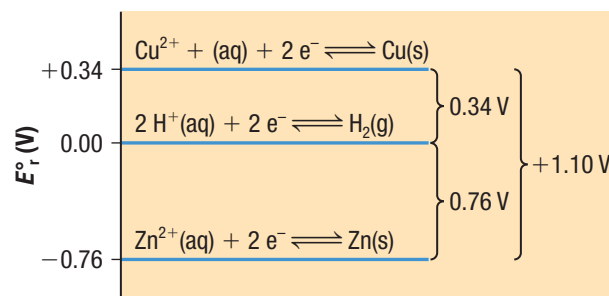
**Figure 4** $Cu^{2+}(aq)$ is a stronger oxidizing agent than both $H^+(aq)$ and $Zn^{2+}(aq)$. A standard cell constructed using copper and zinc half-cells has a predicted cell potential of +1.10 V.
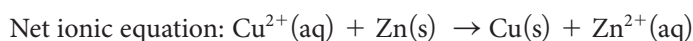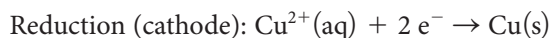
We can therefore add copper to our reduction potentials list:

$$Cu^{2+}(aq) + 2\,e^- \rightleftharpoons Cu(s)\ E^\circ = +0.34\text{ V}$$

$$2\,H^+(aq) + 2\,e^- \rightleftharpoons H_2(g)\ E^\circ = 0.00\text{ V}$$

$$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)\ E^\circ = -0.76\text{ V}$$

Figure 5 A standard galvanic cell that involves the half-reactions $H_2(g) \rightarrow 2\,H^+(aq) + 2\,e^-$ (at the anode) and $Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$ (at the cathode) has a potential of 0.34 V.
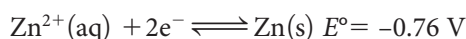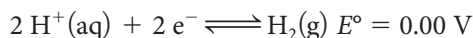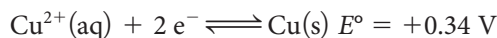
These data can be used to predict the net ionic equation for the cell reaction and cell potential for the zinc–copper cell (**Figure 6**, next page):

Reduction (cathode): $Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$

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

Net ionic equation: $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$

The standard cell potential ($\Delta E^\circ_{r\,(cell)}$) is the difference between the reduction potentials of the two half-cells:

$$\Delta E^\circ_{r\,(cell)} = E^\circ_{r\,(cathode)} - E^\circ_{r\,(anode)}$$
$$= +0.34\text{ V} - (-0.76\text{ V})$$
$$= +1.10\text{ V}$$

If we continued measuring the standard reduction potentials of many other half-cells against the hydrogen half-cell, we could add them to our ranked list of reduction potentials. We would then be able to predict the cell potentials for a variety of cells.
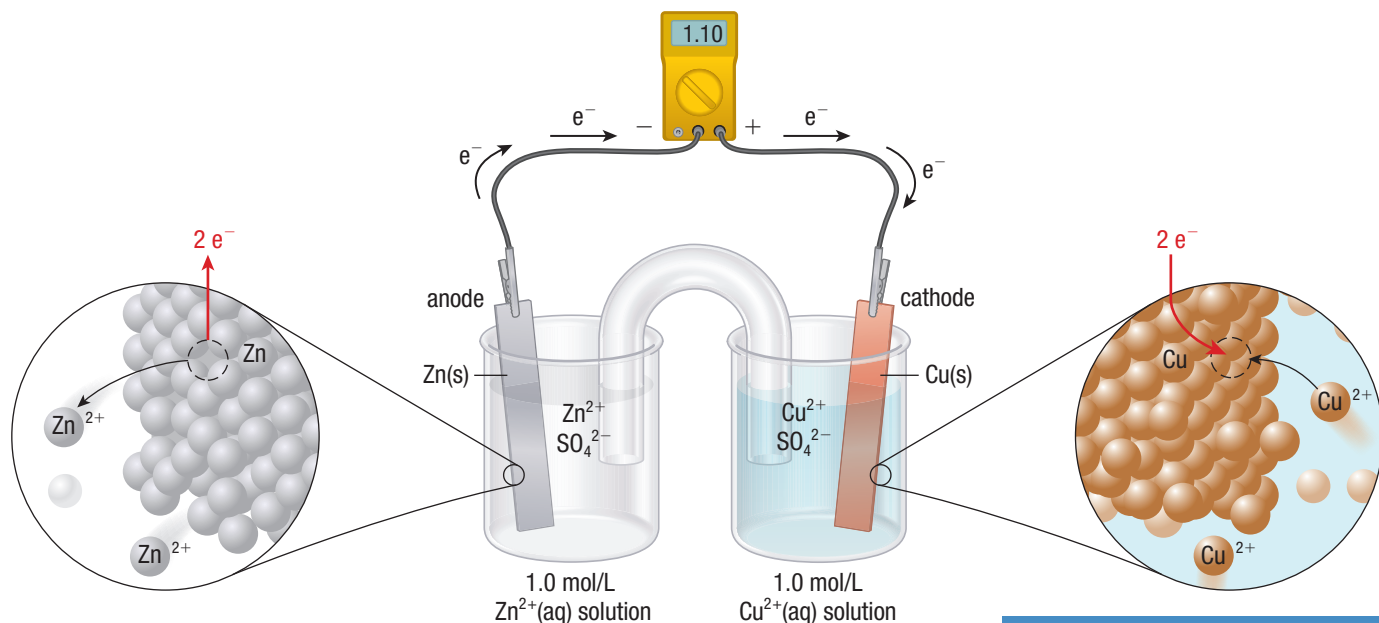
Figure 6 In a galvanic cell that involves the half-reactions $Zn(s) \rightarrow Zn^{2+}(aq) + 2\ e^-$ (at the anode) and $Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$ (at the cathode), $E°_{r\ (cell)} = 1.10$ V.

$\Delta E°_{r\ (cell)}$ and Spontaneity

The sign and magnitude of a predicted cell potential provide important information about the operation of a galvanic cell.

SPONTANEOUS REACTION, $\Delta E°_{r\ (cell)} > 0$

A positive value for the cell potential means that the cell reaction occurs spontaneously. The magnitude of the cell potential, however, does not indicate how quickly the reaction occurs. This can only be determined experimentally. The largest potential is observed when the contents of the cell are at standard conditions. As the cell operates, the observed potential gradually decreases. This occurs because the chemicals within the cell are being used up and are no longer at standard concentrations.

REACTION AT EQUILIBRIUM, $\Delta E°_{r\ (cell)} = 0$

Eventually, the cell potential for an operating cell decreases to zero. At this point, the forward and reverse reactions are occurring at the same rate. In other words, the reaction inside the cell has reached chemical equilibrium. To continue functioning, the cell must be recharged. If this is not possible, the cell can no longer be used and must be discarded.

NON-SPONTANEOUS REACTION, $\Delta E°_{r\ (cell)} < 0$

A negative value for the cell potential means that the cell reaction does not occur spontaneously. As you will learn in Section 10.7, however, the cell reaction can be made to occur by applying energy from an external source.

A Water Analogy for a Galvanic Cell

A water analogy may help explain changes in cell potential as a galvanic cell operates. Consider the water in a lock (**Figure 7(a)**). Each millilitre of surface water on the right side of the gate is higher, and therefore has more gravitational potential energy, than each millilitre of surface water on the left side. When the small sluice gates are opened, potential energy is converted to kinetic energy as water flows to the lower side. As the water level drops (**Figure 7(b)**), the difference between the potential energies of the two sides decreases. The flow of water stops once the potential energies of the two sides are equal (**Figure 7(c)**).

Similarly, the anode is the "high side" of a galvanic cell, while the cathode is the "low side." As the cell operates, electrons flow spontaneously "downhill" from anode to cathode. The cell potential is the measure of the electric potential energy difference (or potential difference) across the two half-cells. Gradually, the cell potential

(a)

(b)

(c)
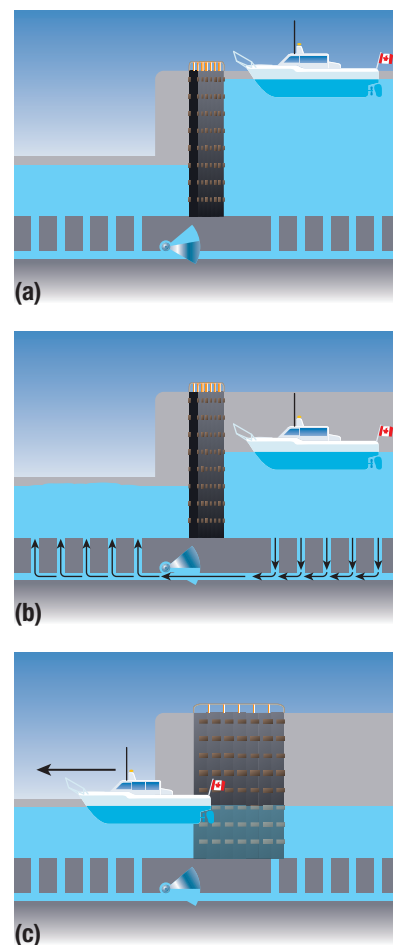
Figure 7 (a) The water behind the main gates in a lock has a certain potential energy, ΔE, relative to the bottom of the canal. (b) When the sluice gates at the bottom are opened, water flows to the lower level on the other side. The potential energy of the stationary water is converted to kinetic energy of flowing water. The flowing water is analogous to electron flow. (c) Water stops flowing when the two levels are equal.

decreases until it reaches zero. At this point, the electron flow has stopped and the cell is considered "dead" because its cell reaction has reached chemical equilibrium.

Predicting Standard Cell Potentials

Table 1 below and Table 1 in Appendix B7 both list the standard reduction potentials of many half-cells, measured relative to the standard hydrogen half-cell. By convention, we always write the equations for these half-cell reactions from left to right as reduction reactions. The reduction potentials can be used to predict the cell potential for a galvanic cell that involves any pair of these half-cell reactions.

Table 1 Standard Reduction Potentials at 25 °C (298 K) for Many Common Half-Reactions

| Half-reaction	$E°_r$ (V)	Half-reaction	$E°_r$ (V)
Strongest Oxidizing Agent; Weakest Reducing Agent			
$F_2(g) + 2\ e^- \rightleftharpoons 2\ F^-(aq)$	2.87	$Cu^+(aq) + e^- \rightleftharpoons Cu(s)$	0.52
$Ag^{2+}(aq) + e^- \rightleftharpoons Ag^+(aq)$	1.99	$O_2(g) + 2\ H_2O(l) + 4\ e^- \rightleftharpoons 4\ OH^-(aq)$	0.40
$Co^{3+}(aq) + e^- \rightleftharpoons Co^{2+}(aq)$	1.82	$Cu^{2+}(aq) + 2\ e^- \rightleftharpoons Cu(s)$	0.34
$H_2O_2(aq) + 2\ H^+(aq) + 2\ e^- \rightleftharpoons 2\ H_2O(l)$	1.78	$Hg_2Cl_2(aq) + 2\ e^- \rightleftharpoons 2\ Hg(l) + 2\ Cl^-(aq)$	0.34
$Ce^{4+}(aq) + e^- \rightleftharpoons Ce^{3+}(aq)$	1.70	$AgCl(aq) + e^- \rightleftharpoons Ag(s) + Cl^-(aq)$	0.22
$PbO_2(s) + 4\ H^+ + SO_4^{2-}(aq) + 2\ e^- \rightleftharpoons PbSO_4(s) + 2\ H_2O(l)$	1.69	$SO_4^{2-}(aq) + 4\ H^+(aq) + 2\ e^- \rightleftharpoons H_2SO_3(aq) + H_2O(l)$	0.20
$MnO_4^-(aq) + 4\ H^+(aq) + 3\ e^- \rightleftharpoons MnO_2(s) + 2\ H_2O(l)$	1.68	$Cu^{2+}(aq) + e^- \rightleftharpoons Cu^+(aq)$	0.16
$IO_4^-(aq) + 2\ H^+(aq) + 2\ e^- \rightleftharpoons IO_3^-(aq) + H_2O(l)$	1.60	$2\ H^+(aq) + 2\ e^- \rightleftharpoons H_2(g)$	0
$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightleftharpoons Mn^{2+}(aq) + 4\ H_2O(l)$	1.51	$Fe^{3+}(aq) + 3\ e^- \rightleftharpoons Fe(s)$	−0.036
$Au^{3+}(aq) + 3e^- \rightleftharpoons Au(s)$	1.50	$Pb^{2+}(aq) + 2\ e^- \rightleftharpoons Pb(s)$	−0.13
$PbO_2(s) + 4\ H^+(aq) + 2e^- \rightleftharpoons Pb^{2+}(aq) + 2\ H_2O(l)$	1.46	$Sn^{2+}(aq) + 2\ e^- \rightleftharpoons Sn(s)$	−0.14
$Cl_2(g) + 2\ e^- \rightleftharpoons 2\ Cl^-(aq)$	1.36	$Ni^{2+}(aq) + 2\ e^- \rightleftharpoons Ni(s)$	−0.23
$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 6\ e^- \rightleftharpoons 2\ Cr^{3+}(aq) + 7\ H_2O(l)$	1.33	$PbSO_4(aq) + 2\ e^- \rightleftharpoons Pb + SO_4^{2-}(aq)$	−0.35
$O_2(g) + 4\ H^+(aq) + 4\ e^- \rightleftharpoons 2\ H_2O(l)$	1.23	$Cd^{2+}(aq) + 2\ e^- \rightleftharpoons Cd(s)$	−0.40
$MnO_2(s) + 4\ H^+(aq) + 2\ e^- \rightleftharpoons Mn^{2+}(aq) + 2\ H_2O(l)$	1.21	$Fe^{2+}(aq) + 2\ e^- \rightleftharpoons Fe(s)$	−0.44
$IO_3^-(aq) + 6\ H^+(aq) + 5\ e^- \rightleftharpoons \frac{1}{2}\ I_2(s) + 3\ H_2O(l)$	1.20	$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	−0.50
$ClO_4^-(aq) + 2\ H^+(aq) + 2\ e^- \rightleftharpoons ClO_3^-(aq) + H_2O(l)$	1.19	$Cr^{3+}(aq) + 3\ e^- \rightleftharpoons Cr(s)$	−0.73
$Br_2(l) + 2\ e^- \rightleftharpoons 2\ Br^-(aq)$	1.09	$Zn^{2+}(aq) + 2\ e^- \rightleftharpoons Zn(s)$	−0.76
$VO_2^+(aq) + 2\ H^+(aq) + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l)$	1.00	$2\ H_2O(l) + 2\ e^- \rightleftharpoons H_2(g) + 2\ OH^-(aq)$	−0.83
$AuCl_4^-(aq) + 3\ e^- \rightleftharpoons Au(s) + 4\ Cl^-(aq)$	0.99	$Mn^{2+}(aq) + 2\ e^- \rightleftharpoons Mn(s)$	−1.18
$NO_3^-(aq) + 4\ H^+(aq) + 3\ e^- \rightleftharpoons NO(g) + 2\ H_2O(l)$	0.96	$Al^{3+}(aq) + 3\ e^- \rightleftharpoons Al(s)$	−1.66
$ClO_2(g) + e^- \rightleftharpoons ClO_2^-(aq)$	0.954	$H_2(g) + 2\ e^- \rightleftharpoons 2\ H^-(aq)$	−2.23
$2\ Hg^{2+}(aq) + 2\ e^- \rightleftharpoons Hg_2^{2+}(aq)$	0.91	$Mg^{2+}(aq) + 2\ e^- \rightleftharpoons Mg(s)$	−2.37
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	0.80	$La^{3+}(aq) + 3\ e^- \rightleftharpoons La(s)$	−2.37
$Hg_2^{2+}(aq) + 2\ e^- \rightleftharpoons 2\ Hg(l)$	0.80	$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	0.77	$Ca^{2+}(aq) + 2\ e^- \rightleftharpoons Ca(s)$	−2.76
$O_2(g) + 2\ H^+(aq) + 2\ e^- \rightleftharpoons H_2O_2(aq)$	0.68	$Ba^{2+}(aq) + 2\ e^- \rightleftharpoons Ba(s)$	−2.90
$MnO_4^-(aq) + e^- \rightleftharpoons MnO_4^{2-}(aq)$	0.56	$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.92
$I_2(s) + 2e^- \rightleftharpoons 2\ I^-(aq)$	0.54	$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.05
		Weakest Oxidizing Agent; Strongest Reducing Agent	

This tutorial shows how to use a standard reduction potentials table to determine the standard cell potential and the net ionic equation for a cell reaction.

Sample Problem 1: Determining a Standard Cell Potential and Net Ionic Equation

Determine the standard cell potential and the net ionic equation for a redox reaction that involves silver and zinc half-cells.

Given: Look up the two reduction half-cell reactions and their standard reduction potentials in Table 1:

$$Ag^+(aq) + e^- \rightarrow Ag(s) \ E^\circ_r = 0.80 \ V$$

$$Zn^{2+}(aq) + 2 \ e^- \rightarrow Zn(s) \ E^\circ_r = -0.76 \ V$$
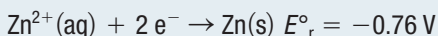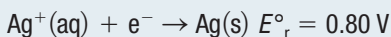
Required: $\Delta E^\circ_{r \ (cell)}$ and the net ionic equation for the cell reaction

Analysis: Use the following equation to calculate the standard cell potential:

$$\Delta E^\circ_{r \ (cell)} = E^\circ_{r \ (cathode)} - E^\circ_{r \ (anode)}$$

Solution:

Step 1. Write the equation for the half-cell reaction that has the more positive reduction potential as a forward reaction. Write the equation for the half-cell reaction that has less positive (or negative) reduction potential as a reverse reaction.

The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of silver ions occurs at the cathode:

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$
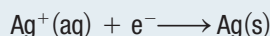
Zinc is oxidized at the anode. Write its equation as an oxidation reaction:

$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$
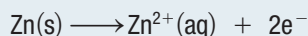
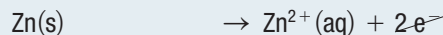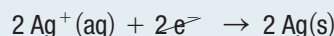**Step 2.** Balance the number of electrons involved in the two half-cell reactions.

The reduction half-reaction requires 1 electron, and the oxidation half-reaction produces 2 electrons. To balance the number of electrons, the silver half-reaction needs to be multiplied by 2:

$$2 \ Ag^+(aq) + 2 \ e^- \rightarrow 2 \ Ag(s)$$
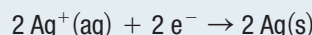
The zinc equation stays the same because it already involves 2 electrons.

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2 \ e^-$$
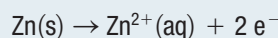
Step 3. Combine the two half-reactions to give the balanced net ionic equation:

$$2 \ Ag^+(aq) + 2 \, \cancel{e^-} \rightarrow 2 \ Ag(s)$$
$$\underline{Zn(s) \rightarrow Zn^{2+}(aq) + 2 \, \cancel{e^-}}$$
$$2 \ Ag^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + 2 \ Ag(s)$$

Step 4. Use the formula $\Delta E^\circ_{r \ (cell)} = E^\circ_{r \ (cathode)} - E^\circ_{r \ (anode)}$ to determine the standard cell potential.

Note that the silver half-cell potential is not multiplied by 2 because the cell potential represents a ratio of energy per coulomb of charge and is therefore unaffected by the number of electrons in the half-cell reaction equation.

$$\Delta E^\circ_{r \ (cell)} = +0.80 \ V - (-0.76 \ V)$$
$$\Delta E^\circ_{r \ (cell)} = +1.56 \ V$$
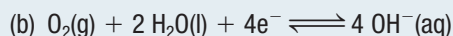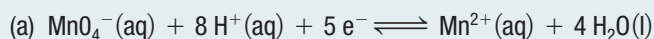
Statement: The net ionic equation and the standard cell potential for a cell involving the zinc and silver half-cells are

$$2 \ Ag^+(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + 2 \ Ag(s)$$
$$\Delta E^\circ_{r \ ((cell)} = +1.56 \ V$$

Practice

1. Galvanic cells that involve the following pairs of half-reactions are constructed. Write the net ionic equation and determine the standard cell potential for each galvanic cell. Assume that all the entities are in their standard states. K/U

   (a) $MnO_4^-(aq) + 8 \ H^+(aq) + 5 \ e^- \rightleftharpoons Mn^{2+}(aq) + 4 \ H_2O(l)$

   $IO_3^-(aq) + 6 \ H^+(aq) + 5 \ e^- \rightleftharpoons \frac{1}{2} \ I_2(s) + 3 \ H_2O(l)$ [ans: 0.31 V]

   (b) $O_2(g) + 2 \ H_2O(l) + 4e^- \rightleftharpoons 4 \ OH^-(aq)$

   $Al^{3+}(aq) + 3 \ e^- \rightleftharpoons Al(s)$ [ans: 2.06 V]

Summary

- All the entities in the cell equation for a standard cell are pure substances under SATP and/or solutions with a concentration of 1.0 mol/L.
- The standard hydrogen half-cell, by definition, has a reduction potential of exactly 0 V. All other potentials are measured relative to this reduction potential.
- The standard cell potential, $\Delta E°_{r\,(cell)}$, can be predicted from the reduction potentials listed in a standard reduction potentials table, using the formula $\Delta E°_{r\,(cell)} = E°_{r\,(cathode)} - E°_{r\,(anode)}$.
- By convention, standard half-cell potentials are written as reduction potentials. This indicates the tendency of a half-cell to undergo reduction.
- The value of the standard cell potential ($\Delta E°_{r\,(cell)}$) is positive if the cell reaction occurs spontaneously.

Questions

1. Why was it necessary to use the standard hydrogen half-cell as a reference? K/U

2. Describe how to predict the spontaneity of the redox reaction in a galvanic cell. K/U

3. Explain why the cell potential for an operating galvanic cell decreases over time but does not change from a positive value to a negative value. T/I A

4. The following half-reactions are combined in a galvanic cell. Write the net ionic equation and determine the standard cell potential for the cell. Assume that all the concentrations are 1.0 mol/L and all the partial pressures are 1.0 atm. K/U

   $Cl_2(g) + 2\,e^- \rightarrow 2\,Cl^-(aq)\ E°_r = 1.36\ V$

   $Br_2(g) + 2\,e^- \rightarrow 2\,Br^-(aq)\ E°_r = 1.09\ V$

5. The following unbalanced equation shows the reactants and products that are involved in a galvanic cell: K/U T/I

   $Fe^{3+}(aq) + Mg(s) \longrightarrow Mg^{2+}(aq) + Fe^{2+}(aq)$

   (a) Write the anode half-reaction equation, the cathode half-reaction equation, and the net ionic equation. Assume that all the concentrations are 1.0 mol/L.

   (b) Calculate the standard cell potential for the overall cell reaction. (Refer to Table 1, page 646)

6. This equation provides information about a reaction that takes place in a galvanic cell:

   $2\,Ag^+(aq) + X(s) \rightarrow 2\,Ag(s) + X^{2+}(aq)\ \Delta E°_r = 1.03\ V$

   Use reduction potentials from Table 1 (page 646) to determine the reduction potential for the half-cell involving X, and to identify X if possible. K/U

7. A student assembled and tested the cell potentials of three different galvanic cells. **Table 2** lists the student's observations. K/U T/I

Table 2 Cell Potential Data

Negative electrode	Positive electrode	$\Delta E°_{r\,(cell)}$(V)
$Pd(s) \mid Pd^{2+}(aq)$	$Cr_2O_7{}^{2-}(aq), H^+(aq) \mid C(s)$	+ 0.28
$Ti(s) \mid Ti^{2+}(aq)$	$Tl^+(aq) \mid Tl(s)$	+1.29
$Tl(s) \mid Tl^+(aq)$	$Pd^{2+}(aq) \mid Pd(s)$	+1.29

   (a) Determine the reduction potential for each half-cell. (You may look up the reduction potential for one half-cell reaction in Table 1, page 646.)

   (b) Using the given evidence, complete a table of relative strengths of oxidizing agents. Justify your prediction.

8. Predict whether each cell will react spontaneously. If it will, write the net ionic equation and determine the standard cell potential for the cell. T/I

   (a) $Mg(s) \mid Mg^{2+}(aq) \parallel Au^{3+}(aq) \mid Au(s)$

   (b) $Cu(s) \mid Cu^+(aq) \parallel Mg^{2+}(aq) \mid Mg(s)$

   (c) $Zn(s) \mid Zn^{2+}(aq) \parallel Sn^{2+}(aq) \mid Sn(s)$