Section 10.2: Standard Reduction Potentials

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1. (a) Given: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:

$$MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \longrightarrow Mn^{2+}(aq) + 4\ H_2O(l) \quad E°_r = +1.51\ V$$

$$IO_3^-(aq) + 6\ H^+(aq) + 5\ e^- \longrightarrow \frac{1}{2}\ I_2(s) + 3\ H_2O(l) \quad E°_r = +1.20\ V$$

Required: $\Delta E°_{r\ (cell)}$ and the net ionic equation for the cell reaction

Analysis: Use the equation $\Delta E°_{r\ (cell)} = E°_{r\ (cathode)} - E°_{r\ (anode)}$ to calculate the standard cell potential.

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

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

Iodine is oxidized at the anode. The equation for this half-cell reaction is written as an oxidization reaction:

$$\frac{1}{2}\ I_2(s) + 3\ H_2O(l) \longrightarrow IO_3^-(aq) + 6\ H^+(aq) + 5\ e^-$$

The number of electrons is equal in both half-reaction equations, so combine the equations for the two half-cell reactions to give the balanced net ionic equation for this reaction:

$$\begin{array}{l} MnO_4^-(aq) + \overset{2}{\cancel{8}}\ H^+(aq) + \cancel{5\ e^-} \rightarrow Mn^{2+}(aq) + 4\ H_2O(l) \\ \qquad\qquad\qquad \frac{1}{2}I_2(s) + 3\ H_2O(l) \rightarrow IO_3^-(aq) + \cancel{6\ H^+(aq)} + \cancel{5\ e^-} \\ \hline MnO_4^-(aq) + \frac{1}{2}I_2(s) + 3\ H_2O(l) + 2\ H^+(aq) \rightarrow Mn^{2+}(aq) + 4\ H_2O(l) + IO_3^-(aq) \end{array}$$

Determine the standard cell potential:

$$\begin{aligned} \Delta E°_{r\ (cell)} &= E°_{r\ (cathode)} - E°_{r\ (anode)} \\ &= +1.51\ V - (+1.20\ V) \\ \Delta E°_{r\ (cell)} &= +0.31\ V \end{aligned}$$

Statement: The net ionic equation and the standard cell potential for a cell involving the permanganate ion and iodine half-cells are

$$MnO_4^-(aq) + \frac{1}{2}\ I_2(s) + 3\ H_2O(l) + 2\ H^+(aq) \rightarrow Mn^{2+}(aq) + 4\ H_2O(l) + IO_3^-(aq)$$

$$\Delta E°_{r\ (cell)} = +0.31\ V$$

(b) Given: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:

$O_2(g) + 2\ H_2O(l) + 4\ e^- \longrightarrow 4\ OH^-(aq) \quad E°_r = +0.40\ V$

$Al^{3+}(aq) + 3\ e^- \longrightarrow Al(s) \quad E°_r = -1.66\ V$

Required: $\Delta E°_{r\ (cell)}$ and the net ionic equation for the cell reaction

Analysis: Use the equation $\Delta E°_{r\ (cell)} = E°_{r\ (cathode)} - E°_{r\ (anode)}$ to calculate the standard cell potential.

Solution: The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of oxygen and water occurs at the cathode:

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

Aluminum is oxidized at the anode. The equation for this half-cell reaction is written as an oxidization reaction:

$Al(s) \longrightarrow Al^{3+}(aq) + 3\ e^-$

The reduction half-reaction requires 4 electrons, and the oxidation half-reaction produces 3 electrons. To balance the number of electrons in the equations for the two half-cell reactions, multiply the equation for the reduction half-reaction by 3 and the equation for the oxidation half-reaction by 4:

$3\ O_2(g) + 6\ H_2O(l) + 12\ e^- \rightarrow 12\ OH^-(aq)$

$4\ Al(s) \rightarrow 4\ Al^{3+}(aq) + 12\ e^-$

Combine the equations for the two half-cell reactions to give the balanced net ionic equation for this reaction:

$$\begin{array}{r} 3\ O_2(g) + 6\ H_2O(l) + \cancel{12\ e^-} \rightarrow 12\ OH^-(aq) \\ 4\ Al(s) \rightarrow 4\ Al^{3+}(aq) + \cancel{12\ e^-} \\ \hline 3\ O_2(g) + 6\ H_2O(l) + 4\ Al(s) \rightarrow 12\ OH^-(aq) + 4\ Al^{3+}(aq) \end{array}$$

Determine the standard cell potential:

$$\begin{aligned} \Delta E°_{r\ (cell)} &= E°_{r\ (cathode)} - E°_{r\ (anode)} \\ &= +0.40\ V - (-1.66\ V) \\ \Delta E°_{r\ (cell)} &= +2.06\ V \end{aligned}$$

Statement: The net ionic equation and the standard cell potential for a cell involving oxygen and water and aluminum half-cells are

$3\ O_2(g) + 6\ H_2O(l) + 4\ Al(s) \rightarrow 12\ OH^-(aq) + 4\ Al^{3+}(aq)$

$\Delta E°_{r\ (cell)} = +2.06\ V$

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1. The standard hydrogen half-cell is used as a reference because it has a reduction potential of exactly 0 V.

2. The spontaneity of the redox reaction in a galvanic cell can be predicted as follows: If the redox reaction in a galvanic cell occurs spontaneously, the value of the standard cell potential ($\Delta E^\circ_{\text{r (cell)}}$) is positive.

3. The cell potential for an operating galvanic cell decreases over time because the electric potential energy difference across the two half-cells decreases until it reaches zero, at chemical equilibrium. The cell potential never changes from a positive value to a negative value because a spontaneous reaction is always spontaneous. A negative value for the cell potential means that the cell reaction is not spontaneous.

4. Given: $Cl_2(g) + 2\ e^- \longrightarrow 2\ Cl^-(aq) \quad E^\circ_r = 1.36\ V$

$Br_2(g) + 2\ e^- \longrightarrow 2\ Br^-(aq) \quad E^\circ_r = 1.09\ V$

Required: $\Delta E^\circ_{\text{r (cell)}}$ and the net ionic equation for the cell reaction

Analysis: Use the equation $\Delta E^\circ_{\text{r (cell)}} = E^\circ_{\text{r (cathode)}} - E^\circ_{\text{r (anode)}}$ to calculate the standard cell potential.

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

$Cl_2(g) + 2\ e^- \longrightarrow 2\ Cl^-(aq)$

Bromine is oxidized at the anode. The equation for this half-cell reaction is written as an oxidization reaction:

$2\ Br^-(aq) \longrightarrow Br_2(g) + 2\ e^-$

The number of electrons is equal in both half-reaction equations, so combine the equations for the two half-cell reactions to give the balanced net ionic equation for this reaction:

$$\begin{array}{r} Cl_2(g) + \cancel{2\ e^-} \rightarrow 2\ Cl^-(aq) \qquad\qquad \\ 2\ Br^-(aq) \rightarrow Br_2(g) + \cancel{2\ e^-} \\ \hline Cl_2(g) + 2\ Br^-(aq) \rightarrow 2\ Cl^-(aq) + Br_2(g) \end{array}$$

Determine the standard cell potential:

$$\begin{aligned} \Delta E^\circ_{\text{r (cell)}} &= E^\circ_{\text{r (cathode)}} - E^\circ_{\text{r (anode)}} \\ &= +1.36\ V - (+1.09\ V) \\ \Delta E^\circ_{\text{r (cell)}} &= +0.27\ V \end{aligned}$$

Statement: The net ionic equation and the standard cell potential for a cell involving oxygen and water and aluminum half-cells are

$Cl_2(g) + 2\ Br^-(aq) \rightarrow 2\ Cl^-(aq) + Br_2(g)$

$\Delta E^\circ_{\text{r (cell)}} = +0.27\ V$

5. (a) Given: unbalanced equation for the reactants and products in a galvanic cell, with all concentrations 1.0 mol/L:

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

Required: anode half-reaction equation, the cathode half-reaction equation, and the net ionic equation

Solution: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:

$$Mg^{2+}(aq) + 2\ e^- \longrightarrow Mg(s) \qquad E°_r = -2.37\ V$$

$$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq) \qquad E°_r = +0.77\ V$$

The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of iron ions occurs at the cathode. Magnesium is oxidized at the anode, so the equation for this half-cell reaction is written as an oxidization reaction.

Anode half-reaction equation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

Cathode half-reaction equation: $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$

To balance the number of electrons in the equations for the two half-cell reactions, multiply the equation representing the cathode half-reaction by 2.

$2\ Fe^{3+}(aq) + 2\ e^- \rightarrow 2\ Fe^{2+}(aq)$

Combine the two half-reaction equations to obtain the net ionic equation for the reaction:

$$\begin{array}{r} Mg(s) \rightarrow Mg^{2+}(aq) + \cancel{2\ e^-} \\ 2\ Fe^{3+}(aq) + \cancel{2\ e^-} \rightarrow 2\ Fe^{2+}(aq) \\ \hline Mg(s) + 2\ Fe^{3+}(aq) \rightarrow Mg^{2+}(aq) + 2\ Fe^{2+}(aq) \end{array}$$

Statement: The anode half-reaction equation is

$Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

The cathode half-reaction equation is

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$

and the net ionic equation is

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

(b) Given: $Mg^{2+}(aq) + 2\ e^- \longrightarrow Mg(s) \qquad E°_r = -2.37\ V$

$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq) \qquad E°_r = +0.77\ V$

Required: standard cell potential for the overall cell reaction

Solution: The standard cell potential is

$$\begin{aligned} \Delta E°_{r\,(cell)} &= E°_{r\,(cathode)} - E°_{r\,(anode)} \\ &= +0.77\ V - (-2.37\ V) \\ \Delta E°_{r\,(cell)} &= +3.14\ V \end{aligned}$$

Statement: The standard cell potential for the overall cell reaction is 3.14V.

6. Given: $2\ Ag^+(aq) + X(s) \rightarrow 2\ Ag(s) + X^{2+}(aq)$ $\quad \Delta E°_{r\ (cell)} = 1.03\ V$

Required: X, and reduction potential for half-cell involving X
Analysis: Determine the oxidation and reduction half-reactions and their reduction potentials using the equation $\Delta E°_{r\ (cell)} = E°_{r\ (cathode)} - E°_{r\ (anode)}$, then determine X from the standard cell potential for X.
Solution: The half-reaction equation for silver is

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$

This half-reaction equation represents reduction, which occurs at the cathode. From Table 1 on page 646, the reduction potential for silver is $E°_{r\ (cathode)} = +0.80\ V$.
The half-reaction equation and reduction potential for X is:

$$X(s) \longrightarrow X^{2+}(aq) + 2\ e^-$$

This half-reaction equation represents oxidation, which occurs at the anode. To calculate the reduction potential for this reaction, use

$$\begin{aligned}
\Delta E°_{r\ (cell)} &= E°_{r\ (cathode)} - E°_{r\ (anode)} \\
E°_{r\ (anode)} &= E°_{r\ (cathode)} - \Delta E°_{r\ (cell)} \\
&= +0.80\ V - (+1.03\ V) \\
E°_{r\ (anode)} &= -0.23\ V
\end{aligned}$$

From Table 1, X is nickel.
Statement: X is nickel, and its reduction potential is $-0.23\ V$.
7. (a) Given: for $Pd(s)\ |\ Pd^{2+}(aq)\ ||\ Cr_2O_7^{2-}(aq), H^+(aq)\ |\ C(s)$, $\Delta E°_{r\ (cell)} = +0.28\ V$; for $Ti(s)\ |\ Ti^{2+}(aq)\ ||\ Tl^+(aq)\ |\ Tl(s)$, $\Delta E°_{r\ (cell)} = +1.29\ V$; and for $Tl(s)\ |\ Tl^+(aq)\ ||\ Pd^{2+}(aq)\ |\ Pd(s)$, $\Delta E°_{r\ (cell)} = +1.29\ V$

Required: reduction potentials for dichromate ion, palladium, thallium, and titanium half-cells
Analysis: Use the equation $\Delta E°_{r\ (cell)} = E°_{r\ (cathode)} - E°_{r\ (anode)}$ to calculate the reduction potentials for the unknown half-cells. Look up the equations for the half-cell reactions and their standard reduction potentials in Table 1 on page 646.
Solution: The equation for the reduction half-reaction and the reduction potential for a dichromate ion–carbon half-cell is

$$Cr_2O_7^{2-}(aq) + 14\ H^+(aq) + 6\ e^- \longrightarrow 2\ Cr^{3+}(aq) + 7\ H_2O(l) \quad E°_r = +1.33\ V$$

The standard cell potential for the cell is +0.28 V (given), so the reduction potential for the palladium half-cell is

$$\begin{aligned}
\Delta E°_{r\ (cell)} &= E°_{r\ (cathode)} - E°_{r\ (anode)} \\
E°_{r\ (anode)} &= E°_{r\ (cathode)} - \Delta E°_{r\ (cell)} \\
&= +1.33\ V - (+0.28\ V) \\
E°_{r\ (anode)} &= +1.05\ V
\end{aligned}$$

To determine the reduction potential for the thallium half-cell, substitute the standard cell potential and the reduction potential for palladium into the equation:

$$\begin{aligned} E^\circ_{\text{r (anode)}} &= E^\circ_{\text{r (cathode)}} - \Delta E^\circ_{\text{r (cell)}} \\ &= +1.05\ \text{V} - (+1.29\ \text{V}) \\ E^\circ_{\text{r (anode)}} &= -0.24\ \text{V} \end{aligned}$$

To determine the reduction potential for the titanium half-cell, substitute the standard cell potential and the reduction potential for thallium into the equation:

$$\begin{aligned} E^\circ_{\text{r (anode)}} &= E^\circ_{\text{r (cathode)}} - \Delta E^\circ_{\text{r (cell)}} \\ &= -0.24\ \text{V} - (+1.29\ \text{V}) \\ E^\circ_{\text{r (anode)}} &= -1.53\ \text{V} \end{aligned}$$

Statement: The reduction potential for the dichromate ion–carbon half-cell is 1.33 V; for the palladium half-cell is 1.05 V; for the thallium half-cell is -0.24 V ; and for the titanium half-cell is -1.53 V .

(b) To compile a table of relative strengths of oxidizing agents, write the reduction half-reaction equation for each cell along with the reduction potential calculated in (a). Arrange the equations in order from highest to lowest reduction potentials.

Reduction half-reaction equation	Reduction potential
$Cr_2O_7^{2-}$(aq) + 14 H^+(aq) + 6 e^- $\rightleftharpoons$ 2 Cr^{3+}(aq) + 7 H_2O(l)	$E^\circ_{\text{r}} = +1.33$ V
Pd^{2+}(aq) + 2 e^- $\rightleftharpoons$ Pd(s)	$E^\circ_{\text{r}} = +1.05$ V
Tl^+(aq) + e^- $\rightleftharpoons$ Tl(s)	$E^\circ_{\text{r}} = -0.24$ V
Ti^{2+}(aq) + 2 e^- $\rightleftharpoons$ Ti(s)	$E^\circ_{\text{r}} = -1.53$ V

8. (a) Given: Mg(s) | Mg^{2+}(aq) || Au^{3+}(aq) | Au(s)

Required: whether the cell will react spontaneously; if yes, the net ionic equation and the standard cell potential

Analysis: Determine whether $\Delta E^\circ_{\text{r}} > 0$. Use the equation $\Delta E^\circ_{\text{r (cell)}} = E^\circ_{\text{r (cathode)}} - E^\circ_{\text{r (anode)}}$ to calculate the standard cell potential.

Solution: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:

$$Mg^{2+}(aq) + 2\ e^- \longrightarrow Mg(s) \qquad E^\circ_{\text{r}} = -2.37\ \text{V}$$

$$Au^{3+}(aq) + 3\ e^- \longrightarrow Au(s) \qquad E^\circ_{\text{r}} = +1.50\ \text{V}$$

The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of gold ions occurs at the cathode. Magnesium is oxidized at the anode.

$$\begin{aligned} \Delta E^\circ_{\text{r (cell)}} &= E^\circ_{\text{r (cathode)}} - E^\circ_{\text{r (anode)}} \\ &= +1.50\ \text{V} - (-2.37\ \text{V}) \\ \Delta E^\circ_{\text{r (cell)}} &= 3.87\ \text{V} \end{aligned}$$

Since $\Delta E^\circ_r > 0$, this cell will react spontaneously.

Rewrite the equation for the oxidation half-reaction:

$Mg(s) \rightarrow Mg^{2+}(aq) + 2\ e^-$

To balance the number of electrons in the equations for the two half-cell reactions, multiply the equation for the oxidation reaction by 3 and the equation for the reduction reaction by 2:

$3\ Mg(s) \rightarrow 3\ Mg^{2+}(aq) + 6\ e^-$

$2\ Au^{3+}(aq) + 6\ e^- \rightarrow 2\ Au(s)$

Combine the equations for the two half-cell reactions to obtain the net ionic equation for the cell:

$$\begin{array}{r} 3\ Mg(s) \rightarrow 3\ Mg^{2+}(aq) + \cancel{6\ e^-} \\ 2\ Au^{3+}(aq) + \cancel{6\ e^-} \rightarrow 2\ Au(s) \\ \hline 3\ Mg(s) + 2\ Au^{3+}(aq) \rightarrow 3\ Mg^{2+}(aq) + 2\ Au(s) \end{array}$$

Statement: The cell will react spontaneously. The net ionic equation is

$3\ Mg(s) + 2\ Au^{3+}(aq) \rightarrow 3\ Mg^{2+}(aq) + 2\ Au(s)$

and the standard cell potential is 3.87 V.

(b) Given: $Cu(s)\ |\ Cu^+(aq)\ ||\ Mg^{2+}(aq)\ |\ Mg(s)$

Required: whether the cell will react spontaneously; if yes, net ionic equation and the standard cell potential

Analysis: Determine whether $\Delta E^\circ_r > 0$. Use the equation $\Delta E^\circ_{r\ (cell)} = E^\circ_{r\ (cathode)} - E^\circ_{r\ (anode)}$ to calculate the standard cell potential.

Solution: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646. The anode goes on the left side in a line notation and the cathode goes on right side. Since copper is on the left in the line equation and magnesium is on the right, the reduction half-reaction equation for the anode will be

$Cu^+(aq) + e^- \longrightarrow Cu(s) \quad E^\circ_r = +0.52\ V$

and the reduction half-reaction for the cathode will be

$Mg^{2+}(aq) + 2\ e^- \longrightarrow Mg(s) \quad E^\circ_r = -2.37\ V$

So, the standard cell potential is

$$\begin{aligned} \Delta E^\circ_{r\ (cell)} &= E^\circ_{r\ (cathode)} - E^\circ_{r\ (anode)} \\ &= -2.37\ V - (+0.52\ V) \\ \Delta E^\circ_{r\ (cell)} &= -1.85\ V \end{aligned}$$

Since $\Delta E^\circ_r < 0$, this cell will not react spontaneously.

Statement: The cell will not react spontaneously.

(c) Given: $Zn(s)\ |\ Zn^{2+}(aq)\ ||\ Sn^{2+}(aq)\ |\ Sn(s)$

Required: whether the cell will react spontaneously; if yes, net ionic equation and the standard cell potential

Analysis: Determine whether $\Delta E°_{r} > 0$. Use the equation $\Delta E°_{r\,(cell)} = E°_{r\,(cathode)} - E°_{r\,(anode)}$ to calculate the standard cell potential.

Solution: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:

$$Zn^{2+}(aq) + 2\ e^- \longrightarrow Zn(s) \qquad E°_{r} = -0.76\ V$$

$$Sn^{2+}(aq) + 2\ e^- \longrightarrow Sn(s) \qquad E°_{r} = -0.14\ V$$

The standard cell potential is:

$$\begin{aligned}\Delta E°_{r\,(cell)} &= E°_{r\,(cathode)} - E°_{r\,(anode)} \\ &= -0.14\ V - (-0.76\ V) \\ \Delta E°_{r\,(cell)} &= 0.62\ V\end{aligned}$$

Since $\Delta E°_{r} > 0$, this cell will react spontaneously.

Combine the equations for the two half-cell reactions to obtain the net ionic equation for the cell:

$$\begin{array}{r} Sn^{2+}(aq) + \cancel{2\ e^-} \rightarrow Sn(s) \\ Zn(s) \rightarrow Zn^{2+}(aq) + \cancel{2\ e^-} \\ \hline Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s) \end{array}$$

Statement: The cell will react spontaneously. The net ionic equation is

$$Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s)$$

and the standard cell potential is 2.89 V.