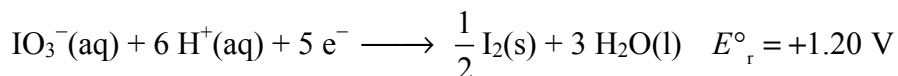
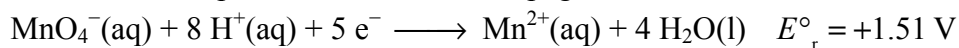


Section 10.2: Standard Reduction Potentials

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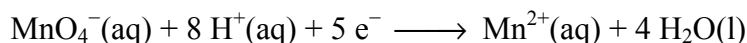
1. (a) **Given:** Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:



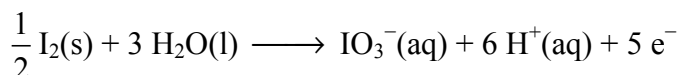
Required: $\Delta E^\circ_{\text{r}(\text{cell})}$ and the net ionic equation for the cell reaction

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

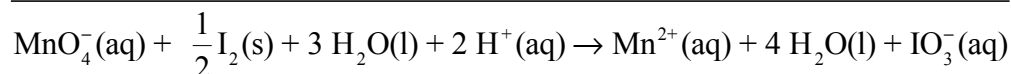
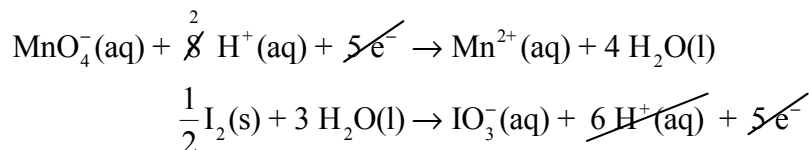
Solution: The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of permanganate ions occurs at the cathode:



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



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:

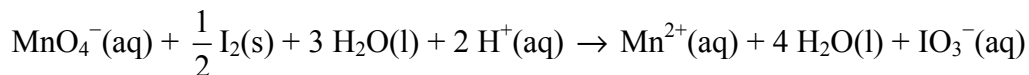


Determine the standard cell potential:

$$\begin{aligned} \Delta E^\circ_{\text{r}(\text{cell})} &= E^\circ_{\text{r}(\text{cathode})} - E^\circ_{\text{r}(\text{anode})} \\ &= +1.51 \text{ V} - (+1.20 \text{ V}) \end{aligned}$$

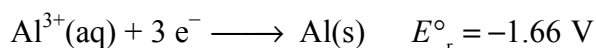
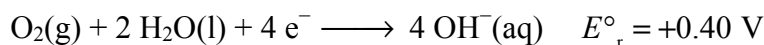
$$\Delta E^\circ_{\text{r}(\text{cell})} = +0.31 \text{ V}$$

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



$$\Delta E^\circ_{\text{r}(\text{cell})} = +0.31 \text{ V}$$

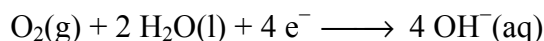
(b) Given: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:



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

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

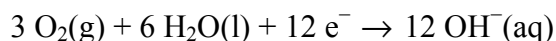
Solution: The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of oxygen and water occurs at the cathode:



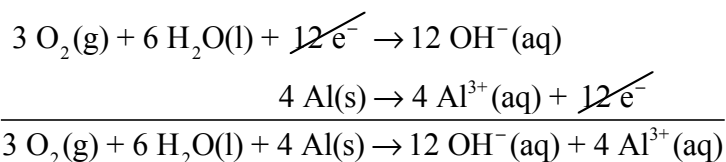
Aluminum is oxidized at the anode. The equation for this half-cell reaction is written as an oxidization reaction:



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:



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

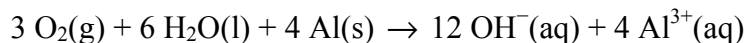


Determine the standard cell potential:

$$\begin{aligned} \Delta E^\circ_{\text{r}(\text{cell})} &= E^\circ_{\text{r}(\text{cathode})} - E^\circ_{\text{r}(\text{anode})} \\ &= +0.40 \text{ V} - (-1.66 \text{ V}) \end{aligned}$$

$$\Delta E^\circ_{\text{r}(\text{cell})} = +2.06 \text{ V}$$

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



$$\Delta E^\circ_{\text{r}(\text{cell})} = +2.06 \text{ V}$$

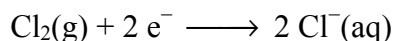
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1. The standard hydrogen half-cell is used as a reference because it has a reduction potential of exactly 0 V.
2. The spontaneity of the redox reaction in a galvanic cell can be predicted as follows: If the redox reaction in a galvanic cell occurs spontaneously, the value of the standard cell potential ($\Delta E^\circ_{r(\text{cell})}$) 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:** $\text{Cl}_2(\text{g}) + 2 \text{e}^- \longrightarrow 2 \text{Cl}^-(\text{aq}) \quad E^\circ_r = 1.36 \text{ V}$
 $\text{Br}_2(\text{g}) + 2 \text{e}^- \longrightarrow 2 \text{Br}^-(\text{aq}) \quad E^\circ_r = 1.09 \text{ V}$

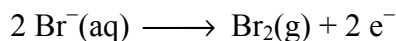
Required: $\Delta E^\circ_{r(\text{cell})}$ and the net ionic equation for the cell reaction

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

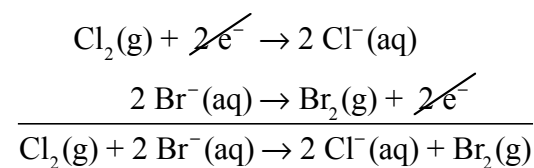
Solution: The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of chlorine occurs at the cathode:



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



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:

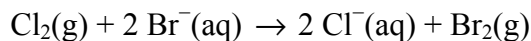


Determine the standard cell potential:

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

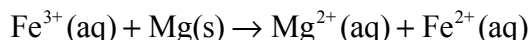
$$\Delta E^\circ_{r(\text{cell})} = +0.27 \text{ V}$$

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



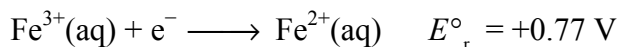
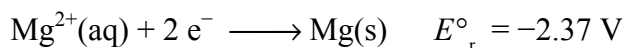
$$\Delta E^\circ_{r(\text{cell})} = +0.27 \text{ V}$$

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

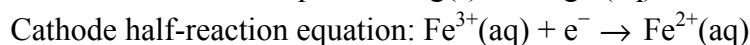
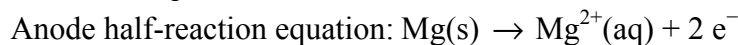


Required: anode half-reaction equation, the cathode half-reaction equation, and the net ionic equation

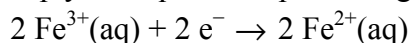
Solution: Look up the equations for the two reduction half-cell reactions and their standard reduction potentials in Table 1 on page 646:



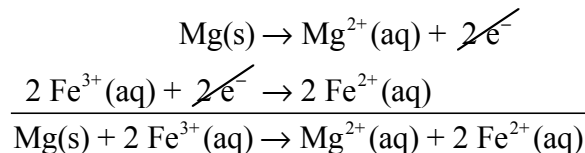
The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of iron ions occurs at the cathode. Magnesium is oxidized at the anode, so the equation for this half-cell reaction is written as an oxidization reaction.



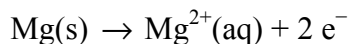
To balance the number of electrons in the equations for the two half-cell reactions, multiply the equation representing the cathode half-reaction by 2.



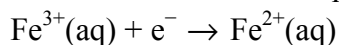
Combine the two half-reaction equations to obtain the net ionic equation for the reaction:



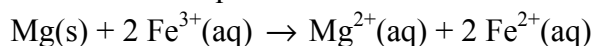
Statement: The anode half-reaction equation is



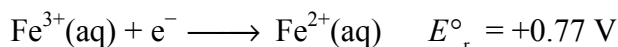
The cathode half-reaction equation is



and the net ionic equation is



(b) Given: $\text{Mg}^{2+}(\text{aq}) + 2 \text{e}^{-} \longrightarrow \text{Mg}(\text{s}) \quad E_{\text{r}}^{\circ} = -2.37 \text{ V}$



Required: standard cell potential for the overall cell reaction

Solution: The standard cell potential is

$$\begin{aligned} \Delta E_{\text{r}(\text{cell})}^{\circ} &= E_{\text{r}(\text{cathode})}^{\circ} - E_{\text{r}(\text{anode})}^{\circ} \\ &= +0.77 \text{ V} - (-2.37 \text{ V}) \end{aligned}$$

$$\Delta E_{\text{r}(\text{cell})}^{\circ} = +3.14 \text{ V}$$

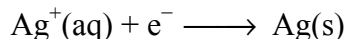
Statement: The standard cell potential for the overall cell reaction is 3.14V.



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 $\Delta E_{\text{r}(\text{cell})}^{\circ} = E_{\text{r}(\text{cathode})}^{\circ} - E_{\text{r}(\text{anode})}^{\circ}$, then determine X from the standard cell potential for X.

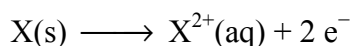
Solution: The half-reaction equation for silver is



This half-reaction equation represents reduction, which occurs at the cathode. From

Table 1 on page 646, the reduction potential for silver is $E_{\text{r}(\text{cathode})}^{\circ} = +0.80 \text{ V}$.

The half-reaction equation and reduction potential for X is:



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

$$\Delta E_{\text{r}(\text{cell})}^{\circ} = E_{\text{r}(\text{cathode})}^{\circ} - E_{\text{r}(\text{anode})}^{\circ}$$

$$\begin{aligned} E_{\text{r}(\text{anode})}^{\circ} &= E_{\text{r}(\text{cathode})}^{\circ} - \Delta E_{\text{r}(\text{cell})}^{\circ} \\ &= +0.80 \text{ V} - (+1.03 \text{ V}) \end{aligned}$$

$$E_{\text{r}(\text{anode})}^{\circ} = -0.23 \text{ V}$$

From Table 1, X is nickel.

Statement: X is nickel, and its reduction potential is -0.23 V .

7. (a) Given: for $\text{Pd}(\text{s}) \mid \text{Pd}^{2+}(\text{aq}) \parallel \text{Cr}_2\text{O}_7^{2-}(\text{aq}), \text{H}^+(\text{aq}) \mid \text{C}(\text{s})$, $\Delta E_{\text{r}(\text{cell})}^{\circ} = +0.28 \text{ V}$; for

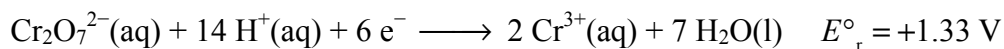
$\text{Ti}(\text{s}) \mid \text{Ti}^{2+}(\text{aq}) \parallel \text{Ti}^+(\text{aq}) \mid \text{Ti}(\text{s})$, $\Delta E_{\text{r}(\text{cell})}^{\circ} = +1.29 \text{ V}$; and for

$\text{Tl}(\text{s}) \mid \text{Ti}^+(\text{aq}) \parallel \text{Pd}^{2+}(\text{aq}) \mid \text{Pd}(\text{s})$, $\Delta E_{\text{r}(\text{cell})}^{\circ} = +1.29 \text{ V}$

Required: reduction potentials for dichromate ion, palladium, thallium, and titanium half-cells

Analysis: Use the equation $\Delta E_{\text{r}(\text{cell})}^{\circ} = E_{\text{r}(\text{cathode})}^{\circ} - E_{\text{r}(\text{anode})}^{\circ}$ 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



The standard cell potential for the cell is $+0.28 \text{ V}$ (given), so the reduction potential for the palladium half-cell is

$$\Delta E_{\text{r}(\text{cell})}^{\circ} = E_{\text{r}(\text{cathode})}^{\circ} - E_{\text{r}(\text{anode})}^{\circ}$$

$$\begin{aligned} E_{\text{r}(\text{anode})}^{\circ} &= E_{\text{r}(\text{cathode})}^{\circ} - \Delta E_{\text{r}(\text{cell})}^{\circ} \\ &= +1.33 \text{ V} - (+0.28 \text{ V}) \end{aligned}$$

$$E_{\text{r}(\text{anode})}^{\circ} = +1.05 \text{ V}$$

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

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

$$E_{r(\text{anode})}^{\circ} = -0.24 \text{ V}$$

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

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

$$E_{r(\text{anode})}^{\circ} = -1.53 \text{ V}$$

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
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14 \text{H}^+(\text{aq}) + 6 \text{e}^- \rightleftharpoons 2 \text{Cr}^{3+}(\text{aq}) + 7 \text{H}_2\text{O}(\text{l})$	$E_r^{\circ} = +1.33 \text{ V}$
$\text{Pd}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Pd}(\text{s})$	$E_r^{\circ} = +1.05 \text{ V}$
$\text{Tl}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Tl}(\text{s})$	$E_r^{\circ} = -0.24 \text{ V}$
$\text{Ti}^{2+}(\text{aq}) + 2 \text{e}^- \rightleftharpoons \text{Ti}(\text{s})$	$E_r^{\circ} = -1.53 \text{ V}$

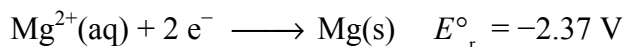
8. (a) Given: $\text{Mg}(\text{s}) \mid \text{Mg}^{2+}(\text{aq}) \parallel \text{Au}^{3+}(\text{aq}) \mid \text{Au}(\text{s})$

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

Analysis: Determine whether $\Delta E_r^{\circ} > 0$. Use the equation

$$\Delta E_{r(\text{cell})}^{\circ} = E_{r(\text{cathode})}^{\circ} - E_{r(\text{anode})}^{\circ} \text{ to calculate the standard cell potential.}$$

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



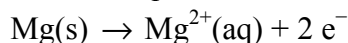
The half-cell reaction with more positive potential is the reduction half-reaction. In this case, the reduction of gold ions occurs at the cathode. Magnesium is oxidized at the anode.

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

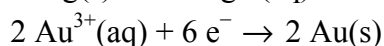
$$\Delta E_{r(\text{cell})}^{\circ} = 3.87 \text{ V}$$

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

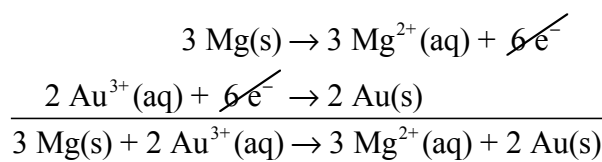
Rewrite the equation for the oxidation half-reaction:



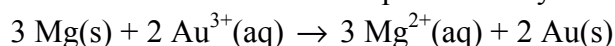
To balance the number of electrons in the equations for the two half-cell reactions, multiply the equation for the oxidation reaction by 3 and the equation for the reduction reaction by 2:



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



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



and the standard cell potential is 3.87 V.

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

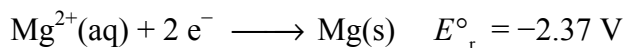
Required: whether the cell will react spontaneously; if yes, net ionic equation and the standard cell potential

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



and the reduction half-reaction for the cathode will be



So, the standard cell potential is

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

$$\Delta E^\circ_{r(\text{cell})} = -1.85 \text{ V}$$

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

Statement: The cell will not react spontaneously.

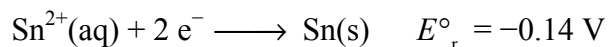
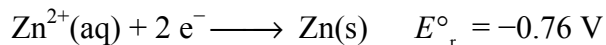
(c) Given: $\text{Zn(s)} | \text{Zn}^{2+}(\text{aq}) || \text{Sn}^{2+}(\text{aq}) | \text{Sn(s)}$

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

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

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



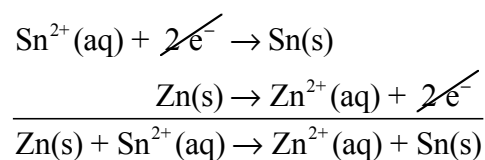
The standard cell potential is:

$$\begin{aligned} \Delta E^\circ_{r(\text{cell})} &= E^\circ_{r(\text{cathode})} - E^\circ_{r(\text{anode})} \\ &= -0.14 \text{ V} - (-0.76 \text{ V}) \end{aligned}$$

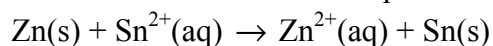
$$\Delta E^\circ_{r(\text{cell})} = 0.62 \text{ V}$$

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

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



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



and the standard cell potential is 2.89 V.