**CHEMISTRY 109 – Help Sheet #41**  
**Electrochemistry (Part II)**

*Do the topics appropriate for your lecture*

Prepared by Dr. Tony Jacob  
https://clc.chem.wisc.edu (Resources page)

**Nuggets:** Stronger Reducing/Oxidizing Agent; \( \Delta G_{\text{rxn}}^0 \), \( K \), and \( E^0 \) cell; Nernst Equation; Concentration Cells; Electrolysis; Aqueous and Molten Systems; BAN APE; Corrosion

---

**STRONGER OXIDIZING /REDUCING AGENT:** use Standard Reduction Table

The more positive the \( E^0 \) \( \rightarrow \) the more spontaneous

---

**Example 1:** a. Using the standard reduction table at the end of this Help Sheet, place the chemicals in order from strongest to weakest oxidizing agent.

\[ \text{Au}^{+3}, \text{I}_2(s), \text{Ag}^+ \]

(stronger oxidizing agent) \( \quad > \quad > \) (weaker oxidizing agent)

b. Using the standard reduction table at the end of this Help Sheet, place the substances in order from strongest to weakest reducing agent.

\[ \text{Pb}^{+2}, \text{Br}^-, \text{Ni}(s) \]

(stronger reducing agent) \( \quad > \quad > \) (weaker reducing agent)

**Answer 1:** a. (stronger oxidizing agent) \( \text{Au}^{+3} > \text{Ag}^+ > \text{I}_2(s) \) (weaker oxidizing agent)

An oxidizing agent is reduced. Using the standard reduction table at the end of this Help Sheet, the reactions should be read in the forward direction (left to right) because in the forward direction they are reduction reactions. The following reduction reactions are for the chemicals listed above:

\[ \text{Au}^{+3} + 3e^- \rightarrow \text{Au}(s) \quad E^0 = 1.52V \]
\[ \text{I}_2(s) + 2e^- \rightarrow 2\text{I}^- \quad E^0 = 0.535V \]
\[ \text{Ag}^+ + e^- \rightarrow \text{Ag}(s) \quad E^0 = 0.799V \]

The *more positive* the potential the more likely the reaction will occur. Hence, \( \text{Au}^{+3} \) has the most positive \( E^0 \) and is the *most likely to be reduced \( \rightarrow \) strongest oxidizing agent*. The next is \( \text{Ag}^+ \) and then \( \text{I}_2(s) \). When working with oxidizing agents, the chemicals must be on the *left side of the half-reaction* in the Standard Reduction Table.

b. (stronger reducing agent) \( \text{Ni}(s) > \text{Br}^- > \text{Pb}^{+2} \) (weaker reducing agent)

A reducing agent is oxidized. Using the standard reduction table at the end of this Help Sheet, the reactions should be read in the reverse direction (right to left) because in the reverse direction they are oxidation reactions. The following oxidation reactions are for the chemicals listed above:

\[ \text{Pb}^{+2} \rightarrow \text{Pb}^{+4} + 2e^- \quad E^0 = -1.69V \]
\[ 2\text{Br}^- \rightarrow \text{Br}_2(l) + 2e^- \quad E^0 = -1.066V \]
\[ \text{Ni} \rightarrow \text{Ni}^{+2} + 2e^- \quad E^0 = 0.25V \]

The *more positive* the potential the more likely the reaction will occur. Hence, \( \text{Ni} \) has the positive \( E^0 \) and is the *most likely to be oxidized \( \rightarrow \) strongest reducing agent*. The next is \( \text{Br}^- \) and then \( \text{Pb}^{+2} \). Note: Another reaction with \( \text{Pb}^{+2} \) is shown in the Standard Reduction Table: \( \text{Pb}^{+2} + 2e^- \rightarrow \text{Pb}(s) \); this reaction can *NOT* be used because \( \text{Pb}^{+2} \) is *reduced* in this reaction and is therefore acting as an oxidizing agent – the question asked for reducing agents! Be careful choosing the correct reaction. When working with reducing agents, the chemicals of interest must be on the *right side of the half-reaction* in the Standard Reduction Table.
\[ K, E^0, \Delta G_{rxn}^0 \]
Electrical work = \( q(E_{cell}) \) where \( q \) = charge and \( E \) = cell potential

The charge on an \( e^- = 1.602 \times 10^{-19} \text{C} \)

for 1mol \( e^-: (1.6022 \times 10^{-19} \text{C/e}) (6.022 \times 10^{23} \text{e}/\text{mol e}^-) = 96,485 \text{C/mol e}^- \)

96,485\text{C/mol e}^- \text{ is called the Faraday constant = } F

\[ q = nF \text{ where } n \text{ = number of mol e}^- \text{ and } F \text{ = Faraday’s constant; going back to the above equation:} \]
Electrical work = \( q(E_{cell}) \Rightarrow \text{Electrical work = } nFE_{cell} \)

Electrical work = \( nFE_{cell} \) and Chemical work = \( -\Delta G \); setting the work equations equal to one another yields \( \Delta G = -nFE_{cell} \) and under standard conditions: \( \Delta G^0 = -nFE^0_{cell} \)

If \( \Delta G < 0 \) \( (-) \) \text{ a spontaneous reaction and this implies } \( E_{cell} > 0 \) \( (+) \) \text{ a spontaneous reaction since } \( \Delta G^0 = -RTlnK \) \text{ we can substitute this into } \( \Delta G^0 = -nFE^0_{cell} \) \text{ to yield } -RTlnK = -nFE^0_{cell} ;

solving for \( E^0 \) yields \( \Rightarrow E^0_{cell} = \frac{RT}{nF} \ln K \) \text{ (} T \text{ can be any value) } \text{ or } \ln K = \frac{nE^0_{cell}}{0.0257} \text{ (at } T = 25.00^0C) \text{ }

\[ \text{Relationship between } K, E^0, \Delta G^0 \]

\[ \Delta G^0 = -nFE^0_{cell} \]
\[ \Delta G^0 = -RTlnK \]
\[ E^0_{cell} = \frac{RT}{nF} \ln K \]

Example 2: An electrochemical cell is designed as follows: \( \text{Zn(s) | Zn}^{2+} \text{(aq) || Pb}^{2+} \text{(aq) | Pb(s)} \)

\( \text{Pb}^{2+} \text{(aq) + 2e}^- \rightarrow \text{Pb(s)} \quad E^0 = -0.13 \text{V} \)

\( \text{Zn}^{2+} \text{(aq) + 2e}^- \rightarrow \text{Zn(s)} \quad E^0 = -0.76 \text{V} \)

If \( T = 25.00^0C \) what is the value of the equilibrium constant \( K \)?

Answer 2: \( 2.0 \times 10^{21} \)

Step 1: Find \( E^0_{cell} \) from \( E^0_{cell} = E^0 \text{ cathode} - E^0 \text{ anode} \) (recall line notation: anode \parallel cathode); \( E^0_{cell} = -0.13 \text{ } (-0.76) = 0.63 \text{V}; \)

Step 2: Find \( n \); since the balanced overall reaction will have 2\( e^- \) canceled, that means \( n = 2; \)

Step 3: At \( T = 25.00^0C \), \( \ln K = \frac{nE^0_{cell}}{0.0257} \); \( \ln K = \frac{(2)(0.63)}{0.0257} = 49.03 \); \( e^{\ln K} = e^{49.03} \); \( K = e^{49.03} = 1.96 \times 10^{21} \)
NERNST EQUATION: determine $E_{\text{cell}}$ under non-standard conditions (e.g., at 0.1M instead of 1M)

**Deriving the Nernst Equation:**

$$\Delta G = \Delta G^\circ + RT \ln Q$$  substituting $\Delta G = -nF E_{\text{cell}}$ and $\Delta G^\circ = -nF E_{\text{cell}}^\circ$ yields ⇒

$$-nF E_{\text{cell}} = -nF E_{\text{cell}}^\circ + RT \ln Q;$$  divide by $-nF$ and solve for $E_{\text{cell}}$ yields ⇒

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{RT}{nF} \ln Q \quad \text{(Nernst Equation)}$$

*use this equation if there are non-standard concentrations* or if $T \neq 25^\circ C$ (e.g., $[\ ] \neq 1M$ or $T = 50^\circ C$);

$T$ in K; $R = 8.314J/molK; n = \# \text{mol e}^-$ transferred in balanced reaction; $F =$ Faraday’s constant $= 96,485C/mol e^-$;

$Q =$ reaction quotient using M or atm

*if $T = 25.0^\circ C$ (298K) the Nernst equation can be simplified:*

$$\frac{RT}{F} = \left(\frac{8.314J}{molK}\right)\left(\frac{298.15K}{mol}\right) = 0.0257V \Rightarrow$$

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0257}{n} \ln Q \quad \text{use this with non-standard concentrations (} [\ ] \neq 1M \text{ but } T = 25.0^\circ C)$$

**Example 3:** Calculate $E_{\text{cell}}$ of a cell with the following reaction with $T = 25^\circ C$, $[\text{ClO}_4^-] = 0.010 \text{ M}$, $[\text{ClO}_3^-] = 0.010 \text{ M}$, $[\Gamma^-] = 0.25 \text{ M}$, pH = 3.00.

$$2\Gamma^- \text{(aq)} + \text{ClO}_4^- \text{(aq)} + 2\text{H}^+ \text{(aq)} \rightarrow \text{I}_2(\text{s}) + \text{ClO}_3^- \text{(aq)} + \text{H}_2\text{O(l)}$$

**Answer 3:** $E_{\text{cell}} = 0.45V$

Since the concentrations $\neq 1M$ (non-standard conditions) this will require the Nernst equation:

**Step 1:** Calculate $E_{\text{cell}}^\circ$: Find the 2 half-reactions in the above overall reaction and look up their Standard Reduction Potentials:

$$\text{ClO}_4^- \text{(aq)} + \text{H}^+ \text{(aq)} + 2\text{e}^- \rightarrow \text{ClO}_3^- \text{(aq)} + \text{H}_2\text{O(l)} \quad (E^\circ = 1.201V) \quad \text{and} \quad \text{I}_2(\text{s}) + 2\text{e}^- \rightarrow 2\Gamma^- \text{(aq)} \quad (E^\circ = 0.535V)$$

$$E_{\text{cell}}^\circ = E_{\text{cathode}}^\circ - E_{\text{anode}}^\circ = 1.20 - 0.535 = 0.666V$$

**Step 2:** Convert pH $\rightarrow [\text{H}^+]$; $[\text{H}^+] = 10^{-\text{pH}} = 10^{-3.00} = 0.001M$

**Step 3:** Calculate $E_{\text{cell}}$: At $T = 25.0^\circ C$: $E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0257}{n} \ln Q$ where $Q = \frac{[\text{ClO}_3^-]}{[\Gamma^-][\text{ClO}_4^-][\text{H}^+]^2}$;

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0257}{n} \ln \left(\frac{[\text{ClO}_3^-]}{[\Gamma^-][\text{ClO}_4^-][\text{H}^+]^2}\right);$$

plugging in the values: $E_{\text{cell}} = 0.666 - \frac{0.0257}{2} \ln \left(\frac{0.010}{0.25^2[0.001][0.001]^2}\right)$;

$E_{\text{cell}} = 0.666 - 0.213 = 0.453V$

At equilibrium → the cell is dead and $E_{\text{cell}} = 0$ ($E_{\text{cell}}^\circ \neq 0$ when cell is dead!); $\ln Q$ becomes $\ln K$; re-arranging the Nernst equation:

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0257}{n} \ln Q; \quad 0 = E_{\text{cell}}^\circ - \frac{0.0257}{n} \ln K; \quad \ln K = \frac{nE_{\text{cell}}^\circ}{0.0257} \quad \text{(this is the same equation shown earlier)}$$

**OVER TIME**

At the beginning of the process for a voltaic cell: $E_{\text{cell}} > 0$ (+) and $\Delta G < 0$ (-) and $Q < K$

All of these describe a product-favored spontaneous reaction and the reaction will go towards the products.

At the end of the process: $E_{\text{cell}} = 0$ and $\Delta G = 0$ and $Q = K$

The electrochemical has reached equilibrium and it is dead.

For these changes to occur, over time as the electrochemical cell runs: $E_{\text{cell}} \downarrow$ (from (+) to zero), $\Delta G \uparrow$ (from (−) to zero), and $Q \uparrow$ (from $Q < K$ to $Q = K$)
**CONCENTRATION CELL:** An electrochemical cell in which the half-reactions are the same but the chemical concentrations in each half-cell are different. To solve $E_{cell}$ use the Nernst equation with

$$ Q = \frac{[\text{anode}]}{[\text{cathode}]} = \frac{[M^{+2}]_{\text{dilute}}}{[M^{+2}]_{\text{concentrated}}} $$

for a spontaneous reaction the anode concentration will always be the smaller/more dilute concentration

**Example 4:** Calculate the $E_{cell}$ for the electrochemical cell: $\text{Cu(s)} | \text{Cu}^{+2}(aq) \ (0.0010 \text{M}) \ || \text{Cu}^{+2}(aq) \ (1.0 \text{M}) | \text{Cu(s)}$

$\text{Cu(s)} \rightarrow \text{Cu}^{+2}(aq) + 2e^- \quad E^0 = -0.34 \text{V}$

**Answer 4:** $E_{cell} = 0.089 \text{V}$

Since the concentrations ≠ 1M (non-standard conditions) this will require the Nernst equation:

**Step 1.** Find $E^0_{cell}$:

- $\text{Cu(s)} \rightarrow \text{Cu}^{+2}(aq) + 2e^- \quad E^0 = -0.34 \text{V}$
- $\text{Cu}^{+2}(aq) + 2e^- \rightarrow \text{Cu(s)} \quad E^0 = 0.34 \text{V}$

$E_{cell} = E^0_{cell} - \frac{0.0257}{n} \ln Q; n = 2; \quad E_{cell}^o = 0.00 - \frac{0.0257}{2} \ln \left( \frac{[\text{Cu}^{+2}]_{\text{dilute}}}{[\text{Cu}^{+2}]_{\text{concentrated}}} \right); \quad E_{cell} = 0.00 - \frac{0.0257}{2} \ln \left( \frac{0.0010 \text{M}}{1.0 \text{M}} \right) = 0.0888V$

**ELECTROLYSIS:** $E_{cell} < 0$ (-), reaction is non-spontaneous and is called an electrolytic cell; the reaction is driven by an outside energy source

**Example 5:** An electrochemical cell is designed as follows: $\text{Cu(s)} | \text{Cu}^{+2}(aq) \ (1.0 \text{M}) \ || \text{Pb}^{+2}(aq) \ (1.0 \text{M}) | \text{Pb(s)}$

$\text{Cu}^{+2}(aq) + 2e^- \rightarrow \text{Cu(s)} \quad E^0 = 0.34 \text{V}$

$\text{Pb}^{+2}(aq) + 2e^- \rightarrow \text{Pb(s)} \quad E^0 = -0.13 \text{V}$

If $T = 25.0^\circ \text{C}$ what is the equilibrium constant $K$?

**Answer 5:** $1.3 \times 10^{-16}$

**Step 1:** Find $E^0_{cell}$ from $E^0_{cell} = E^0_{\text{cathode}} - E^0_{\text{anode}}$ (recall line notation: anode \ || cathode); $E^0_{cell} = -0.13 - (0.34) = -0.47V$;

**Step 2:** Find $n$; since the balanced overall reaction will have $2e^-$ canceled, that means $n = 2$;

**Step 3:** At $T = 25.0^\circ \text{C}$, $\ln K = \frac{nE^0_{cell}}{0.0257}; \quad \ln K = \frac{2(-0.47)}{0.0257} = -36.58; \quad e^{\ln K} = e^{-36.58}; \quad K = e^{-36.58} = 1.30 \times 10^{-16}$

**Quantitative Electrolysis Calculations: Plate out (precipitating) a metal**

$$ \text{mol e}^- \ \text{supplied} = \text{(Amps)} \times \text{(time)} \times \left( \frac{1}{F} \right) \ \text{with time in s; F = Faraday’s constant = 96,485 C/mol e}^-; $$

$$ \text{mol e}^- \ \text{required} \ \text{to get a specific mass to plate out} = (\text{mass M}_{\text{plated out}}) \times \left( \frac{1 \text{mol M}}{\text{g M}} \right) \times \left( \frac{\text{mol e}^-}{\text{mol M}} \right) \ \text{(M = metal)} $$

mass $M_{\text{plated out}}$ in grams; $\frac{1 \text{mol M}}{\text{g M}} = \text{molar mass of M}$; and $\left( \frac{\text{mol e}^-}{\text{mol M}} \right)$ determined from the metal oxidation numbers

set the equations equal to one another: $\text{mol e}^-_{\text{supplied}} = \text{mol e}^-_{\text{required}}$

$$ (\text{Amps})(\text{time}) \times \left( \frac{1}{F} \right) = (\text{mass M}_{\text{plated out}}) \times \left( \frac{1 \text{mol M}}{\text{g M}} \right) \times \left( \frac{\text{mol e}^-}{\text{mol M}} \right) $$

Unit analysis: $\left( \frac{C}{s} \right) \times \left( \frac{\text{mol e}^-}{C} \right) = (\text{gM}) \left( \frac{\text{mol M}}{\text{gM}} \right) \left( \frac{\text{mol e}^-}{\text{mol M}} \right)$; note: *time must be in seconds*

**Variation:** To calculate Charge where charge = $C$ use: $C = (\text{Amps})(\text{time})$; rearranging above equation yields:

$$ \text{Charge} = C = (\text{Amps})(\text{time}) = (\text{mass M}_{\text{plated out}}) \times \left( \frac{1 \text{mol M}}{\text{g M}} \right) \times \left( \frac{\text{mol e}^-}{\text{mol M}} \right) \times (F) $$
Example 6: How many grams of Al will be deposited from molten AlCl₃(l) using a current of 15.0 ampere flowing for 24.0 hours?

Answer 6: 121g Al

Step 1: Write reaction: $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al(s)}$  (Note ratio of mol e\(^-\) to mol Al(s); 3mol e\(^-\) : 1mol Al)

Step 2: Use: $(\text{Amp})(\text{time}) \left( \frac{1}{F} \right) = (\text{mass}_{\text{plate}}} (\text{g M}) \left( \frac{1\text{mol M}}{\text{mol e}^-} \right)$

Plug in values: $(15.0 \text{A}) \left( \frac{24.0\text{hrs}}{1\text{hr}} \right) \left( \frac{60\text{min}}{1\text{min}} \right) \left( \frac{1\text{mol e}^-}{96,485\text{C}} \right) = (\text{g Al}) \left( \frac{1\text{mol Al}}{26.98\text{g Al}} \right) \left( \frac{3\text{mol e}^-}{1\text{mol Al}} \right) ; \text{gAl} = 120.80\text{g Al}

Molten systems (multi-component systems)

- **molten** (means liquified); electrolysis of molten NaBr(l) will **yield its elements**:
  - 2Na (cathode-reduction: $2\text{Na}^+ + 2\text{e}^- \rightarrow 2\text{Na}$) and Br\(_2\) (anode-oxidation: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$); *(see example 7 below)*
- **aqueous** NaBr(aq) means dissolved into water from the “aq” with the water also being a possible reactant and potentially being electrolyzed; in this example, in an **aqueous solution** there is the salt (NaBr) and H\(_2\)O(l) *(see example 7 below on how the answer is determined)*; for this aqueous reaction the products end up to be:
  
  $$\text{H}_2(\text{g}) \text{ (cathode-reduction: } 2\text{H}_2\text{O(l)} + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-\text{(aq)} \text{) and } \text{Br}_2 \text{ (anode-oxidation: } 2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^- \text{)}$$

choose the **more positive** $E^0$ **reduction reaction** and the **more positive** $E^0$ **oxidation reaction**

Example 7: a. When molten NaBr(l) is electrolyzed, what chemicals are formed at the anode and cathode?

b. When NaBr(aq) is electrolyzed, what chemicals are formed at the anode and cathode?

**Answer 7: Part a.** An electrolytic reaction of **molten (liquid)** NaBr(l) has 2 possible reactions:

| Cathode (reduction): $2\text{Na}^+ + 2\text{e}^- \rightarrow 2\text{Na} \quad E^0 = -2.71\text{V}$   |
| --- and ---                                                                                           |
| Anode (oxidation): $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^- \quad E^0 = -1.08\text{V}$   |
| (recall: when you electrolyze a molten salt or an aqueous solution the elements are produced)        |

**The products that form: Cathode: Na; Anode: Br\(_2\)**

| Cathode (reduction): $2\text{Na}^+ (\text{aq}) + 2\text{e}^- \rightarrow 2\text{Na} \quad E^0 = -2.71\text{V}$   |
| --- or ---                                                                                           |
| Cathode (reduction): $2\text{H}_2\text{O(l)} + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-\text{(aq)} \quad E^0 = -0.83\text{V}$   |
| --- or ---                                                                                           |
| Anode (oxidation): $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^- \quad E^0 = -1.08\text{V}$   |
| (recall: when you electrolyze a molten salt or an aqueous solution the elements are produced)        |

| The more positive $E^0$ (least “uphill”) reaction at the cathode is the $\text{H}_2\text{O(l)}$ reaction to produce $\text{H}_2(\text{g})$. $E^0$ is not positive so it is not product-favored but it is more product-favored than the $\text{Na}^+$ reaction which has a more negative $E^0$. |
| The more positive $E^0$ (least “uphill”) reaction at the anode is the $2\text{Br}^-$ reaction to produce $\text{Br}_2$. $E^0$ is not positive so it is not product-favored but it is more product-favored than the $\text{H}_2\text{O(l)}$ reaction which has a more negative $E^0$. |

Must choose one reaction from the two reduction reactions and one reaction from the two oxidation reactions. (That is, you cannot choose 2 reduction reactions or 2 oxidation reactions even if those 2 reactions have the most positive $E^0$.)

**The products that form: Cathode: $\text{H}_2$; Anode: $\text{Br}_2$**
TERMINAL SIGN CONVENTIONS: BAN APE

Battery \( (E_{\text{cell}} > 0) \) – BAN: Battery Anode Negative (Anode assigned (-); Cathode assigned (+))

Electrolytic cell \( (E_{\text{cell}} < 0) \) – APE: Anode Positive Electrolysis (Anode assigned (+); Cathode assigned (-))

For a battery: When a voltmeter is attached with the black lead (– post) attached to the anode of the electrochemical cell and the red lead (+ post) of the voltmeter attached to the cathode of the electrochemical cell the volts measured will be positive (+1.10V in this example).

For a battery: When a voltmeter is attached backwards to the prior arrangement with the red lead (+ post) attached to the anode of the electrochemical cell and the black lead (– post) of the voltmeter attached to the cathode of the electrochemical cell the volts measured will be negative (-1.10V in this example).

In this manner one can determine which half-cell of an electrochemical cell is the anode and which is the cathode.

CORROSION (skip if not covered!) – unwanted spontaneous redox reactions of metals usually with \( O_2(g) \) or \( H_2O(l) \)

Corrosion protection – use a sacrificial, alternative, more reactive metal to protect the desired metal from corrosion. For example, Zn is often used with Fe to protect the Fe.

Corrosion (oxidation) of Fe (steel):
\[
\begin{align*}
\text{Fe(s)} & \rightarrow \text{Fe}^{+2}(aq) + 2e^- & E^0 = 0.44V \\
\text{O}_2(g) & + 4\text{H}^+(aq) + 4e^- \rightarrow 2\text{H}_2\text{O}(l) & E^0 = 1.23V
\end{align*}
\]

In subsequent reactions, the \( \text{Fe}^{+2} \) is oxidized further to \( \text{Fe}^{+3} \) and forms rust, \( \text{Fe}_2\text{O}_3(s) \).
\[
2\text{Fe}^{+2}(aq) + \frac{1}{2}\text{O}_2(g) + (2+n)\text{H}_2\text{O}(l) \rightarrow \text{Fe}_2\text{O}_3\cdot n\text{H}_2\text{O}(s) + 4\text{H}^+(aq)
\]

To interrupt this corrosion process Zn is add as a sacrificial anode because of its higher reactivity than Fe:
\[
\text{Zn(s)} \rightarrow \text{Zn}^{+2}(aq) + 2e^- & E^0 = 0.76V
\]

Galvanizing often refers to Fe dipped into molten Zn thereby coating the steel with Zn. A common item found in hardware stores are galvanized nails: The steel nails are coated with Zn. Galvanizing works twofold: It coats the Fe protecting it from direct \( O_2 \) exposure (preventing corrosion) and if the Zn is scratched off the surface, the Zn cathodically protects the Fe as well. For example:

\[
\begin{align*}
anode (\text{Zn(s)}): & \text{Zn(s)} \rightarrow \text{Zn}^{+2}(aq) + 2e^- & E^0 = 0.76V \\
cathode (\text{Fe(s)}): & \text{O}_2(g) + 2\text{H}_2\text{O}(l) + 4e^- \rightarrow 4\text{OH}^-(aq) & E^0 = 0.40V
\end{align*}
\]

\[
2\text{Zn(s)} + \text{O}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{Zn}^{+2}(aq) + 4\text{OH}^-(aq) & E^0 = 1.16V
\]

Note how the Fe is now the cathode electrode; hence, cathodic protection.

Aluminum: Al is interesting in that it is highly reactive with \( O_2(g) \) due to its high potential and yet corrosion resistant under some conditions.
\[
\text{Al(s)} \rightarrow \text{Al}^{+3}(aq) + 3e^- & E^0 = 1.68V
\]

After reacting with \( O_2(g) \), a layer of gray-colored aluminum oxide, \( \text{Al}_2\text{O}_3(s) \), forms. This oxide coating forms a thin, hard layer over the remaining Al(s) thus protecting it from further corrosion (e.g., aluminum soft drink cans never appear “rusted”). In contrast, the rust, \( \text{Fe}_2\text{O}_3(s) \), formed when Fe corrodes is powdery and flaky, and typically falls off thereby exposing new Fe(s) to further corrosion.
**BATTERIES**

- **Primary cells**: not reversible; can't be re-charged; dry cell; alkaline battery; mercury battery; lithium-ion cell
- **Secondary cells**: re-chargeable; storage batteries; lead battery (inexpensive, reliable, simple); NiCad battery
- **Fuel cells**: fuel supplied from an external source; advantages: efficient; clean; H₂/O₂ cell
- **Series**: Placing batteries in series increases the voltage. The lead acid battery found in cars usually have a potential of ~12V yet the electrochemical cell used in these batteries has a voltage of 2.04V. The battery’s higher voltage is achieved by placing 6 electrochemical cells in series as shown below (6 cells x 2V/cell = 12V):

![Diagram of series connection of batteries]

**TYPES OF BATTERIES (skip if not covered!)**

**Primary cells**: Non-re-chargeable; not reversible; examples: dry cell; alkaline; mercury; lithium metal

- **Dry cell (zinc-carbon dry cell)**: disadvantages: if current is withdrawn too rapidly, buildup of H₂(g) and NH₃(g) leads to drop in voltage; Zn and NH₄⁺ react slowly even when not in use so poor shelf life
  
  (E⁰\text{cell-theoretical} = \sim 1.5V)

  \hspace{1cm}
  \begin{align*}
  \text{Anode: } & \text{Zn(s) } \rightarrow \text{Zn}^{2+}(aq) + 2e^- \\
  \text{Cathode: } & 2\text{MnO}_2(s) + 2\text{NH}_4\text{Cl}(aq) + 2e^- \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{NH}_3(aq) + \text{H}_2\text{O}(l) + 2\text{Cl}^-\text{(aq)} \\
  \text{Net reaction: } & 2\text{MnO}_2(s) + 2\text{NH}_4\text{Cl}(aq) + \text{Zn(s) } \rightarrow \text{Mn}_2\text{O}_3(s) + \text{H}_2\text{O}(l) + 2\text{NH}_3(aq) + 2\text{Cl}^-(aq) + \text{Zn}^{2+}(aq)
  \end{align*}

- **Alkaline battery**: advantages: no gases form; longer shelf life (E⁰\text{cell-theoretical} = 1.43V)

  \hspace{1cm}
  \begin{align*}
  \text{Anode: } & \text{Zn(s) } + 2\text{OH}^-\text{(aq) } \rightarrow \text{ZnO(s) } + \text{H}_2\text{O}(l) + 2e^- \\
  \text{Cathode: } & 2\text{MnO}_2(s) + \text{H}_2\text{O}(l) + 2e^- \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{OH}^-\text{(aq)} \\
  \text{Net reaction: } & 2\text{MnO}_2(s) + \text{Zn(s) } \rightarrow \text{Mn}_2\text{O}_3(s) + \text{ZnO(s)}
  \end{align*}

**Secondary cells**: Re-chargeable; storage batteries; examples: lead storage; lithium ion

- **Lead storage battery**: advantages: inexpensive, reliable, simple, rechargeable because the products are insoluble sulfate that remain at electrode surface; disadvantages: heavyweight; lead is toxic

  (E⁰\text{cell-theoretical} = 2.041V)

  \hspace{1cm}
  \begin{align*}
  \text{Anode: } & \text{Pb(s) } + \text{HSO}_4^-\text{2(aq) } \rightarrow \text{PbSO}_4(s) + \text{H}^+(aq) + 2e^- \\
  \text{Cathode: } & \text{PbO}_2(s) + \text{HSO}_4^-\text{2(aq) } + 3\text{H}^+(aq) + 2e^- \rightarrow \text{PbSO}_4(s) + 2\text{H}_2\text{O}(l) \\
  \text{Net reaction: } & \text{PbO}_2(s) + \text{Pb(s) } + 2\text{H}_2\text{SO}_4(aq) \rightarrow 2\text{PbSO}_4(s) + 2\text{H}_2\text{O}(l)
  \end{align*}

- **Lithium-ion battery**: advantages: high voltage; lightweight from low density Li (E⁰\text{cell-theoretical} = 3.7V)

  \hspace{1cm}
  \begin{align*}
  \text{Anode: } & \text{LiCoO}_2(s) \rightarrow \text{Li}_1-x\text{CoO}_2(s) + x\text{Li}^+ + xe^- \\
  \text{Cathode: } & x\text{C}_6(s) + x\text{Li}^+ + xe^- \rightarrow x\text{LiC}_6 \\
  \text{Net reaction: } & \text{LiCoO}_2(s) + x\text{C}_6(s) \rightarrow \text{Li}_1-x\text{CoO}_2(s) + x\text{LiC}_6
  \end{align*}

**Fuel cells**: chemicals are continually supplied from an external source; Advantages: efficient; clean

- **H₂/O₂ cell**: spontaneous if P_H₂ and P_O₂ are sufficiently high; clean (E⁰\text{cell-theoretical} = 1.23V; E⁰\text{cell-actual} = < 1V)

  \hspace{1cm}
  \begin{align*}
  \text{Anode: } & 2\text{H}_2(g) \rightarrow 4\text{H}^+(aq) + 4e^- \\
  \text{Cathode: } & \text{O}_2(g) + 4\text{H}^+(aq) + 4e^- \rightarrow 2\text{H}_2\text{O}(l) \\
  \text{Net reaction: } & 2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l)
  \end{align*}
1. What is the correct ranking when the following chemicals are placed in order from strongest to weakest oxidizing agents? (Note: The standard reduction values listed below are not in order from most positive \( E^0 \) to most negative \( E^0 \).)

<table>
<thead>
<tr>
<th>Chemical</th>
<th>( E^0 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Ag}^+ (aq) )</td>
<td>+0.80V</td>
</tr>
<tr>
<td>( \text{I}_2 (s) )</td>
<td>+0.54V</td>
</tr>
<tr>
<td>( \text{Pb}^{2+} (aq) )</td>
<td>-0.13V</td>
</tr>
<tr>
<td>( \text{Sn}^{2+} (aq) )</td>
<td>-0.14V</td>
</tr>
<tr>
<td>( \text{Sn}^{4+} (aq) )</td>
<td>+0.15V</td>
</tr>
</tbody>
</table>

**strongest oxidizing agent**

**weakest oxidizing agent**

a. \( \text{Ag}^+ (aq) \) > \( \text{I}_2 (s) \) > \( \text{Sn}^{2+} (aq) \) > \( \text{Pb}^{2+} (aq) \)

b. \( \text{Ag}^+ (aq) \) > \( \text{I}_2 (s) \) > \( \text{Sn}^{2+} (aq) \) > \( \text{Pb}^{2+} (aq) \)

c. \( \text{Sn}^{2+} (aq) \) > \( \text{Pb}^{2+} (aq) \) > \( \text{I}_2 (s) \) > \( \text{Ag}^+ (aq) \)

d. \( \text{Pb}^{2+} (aq) \) > \( \text{Sn}^{2+} (aq) \) > \( \text{I}_2 (s) \) > \( \text{Ag}^+ (aq) \)

e. \( \text{Pb}^{2+} (aq) \) > \( \text{Sn}^{4+} (aq) \) > \( \text{I}_2 (s) \) > \( \text{Ag}^+ (aq) \)

2. What is the correct order when the following chemicals are placed in order from strongest to weakest reducing agents?

<table>
<thead>
<tr>
<th>Chemical</th>
<th>( E^0 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Ni}^{2+} (aq) )</td>
<td>-0.28V</td>
</tr>
<tr>
<td>( \text{Ag}^+ (aq) )</td>
<td>+0.80V</td>
</tr>
<tr>
<td>( \text{Fe}^{2+} (aq) )</td>
<td>-0.44V</td>
</tr>
<tr>
<td>( \text{Cl}_2 (g) )</td>
<td>+1.36V</td>
</tr>
<tr>
<td>( \text{Fe}^{3+} (aq) )</td>
<td>+0.77V</td>
</tr>
</tbody>
</table>

a. \( \text{Cl}^- (aq) \) > \( \text{Ag} (s) \) > \( \text{Fe}^{2+} (aq) \) > \( \text{Ni} (s) \)

b. \( \text{Cl}^- (aq) \) > \( \text{Fe}^{2+} (aq) \) > \( \text{Ag} (s) \) > \( \text{Ni} (s) \)

c. \( \text{Ni} (s) \) > \( \text{Ag} (s) \) > \( \text{Fe}^{2+} (aq) \) > \( \text{Cl}^- (aq) \)

d. \( \text{Ni} (s) \) > \( \text{Fe}^{2+} (aq) \) > \( \text{Ag} (s) \) > \( \text{Cl}^- (aq) \)

e. \( \text{Fe}^{2+} (aq) \) > \( \text{Ni} (s) \) > \( \text{Ag} (s) \) > \( \text{Cl}^- (aq) \)

3. Which cell will have the greatest \( E_{\text{cell}}^0 \) voltage at 25.0°C? (Hint: Can be done without doing a calculation.)

<table>
<thead>
<tr>
<th>Chemical</th>
<th>( E^0 )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Zn}^{2+} (aq) ) + 2( e^- ) → ( \text{Zn} (s) )</td>
<td>-0.76V</td>
</tr>
<tr>
<td>( \text{Cu}^{2+} (aq) ) + 2( e^- ) → ( \text{Cu} (s) )</td>
<td>+0.34V</td>
</tr>
</tbody>
</table>

a. \( \text{Zn} (s) | \text{Zn}^{2+} (aq), (0.20M) \| \text{Cu}^{2+} (aq), (1.00M) \| \text{Cu} (s) \)

b. \( \text{Zn} (s) | \text{Zn}^{2+} (aq), (0.40M) \| \text{Cu}^{2+} (aq), (0.80M) \| \text{Cu} (s) \)

c. \( \text{Zn} (s) | \text{Zn}^{2+} (aq), (0.60M) \| \text{Cu}^{2+} (aq), (0.60M) \| \text{Cu} (s) \)

d. \( \text{Zn} (s) | \text{Zn}^{2+} (aq), (0.80M) \| \text{Cu}^{2+} (aq), (0.40M) \| \text{Cu} (s) \)

e. \( \text{Zn} (s) | \text{Zn}^{2+} (aq), (1.00M) \| \text{Cu}^{2+} (aq), (0.20M) \| \text{Cu} (s) \)
4. Calculate the standard free-energy change (ΔG°) for the following reaction. Use the standard reduction potentials shown below.

\[ 3\text{Cu}(s) + 2\text{NO}_3^-(aq) + 8\text{H}^+(aq) \rightarrow 3\text{Cu}^{2+}(aq) + 2\text{NO}(g) + 4\text{H}_2\text{O}(l) \]

\[
\begin{align*}
\text{Cu}^{2+}(aq) + 2e^- & \rightarrow \text{Cu}(s) & E^0 &= 0.34V \\
\text{NO}_3^-(aq) + 4\text{H}^+(aq) + 3e^- & \rightarrow \text{NO}(g) + 2\text{H}_2\text{O}(l) & E^0 &= 0.96V
\end{align*}
\]

5. Given the two reactions under standard conditions, which of the statements is correct?

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

\[
\begin{align*}
\text{Ag}^+(aq) + e^- & \rightarrow \text{Ag}(s) & E^0 &= 0.80V \\
\text{Cu}^{2+}(aq) + 2e^- & \rightarrow \text{Cu}(s) & E^0 &= 0.34V
\end{align*}
\]

a. E°_cell < 0; ΔG° < 0; K < 1
b. E°_cell > 0; ΔG° > 0; K > 1
c. E°_cell < 0; ΔG° > 0; K > 1
d. E°_cell > 0; ΔG° < 0; K < 1
e. E°_cell < 0; ΔG° > 0; K < 1

6. a. Given the chemical reaction and standard electrode potentials below, determine the E°_cell:

\[ 2\text{Ag}(s) + \text{Ni}^{2+}(aq) \rightarrow 2\text{Ag}^{+}(aq) + \text{Ni}(s) \]

\[
\begin{align*}
\text{Ag}^+(aq) + e^- & \rightarrow \text{Ag}(s) & E^0 &= 0.80V \\
\text{Ni}^{2+}(aq) + 2e^- & \rightarrow \text{Ni}(s) & E^0 &= -0.28V
\end{align*}
\]

b. Would this reaction be product or reactant favored? Assume all solutions are 1M.
c. Determine the equilibrium constant K.
d. Determine ΔG°.

7. Calculate the E_{cell} of a cell operating with the following reaction at 25°C, in which [MnO_4^-] = 0.010 M, [Br^-] = 0.010 M, [Mn^{2+}] = 0.15 M, and [H^+] = 1.0 M.

\[ 2\text{MnO}_4^-(aq) + 10\text{Br}^-(aq) + 16\text{H}^+(aq) \rightarrow 2\text{Mn}^{2+}(aq) + 5\text{Br}_2(l) + 8\text{H}_2\text{O}(l) \]

\[
\begin{align*}
\text{MnO}_4^-(aq) + 8\text{H}^+(aq) + 5e^- & \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) & E^0 &= 1.51V \\
\text{Br}_2(l) + 5e^- & \rightarrow 2\text{Br}^-(aq) & E^0 &= 1.08
\end{align*}
\]

8. If the electrochemical cell below has E_{cell} = 0.597V with [Mn^{2+}] = [Cu^{2+}] = 1.00M, what is the pH of the solution at 25°C? (Hint: Use Nernst equation and find [H^+].)

\[
\begin{align*}
\text{MnO}_2(s) + 4\text{H}^+(aq) + 2e^- & \rightarrow \text{Mn}^{2+}(aq) + 2\text{H}_2\text{O}(l) & E^0 &= 1.230V \\
\text{Cu}^{2+}(aq) + 2e^- & \rightarrow \text{Cu}(s) & E^0 &= 0.337V
\end{align*}
\]
9. Answer the questions below given the voltaic electrochemical cell.

\[
\text{Pt(s) | H}^+(\text{aq}) \ (0.010\text{M}) \ | \text{H}_2(\text{g}) \ (2.0\text{atm}) \ | | \text{Cu}^{2+}(\text{aq}) \ (0.10\text{M}) \ | \text{Cu(s)}
\]

I. What is the value of \( E^\circ_{\text{cell}} \)? Use the Standard Reduction Table at the back of this Help Sheet for \( E^\circ \) values.
   a. 0.34V  
   b. -0.34V  
   c. 0.00V  
   d. 0.68V  
   e. Need \( E^\circ \) for Pt reaction

II. Which statement below is correct? As the reaction proceeds
   a. the concentration of Cu\(^{2+}\) will increase.
   b. bubbles will form in one solution.
   c. the pH will drop in one solution.
   d. the Pt electrode will increase in weight.
   e. the Cu electrode will decrease in weight.

III. What is the value of \( E_{\text{cell}} \) for this electrochemical cell at 25.0°C?
   (Hint: Both atm and M can be included in the Q expression)
   a. 0.24V  
   b. 0.34V  
   c. 0.38V  
   d. 0.44V  
   e. 0.56V

IV. What occurs to \( \Delta G \) as the electrochemical reaction occurs?
   a. \( \Delta G \) would start out positive and would increase as the reaction proceeds.
   b. \( \Delta G \) would start out positive and would decrease as the reaction proceeds.
   c. \( \Delta G \) would start out negative and would increase (become less negative) as the reaction proceeds.
   d. \( \Delta G \) would start out negative and would decrease (become more negative) as the reaction proceeds.
   e. \( \Delta G \) would start out as zero and would stay at zero.

10. Given the electrochemical cell
   \[
   \text{Cr(s) | Cr}^{+3}(\text{aq}) \ (1.0 \times 10^{-4}\text{M}) \ | | \text{Cr}^{+3}(\text{aq}) \ (1.5\text{M}) \ | \text{Cr(s)}
   \]

I. What is the value of \( E^\circ_{\text{cell}} \)?
   a. -0.74V  
   b. 0.00V  
   c. 0.74V  
   d. 1.11V  
   e. 2.22V

II. What is the value of \( E_{\text{cell}} \) at 25.0°C
   a. -0.082V  
   b. 0.00V  
   c. 0.082V  
   d. 0.19V  
   e. 0.25V

11. Which of the following statements is incorrect for an electrochemical cell with \( E_{\text{cell}} > 0 \)? If all the statements are correct, choose answer “e”.
   a. \( E^\circ_{\text{cell}} \) can be zero or non-zero.
   b. \( \Delta G \) would be less than zero.
   c. \( Q \) will be greater than \( K \).
   d. A change in temperature would change \( E_{\text{cell}} \).
   e. All statements are correct.

12. I. Molten CaCl\(_2\)(l) is electrolyzed.  a. What is produced at the cathode? Write the reaction.
   b. What is produced at the anode? Write the reaction.
II. A solution of CaCl\(_2\)(aq) is electrolyzed.  a. What is produced at the cathode? Write the reaction.
   b. What is produced at the anode? Write the reaction.
13. I. Molten NaBr(l) is electrolyzed.
   a. What is produced at the cathode? Write the reaction.
      b. What is produced at the anode? Write the reaction.

II. A solution of NaBr(aq) is electrolyzed.
   a. What is produced at the cathode? Write the reaction.
      b. What is produced at the anode? Write the reaction.

14. a. A Cr$^{+3}$ solution is electrolyzed with a 1.50A current. What mass of Cr(s) is plated out after 24.0 hrs?
   b. What current (A) would be required to plate out 2.85 g of Cr from a Cr$^{+3}$ solution in a period of 2.50 hours?

15. An antique automobile bumper is to be chrome plated. The bumper is dipped into an acidic Cr$_2$O$_7^{2-}$(aq) solution and serves as the cathode of an electrolytic cell. If the current is 10.0 Amps **how long will it take** (in seconds) to deposit 1.00 x $10^2$ grams of Cr(s) on the bumper?
   a. 1.12 x $10^{-5}$ s    b. 1.11 x $10^{5}$ s    c. 1.30 x $10^{5}$ s    d. 2.23 x $10^{5}$ s    e. 3.01 x $10^{8}$ s

16. How **many coulombs** (C) are needed to plate out 10.0 grams of copper metals from a Cr(NO$_3$)$_3$(aq) solution?
   a. 0.577C    b. 5.57 x $10^4$C    c. 5.98 x $10^{-6}$C    d. 1.86 x $10^4$C    e. 6.18 x $10^3$C

17. Given the spontaneous reaction, 2NO$_3^-$$(aq) + 8H^+$$(aq) + 3S$^{2-}$(aq) $\rightarrow$ 2NO(g) + 4H$_2$O(l) + 3S(s), when the **acidity of the solution is increased** the electrochemical potential, $E_{\text{cell}}$ will
   a. increase.    b. decrease.    c. remain unchanged.
   d. be favored more kinetically.    e. need more information.

18. The following reaction occurs. Initially, all of the solutions are 1.0M and $T = 25.0^\circ C$.
    \[
    \text{Mn(s) + Sn}^{+2}$(aq) $\rightarrow$ Mn$^{+2}$(aq) + Sn(s) \quad E_{\text{cell}}^{0} = 1.05V
    \]
Which statement is **incorrect**? If all statements are correct, select answer “e”.
   a. Initially, the value of $E_{\text{cell}}$ will equal $E_{\text{cell}}^{0}$.
   b. The value of Q will start out as 1.0, and as the reaction proceeds Q will increase.
   c. The value of $E_{\text{cell}}$ will decrease as the reaction proceeds and $E_{\text{cell}}$ will approach zero.
   d. The value of K will remain constant as the reaction proceeds.
   e. All of the statements are correct.
19. Given the reaction shown below that starts at standard conditions, answer the questions below.

\[ 2\text{Ag}^{+}(aq) + \text{Cu(s)} \rightarrow 2\text{Ag}(s) + \text{Cu}^{2+}(aq) \quad E^{\circ}_{\text{cell}} = 0.46\text{V} \]

I. As the reaction proceeds what occurs to \( E^{\circ}_{\text{cell}} \)?
   a. There is no change in \( E^{\circ}_{\text{cell}} \).
   b. \( E^{\circ}_{\text{cell}} \) increases.
   c. \( E^{\circ}_{\text{cell}} \) decreases.
   d. The sign of \( E^{\circ}_{\text{cell}} \) changes.

II. As the reaction proceeds what occurs to \( E_{\text{cell}} \)?
   a. There is no change in \( E_{\text{cell}} \).
   b. \( E_{\text{cell}} \) increases.
   c. \( E_{\text{cell}} \) decreases.
   d. The sign of \( E_{\text{cell}} \) changes.

III. As the reaction proceeds what occurs to \( K \)?
   a. There is no change in \( K \).
   b. \( K \) increases.
   c. \( K \) decreases.
   d. The sign of \( K \) changes.

IV. As the reaction proceeds what occurs to \( Q \)?
   a. There is no change in \( Q \).
   b. \( Q \) increases.
   c. \( Q \) decreases.
   d. The sign of \( Q \) changes.

20. Use the reaction shown to answer the questions below.

\[ 2\text{I}^{-}(aq) + \text{ClO}_4^{-}(aq) + 2\text{H}^{+}(aq) \rightarrow \text{I}_2(s) + \text{ClO}_3^{-}(aq) + \text{H}_2\text{O}(l) \quad E^{\circ}_{\text{cell}} = 0.67\text{V} \]

I. What happens to \( E_{\text{cell}} \) when the concentration of \( \text{ClO}_3^{-} \) is increased?
   a. There is no change in \( E_{\text{cell}} \).
   b. \( E_{\text{cell}} \) increases.
   c. \( E_{\text{cell}} \) decreases.
   d. The sign of \( E_{\text{cell}} \) changes.
   e. The \( E_{\text{cell}} \) is cut in half.

II. What happens to \( E_{\text{cell}} \) when the pH decreases?
   a. There is no change in \( E_{\text{cell}} \).
   b. \( E_{\text{cell}} \) increases.
   c. \( E_{\text{cell}} \) decreases.
   d. The sign of \( E_{\text{cell}} \) changes.

III. What happens to \( E_{\text{cell}} \) if all the solution concentrations are 1.0M and the temperature (T) is increased?
   a. There is no change in \( E_{\text{cell}} \).
   b. \( E_{\text{cell}} \) increases.
   c. \( E_{\text{cell}} \) decreases.
   d. The sign of \( E_{\text{cell}} \) changes.
   e. The \( E_{\text{cell}} \) is cut in half.
ANSWERS

1. b  \[ \text{[re-write standard reductions in order with most positive potential on top:} \]

\[
\begin{array}{c|c|c|c|c|c}
\text{Reduction Agent} & \text{Reaction} & \text{E}^\circ_{\text{cell}} & \text{Stronger reducing agents} \\
\hline
\text{Ag}^+(aq) + e^- & \text{Ag(s)} & 0.80V & \text{I}_2(s) + 2e^- & \text{2I}^-(aq) & 0.54V \\
\text{Sn}^{4+}(aq) + 2e^- & \text{Sn}^{2+}(aq) & 0.15V & \text{Pb}^{2+}(aq) + 2e^- & \text{Pb(s)} & -0.13V \\
\text{Sn}^{2+}(aq) + 2e^- & \text{Sn(s)} & -0.14V & \text{Ag}^+(aq) + e^- & \text{Ag(s)} & 0.80V \\
\hline
\end{array}
\]

\text{to be an oxidizing agent the chemical must be reduced so look at the chemicals on the left side of the reaction; these are the reactions read left-to-right in which the chemical is reduced; hence, Ag(s) has the most positive potential (+0.80V) and Sn}^{2+}(aq) \text{ has the most negative potential (-0.14V); Ag(s) > I}_2(s) > \text{Pb}^{2+}(aq) > \text{Sn}^{2+}(aq) \to 0.80V > 0.54V > -0.13V > -0.14V; cannot use the Sn}^{4+}(aq) + 2e^- \text{ Sn}^{2+}(aq) \text{ reaction because if read left-to-right then Sn}^{4+} \text{ is being reduced but Sn}^{4+} \text{ was not given as a possible chemical in the question; if the reaction is read right-to-left then Sn}^{2+} \text{ is being oxidized but the question is asking about oxidizing agents so the Sn}^{2+} \text{ must be reduced not oxidized)}

2. d  \[ \text{[re-write standard reductions in order with most positive potential on top:} \]

\[
\begin{array}{c|c|c|c|c|c}
\text{Oxidation Agent} & \text{Reaction} & \text{E}^\circ_{\text{cell}} & \text{Stronger oxidizing agents} \\
\hline
\text{Cl}_2(aq) + 2e^- & \text{2Cl}^-(aq) & 1.36V & \text{Ag}^+(aq) + e^- & \text{Ag(s)} & 0.80V \\
\text{Ag}^+(aq) + e^- & \text{Ag(s)} & 0.80V & \text{Fe}^{3+}(aq) + e^- & \text{Fe}^{2+}(aq) & 0.77V \\
\text{Ni}^{2+}(aq) + 2e^- & \text{Ni(s)} & -0.28V & \text{Fe}^{3+}(aq) + e^- & \text{Fe}^{2+}(aq) & 0.44V \\
\hline
\end{array}
\]

\text{to be a reducing agent the chemical must be oxidized so look at the chemicals on the right side of the reaction; these are the reactions read right-to-left in which the chemical is oxidized; since the reactions are read backwards, the sign on E changes; hence, Ni(s) has the most positive potential (+0.28V) and Cl}^-(aq) \text{ has the most negative potential (-1.36V); Ni(s) > Fe}^{3+}(aq) > \text{Ag(s)} > \text{Cl}^-(aq) \to 0.28V > -0.77V > -0.80V > -1.36V; the reaction of Fe}^{3+}(aq) + 2e^- \text{ Fe(s) if read forwards is the reduction of Fe}^{2+} \text{ rather than the oxidation and is therefore ignored)}

3. a  \[ \text{[since the reactions are not under standard conditions, to calculate the E}^\circ_{\text{cell}} \text{ requires the Nernst equation. In all of the} \]

\text{calculations the E}^\circ_{\text{cell}} \text{ (1.10V) and the n (n = 2) value will be the same, and what differs is the Q value. The overall reaction is:} \]

\[
\text{Cu}^{2+}(aq) + \text{Zn(s)} \to \text{Cu(s)} + \text{Zn}^{2+}(aq) \text{ and from this, Q} = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} \text{. To maximize the E}^\circ_{\text{cell}} \text{ requires that the component,} \\
\]

\[
\left( \frac{-0.0257}{n} \right) \ln Q \text{ to be as large and positive as possible. To be positive, } \ln Q \text{ needs to be negative (-) which when multiplied by} \\
\]

\[
\frac{-0.0257}{n} \text{ will yield a positive value. For } \ln Q \text{ to be negative means } Q < 1, \text{ and the smaller the value of } Q \text{ the larger negative number it becomes. This occurs when } [\text{Zn}^{2+}] = 0.20 \text{ and } [\text{Cu}^{2+}] = 1.00; \ln Q = \ln \left( \frac{0.20}{1.00} \right) = -1.609 ; \text{ answer “a”} \}

4. -359 kJ  \[ \text{[Find E}^\circ_{\text{cell}} \text{: } 3\text{Cu}^{2+}(aq) + 6e^- \to 3\text{Cu(s)}, \text{E}^\circ = 0.34V; 2\text{NO}_3^-(aq) + 8H^+(aq) + 6e^- \to 2\text{NO(g)} + 4\text{H}_2\text{O(l)}, \text{E}^\circ = 0.96V; \text{E}^\circ_{\text{cell}} = 0.96 - (0.34) = 0.62V; \Delta G^\circ = -nFE^\circ_{\text{cell}} = -(6\text{ mol e}^-)(96,500\text{ C/mol e}^-)(0.62\text{ V}) = -358,980\text{ VC;} \]

\[
-358,980\text{VC} \left( \frac{1\text{ J}}{1\text{VC}} \right) \left( \frac{1\text{ kJ}}{1000\text{ J}} \right) = -359\text{kJ} \}
\]

5. e  \[ \text{[E}^\circ_{\text{cell} = E}^\circ_{\text{cathode} - E}^\circ_{\text{anode}}; \text{E}^\circ_{\text{cell} = 0.34 - 0.80 = -0.46V; hence, E}^\circ_{\text{cell} < 0; this means the reaction is not} \]

\text{product-favored and hence, } \Delta G^\circ > 0 (+) \text{ which is non-spontaneous to the product, and K < 1 (reactant-favored)} \]
6. a. -1.1V \{2Ag^{+}(aq) + 2e^{-} \rightarrow 2Ag(s); E^{0} = 0.80V; Ni^{2+}(aq) + 2e^{-} \rightarrow Ni(s); E^{0} = -0.28V; E^{\circ}_{\text{cell}} = 0.28 - (0.80) = -1.08V\}

b. reactant favored \{E^{\circ}_{\text{cell}} < 0\}

c. K = 3.3 \times 10^{-37} \{ \ln K = \frac{-nE^{0}}{0.0257}; \ln K = \frac{-(2)(-1.08)}{0.0257} = -84.05; K = e^{-84.05} = 3.15 \times 10^{-37} \}

d. 208kJ \{\Delta G^{0} = -nF E^{\circ}_{\text{cell}} = -(2\text{mol e}^{-})(96,500\text{C/mol e}^{-})(-1.08V) = 208,440\text{J}(1\text{kJ}/1000\text{J}) = 208.4\text{kJ}\}

7. 0.30V \{\text{MnO}_{4}^{-}(aq) + 8H^{+}(aq) + 5e^{-} \rightarrow \text{Mn}^{2+}(aq) + 4H_{2}O(l); E^{0} = 1.51V; \text{multiply rxn by 2; for: Br}_{2}(l) + 2e^{-} \rightarrow 2Br^{-}(aq); E^{0} = 1.08; \text{multiply by 5; E}^{\circ}_{\text{cell}} = 1.51 - (1.08) = 0.43V; E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln Q; E_{\text{cell}} = 0.43 - 0.00257\ln \left(\frac{[\text{Mn}^{2+}][\text{Br}^{-}]^{2}}{[\text{MnO}_{4}^{-}]^{2}[\text{H}^{+}]^{0}}\right); E_{\text{cell}} = 0.43 - \frac{0.00257}{51.46} = 0.2989\}

8. pH = 2.500 \{\text{Overall rxn: MnO}_{2}(s) + 4H^{+}(aq) + Cu(s) \rightarrow \text{Mn}^{2+}(aq) + 2H_{2}O(l) + Cu^{2+}(aq); E^{\circ}_{\text{cell}} = 1.230 - 0.337 = 0.939V; E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln Q; E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln \left(\frac{[\text{Mn}^{2+}][\text{Cu}^{2+}]}{[\text{H}^{+}]^{4}}\right); 0.597 = 0.893\left(\frac{0.0257}{n}\right)\ln \left(\frac{1}{[\text{H}^{+}]^{4}}\right); -0.296 = -0.01285\ln \left(\frac{1}{[\text{H}^{+}]^{4}}\right); 23.035 = \ln \left(\frac{1}{[\text{H}^{+}]^{4}}\right); e^{23.035} = e^{\ln \left(\frac{1}{[\text{H}^{+}]^{4}}\right)}; 1.0092 \times 10^{10} = \left(\frac{1}{[\text{H}^{+}]^{4}}\right); [\text{H}^{+}]^{4} = 9.9087 \times 10^{-11}; [\text{H}^{+}] = 0.003155; \text{pH} = -\log(0.003155) = 2.500\}

9. I. a {line notation written anode // cathode; E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}; E^{\circ}_{\text{cell}} = 0.34 - (0.00) = 0.34V}

II. c {from the overall reaction: H_{2}(g) + Cu^{2+}(aq) \rightarrow 2H^{+}(aq) + Cu(s):

“a” \rightarrow more Cu^{+2} is dissolving – false; “b” \rightarrow H_{2}(g) is forming – false;

“d” \rightarrow a precipitate forms on the Pt(s) but Pt(s) is an inert electrode and the only things that can form in the SHE are H_{2}(g) and

H^{+}(aq) which will not increase the weight of the Pt; “e” \rightarrow the Cu(s) is dissolving – false}

III. d {E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln Q; E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln \left(\frac{[\text{H}^{+}]^{2}}{[\text{Cu}^{2+}]_{\text{H}_{2}}}\right); E_{\text{cell}} = 0.34 - \left(\frac{0.0257}{2}\right)\ln \left(\frac{(0.01)^{2}}{1(2.0)}\right); E_{\text{cell}} = 0.34 - \left(\frac{0.0257}{2}\right)\ln(0.0005) = 0.44V}

IV. c {from the equation: \Delta G = -nFE_{\text{cell}}, since E_{\text{cell}} is positive then \Delta G would start out negative; when E_{\text{cell}} is positive \rightarrow spontaneous and therefore \Delta G < 0 \rightarrow which is also spontaneous; as the reaction proceeds and starts to approach equilibrium, E_{\text{cell}} decreases (approaches zero) and therefore \Delta G increases (become less negative) as it also approaches zero and equilibrium}

10. I. b \{Cr^{3+}(aq) + 3e^{-} \rightarrow Cr(s), E^{0} = -0.74V and Cr(s) \rightarrow Cr^{3+}(aq) + 3e^{-}, E^{0} = 0.74; overall rxn = Cr^{3+}(1.5) + Cr(s) \rightarrow Cr(s) + Cr^{3+} (1.0 \times 10^{-4}); E^{0}_{\text{cell}} = -0.74 + 0.74 = 0V which is true for all concentration cells\}

II. c {E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln Q; E_{\text{cell}} = E^{\circ}_{\text{cell}} - \left(\frac{0.0257}{n}\right)\ln \left(\frac{[Cr^{3+}]_{\text{dilute}}}{[Cr^{3+}]_{\text{concentrated}}}\right); concentration cell \rightarrow E^{0}_{\text{cell}} = 0; E_{\text{cell}} = 0 - \left(\frac{0.0257}{3}\right)\ln \left(\frac{1.0 \times 10^{-4}}{1.5}\right) = 0.0824V\}
II. 

a. cathode: $H$  

$Ca^{2+} + 2e^- \rightarrow Ca$  

b. anode: $Br$  

$2Cl^-$ $\rightarrow$ $Cl_2 + 2e^-$  

II. 

a. cathode: $H_2(g)$  

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-$  

b. anode: $O_2(g)$  

$2O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

II. 

a. cathode: $H_2(g)$  

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$  

b. anode: $O_2(g)$  

$2O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$

13. 

I. 

a. cathode: $Na$  

$2Na^+ + 2e^- \rightarrow 2Na$  

b. anode: $Br_2$  

$2Br^- \rightarrow Br_2 + 2e^-$

II. 

a. cathode: $H_2(g)$  

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$  

b. anode: $Br_2(l)$  

$2Br^- \rightarrow Br_2 + 2e^-$

14. 

a. $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$  

$(\text{Amps})(\text{time}) \left(\frac{1}{F}\right) = (\text{mass_{metal}}) \left(\frac{1}{\text{molar mass_{metal}}}\right) \left(\frac{\text{mole}^-}{\text{mol metal}}\right)$

$b. 1.76A$  

$(\text{Amps})(\text{time}) \left(\frac{1}{F}\right) = (\text{mass_{metal}}) \left(\frac{1}{\text{molar mass_{metal}}}\right) \left(\frac{\text{mole}^-}{\text{mol metal}}\right)$
16. b \[ Cr^{3+}(aq) + 3e^- \rightarrow Cr(s); \text{ the } Cr^{3+} \text{ came from } Cr(NO_3)_3 \text{ and the oxidation numbers changed: } +3 \rightarrow 0; \]

\[
\text{(Amps)(time)} \frac{1}{F} = \left( \text{mass}_{\text{metal}} \right) \frac{1}{\text{molar mass}_{\text{metal}}} \frac{\text{mole}^-}{\text{molmetal}}; \text{ recall that } A \times \text{time} = (C/s) \times (s) = C \text{ so solve for } A \times \text{time};
\]

\[
\text{(Amps)(time)} = (10.00g Cr) \left( \frac{1\text{molCr}}{52.00g\text{Cr}} \right) \left( \frac{3\text{mole}^-}{1\text{molCr}} \right) \left( \frac{96,485C}{1\text{mole}^-} \right) = 55,664C \}
\]

17. a \{As } H^+ \text{ is added, the } [H^+] \text{ increases which from } Q = \frac{P_{\text{NO}_2}}{[\text{NO}_3^-][H^+]^2[S^{2-}]^3} \text{ will cause } Q \text{ to decrease; as } Q \text{ decreases } \Rightarrow \ln Q \text{ will decrease } \Rightarrow E_{cell} \text{ will increase.}

18. e \{“a”: } Q = \frac{[\text{Mn}^{2+}]}{[\text{Sn}^{2+}]} = \frac{1.0}{1.0} = 1; \ln Q = \ln(1) = 0; \text{ hence, initially, since everything is at standard concentrations,}

\[ E_{cell} = E^0_{cell} - 0; E_{cell} = E^0_{cell}; \]

“b”: \[ Q = \frac{[\text{Mn}^{2+}]}{[\text{Sn}^{2+}]} = \frac{1.0}{1.0} = 1; \ln Q = \ln(1) = 0; \] as the reaction proceeds, [Mn^{2+}] will increase (it’s a product) and [Sn^{2+}] will decrease (it’s a reactant), and hence, \[ Q = \frac{[\text{Mn}^{2+}]}{[\text{Sn}^{2+}]} \] will increase;

“c”: As the reaction proceeds, Q will increase, lnQ will increase, and E_{cell} will decrease. As the reaction proceeds E_{cell} \rightarrow 0 (when E_{cell} = 0 the battery is dead); e.g., after the reaction has run for a while, then let’s assume that [Mn^{2+}] = 1.9M and [Sn^{2+}] = 0.1M, therefore \[ \ln Q = \ln(1.9/0.1) = \ln(19) = 2.94, \text{ and } E_{cell} = 1.05 - \left( \frac{0.0257}{2} \right) (2.94) = 0.99V; \]

“d”: \[ \ln K = \frac{nE^0_{cell}}{0.0257}; \text{ since } E^0_{cell} \text{ is constant } K \text{ will be constant as well} \}

19. I. a \{E^0_{cell} \text{ is constant since it is a value with the concentrations defined as 1.0M; it is independent of what is occurring in the reaction} \}

II. c \{E_{cell} \text{ decreases as the concentrations change; think of it this way: When the battery is “dead” the voltage (E_{cell}) = 0; hence, it starts at E_{cell} = 0.46V and goes to E_{cell} = 0; hence, it must be decreasing} \}

III. a \{K \text{ is constant and only changes with } T} \}

IV. b \{since } E^0_{cell} > 0 \text{ reaction is spontaneous (goes to the products) } \Rightarrow Q < K; \text{ as the reaction proceeds Q approaches } K \rightarrow Q \uparrow; \text{ another way to consider this: } Q = [\text{products}]/[\text{reactants}] \text{ and if the reaction is going to products (spontaneous) then } [\text{products}] \uparrow \text{ and } [\text{reactants}] \downarrow \Rightarrow Q \uparrow \}

20. I. c  
{since the concentrations are changing the system is no longer under standard conditions and you should consider the

Nernst equation, \( E_{\text{cell}} = E^\circ_{\text{cell}} - \left( \frac{0.0257}{n} \right) \ln Q \), to determine how \( E_{\text{cell}} \) is changing. Write Q; \( Q = \frac{[\text{ClO}_3^-]}{[\Gamma]^2[\text{ClO}_4^-][\text{H}^+]^2} \);

if \([\text{ClO}_3^-] \uparrow \rightarrow Q \uparrow \) (since \([\text{ClO}_3^-] \) is in the numerator) and since the equation has \(- \left( \frac{RT}{nF} \right) \ln Q \) with a negative sign in

front \( \rightarrow E_{\text{cell}} \downarrow \)}

II. b  
{since the concentrations are changing the system is no longer under standard conditions and you should consider the

Nernst equation, \( E_{\text{cell}} = E^\circ_{\text{cell}} - \left( \frac{0.0257}{n} \right) \ln Q \), to determine how \( E_{\text{cell}} \) is changing. Write Q; \( Q = \frac{[\text{ClO}_5^-]}{[\Gamma]^2[\text{ClO}_4^-][\text{H}^+]^2} \);

if pH \( \downarrow \rightarrow [\text{H}^+] \uparrow \rightarrow Q \downarrow \) (since \([\text{H}^+] \) is in the denominator) and since the equation has \(- \left( \frac{RT}{nF} \right) \ln Q \) with a negative sign in

front \( \rightarrow E_{\text{cell}} \uparrow \)}

III. a  
{since the concentrations are changing the system is no longer under standard conditions and you should consider the

Nernst equation, \( E_{\text{cell}} = E^\circ_{\text{cell}} - \left( \frac{0.0257}{n} \right) \ln Q \), to determine how \( E_{\text{cell}} \) is changing. Write Q; \( Q = \frac{[\text{ClO}_5^-]}{[\Gamma]^2[\text{ClO}_4^-][\text{H}^+]^2} \) since all

the concentrations \( = 1.0M \rightarrow Q = 1 \rightarrow \ln Q = 0 \) and the term \(- \left( \frac{RT}{nF} \right) \ln Q = 0 \) and \( T \) has no effect on the reaction; note: if the

concentrations \( \neq 1.0M \) then if \( T \uparrow \rightarrow E_{\text{cell}} \downarrow \)}
<table>
<thead>
<tr>
<th>Reaction</th>
<th>E° (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$F_2(g) + 2e^- \rightarrow 2F^-(aq)$</td>
<td>+2.87V</td>
</tr>
<tr>
<td>$Co^{3+}(aq) + e^- \rightarrow Co^{2+}(aq)$</td>
<td>+1.92V</td>
</tr>
<tr>
<td>$Pb^{2+}(aq) + 2e^- \rightarrow Pb^{0}(aq)$</td>
<td>+1.69V</td>
</tr>
<tr>
<td>$PbO_2(s) + HSO_4^-(aq) + 3H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$</td>
<td>-1.099V</td>
</tr>
<tr>
<td>$2BrO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Br_2(l) + 6H_2O(l)$</td>
<td>+1.52V</td>
</tr>
<tr>
<td>$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$</td>
<td>+1.52V</td>
</tr>
<tr>
<td>$MnO_4^2-(g) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$</td>
<td>+1.51V</td>
</tr>
<tr>
<td>$2ClO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow Cl_2(g) + 6H_2O(l)$</td>
<td>+1.47V</td>
</tr>
<tr>
<td>$Cl_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$</td>
<td>+1.36V</td>
</tr>
<tr>
<td>$Cl_3(g) + 2e^- \rightarrow 2Cl^-(aq)$</td>
<td>+1.358V</td>
</tr>
<tr>
<td>$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$</td>
<td>+1.229V</td>
</tr>
<tr>
<td>$ClO_4^-(aq) + 2H^+(aq) + 2e^- \rightarrow ClO_3^-(aq) + H_2O(l)$</td>
<td>+1.201V</td>
</tr>
<tr>
<td>$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$</td>
<td>+1.066V</td>
</tr>
<tr>
<td>$NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$</td>
<td>+0.96V</td>
</tr>
<tr>
<td>$Ag^+(aq) + e^- \rightarrow Ag(s)$</td>
<td>+0.799V</td>
</tr>
<tr>
<td>$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$</td>
<td>+0.77V</td>
</tr>
<tr>
<td>$O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O_2(aq)$</td>
<td>+0.695V</td>
</tr>
<tr>
<td>$I_2(s) + 2e^- \rightarrow 2I^-(aq)$</td>
<td>+0.535V</td>
</tr>
<tr>
<td>$Cu^{2+}(aq) + e^- \rightarrow Cu(s)$</td>
<td>+0.521V</td>
</tr>
<tr>
<td>$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$</td>
<td>+0.40V</td>
</tr>
<tr>
<td>$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$</td>
<td>-0.34V</td>
</tr>
<tr>
<td>$AgCl(s) + e^- \rightarrow Ag(s) + Cl^-(aq)$</td>
<td>+0.222V</td>
</tr>
<tr>
<td>$Cu^{2+}(aq) + e^- \rightarrow Cu^{+}(aq)$</td>
<td>+0.159V</td>
</tr>
<tr>
<td>$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$</td>
<td>+0.15V</td>
</tr>
<tr>
<td>$AgBr(s) + e^- \rightarrow Ag(s) + Br^-(aq)$</td>
<td>+0.071V</td>
</tr>
<tr>
<td>$2H^+(aq) + 2e^- \rightarrow H_2(g)$</td>
<td>0.00V</td>
</tr>
<tr>
<td>$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$</td>
<td>-0.125V</td>
</tr>
<tr>
<td>$Sn^{4+}(aq) + 2e^- \rightarrow Sn(s)$</td>
<td>-0.138V</td>
</tr>
<tr>
<td>$Cu(s) + e^- \rightarrow Cu(s) + \Gamma^{aq}$</td>
<td>-0.185V</td>
</tr>
<tr>
<td>$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$</td>
<td>-0.25V</td>
</tr>
<tr>
<td>$Co^{3+}(aq) + 2e^- \rightarrow Co(s)$</td>
<td>-0.277V</td>
</tr>
<tr>
<td>$PbSO_4(s) + H^+(aq) + 2e^- \rightarrow Pb(s) + HSO_4^-(aq)$</td>
<td>-0.351V</td>
</tr>
<tr>
<td>$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$</td>
<td>-0.40V</td>
</tr>
<tr>
<td>$Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$</td>
<td>-0.424V</td>
</tr>
<tr>
<td>$Fe^{3+}(aq) + 2e^- \rightarrow Fe(s)$</td>
<td>-0.44V</td>
</tr>
<tr>
<td>$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$</td>
<td>-0.74V</td>
</tr>
<tr>
<td>$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$</td>
<td>-0.76V</td>
</tr>
<tr>
<td>$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$</td>
<td>-0.83V</td>
</tr>
<tr>
<td>$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$</td>
<td>-1.68V</td>
</tr>
<tr>
<td>$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s)$</td>
<td>-2.36V</td>
</tr>
<tr>
<td>$Na^+(aq) + e^- \rightarrow Na(s)$</td>
<td>-2.71V</td>
</tr>
<tr>
<td>$K^+(aq) + e^- \rightarrow K(s)$</td>
<td>-2.925V</td>
</tr>
<tr>
<td>$Li^+(aq) + e^- \rightarrow Li(s)$</td>
<td>-3.05V</td>
</tr>
</tbody>
</table>