Electrolytic cells – the use of electrical energy to drive thermodynamically unfavorable (nonspontaneous) oxidation-reduction reactions.
Electrolysis

- **Electrolysis** is the process of using electrical energy to break a compound apart or to reduced an metal ion to an element.

- Electrolysis is done in an electrolytic cell.

- Electrolytic cells can be used to separate elements from their compounds.

- Electrolysis involves forcing electrons into a cell to cause a nonspontaneous reaction (thermodynamically unfavorable) to occur.
Electrochemical Terminology: Voltage

Voltage: The difference in potential energy between the reactants and products

- Unit is the volt (V).
  - 1 V of force = 1 J of energy/coulomb of charge
  - The voltage is needed to drive electrons through an external circuit
  - Amount of force pushing the electrons through the wire is called the **electromotive force, emf**.

Electrons in an electrochemical cell are “driven” from the anode to the cathode by an electromotive force (emf).
A cell or battery provides the electromotive force that pushes electrons out of the negative terminal and pulls electrons into the positive terminal.
Electrons are pulled into the positive terminal of the battery.

Electrons are pushed out of the negative terminal of the battery.

Direction of e⁻ flow

Electrons are pulled into the positive terminal of the battery.
Resistance

Resistance is a property of materials which opposes the flow of electrons (current) through it.

When electrons flow through any material, they collide with each other which gives rise to opposition to the flow of current.

The unit of resistance is the ohm.
Physics Electrical Terms & Concepts

Ohm's Law: \[ V = iR \]

Ohm's Triangle:

Cover the variable you want to find and perform the resulting calculation (Multiplication/Division) as indicated.
Resistance

The diagram illustrates the concept of electrical resistance. On the left, there is a 1Ω resistor with a current of 1A, resulting in a 1 Volt drop. On the right, a 2Ω resistor with a current of 0.5A also results in a 1 Volt drop. This shows how resistance affects current and voltage in electrical circuits.
An electrode surface area is a factor governing the number of electrons that can flow.

- Larger batteries produce larger currents

4.5 amp hour

18.0 amp hour
The D.C. power source forces electrons to the ring.

**Application of Electrolysis: Electroplating silver metal on a silver spoon**

**Electrolytic Cell for Silver Plating**

- **Anode**
  - \( Ag(s) \rightarrow Ag^+(aq) + e^- \)

- **Cathode**
  - \( Ag^+(aq) + e^- \rightarrow Ag(s) \)
Ag/Ag Electrolysis Experiment
How does the current and time influence the amount of substance collected on an electrode?

GOTO the URL and set-up an electrolysis simulation with two silver electrodes.

URL: http://media.pearsoncmg.com/bc/bc_0media_chem/chem_sim/electrolysis_fc1_gm_11-26-12/main.html
Ag/Ag Electrolysis Experiment
How does the current and time influence the amount of substance collected on an electrode?

Design an experiment to answer the research question.

Identify the independent, dependent, and control variables.
How does increasing the time of the applied current influence the amount of metal plated on an electrode, keeping the current constant?

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Ag$^+ + 1$ e$^- \rightarrow$ Ag

how many moles of electrons were involved?

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Ag/Ag Electrolysis Experiment

Make note of the mass of each electrode. Turn the power supply on and go to the microscopic view for each electrode.

URL: http://media.pearsoncmg.com/bc/bc_0media_chem/chem_sim/electrolysis_fc1_gm_11-26-12/main.html
Ag/Ag Electrolysis particle view at the silver electrode connected to the negative terminal of the power supply.

Write the half-reaction that occurs at this electrode.
Ag/Ag Electrolysis particle view at the silver electrode connected to the negative terminal of the power supply.

\[ 1 \text{Ag}^+(\text{aq}) + 1 \text{e}^- \rightarrow 1 \text{Ag(s)} \]
Ag/Ag Electrolysis particle view at the silver electrode connected to the positive terminal of the power supply.

Write the half-reaction that occurs at this electrode.
Ag/Ag Electrolysis particle view at the silver electrode connected to the positive terminal of the power supply.

\[ \text{Ag(s)} \rightarrow \text{Ag}^+(\text{aq}) + 1 \text{ e}^- \]
Reduction occurs at the cathode

**Key Concept:** how many electrons does it take to reduce two Ag$^+$ ions? ten Ag$^+$ ions? One mole of Ag$^+$ ions?

1 e$^-$ + Ag$^+$ $\rightarrow$ Ag
2e$^-$ + 2 Ag$^+$ $\rightarrow$ ? Ag

10 e$^-$ + 10 Ag$^+$ $\rightarrow$ ?

1 mole e$^-$ + 1 mole Ag$^+$ $\rightarrow$ ?

**Application of Electrolysis: Electroplating**

<table>
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<tr>
<th>Electrolytic Cell for Silver Plating</th>
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<tr>
<td>Voltage source</td>
</tr>
<tr>
<td>Silver electrode</td>
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<tr>
<td>Electrolyte</td>
</tr>
<tr>
<td>Ag$^+$ (aq) + e$^-$ $\rightarrow$ Ag(s)</td>
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<tr>
<td>Anode</td>
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<td>Ag(s) $\rightarrow$ Ag$^+$ (aq) + e$^-$</td>
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<td>Cathode</td>
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Object to be plated
Ag/Ag Electrolysis Experiment

1 e⁻ + Ag⁺ → Ag

How many electrons were forced into this system?

0.0326 mole Ag⁺ + 0.0326 mole e⁻ → 0.0326 mole Ag
Ag/Ag Electrolysis Experiment

\[ \text{Ag}^+ + 1 \text{e}^- \rightarrow \text{Ag} \]

Which electrode gained mass? How many grams?
Which electrode lost mass? How many grams?
Ag/Ag Electrolysis Experiment

1 e\(^{-}\) + Ag\(^{+}\) → Ag

How many moles of electrons were forced into this system?
How many moles of electrons were forced into this system?

0.0326 mole Ag$^+$ + 0.0326 mole e$^-$ $\rightarrow$ 0.0326 mole Ag
Given \( \text{Ag}^+ + 1 \text{ e}^- \rightarrow \text{Ag} \), how many moles of electrons were involved?

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Describe the relationship between the run time and mass collected. Linear? Exponential? Inverse?
Electrochemical Terminology: Current

**Current:** The number of electrons that flow by a point in the circuit per second

- Unit for current is the ampere (A).
- 1 Amp of current = 1 coulomb of charge per second
  - 1 Amp = $6.242 \times 10^{18}$ electrons/second
Electrolytic Cells
Model 5: Setting up a Cu/Cu electrolysis experiment

URL:
http://media.pearsoncmg.com/bc/bc_0media_chem/chem_sim/electrolysis_fc1_gm_11-26-12/main.html
Electrolytic Cell

We need to set the amps to 8.00 and the minutes to 15.00.

URL: http://media.pearsoncmg.com/bc/bc_0media_chem/chem_sim/electrolysis_fc1_gm_11-26-12/main.html
Cu/Cu Electrolysis Experiment: particle view at the electrode connected to the negative terminal of the power supply.

Write the half-reaction that occurs at this electrode.
Cu/Cu Electrolysis Experiment: particle view at the electrode connected to the negative terminal of the power supply.

\[ 1 \text{Cu}^{2+}(aq) + 2 \text{e}^{-} \rightarrow 1 \text{Cu}(s) \]
Key Concept: how many electrons does it take to reduce two Cu$^{2+}$ ions? ten Cu$^{2+}$ ions? One mole of Cu$^{2+}$ ions?

\[2 \text{e}^- + \text{Cu}^{2+} \rightarrow \text{Cu}\]
\[? + 2 \text{Cu}^{2+} \rightarrow 2 \text{Cu}\]
\[?^- + 10 \text{Cu}^{2+} \rightarrow 10 \text{Cu}\]
\[? \text{ mole e}^- + 1 \text{ mole Cu}^{2+} \rightarrow \text{Cu}\]
How does increasing the time of the applied current influence the amount of metal plated on an electrode, keeping the current constant?

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Application of Electrolysis: Electroplating

The power source forces electrons to the copper electrode connected to the negative terminal of the power source.

Key Concept: how many electrons does it take to reduce two \( \text{Cu}^{2+} \) ions? ten \( \text{Cu}^{2+} \) ions? One mole of \( \text{Cu}^{2+} \) ions?

\[ 2 \text{e}^- + \text{Cu}^{2+} \rightarrow \text{Cu} \]
\[ 4 \text{e}^- + 2 \text{Cu}^{2+} \rightarrow 2 \text{Cu} \]
\[ 20 \text{e}^- + 10 \text{Cu}^{2+} \rightarrow 10 \text{Cu} \]

\[ 2 \text{mole e}^- + 1 \text{mole Cu}^{2+} \rightarrow 1 \text{mole Cu} \]

\[ \text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu} \quad E^\circ = +0.34 \text{ V} \]
Electrolytic Cell
3.50 amps were applied to this electrolytic cell for 15.00 minutes, how many mole of e- were involved?
Electrolytic Cell
3.50 amps were applied to this electrolytic cell for 15.00 minutes, how many mole of e- were involved?

0.0163 mole Cu²⁺ + 0.0326 mole e⁻ → 0.0163 mole Cu
Given \( \text{Cu}^{2+} + 2 \text{e}^- \rightarrow \text{Cu} \), how many moles of electrons were involved?

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Describe the relationship between the run time and mass collected. Linear? Exponential? Inverse?
Nine fun things to do with a Cu/Cu electrolysis cell:

1. Identify the anode and cathode
2. Write the two half-reactions
3. Write the net cell reaction
4. Identify what is being oxidized and what is being reduced
5. Diagram the electrochemical cell and show what occurs at each electrode
6. Show movement of ions and electrons
7. Which electrode gains mass?
8. Calculate the cell potential
9. Calculate the mass of copper produced.

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \quad E^\circ = +0.34 \text{ V} \]
Complete the diagram for this Cu/Cu electrolysis cell

Identify the anode and cathode.

Write the two half-reactions.

Identify what is being oxidized.

Identify what is being reduced.

Show movement of ions and e-

Which electrode gains mass?

Cu$^{2+}$ + 2e$^-$ -> Cu

$E^\circ = +0.34 \text{ V}$
Electrolysis

D.C. Power Supply

Battery

Anode: \[ \text{Cu} \rightarrow \text{Cu}^{2+} + 2e^- \]

Cathode: \[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \]

1.0M \( \text{Cu}^{2+} \)

Redox: \[ \text{Cu (s)} \rightarrow \text{Cu}^{2+} (aq) + 2e^- \]
Quantitative aspects of electrolysis

During electrolysis, how does the **number of electrons** passed through a circuit **each second** influence the **amount of substance (moles or mass)** that forms?

**Faraday’s law of electrolysis:**
The amount of substance produced at each electrode is directly proportional to the quantity of charge (electrons per second) flowing through the cell.

If we want to plate silver on a ring or spoon electrolysis is used.
Quantitative aspects of electrolysis

We can measure the **current** pushed into the system by the battery and we can measure **time**. How can we determine **mass** of metal deposited on one of the electrodes from this information?

The first thing to do is to make the correct connections to the battery.

If we want to plate silver on a ring or spoon electrolysis is used.
How many electrons were forced into this system?
Information we will need:

- Balanced half reaction, telling us moles of electrons
- Electrical current is measured in **amperes (A)**
  \[
  1 \text{ A} = 1 \text{ C/s}
  \]
  where a coulomb, C, is the SI unit of electrical charge.
- **Faraday constant:** \( F = 96,500 \text{ C/mole e}^- \)
3.53 g of silver is produced in 15.0 minutes by the electrolysis of AgNO\textsubscript{3}(aq) when the electrical current is 3.50 Amps. How many moles of electrons were passed?

Anode:  Ag\(\rightarrow\) Ag\textsuperscript{+}(aq) + 1 e\textsuperscript{-}

Cathode:  Ag\textsuperscript{+}(aq) + 1e\textsuperscript{-} \(\rightarrow\) Ag(s)

Overall:  Ag(s) + Ag\textsuperscript{+}(aq) \(\rightarrow\) Ag(s) + Ag\textsuperscript{+}(aq)

1 Amp sec = 1 C

- **Faraday constant:**  \( F = 96,500 \text{ C/mole e}^{-}\)
3.53 g of silver is produced in 15.0 minutes by the electrolysis of AgNO₃(aq) when the electrical current is 3.50 Amps. How many moles of electrons were passed?

Ag atom is oxidized

Anode: Ag → Ag⁺(aq) + 1 e⁻

Ag⁺ is reduced

Cathode: Ag⁺(aq) + 1e⁻ → Ag(s)

Overall: Ag(s) + Ag⁺(aq) → Ag(s) + Ag⁺(aq)

1) Calculate charge (Coulombs):

\[ \text{Charge} = (3.50 \text{ Amps}) (15.00 \text{ min}) \left( \frac{60.0 \text{ s}}{1 \text{ min}} \right) \left( \frac{1 \text{ C}}{1 \text{ Amp \cdot s}} \right) = 3.15 \times 10^3 \text{ C} \]

2) Calculate moles of electrons that pass into the cell:

\[ 3.15 \times 10^3 \text{ C} \left( \frac{1 \text{ mole}^-}{96,500 \text{ C}} \right) = 0.0326 \text{ mole}^- \]
What mass of silver is produced in 15.0 minutes by the electrolysis of AgNO₃(aq) if the electrical current is 3.50 Amps?

3) Relate electrons to quantity of Ag being formed using half reactions and stoichiometry

\[
\text{mass Ag} = \left( 0.0326 \text{ mol } e^- \right) \left( \frac{1 \text{ mol Ag}}{1 \text{ mol } e^-} \right) \left( \frac{107.9 \text{ g}}{1 \text{ mol Ag}} \right) = 3.52 \text{ g Ag}
\]

\[
\text{Ag}^+(aq) + 1e^- \rightarrow \text{Ag(s)}
\]
Application of Electrolysis: Electroplating

The anode is made of the plate metal (ions in solution). At the anode, Ag atoms are oxidized to Ag\(^+\) ions (oxidation). The Ag\(^+\) ions replace the Ag\(^+\) ions in the solution that are coating the metal electrode.

Ag\(^+\) cations are reduced at the cathode and plate (coat) to the surface of the metal.

Notice: anode and cathode are in the same breaker, they aren’t in separate beakers.
If a reaction occurs, complete and balance the chemical equation

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

\[ \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) \quad E^\circ = +0.34 \text{ V} \]
\[ \text{Zn}^{2+}(aq) + 2e^- \rightarrow \text{Zn}(s) \quad E^\circ = -0.76 \text{ V} \]
\[ \text{Zn}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s) \]
\[ \text{Cu}(s) + \text{Zn}^{2+}(aq) \rightarrow \text{No reaction} \]

\[ \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) \quad E^\circ = +0.34 \text{ V} \]
\[ \text{Zn}^{2+}(aq) + 2e^- \rightarrow \text{Zn}(s) \quad E^\circ = -0.76 \text{ V} \]
Voltaic versus Electrolytic Cells

Voltaic Cell

Volmeter

SALT BRIDGE

Zn(s) → Zn^{2+}(aq) + 2 e^-

Cu^{2+}(aq) + 2 e^- → Cu(s)

Electrolytic Cell

Voltage Source > 1.10 V

CATHODE

Zn^{2+}(aq) + 2 e^- → Zn(s)

Anode

Cu^{2+}(aq) + 2 e^- → Cu(s)
Quantitative aspects of electrolysis

We would like to be able to relate the quantity of reactant or product to the amount of electricity consumed, or the amount of time it takes for the electrolysis.

Faraday’s law of electrolysis:

- The amount of substance produced at each electrode is directly proportional to the quantity of charge (amps x sec or moles of electrons) flowing through the cell.

We can measure current and time. How can we determine moles of electrons passed and mass from this information?
Diagram of a Cu/Cu electrolysis cell

\[ \text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\, \text{e}^- \]

\[ \text{Cu}^{2+}(\text{aq}) + 2\, \text{e}^- \rightarrow \text{Cu(s)} \]
POGIL Activity 86  Electrolytic Cells
8.00 amps at 15.00 minute, how many mole of e-?
What mass of copper is produced in 15.0 minutes by the electrolysis of CuSO$_4$(aq) if the electrical current is 8.00 Amps?

Cu atom is oxidized: Anode: Cu(s) $\rightarrow$ Cu$^{2+}$(aq) + 2 e$^-$

Cu$^{2+}$ is reduced: Cathode: Cu$^{2+}$(aq) + 2e$^-$ $\rightarrow$ Cu(s)

Overall: Cu(s) + Cu$^{2+}$(aq) $\rightarrow$ Cu(s) + Cu$^{2+}$(aq)

1) Calculate charge (Coulombs):

$$C = (8.00 \text{ Amps})(15.00 \text{ min}) \left( \frac{60.0 \text{ s}}{1 \text{ min}} \right) \left( \frac{1 \text{ C}}{1 \text{ Amp \cdot s}} \right) = 7.20 \times 10^3 \text{ C}$$

2) Calculate moles of electrons that pass into the cell:

$$7.20 \times 10^3 \text{ C} \left( \frac{1 \text{ mol e}^-}{96,500 \text{ C}} \right) = 0.0746 \text{ mol e}^-$$
8.00 amps at 15.00 minute, how many mole of e-?

0.0373 mole Cu^{2+} + 0.0746 mole e^- \rightarrow 0.0373 \text{ mole Cu}
What is the mass of copper that is produced in 15.0 minutes by the electrolysis of CuSO₄(aq) if the electrical current is 8.00 Amps

3) Relate electrons to quantity of Cu being formed using half reactions and stoichiometry

\[ g_{Cu} = (0.0746 \text{ mol } e^-) \left( \frac{1 \text{ mol } Cu}{2 \text{ mol } e^-} \right) \left( \frac{63.55 \text{ g}}{1 \text{ mol } Cu} \right) = 2.37 \text{ g } Cu \]

**Put it all together:**

\[ g_{Cu} = (8.00 \text{ A})(15.00 \text{ min}) \left( \frac{60.0 \text{ s}}{1 \text{ min}} \right) \left( \frac{1 \text{ C}}{1 \text{ A} \cdot \text{s}} \right) \left( \frac{1 \text{ mol } e^-}{96,500 \text{ C}} \right) \left( \frac{1 \text{ mol } Cu}{2 \text{ mol } e^-} \right) \left( \frac{63.5 \text{ g}}{1 \text{ mol } Cu} \right) = 2.37 \text{ g } \]
No reaction occurs when copper metal is placed in aqueous zinc nitrate.

\[
\text{Cu(s)} + \text{Zn}^{2+}(\text{aq}) \rightarrow \text{no reaction}
\]

If this reaction could occur

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

what is the \( E_{\text{rxn}}^{\circ} \)?

Is this reaction spontaneous or non-spontaneous?

\[
\text{Cu}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Cu} (\text{s}) \quad + \quad 0.34 \text{ V}
\]

\[
\text{Zn}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Zn} (\text{s}) \quad -0.76 \text{ V}
\]
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

http://media.pearsoncmg.com/bc/bc_0media_chem/chem_sim/electrolysis_fc1_gm_11-26-12/main.html
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

Which metals will you select for each electrode?
A. Cu anode; Cu cathode
B. Cu anode; Zn cathode
C. Zn cathode; Zn anode
D. Cu cathode; Zn anode
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

The copper spoon should be connected to which terminal of the power supply?

A. Negative terminal  
B. Positive terminal
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

Which aqueous solution will you select? Explain?

A. Zn\(^{2+}\)(aq)
B. Cu\(^{2+}\)(aq)
Activity 86 Consider a copper/zinc electrolytic cell

Zinc and copper electrodes are in Zn(NO$_3$)$_2$(aq) and connected to a DC Power Supply. The goal is to plate zinc metal on the copper electrode.

Which electrode serves as the cathode in this set-up?
A. zinc
B. copper

\[
\text{Cu}^{2+} (aq) + 2 \text{e}^- \rightarrow \text{Cu} (s) \quad + \quad 0.34 \text{ V}
\]

\[
\text{Zn}^{2+} (aq) + 2 \text{e}^- \rightarrow \text{Zn} (s) \quad -0.76 \text{ V}
\]
Which reaction occurs at the anode?

A. Zn(s) → Zn^{2+}(aq) + 2e-
B. Cu^{2+}(aq) + 2e- → Cu(s)
C. Zn^{2+}(aq) + 2e- → Zn(s)
D. Cu → Cu^{2+}(aq) + 2e-

\[ \text{Cu}^{2+} (aq) + 2 \text{e}^- \rightarrow \text{Cu} (s) \quad + \quad 0.34 \text{ V} \]
\[ \text{Zn}^{2+} (aq) + 2 \text{e}^- \rightarrow \text{Zn} (s) \quad -0.76 \text{ V} \]
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

The power supply will be set to how many amps and how many minutes? Justify.
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.
What current and what time is required to plate 2.50 grams of zinc?

1) Relate electrons to quantity of Zn being formed using half reactions and stoichiometry

\[
2.50 \text{g} \text{Zn} \times \left( \frac{1 \text{ mole Zn}}{65.3 \text{ g}} \right) \left( \frac{2 \text{ mol } e^-}{1 \text{ mole Zn}} \right) = (0.0765 \text{ mol } e^-)
\]
What current and what time is required to plate 2.50 grams of zinc?

Zn atom is oxidized

Anode:  $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

Zn$^{2+}$ is reduced

Cathode: $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)}$

Overall: $\text{Zn(s)} + \text{Zn}^{2+}(\text{aq}) \rightarrow \text{Zn(s)} + \text{Zn}^{2+}(\text{aq})$

2) Calculate charge (Coulombs):

$$0.0765\text{ mole}^{-1} \times \left(\frac{96,500 \text{ C}}{1\text{ mole}^{-1}}\right) = 7.38 \times 10^3 \text{ C}$$

3) Calculate the time it takes with 8.00 Amps:

$$7.38 \times 10^3 \text{ C} \times \left(\frac{1\text{ Amp} \cdot \text{s}}{1\text{ C}}\right) \times \left(\frac{1}{8.00 \text{ Amp}}\right) \times \left(\frac{1\text{ min}}{60.0 \text{ sec}}\right) = 15.3\text{ min}$$
The reaction at the anode

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

Zinc metal plates onto the copper electrode.

The reaction at the cathode

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

How much mass of Zn is deposited on the Cu electrode?
The reaction at the cathode: $\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Zn}(s)$
The reaction at the cathode: \( \text{Zn}^{2+}(aq) + 2e^- \rightarrow \text{Zn}(s) \)
The reaction at the anode: $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$
The reaction at the anode: 

\[ \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2e^- \]
The reaction at the anode

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

Zinc metal plates onto the copper electrode.

2.55 g of Zn deposited on the Cu electrode (cathode).

2.55 g of Zn removed from Zn electrode (anode).

The reaction at the cathode

\[ \text{Zn}^{2+} (\text{aq}) + 2\text{e}^- \rightarrow \text{Zn(s)} \]
## Compare Electrolysis Experiments

**Complete the Table**

<table>
<thead>
<tr>
<th>Half-reaction</th>
<th>Current (A)</th>
<th>Time (s)</th>
<th>Moles e-</th>
<th>Moles Metal</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Ag}^+ + e^- \rightarrow \text{Ag}$</td>
<td>8.00</td>
<td>900.0</td>
<td>0.0746</td>
<td>0.0746</td>
<td></td>
</tr>
<tr>
<td>$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$</td>
<td>8.00</td>
<td>900.0</td>
<td>0.0746</td>
<td>0.0373</td>
<td></td>
</tr>
<tr>
<td>$\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}$</td>
<td>8.00</td>
<td>900.0</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{Al}^{3+} + 3e^- \rightarrow \text{Al}$</td>
<td>8.00</td>
<td>900.0</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
### Compare Electrolysis Experiments

#### Complete the Table

<table>
<thead>
<tr>
<th>Half-reaction</th>
<th>Current (A)</th>
<th>Time (s)</th>
<th>Moles e-</th>
<th>Moles Metal</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$Ag^+ + e^- \rightarrow Ag$</td>
<td>8.00</td>
<td>900.0</td>
<td>0.0746</td>
<td>0.0746</td>
<td>8.04 g</td>
</tr>
<tr>
<td>$Cu^{2+} + 2e^- \rightarrow Cu$</td>
<td>8.00</td>
<td>900.0</td>
<td>0.0746</td>
<td>0.0373</td>
<td>2.37 g</td>
</tr>
<tr>
<td>$Zn^{2+} + 2e^- \rightarrow Zn$</td>
<td>8.00</td>
<td>900.0</td>
<td>0.0746</td>
<td>0.0373</td>
<td>2.46 g</td>
</tr>
<tr>
<td>$Al^{3+} + 3e^- \rightarrow Al$</td>
<td>8.00</td>
<td>900.0</td>
<td>0.0746</td>
<td>0.0248</td>
<td>0.669 g</td>
</tr>
</tbody>
</table>
If 1.50 amps flow through a Ag\(^{+}(aq)\) solution for 15.0 minutes during an electrolysis experiment, what is the mass of Ag metal deposited?
Answer:
1.50 amps flow through a Ag\(^{+}(aq)\) solution for 15.0 min. Determine the mass of Ag metal deposited.

a. Determine the charge.
   
   Charge (coulombs) = current (A or C/s) \times time (s)
   
   \[ 1350 \text{ C} = (1.5 \text{ amps})(15.0 \text{ min})(60 \text{ s/min}) \]

b. Calculate moles of electrons needed.
   
   Charge \times (n/F) = moles electrons
   
   \[ 1350 \text{ C} \times (1 \text{ mol e}^{-}/96,458 \text{ C}) = 0.0140 \text{ mole electrons} \]

c. Determine mass of metal (Ag) plated.

   Metal plated (g) = mole e\(^{-}\) \times (1 \text{ mol metal/mole } e\(^{-}\)) \times atomic mass
   
   Metal (g) 0.0140 \text{ mol } e^{-} \times (1 \text{ mol Ag/1 mole } e^{-}) \times 108 \text{ g/mol}
   
   Mass of metal plated (Ag) = 1.51 g