CH 223 Friday Sept. 08, 2017 L14B

Previously:

- Relationships between E_{cell}, K, and ΔG
- Concentration and cell potential
- Nernst equation for non-standard conditions:

$$E_{cell} = E^0_{cell} - \frac{0.0592}{n} \log Q$$

at 298.15 K

- Concentration cells – a concentration difference can generate a potential even when all other components of the cell are the same.

Next

- Electrolytic cells – the use of electrical energy to drive non-spontaneous redox reactions.

But first, a couple of clicker questions…

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 to cause a nonspontaneous reaction (thermodynamically unfavorable) to occur.

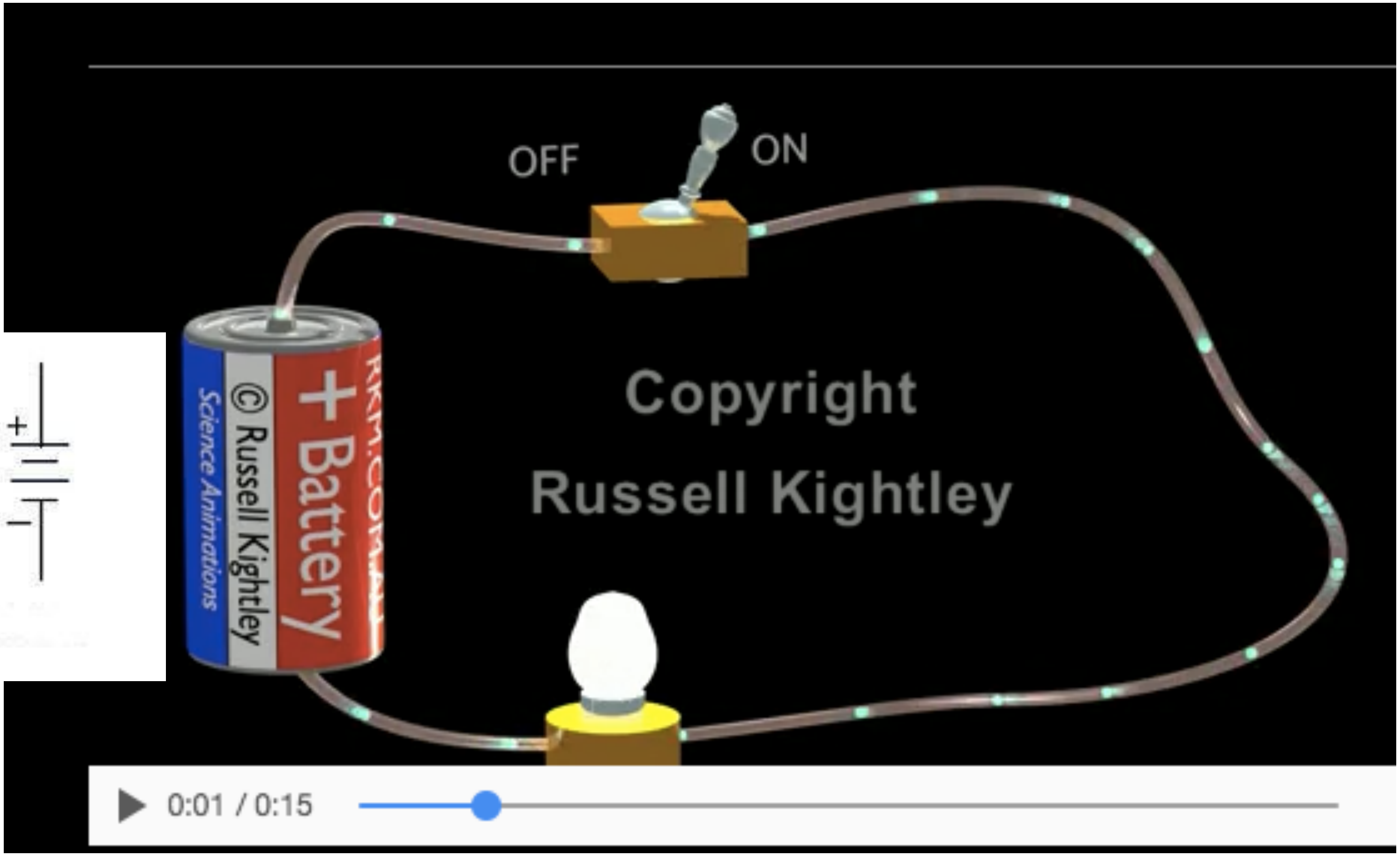
Electrochemical Terminology: Voltage

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).

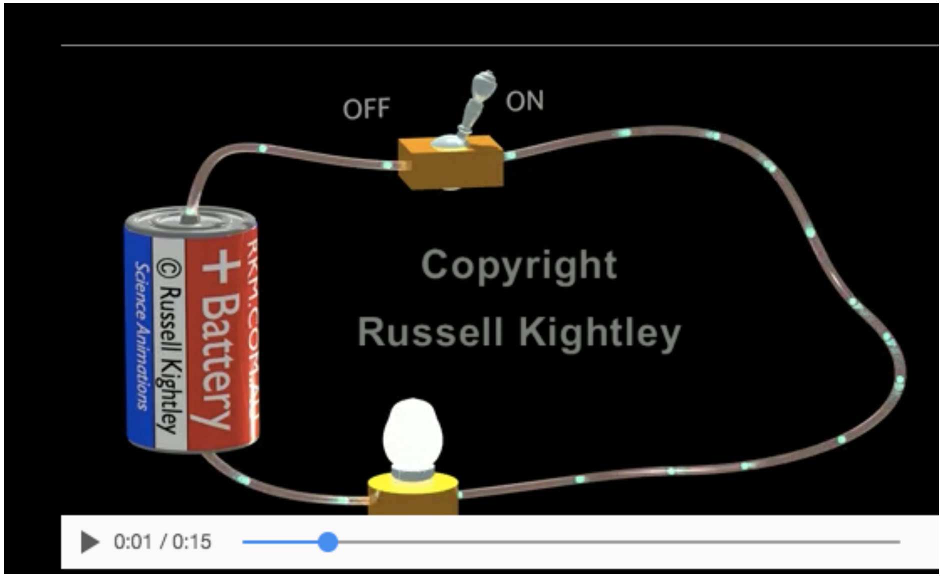
What does a battery do?



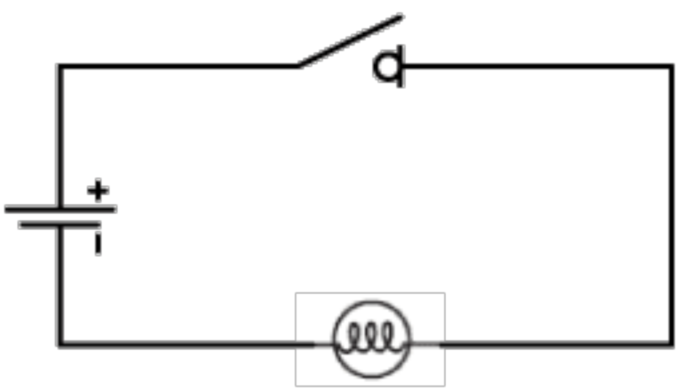
A cell or battery provides the electromotive force that pushes ele
of the negative terminal and pulls electrons into the positive term

https://www.lightrocket.com/russellkightley/galleries/go/12734/elec

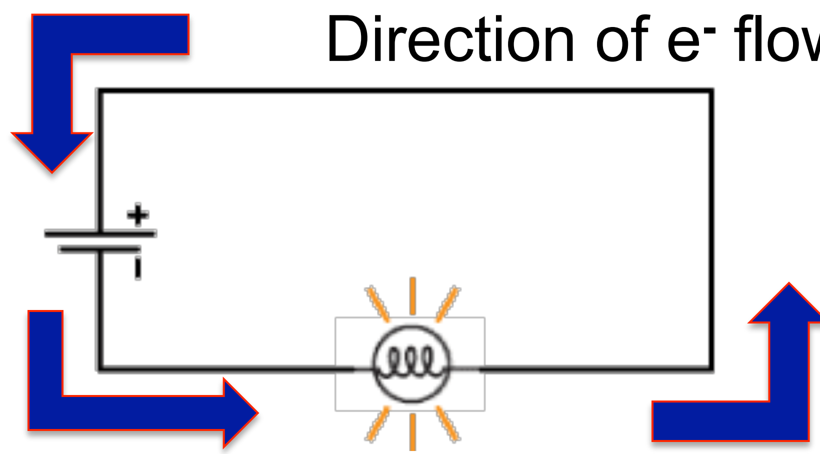
Physics: circuit diagram



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.

Electrochemical Terminology: Current

Current: The number of electrons that flow through the system per second

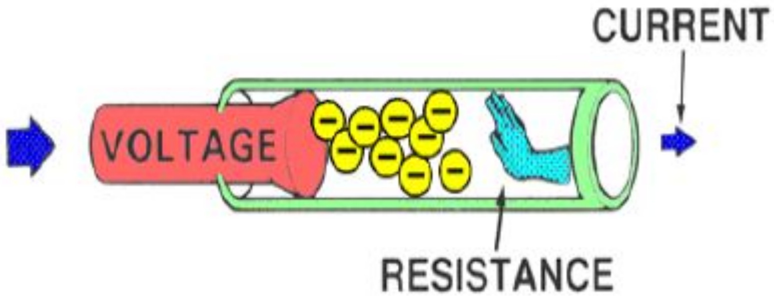
- Unit is the ampere (A).
- 1 Ampere of current = 1 coulomb of charge flowing by each second
  - 1 Amp = 6.242×10^{18} electrons/second

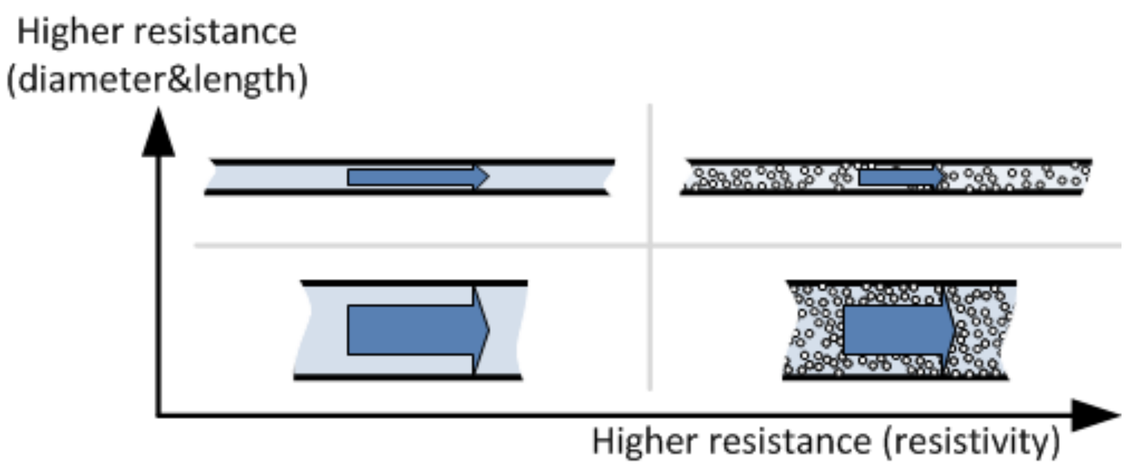
Resistance

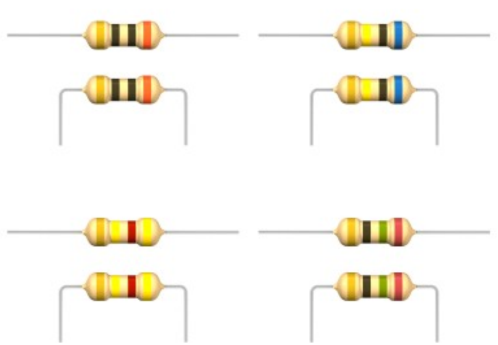
Resistance is a property of materials which opposes the flow of electrons (current) through it.

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

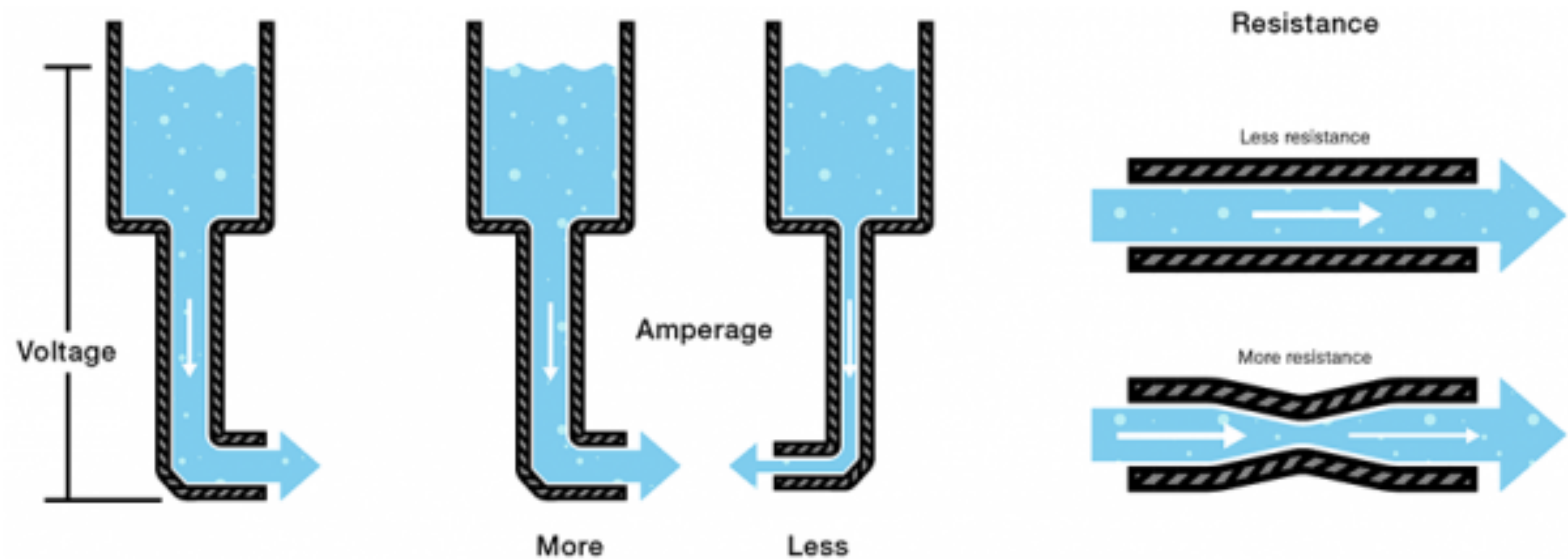
The unit of resistance is the ohm.







Physics Electrical Terms & Concepts

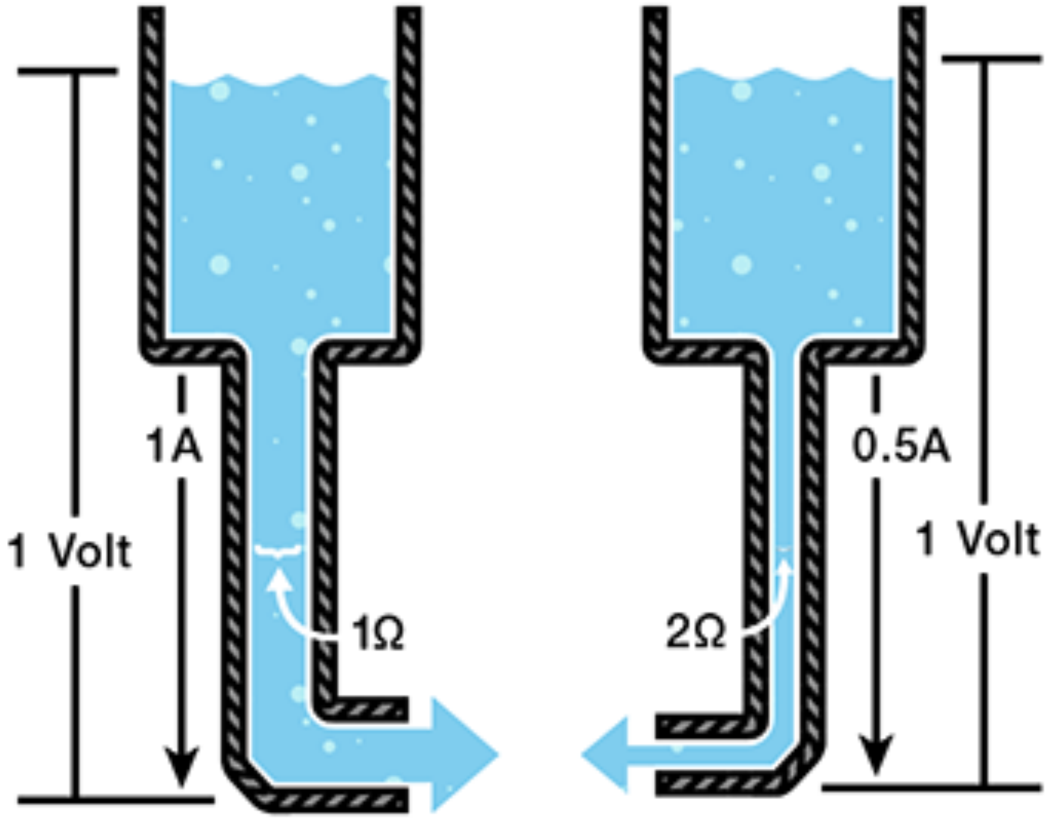


Ohm's Law   V = i R

Ohm's Triangle

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

Resistance is futile



An electrode surface area is a factor governing the number of electrons that can flow.

- Larger batteries produce larger currents

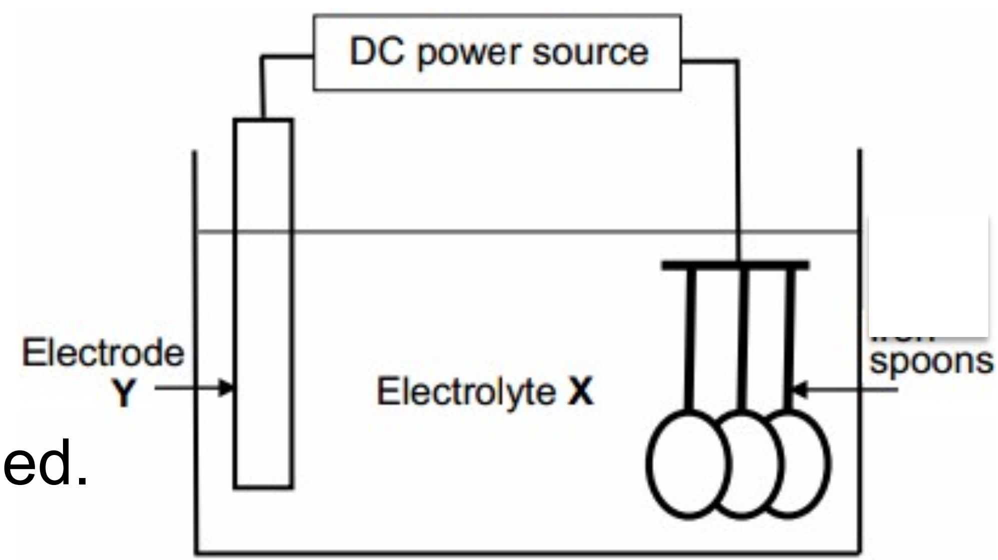
Quantitative aspects of electrolysis

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.

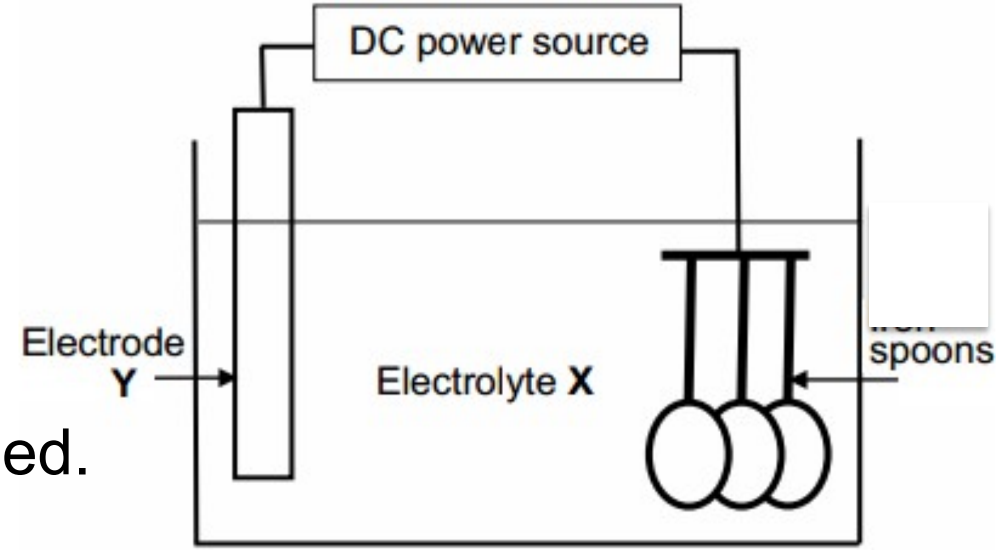


Quantitative aspects of electrolysis

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.

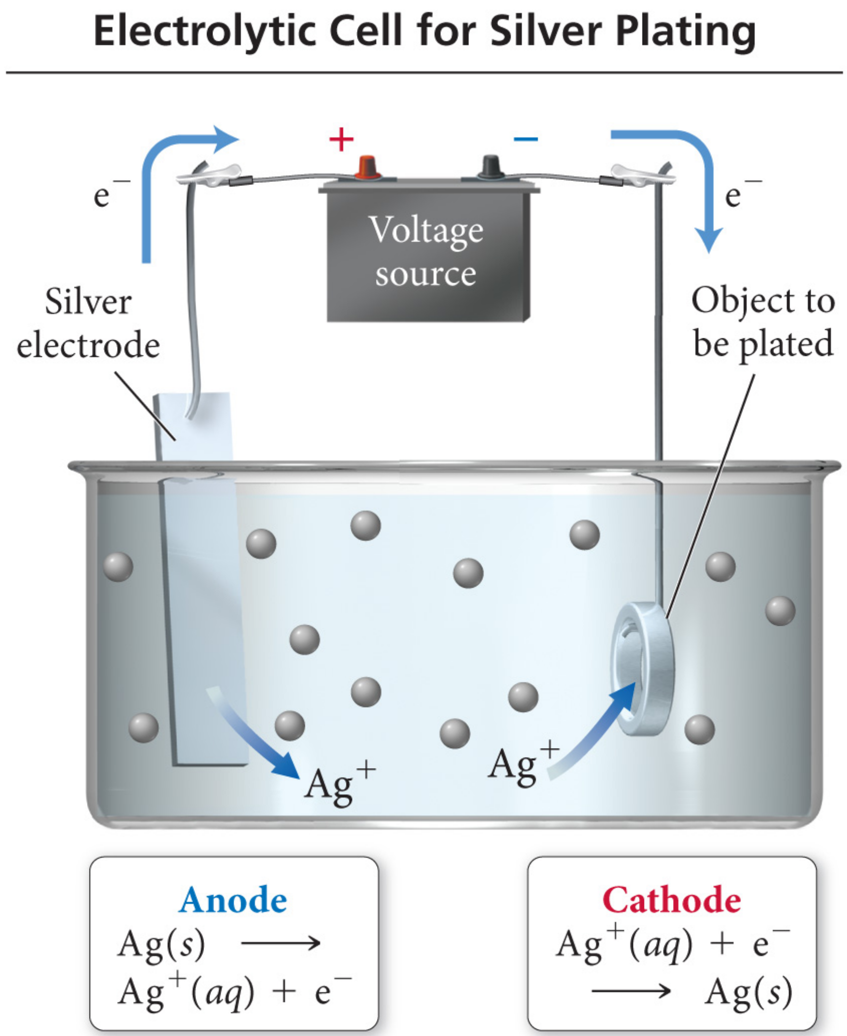


Application of Electrolysis: Electroplating

The power source forces electrons to the ring.

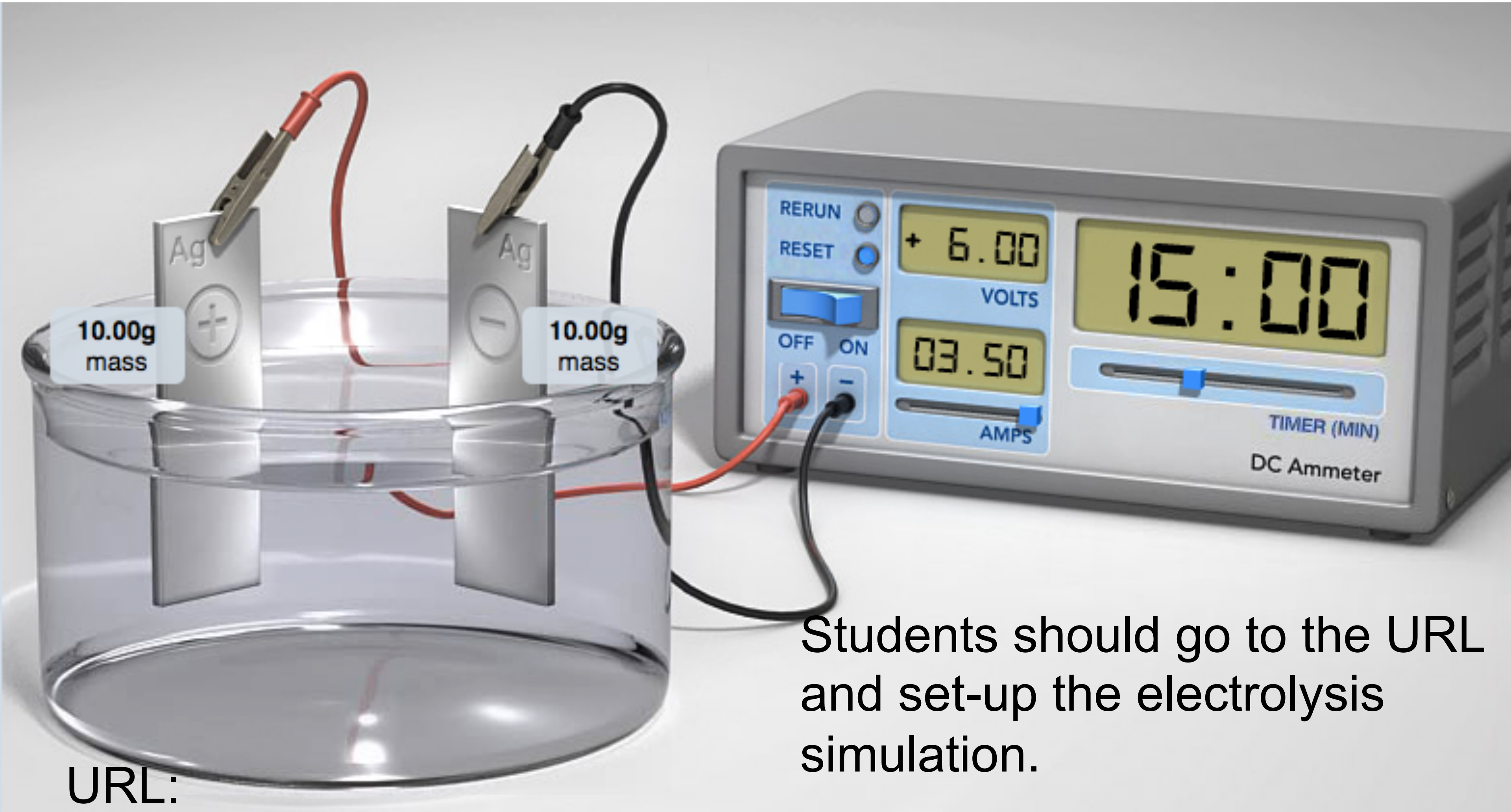
Key Concept: Count electrons

$1\ e^- + Ag^+ \rightarrow Ag$
$10\ e^- + 10\ Ag^+ \rightarrow 10\ Ag$
$1\text{mole } e^- + 1\text{mole } Ag^+ \rightarrow ?$



Notice: anode and cathode aren't separated.

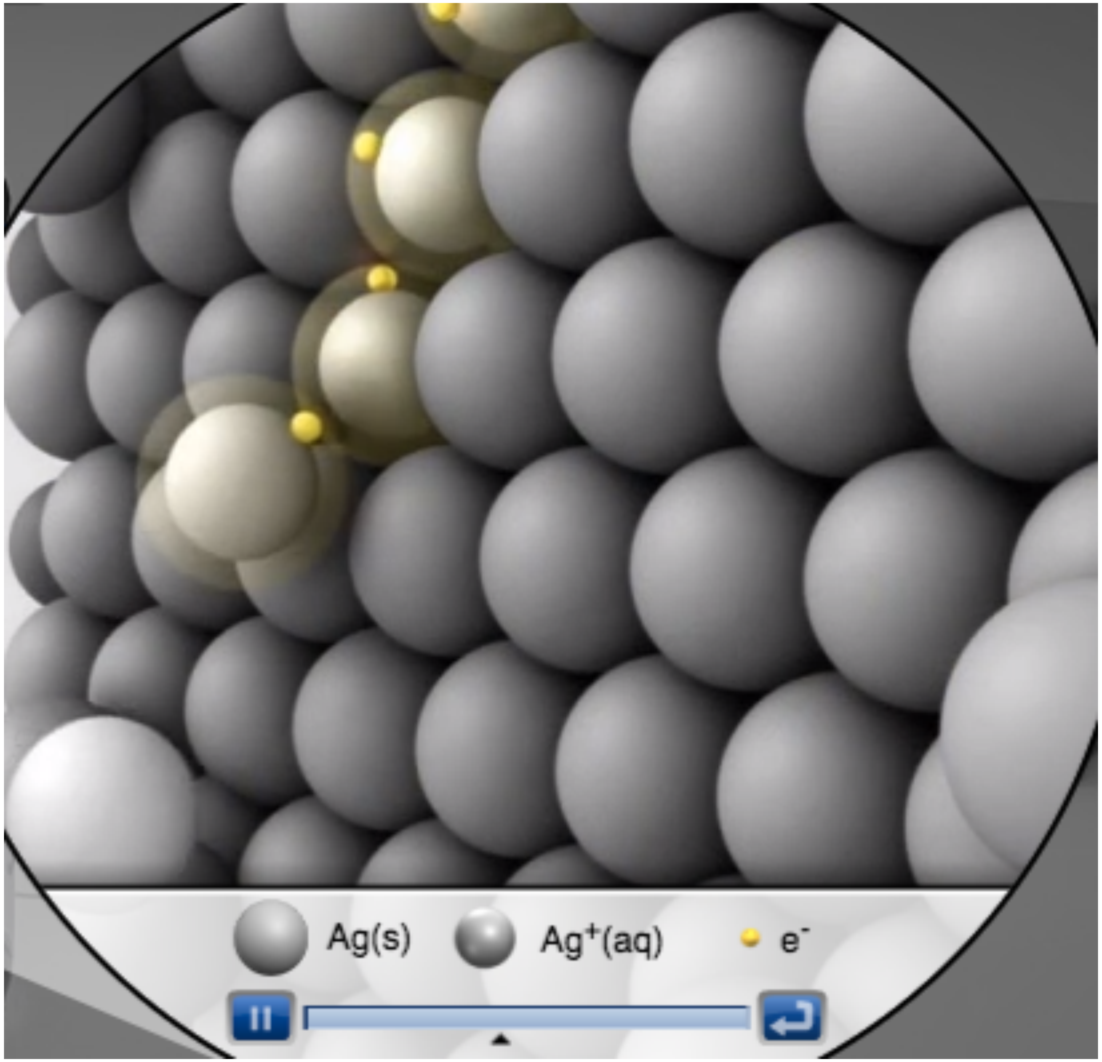
Ag/Ag Electrolysis Experiment



Students should go to the URL and set-up the electrolysis simulation.

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

Ag/Ag Electrolysis Experiment: particle view at the Cathode



$$1\ Ag^{+}(aq) + 1\ e^{-} \rightarrow 1\ Ag(s)$$

Ag/Ag Electrolysis Experiment

How many electrons were forced into this system?

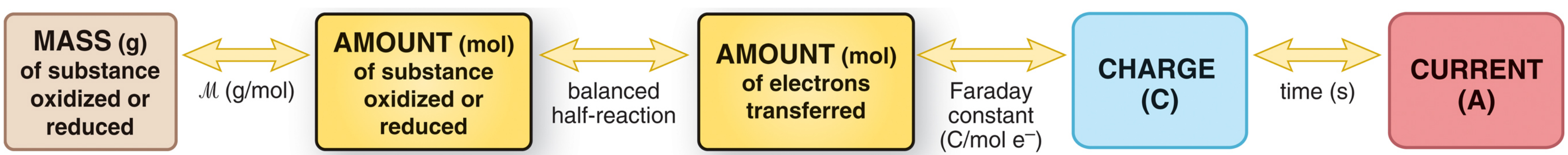
Information we will need:

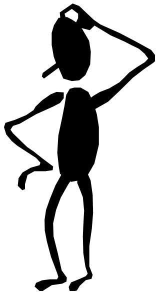
- Balanced half reaction, telling us moles of electrons
- Electrical current is measured in **amperes (A)**

  $$1\ A = 1\ C/s$$

  where a coulomb, C, is the SI unit of electrical charge.

- **Faraday constant:** F = 96,500 C/mole e^-





3.53 g of silver is produced in 15.0 minutes by the electrolysis of $AgNO_3(aq)$ when the electrical current is 3.50 Amps. How many moles of electrons were passed?

Ag atom is oxidized

Anode: $Ag \rightarrow Ag^+(aq) + \mathbf{1}\ \mathbf{e^-}$

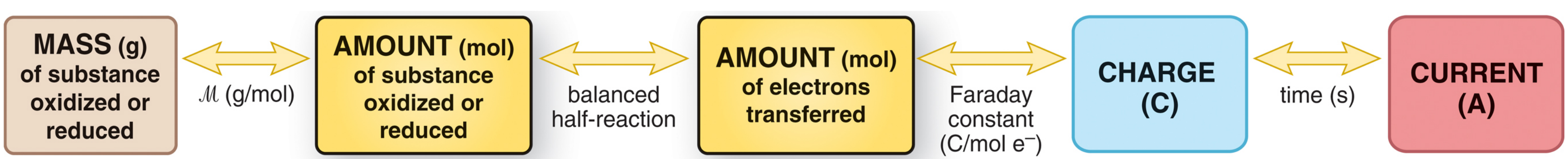
Ag$^+$ is reduced

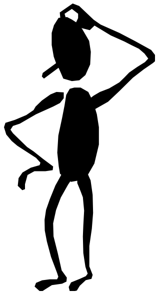
Cathode: $Ag^+(aq) + \mathbf{1e^-} \rightarrow Ag(s)$

Overall: $Ag(s) + Ag^+(aq) \rightarrow Ag(s) + Ag^+(aq)$

1 Amp sec = 1 C

- **Faraday constant:** F = 96,500 C/mole e^-





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Ag^+ is reduced

Cathode: $Ag^+(aq) + \mathbf{1e^-} \rightarrow Ag(s)$

Overall: $Ag(s) + Ag^+(aq) \rightarrow Ag(s) + Ag^+(aq)$

1) Calculate charge (Coulombs):

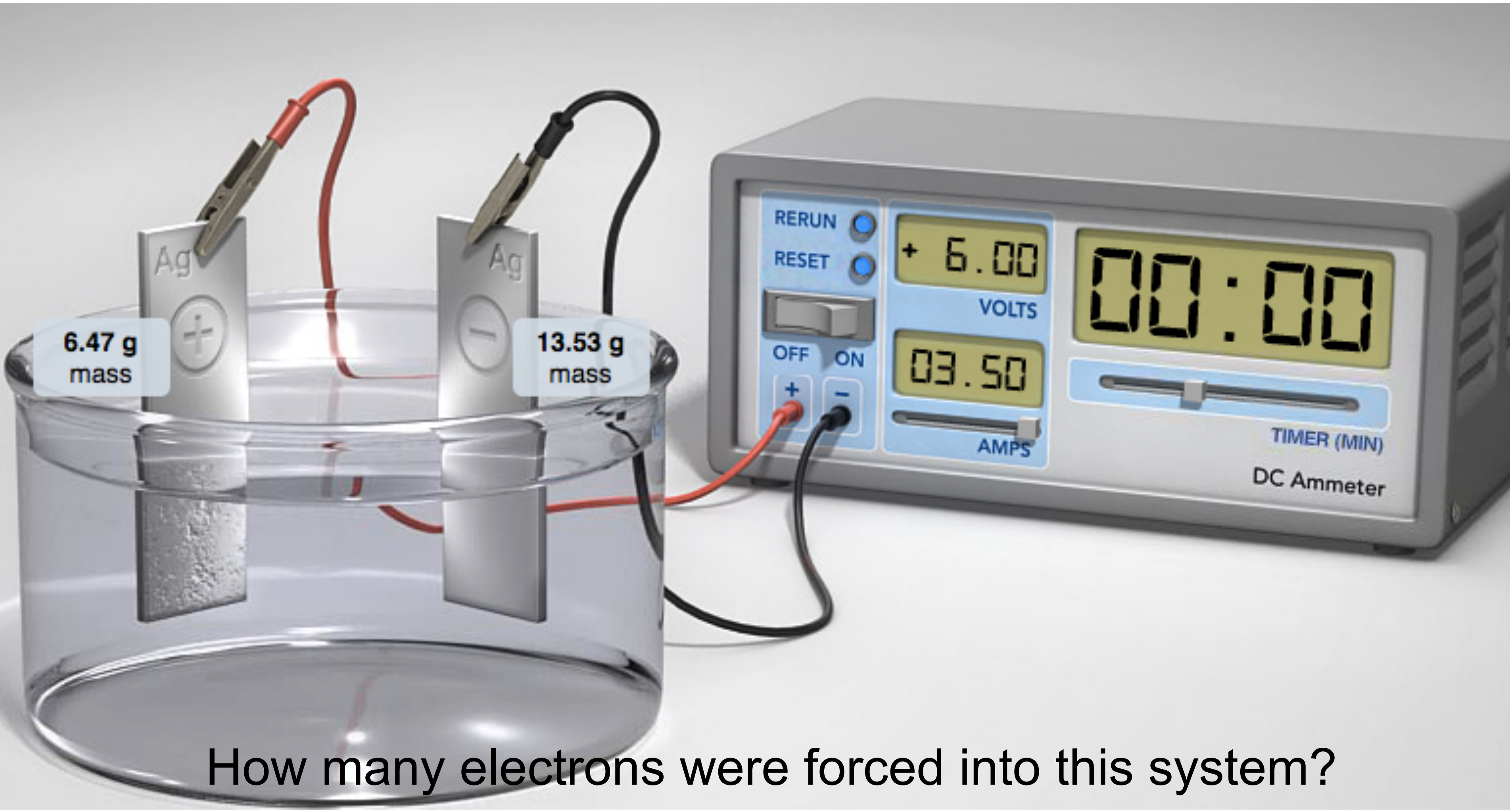
$$Charge = (3.50\,Amps)(15.00\,\min)\left(\frac{60.0\,s}{1\min}\right)\left(\frac{1C}{1\,Amp\cdot s}\right) = 3.15x10^3 C$$

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

$$3.15x10^3 C\left(\frac{1\,mol\,e^-}{96,500\,C}\right) = 0.0326\,mol\,e^-$$

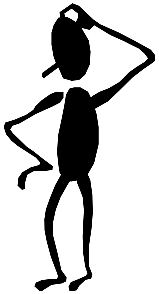
Ag/Ag Electrolysis Experiment

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



How many electrons were forced into this system?

$0.0326\ \text{mole}\ Ag^+ + 0.0326\ \text{mole}\ e^- \rightarrow 0.0326\ \text{mole}\ Ag$

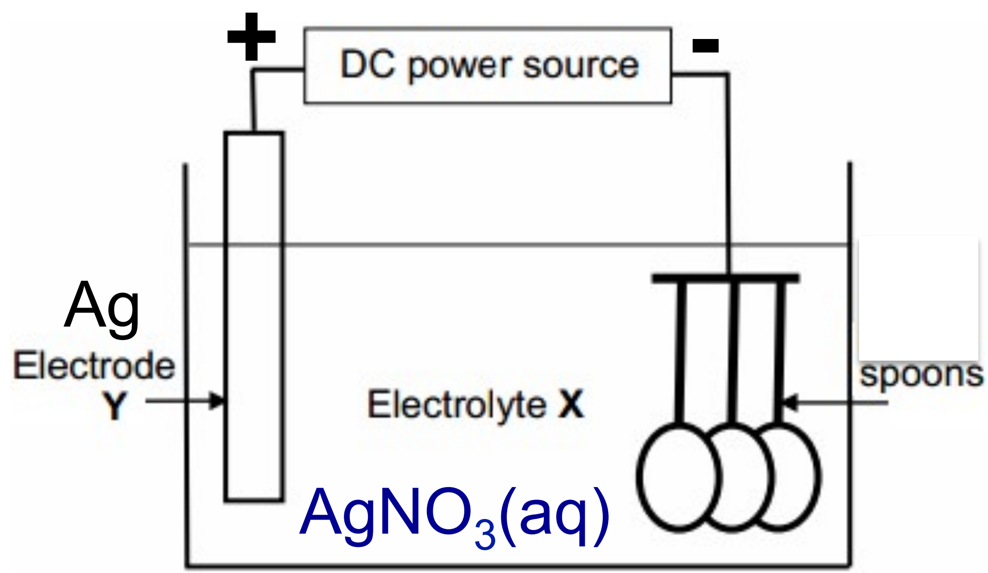


What mass of silver is produced in 15.0 minutes by the electrolysis of $AgNO_3(aq)$ if the electrical current is 3.50 Amps?

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

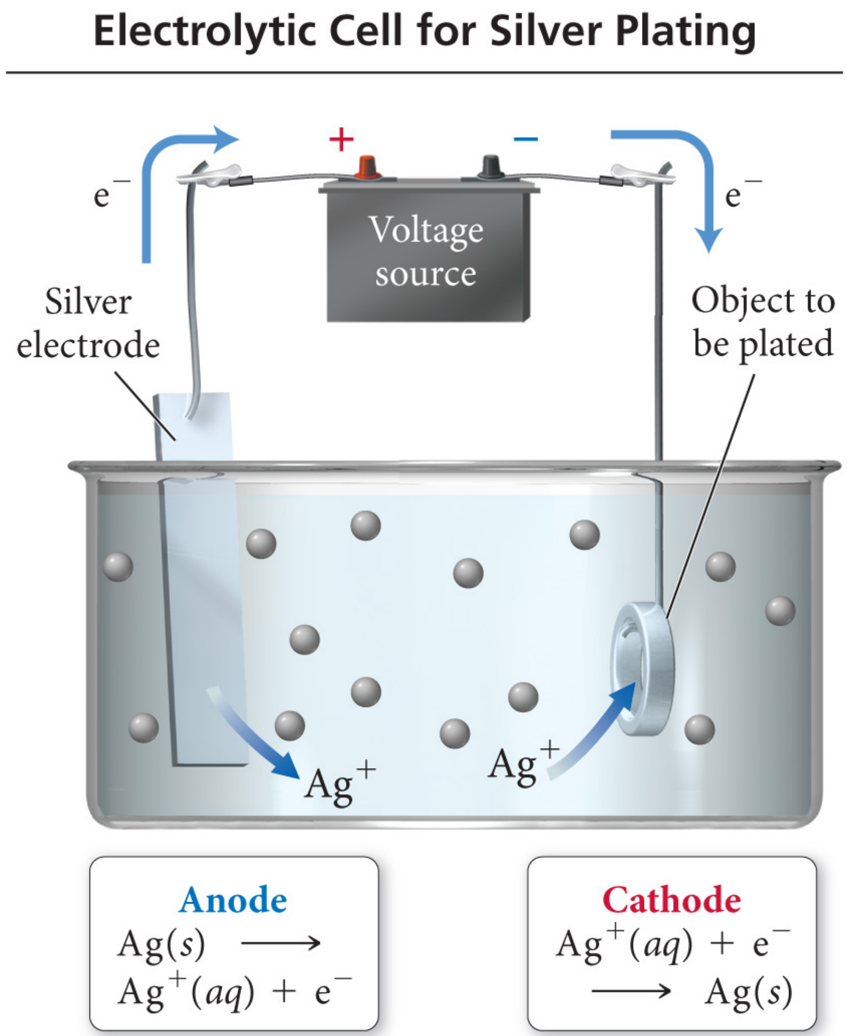
$$mass\ Ag = \left(0.0326\, mol\, e^-\right)\left(\frac{1\, mol\, Ag}{1\, mol\, e^-}\right)\left(\frac{107.9\, g}{1\, mol\, Ag}\right) = 3.52\, g\, Ag$$

$Ag^+(aq) + \mathbf{1e^-} \rightarrow Ag(s)$



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.

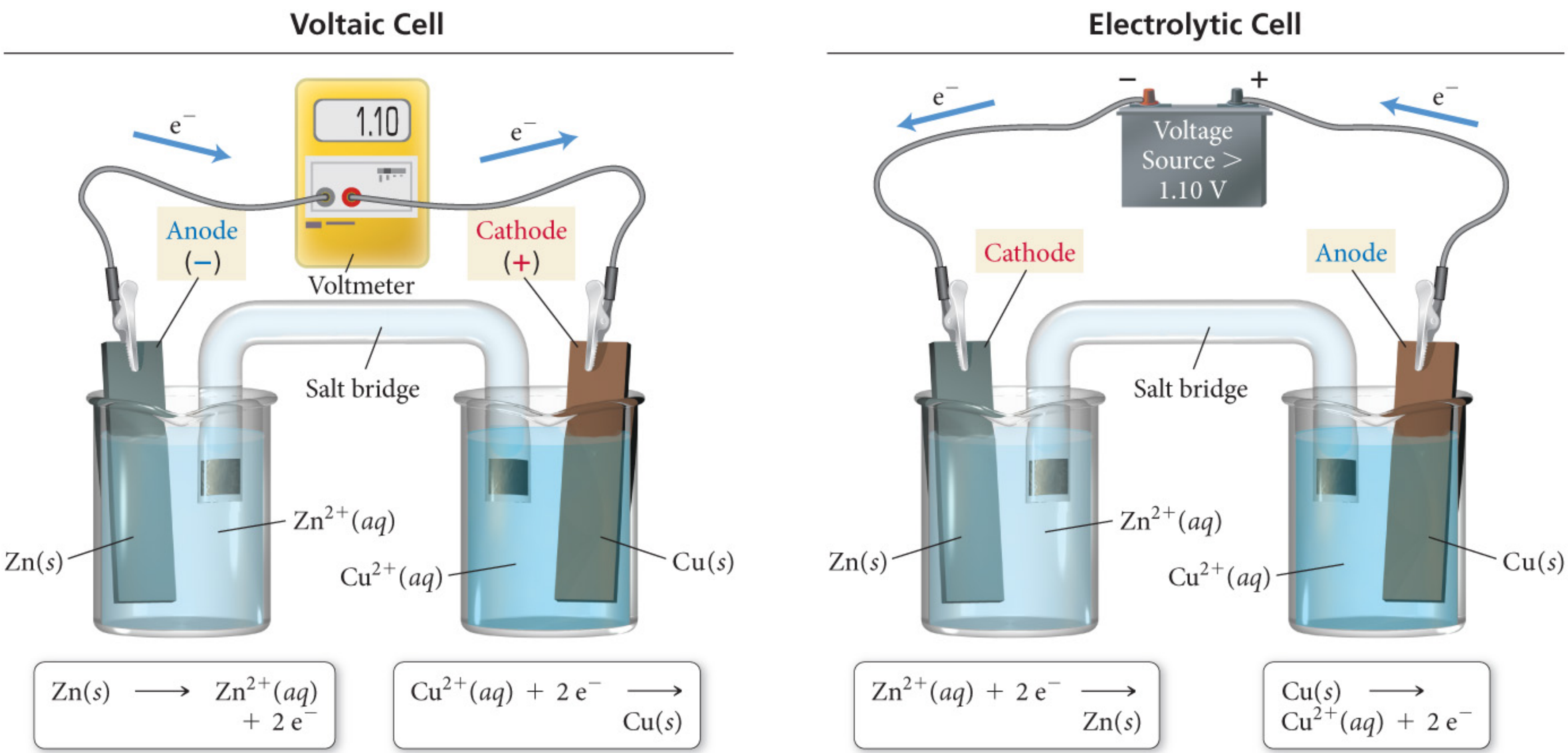
$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

$$Zn^{2+}(aq) + Cu(s) \longrightarrow \text{No reaction}$$

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s) \qquad E^\circ = +0.34\ V$$

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

Voltaic versus Electrolytic Cells



Quantitative aspects of electrolysis

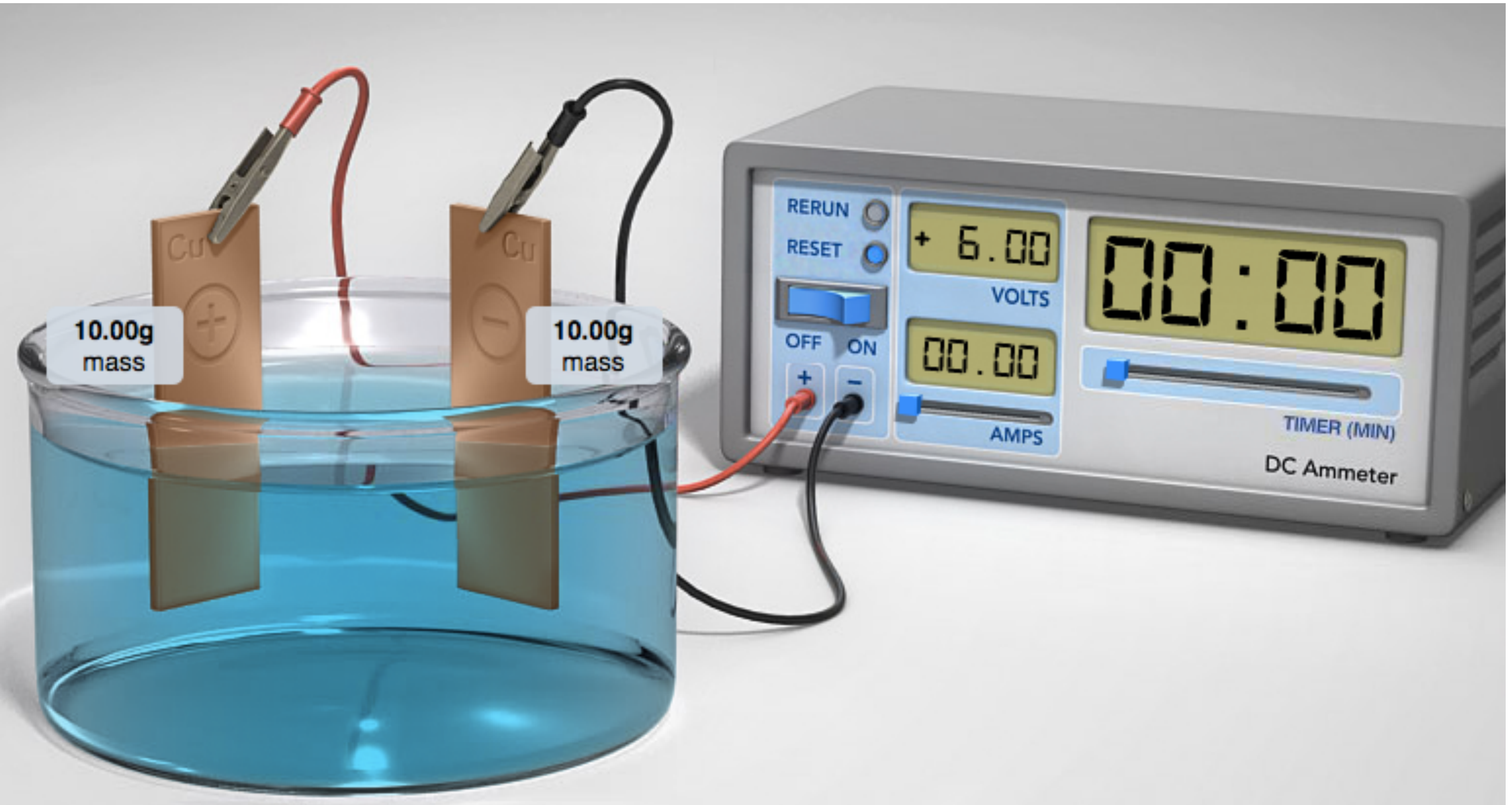
We would like to be able to relate the **quantit**y of reactant or product to the **amount of electricity consumed**, or the **amount of time** it takes for the electrolysis.

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?

POGIL Activity 86 Electrolytic Cells
Model 5: Setting up an electrolysis experiment



Application of Electrolysis: Electroplating

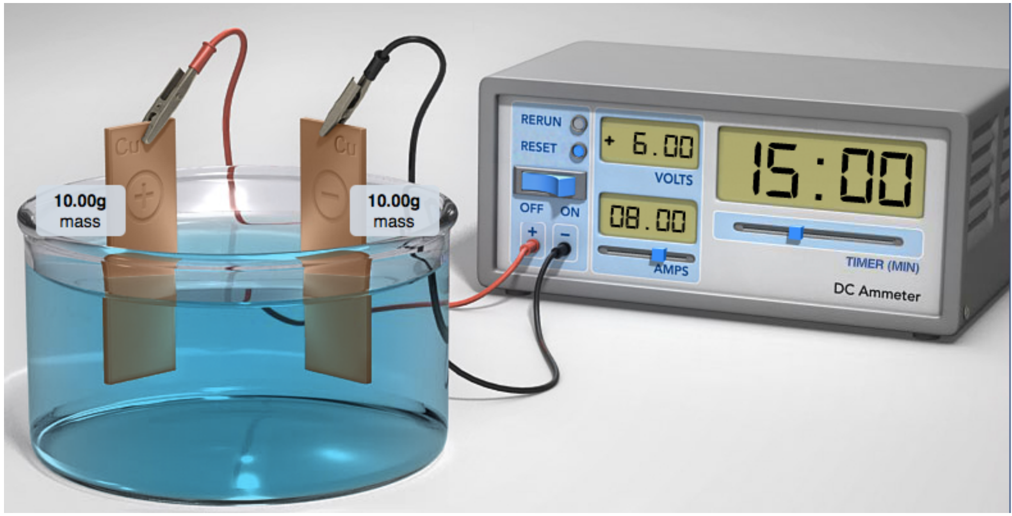
The power source forces electrons to the copper electrode.

Key Concept: Count electrons

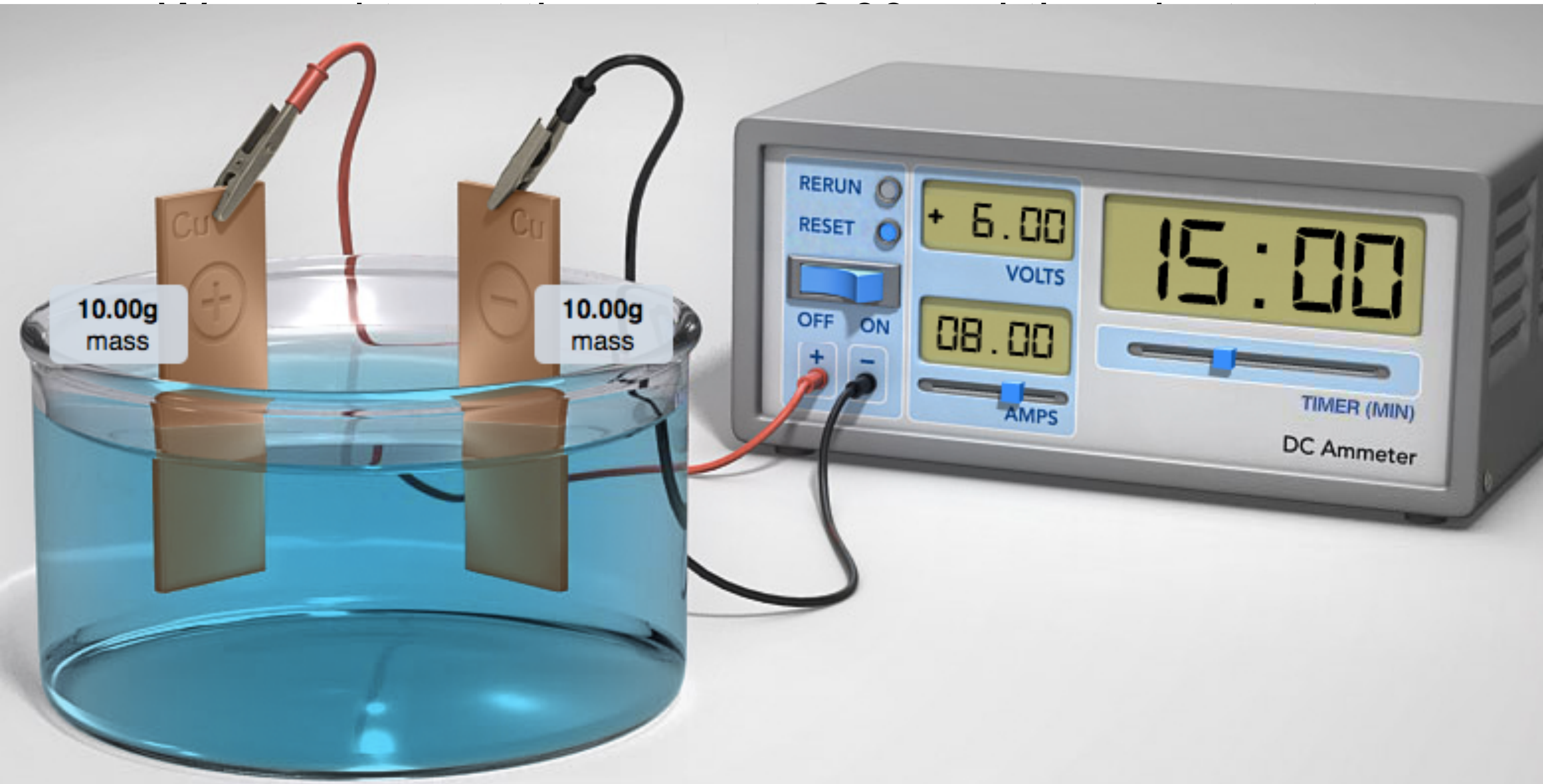
$2\ e^- + Cu^{2+} \rightarrow Cu$

$20\ e^- + 10\ Cu^{2+} \rightarrow 10\ Cu$

2 mole e^- + 1 mole $Cu^{2+} \rightarrow$?

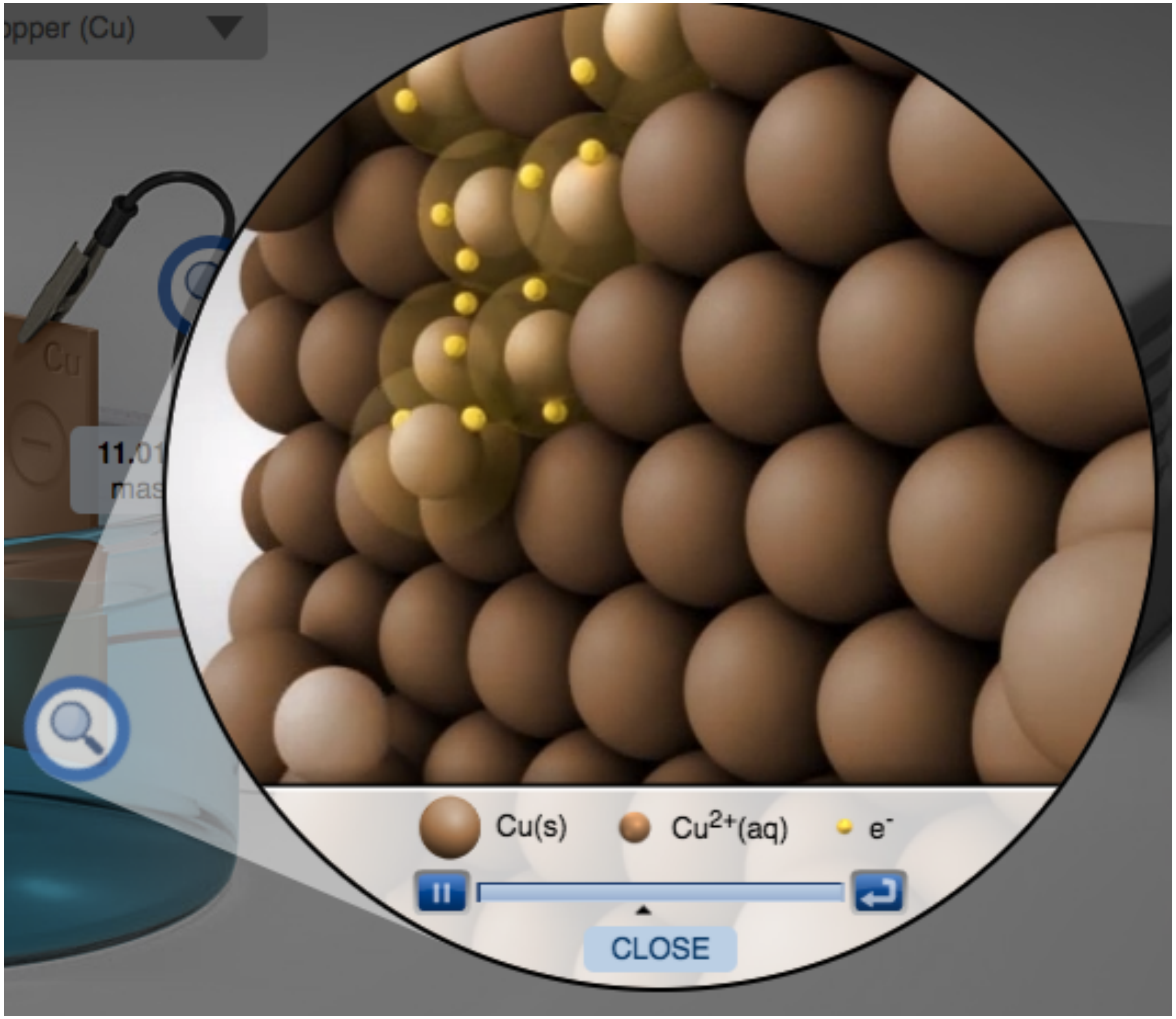


POGIL Activity 86 Electrolytic Cells



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 Cathode



$$1\ Cu^{2+}(aq) + 2\ e^- \rightarrow 1\ Cu(s)$$

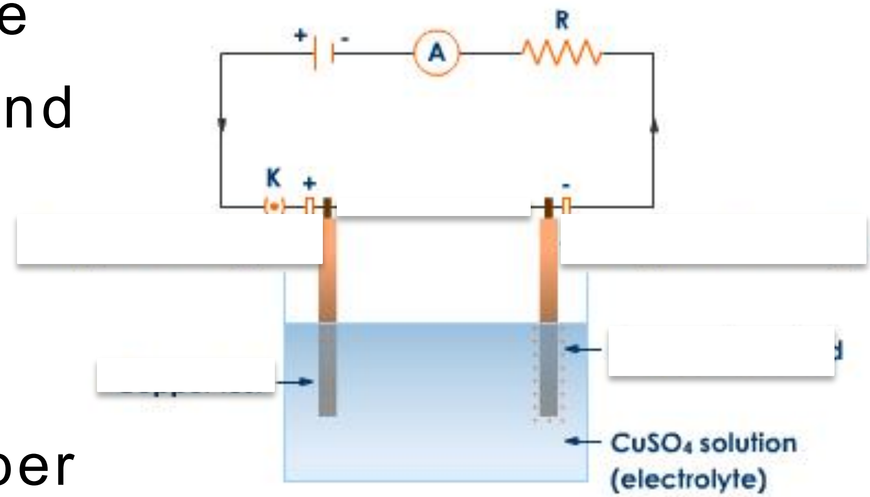
Nine fun things to do with a Cu/Cu electrolysis cell:

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 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.

$Cu^{2+} + 2e^- \rightarrow Cu \qquad E° = +0.34\ V$



Complete the diagram for this Cu/Cu electrolysis cell

$Cu^{2+} + 2e^{-} \rightarrow Cu$ E° = +0.34 V

Identify the anode and cathode.

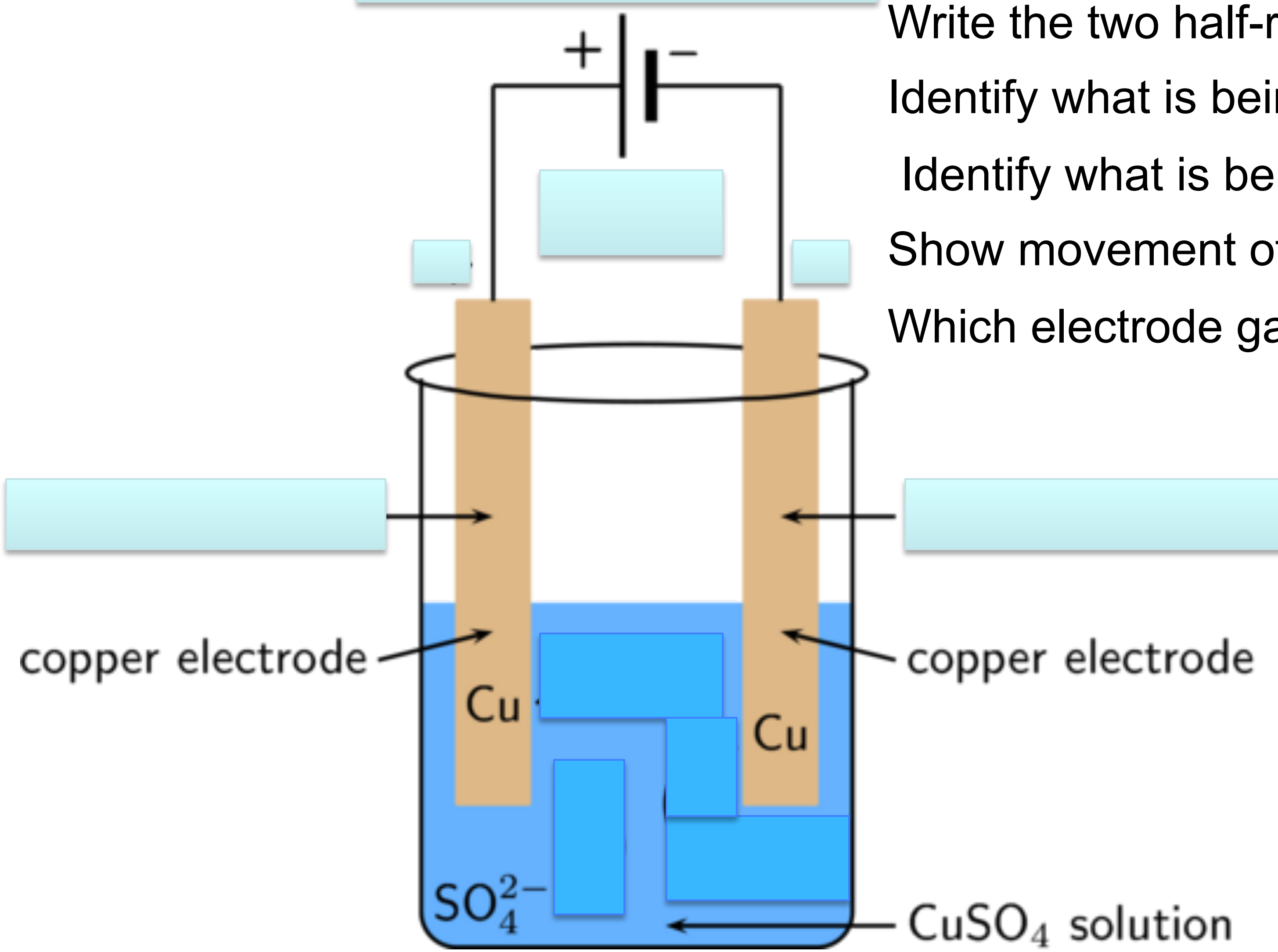
Write the two half-reactions.

Identify what is being oxidized.

Identify what is being reduced.

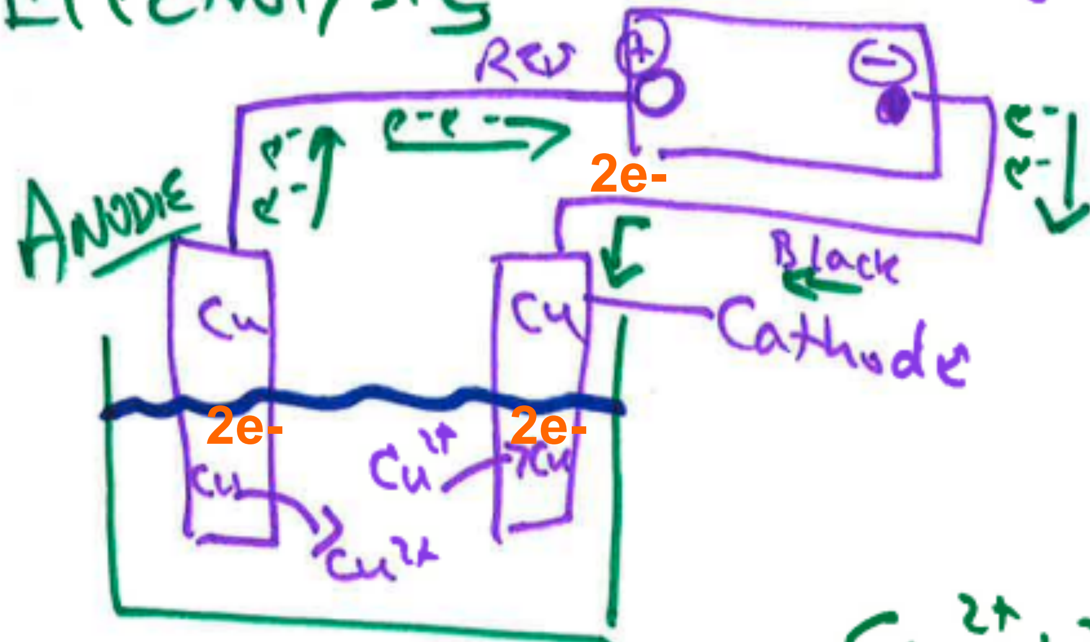
Show movement of ions and e-

Which electrode gains mass?



Electrolysis

D.C. Power Supply or Battery

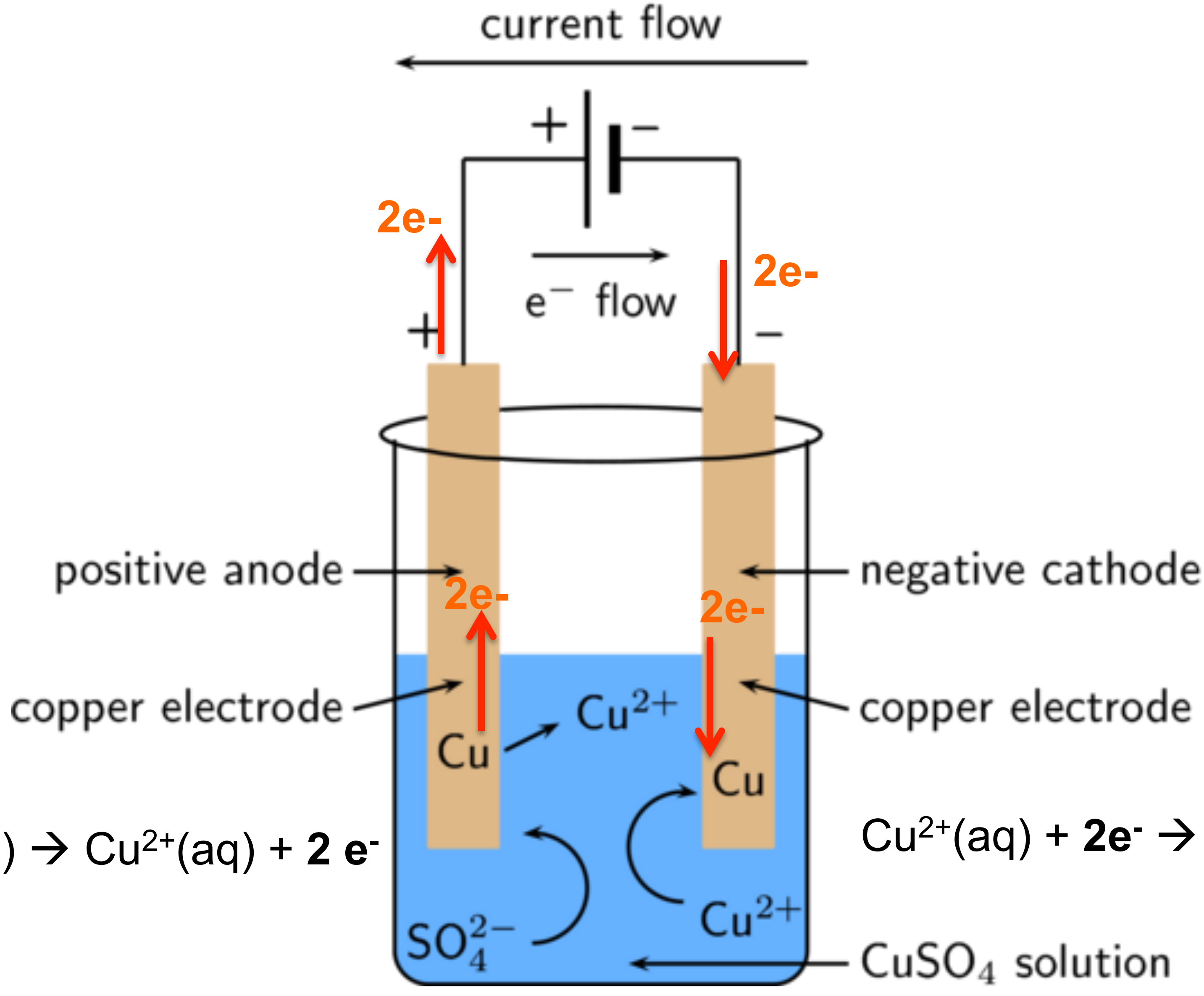


1.0 M Cu^{2+}

$Cu^{2+} + 2e^- \rightarrow Cu$ Red

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

Diagram of a Cu/Cu electrolysis cell

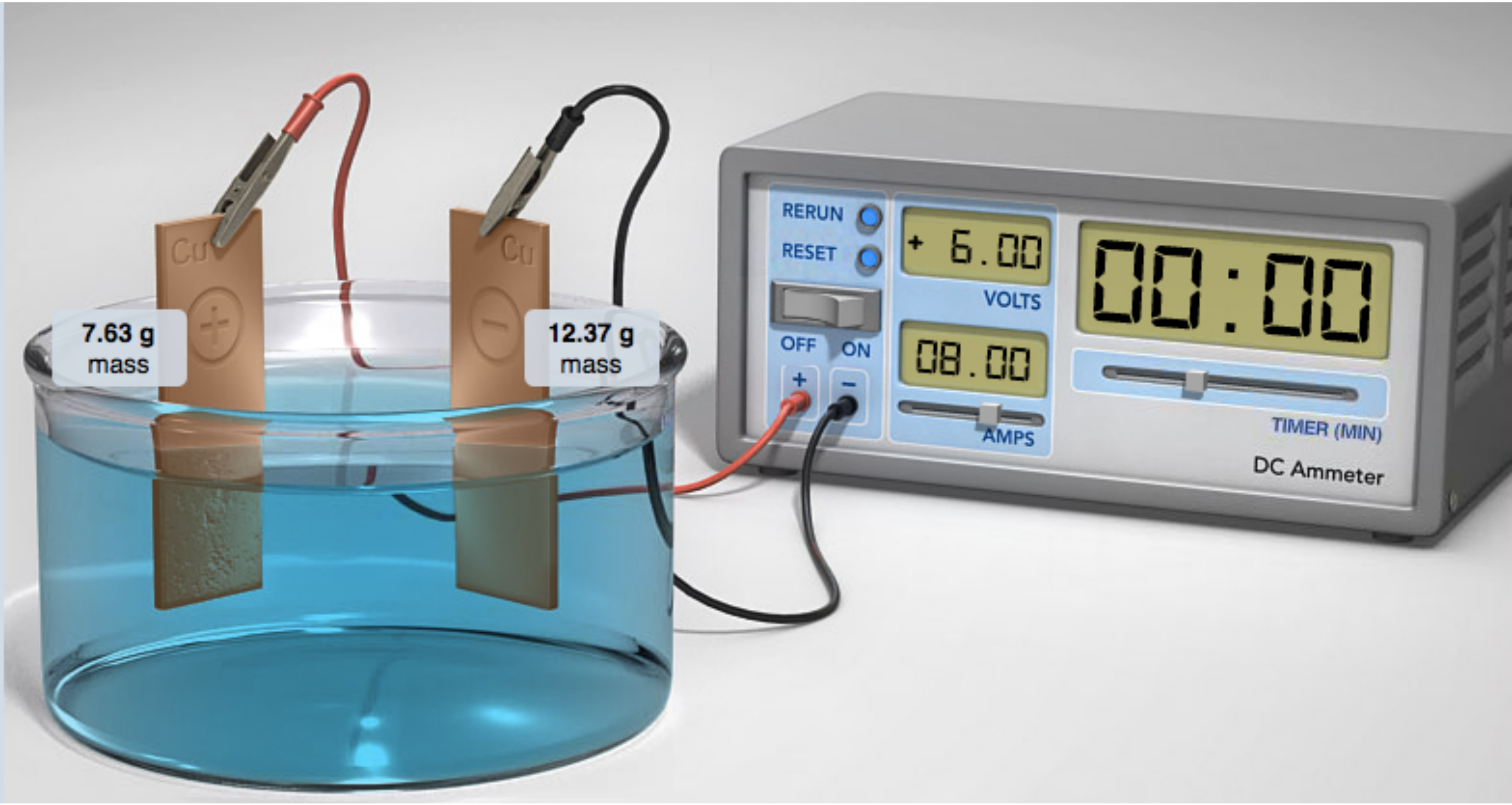


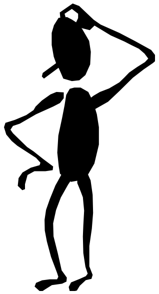
$Cu(s) \rightarrow Cu^{2+}(aq) + \mathbf{2\ e^-}$

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

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) → Cu^{2+}(aq) + **2 e^-**

Cu^{2+} is reduced — Cathode: Cu^{2+}(aq) + **2e^-** → Cu(s)

Overall: Cu(s) + Cu^{2+}(aq) → Cu(s) + Cu^{2+}(aq)

1) Calculate charge (Coulombs):

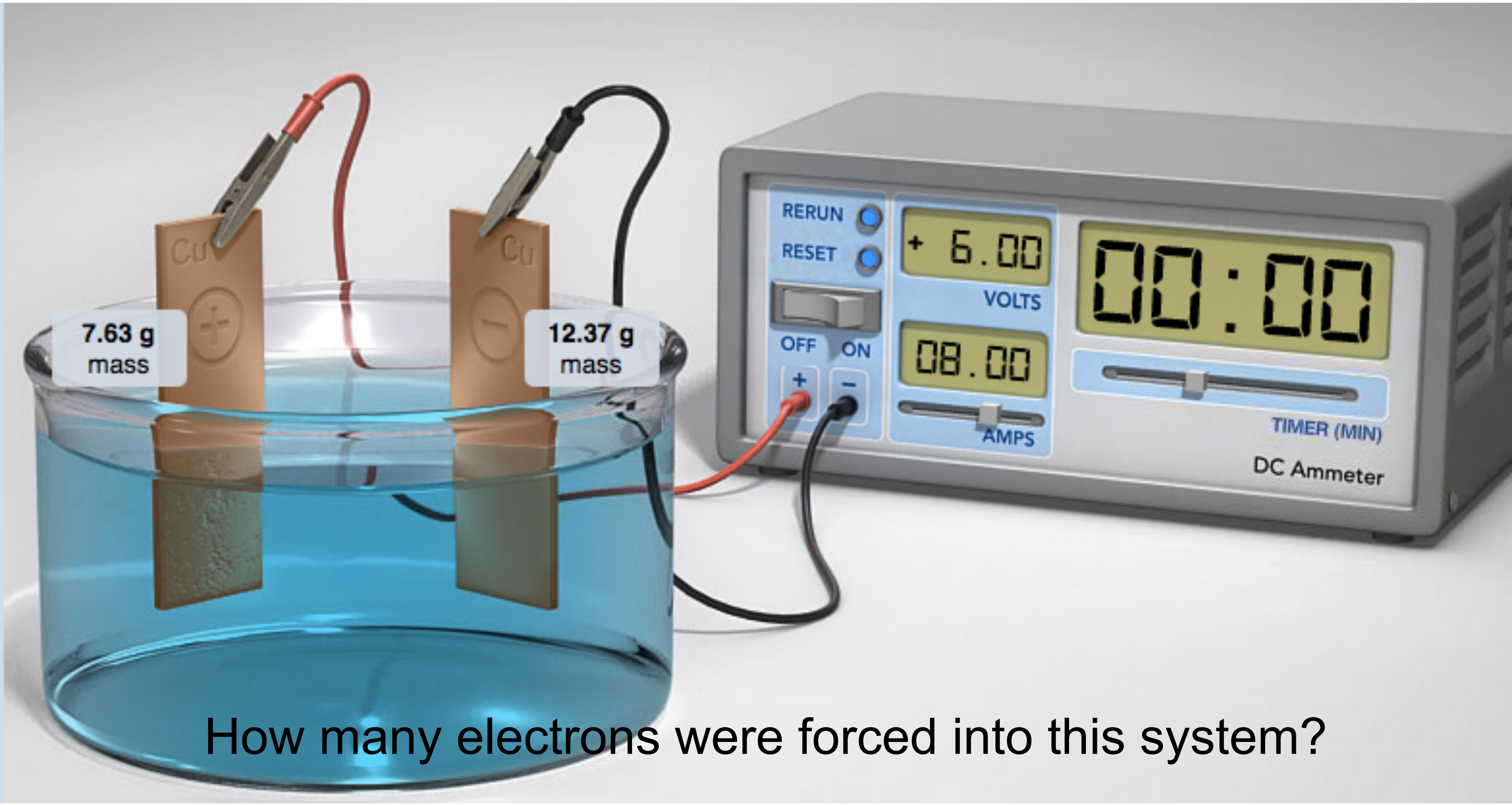
$$C = (8.00\,Amps)(15.00\,\min)\left(\frac{60.0\,s}{1\min}\right)\left(\frac{1\,C}{1\,Amp\cdot s}\right) = 7.20\,x10^3\,C$$

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

$$7.20\,x10^3\,C\left(\frac{1\,mol\,e^-}{96,500\,C}\right) = 0.0746\,mol\,e^-$$

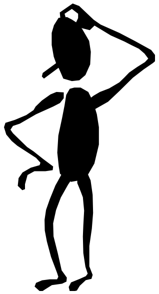
POGIL Activity 86 Electrolytic Cells

8.00 amps at 15.00 minute, how many mole of e-?



How many electrons were forced into this system?

0.0373 mole Cu^{2+} + 0.0746 mole e^- → 0.0373 mole Cu



What is the mass of copper that is produced in 15.0 minutes by the electrolysis of $CuSO_4(aq)$ if the electrical current is 8.00 Amps

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

$$g\,Cu = \left(0.0746\,mol\,e^{-}\right)\left(\frac{1\,mol\,Cu}{2\,mol\,e^{-}}\right)\left(\frac{63.55g}{1\,mol\,Cu}\right) = 2.37\,g\,Cu$$

Put it all together:

$$g\,Cu = (8.00\,\cancel{A})(15.00\,\cancel{\text{min}})\left(\frac{60.0\,\cancel{s}}{1\,\cancel{\text{min}}}\right)\left(\frac{1\,\cancel{C}}{1\,A\cdot\cancel{s}}\right)\left(\frac{1\,\cancel{mol}\,e^{-}}{96,500\,\cancel{C}}\right)\left(\frac{1\,mol\,\cancel{Cu}}{2\,\cancel{mol}\,e^{-}}\right)\left(\frac{63.5\,g}{\cancel{1}\,mol\,Cu}\right) = 2.37\,g$$

No reaction occurs when copper metal is placed in aqueous zinc nitrate.

$Cu(s) + Zn^{2+}(aq)$ -> no reaction

If this reaction could occur

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

what is the $E°_{rxn}$?

Is this reaction spontaneous or non-spontaneous?

$$\mathbf{Cu^{2+}\,(\mathit{aq}) + 2\,e^- \rightarrow Cu\,(\mathit{s})} \qquad \mathbf{+\ 0.34\ V}$$

$$\mathbf{Zn^{2+}\,(\mathit{aq}) + 2\,e^- \rightarrow Zn\,(\mathit{s})} \qquad \mathbf{-0.76\ V}$$

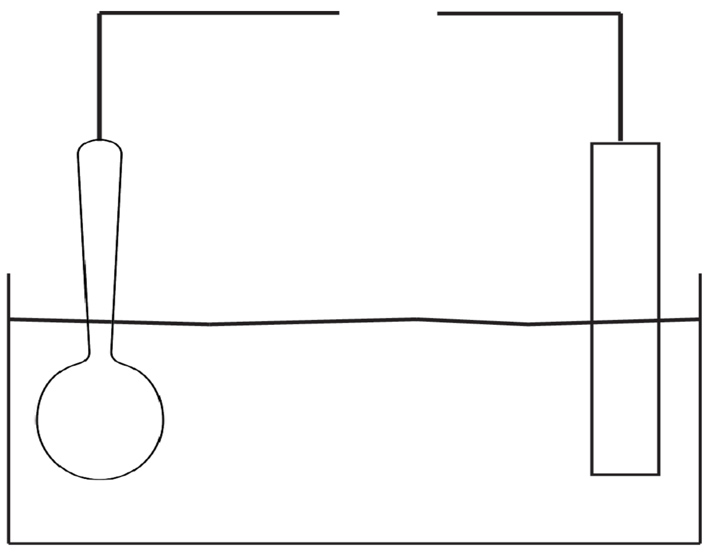
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

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?

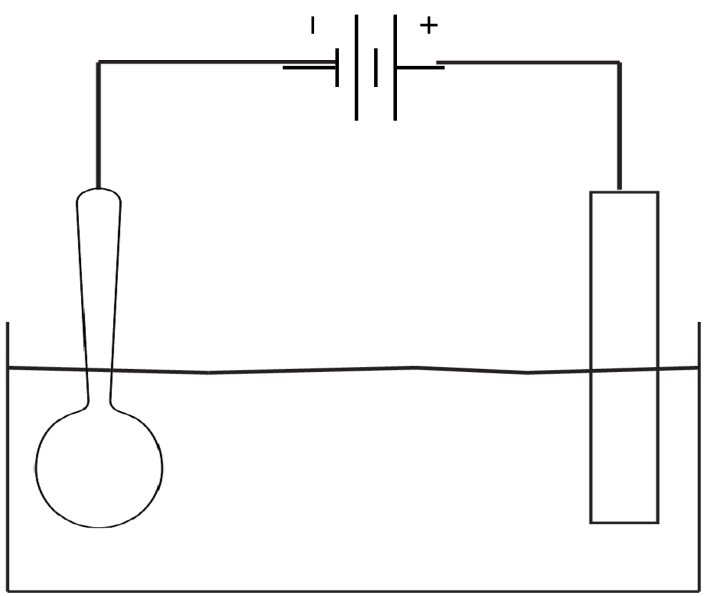
A. Cu anode;   Cu cathode
B. Cu anode;   Zn cathode
C. Zn cathode; Zn anode
D. Cu cathode; Zn anode

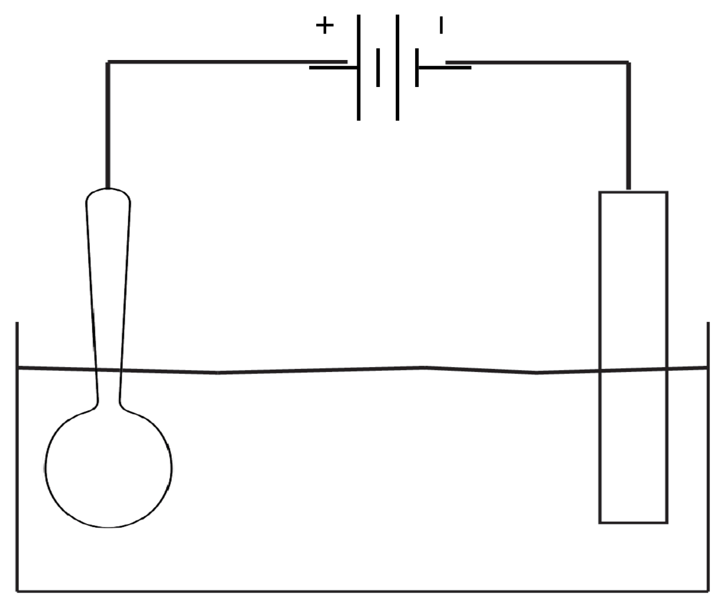


Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

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

A. Negative terminal
B. Positive terminal

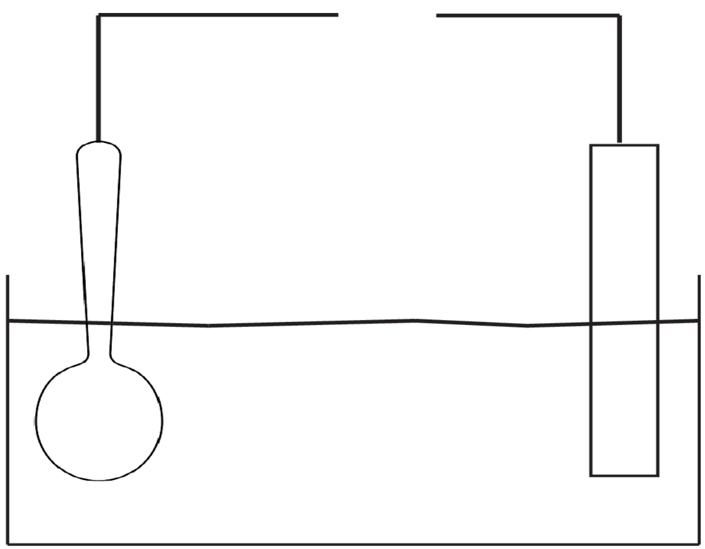




Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

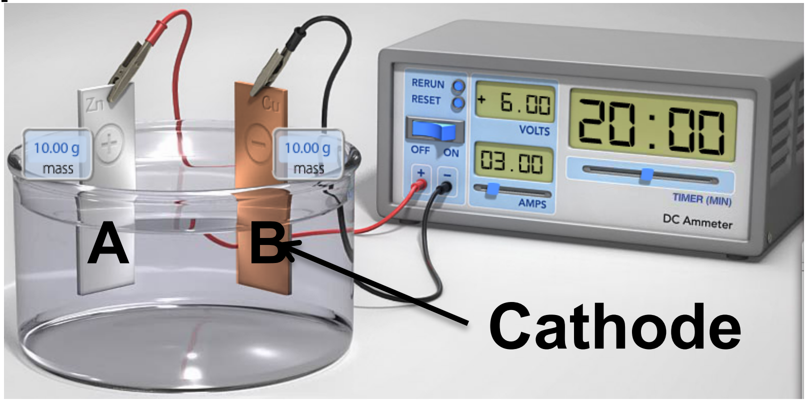
Which aqueous solution will you select? Explain?

A. $Zn^{2+}(aq)$
B. $Cu^{2+}(aq)$



POGIL 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

$$\mathbf{Cu^{2+}\,(\mathit{aq}) + 2\,e^{-} \rightarrow Cu\,(\mathit{s}) \quad +\ 0.34\ V}$$

$$\mathbf{Zn^{2+}\,(\mathit{aq}) + 2\,e^{-} \rightarrow Zn\,(\mathit{s}) \quad -0.76\ V}$$

Which reaction occurs at the anode?

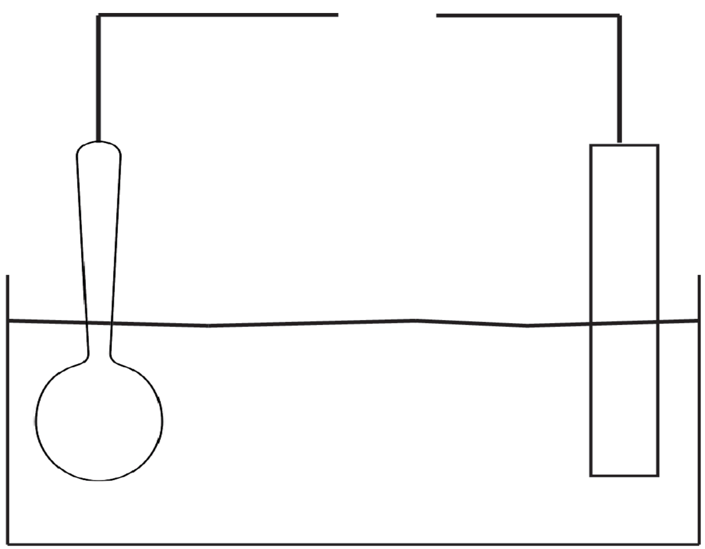
A. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
B. $Cu^{2+}(aq) + + 2e^- \rightarrow Cu(s)$
C. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
D. $Cu \rightarrow Cu^{2+}(aq) + 2e^-$

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \qquad +0.34\ V$$

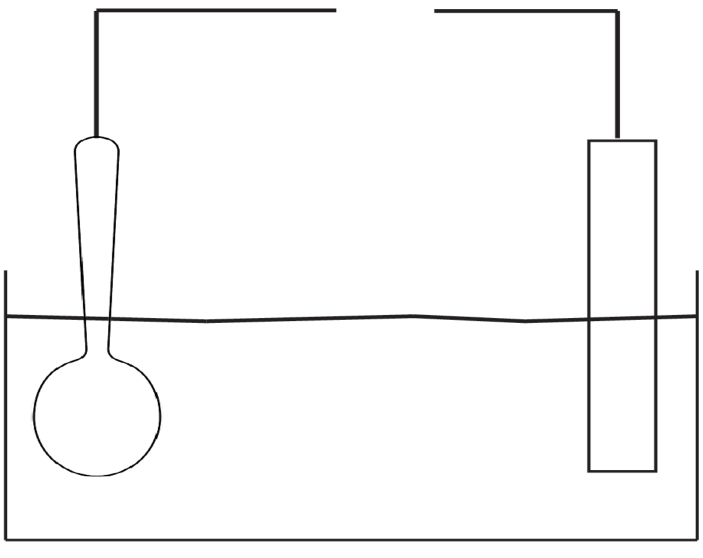
$$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s) \qquad -0.76\ V$$

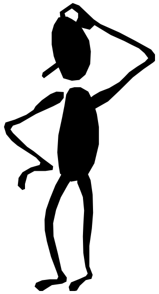
Set-up an electrolysis experiment designed to coat 2.50 grams of zinc on a copper metal spoon.

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.

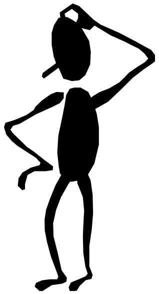




What current and what time is required to plate 2.50 grams of zinc?

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

$$2.50\,gZn\,x\left(\frac{1moleZn}{65.3g}\right)\left(\frac{2\ mol\,e^-}{1mole\,Zn}\right)=\left(0.0765\,mol\,e^-\right)$$



What current and what time is required to plate 2.50 grams of zinc?

Zn atom is oxidized

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

Zn^{2+} is reduced

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

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

2) Calculate charge (Coulombs):

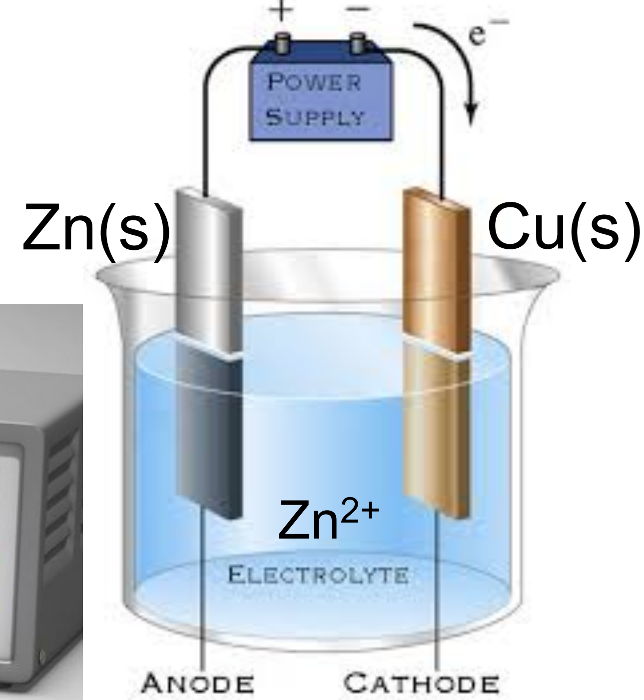
$$0.0765\,mol\,e^- \left(\frac{96{,}500\,C}{1\,mol\,e^-}\right) = 7.38\,x\,10^3\,C$$

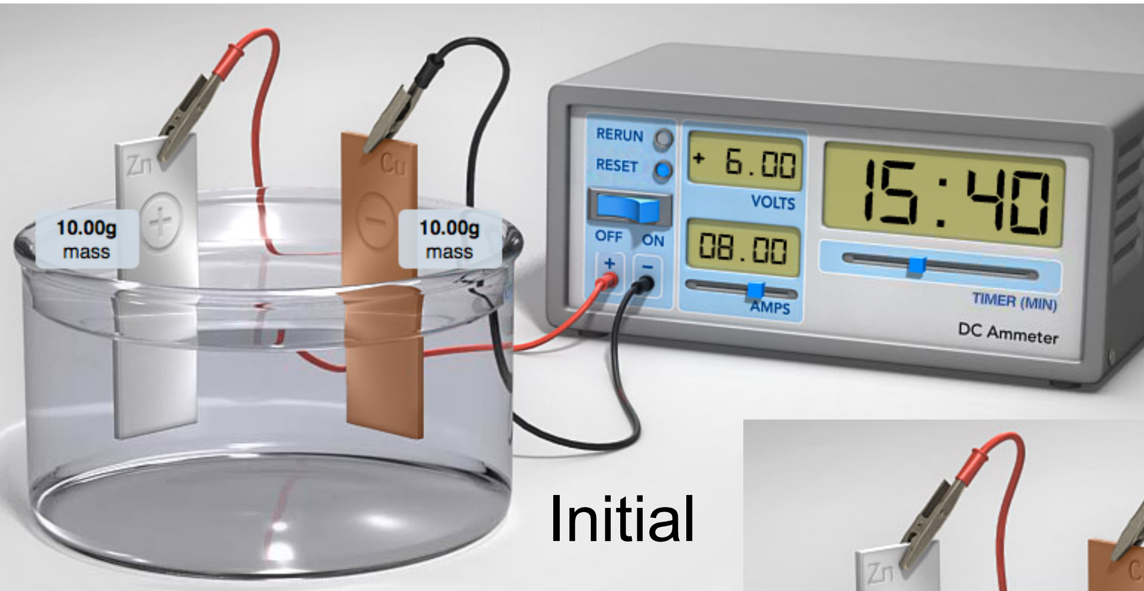
3) Calculate the time it takes with 8.00 Amps:

$$7.38\,x\,10^3\,C\,x\left(\frac{1\,Amp \cdot s}{1C}\right)\left(\frac{1}{8.00\,Amp}\right)\left(\frac{1\,\text{min}}{60.0\,\text{sec}}\right) = 15.3\,\text{min}$$

The reaction at the anode

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





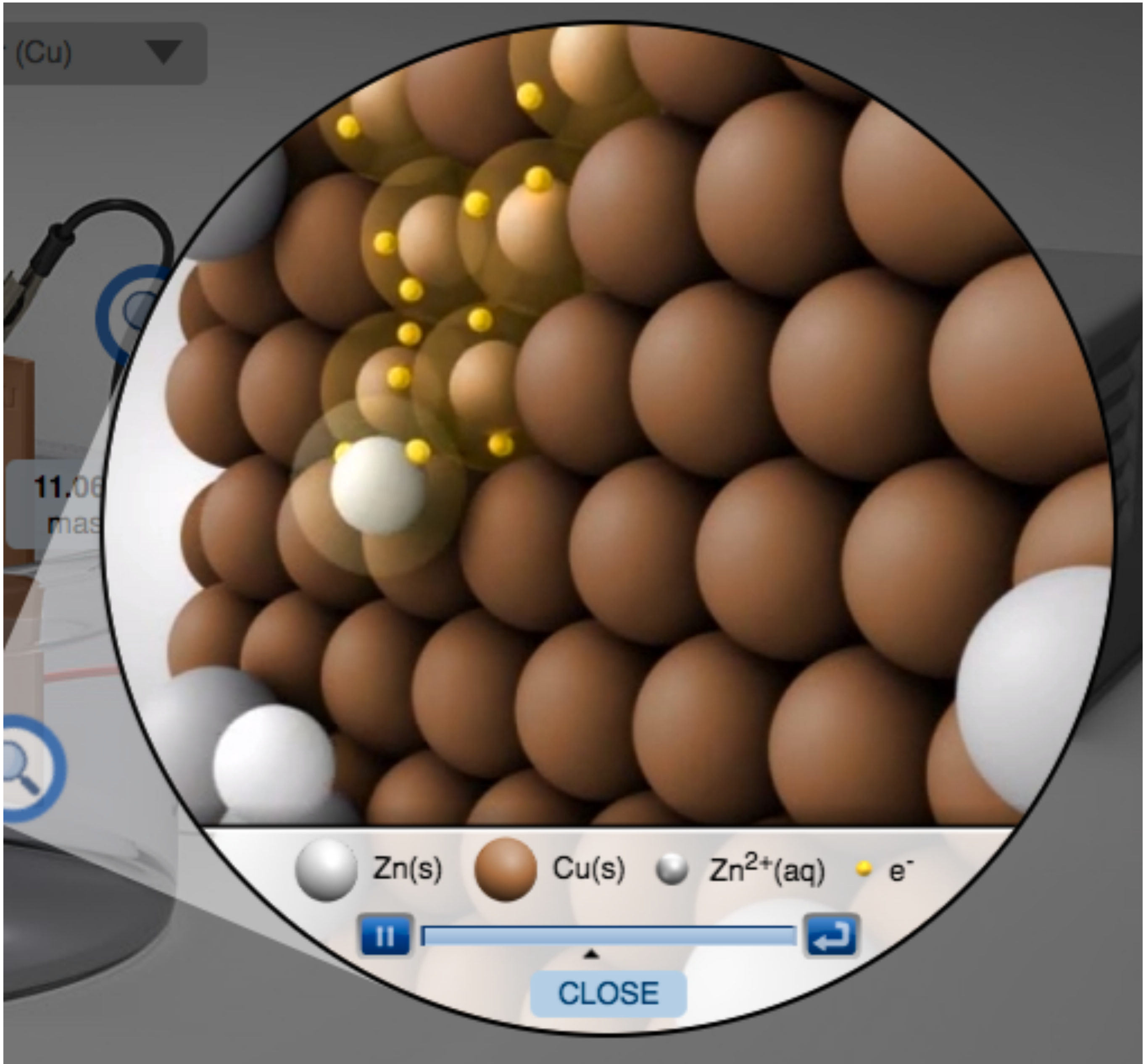
Initial

How much mass of Zn is deposited on the Cu electrode?

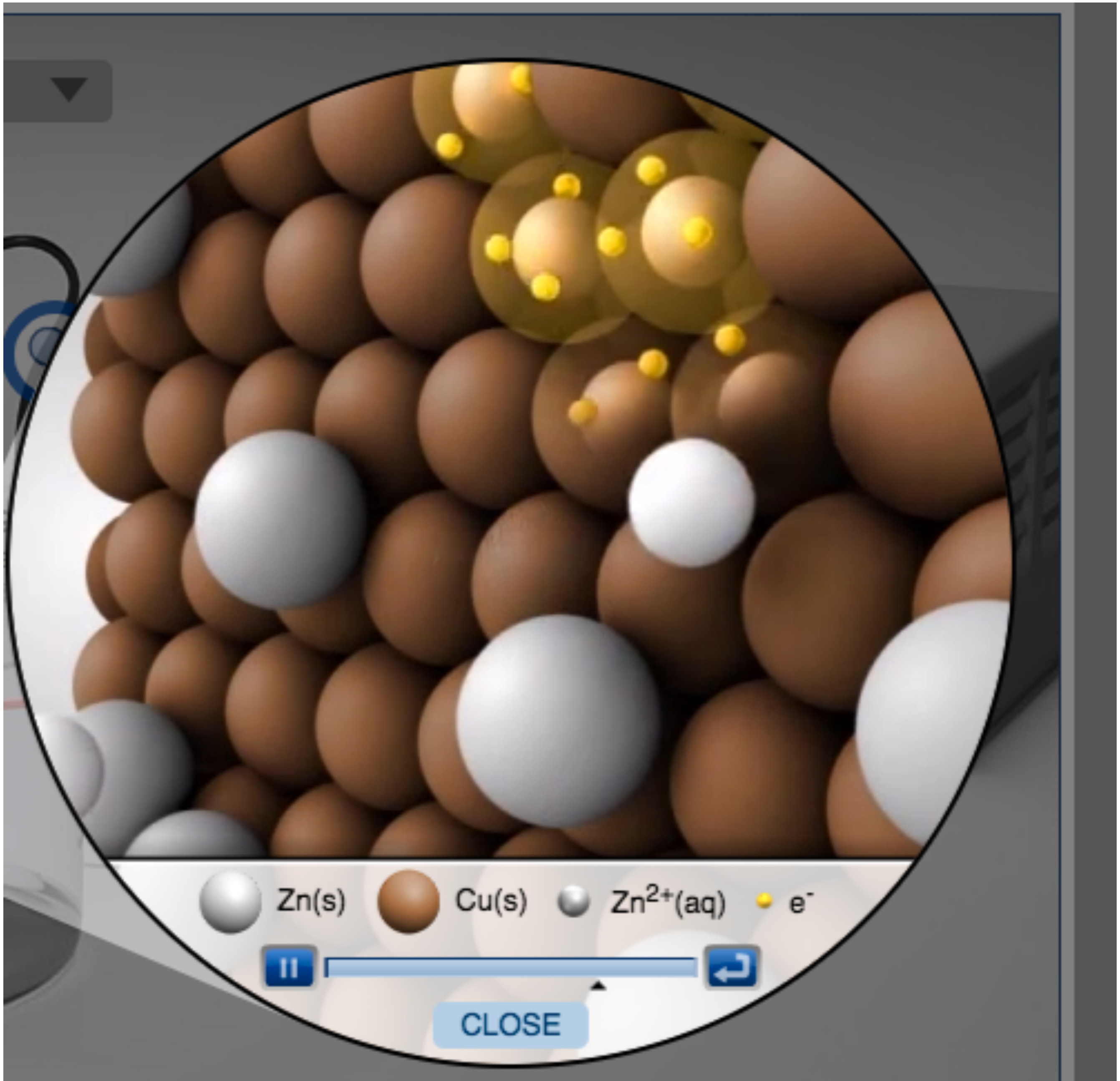
Final

Zinc metal plates onto the copper electrode.

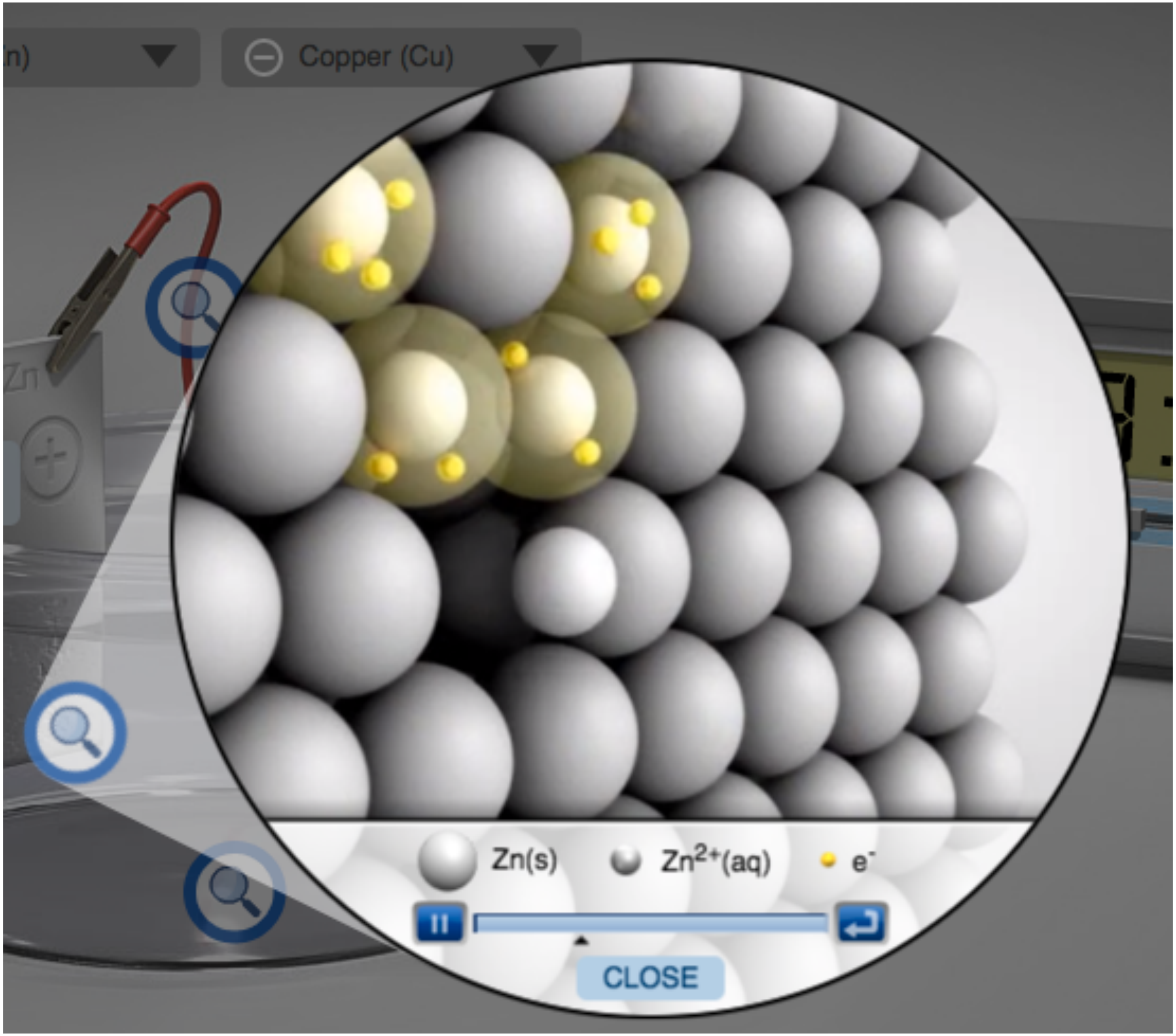
The reaction at the cathode $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$



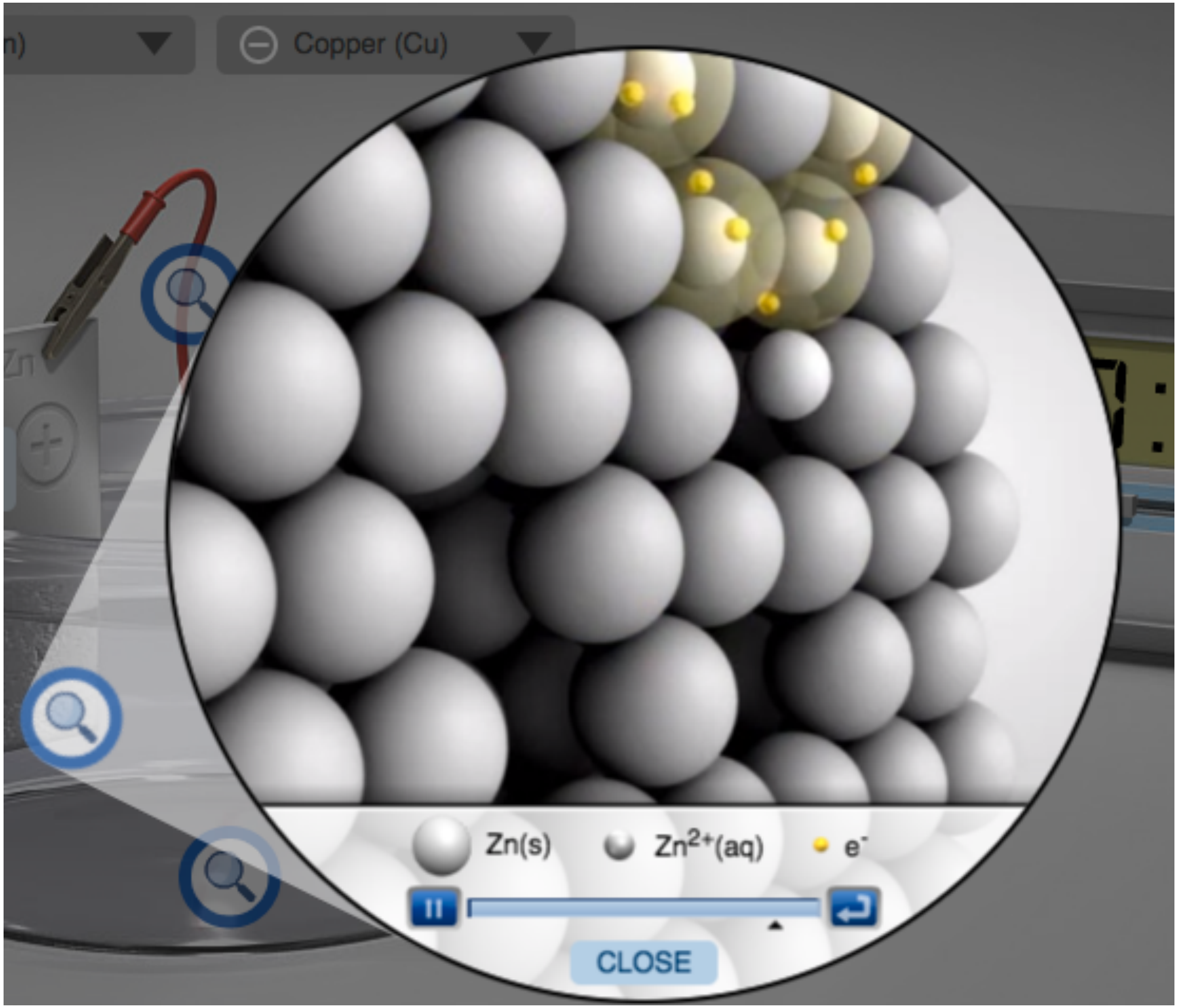
The reaction at the cathode $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$



The reaction at the cathode $Zn^{2+}(aq) + 2e- \rightarrow Zn(s)$



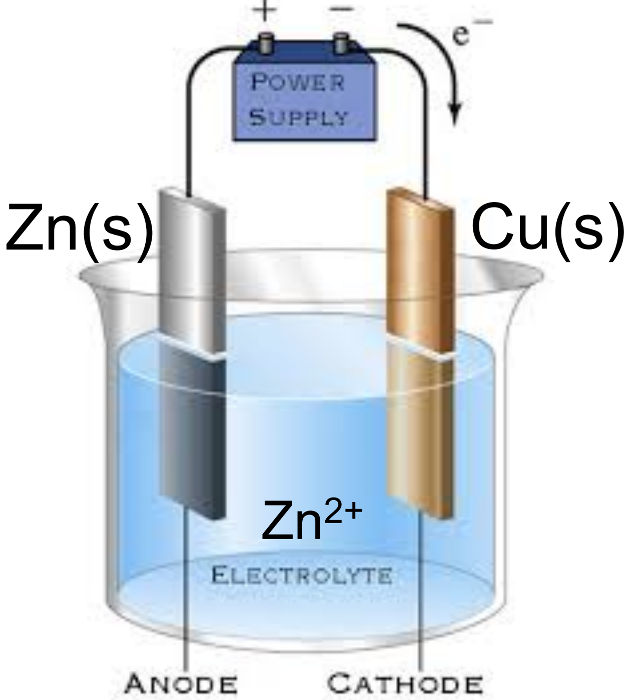
The reaction at the anode $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$



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

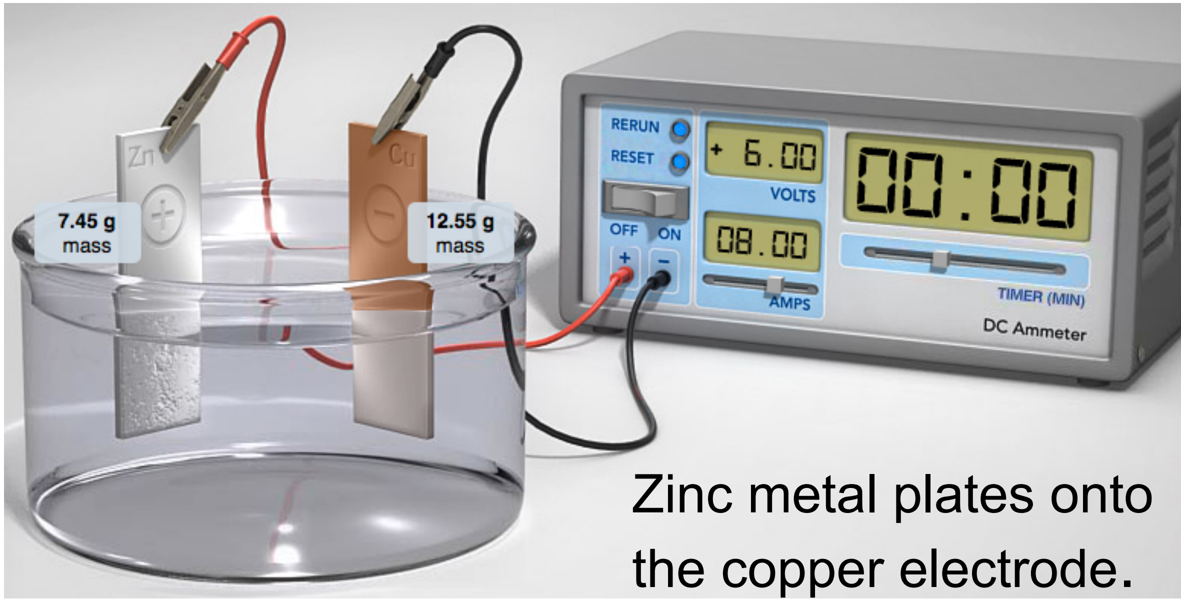
The reaction at the anode

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

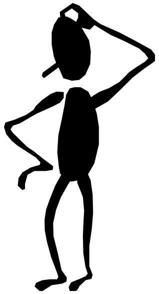


2.55 g of Zn deposited on the Cu electrode (cathode).

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



The reaction at the cathode $Zn^{2+}(aq) + 2e- \rightarrow Zn(s)$



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) → Cu^{2+}(aq) + **2 e$^-$**

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

Overall: Cu(s) + Cu^{2+}(aq) → Cu(s) + Cu^{2+}(aq)

1) Calculate charge (Coulombs):

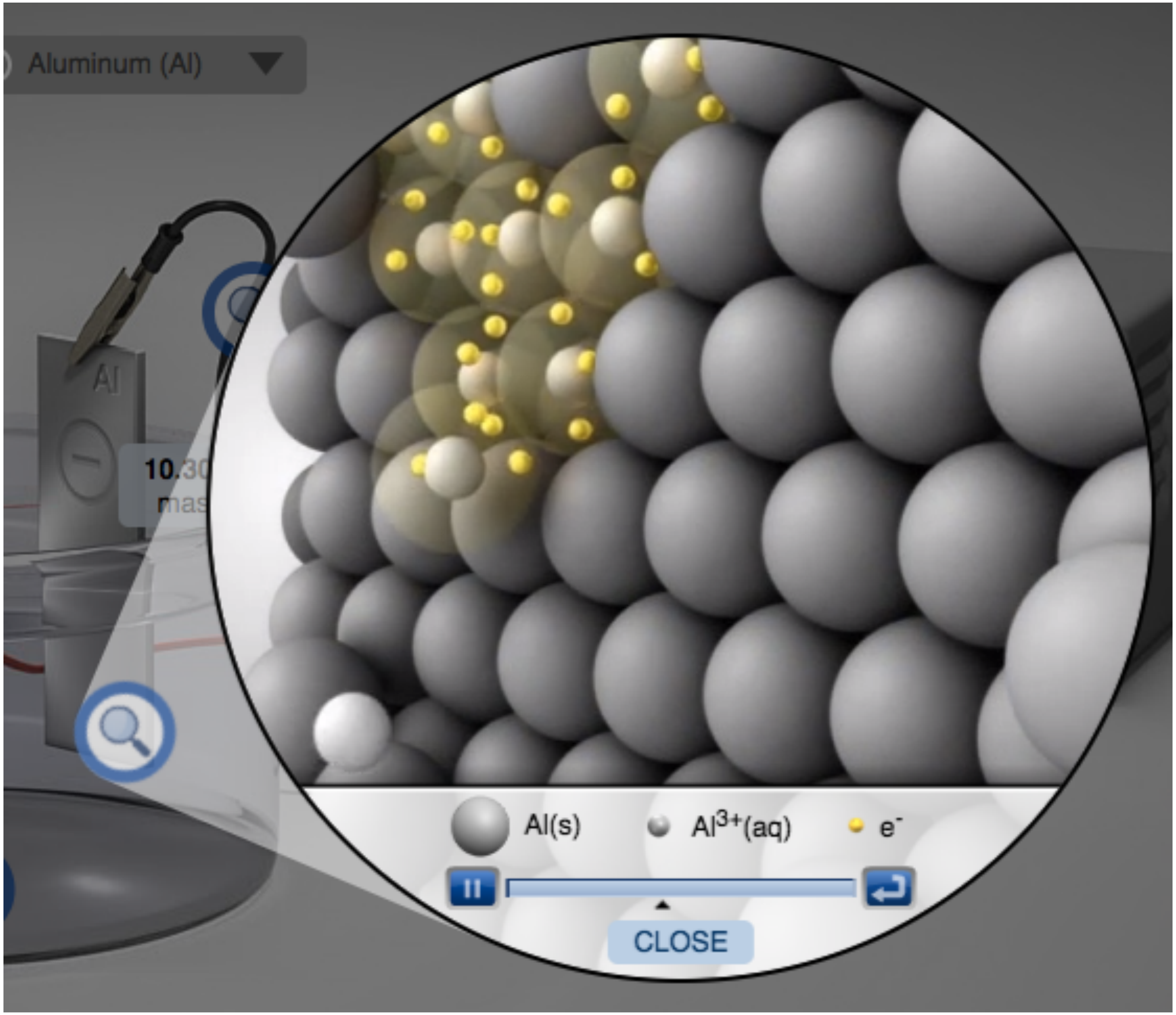
$$C = (8.00\,Amps)(15.00\,\text{min})\left(\frac{60.0\,s}{1\,\text{min}}\right)\left(\frac{1\,C}{1\,Amp\cdot s}\right) = 7.20\,x10^3\,C$$

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

$$7.20\,x10^3\,C\left(\frac{1\,mol\,e^-}{96,500\,C}\right) = 0.0746\,mol\,e^-$$

POGIL Activity 86 Al-Al Electrolytic Cell
8.00 amps at 15.00 minute.

What half-reaction occurs at the cathode?



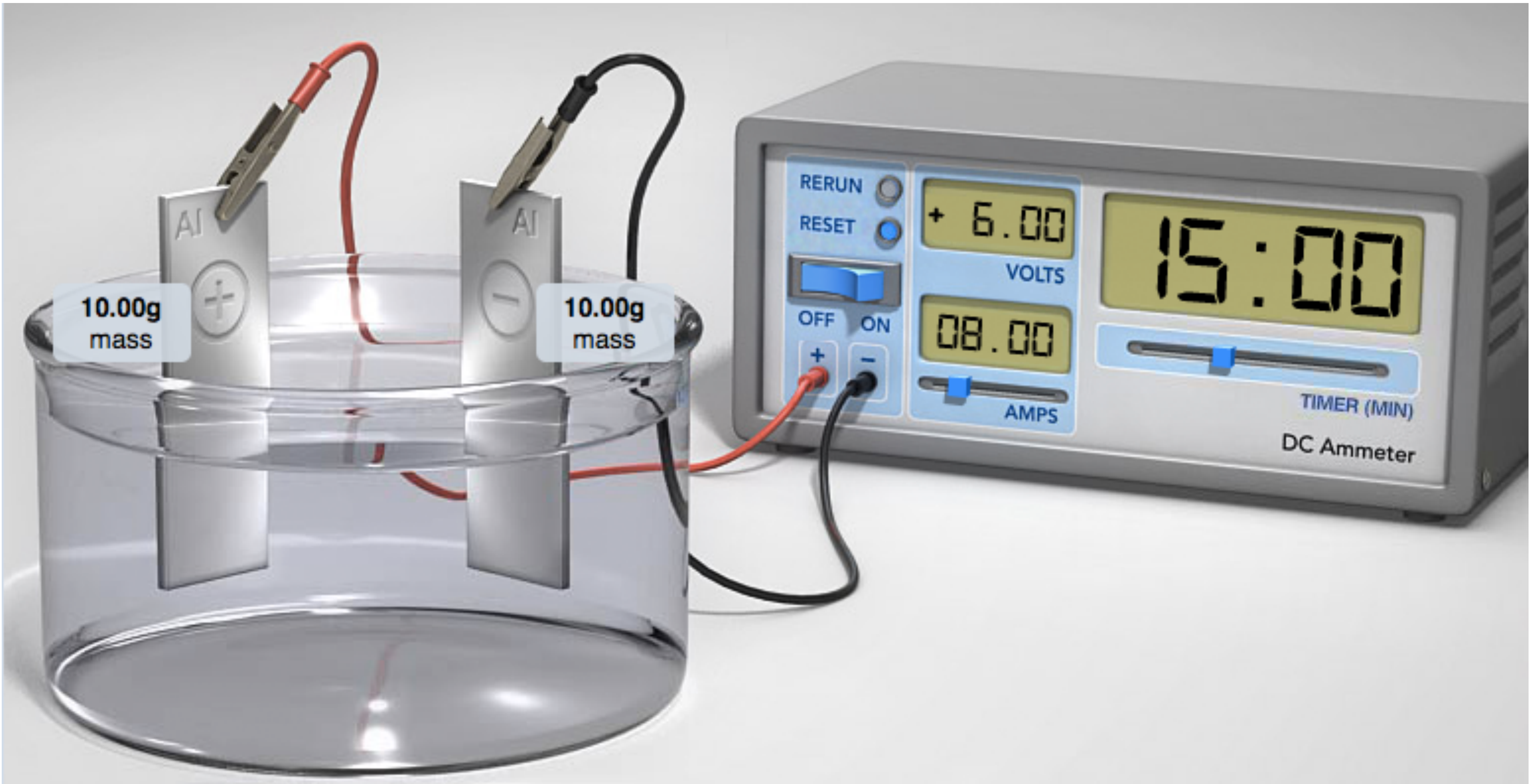
$Al^{3+}(aq) + 3e^{-} \rightarrow Al(s)$

POGIL Activity 86 Al-Al Electrolytic Cell
8.00 amps at 15.00 minute.

What half-reaction occurs at the cathode?

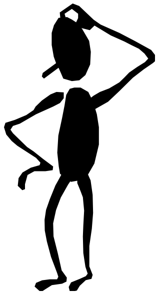
$Al^{3+}(aq) + 3e^{-} \rightarrow Al(s)$

POGIL Activity 86 Al-Al Electrolytic Cell
8.00 amps at 15.00 minute



How many moles of electrons will be transferred into this system?

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



What mass of aluminum is produced in 15.0 minutes by the electrolysis of Al^{3+}(aq) if the electrical current is 8.00 Amps?

Cu atom is oxidized

Cu^{2+} is reduced

Anode: Al(s) → Al^{3+}(aq) + **3 e$^-$**

Cathode: Al^{3+}(aq) + **3e$^-$** → Al(s)

Overall: Al(s) + Al^{3+}(aq) → Al(s) + Al^{3+}(aq)

1) Calculate charge (Coulombs):

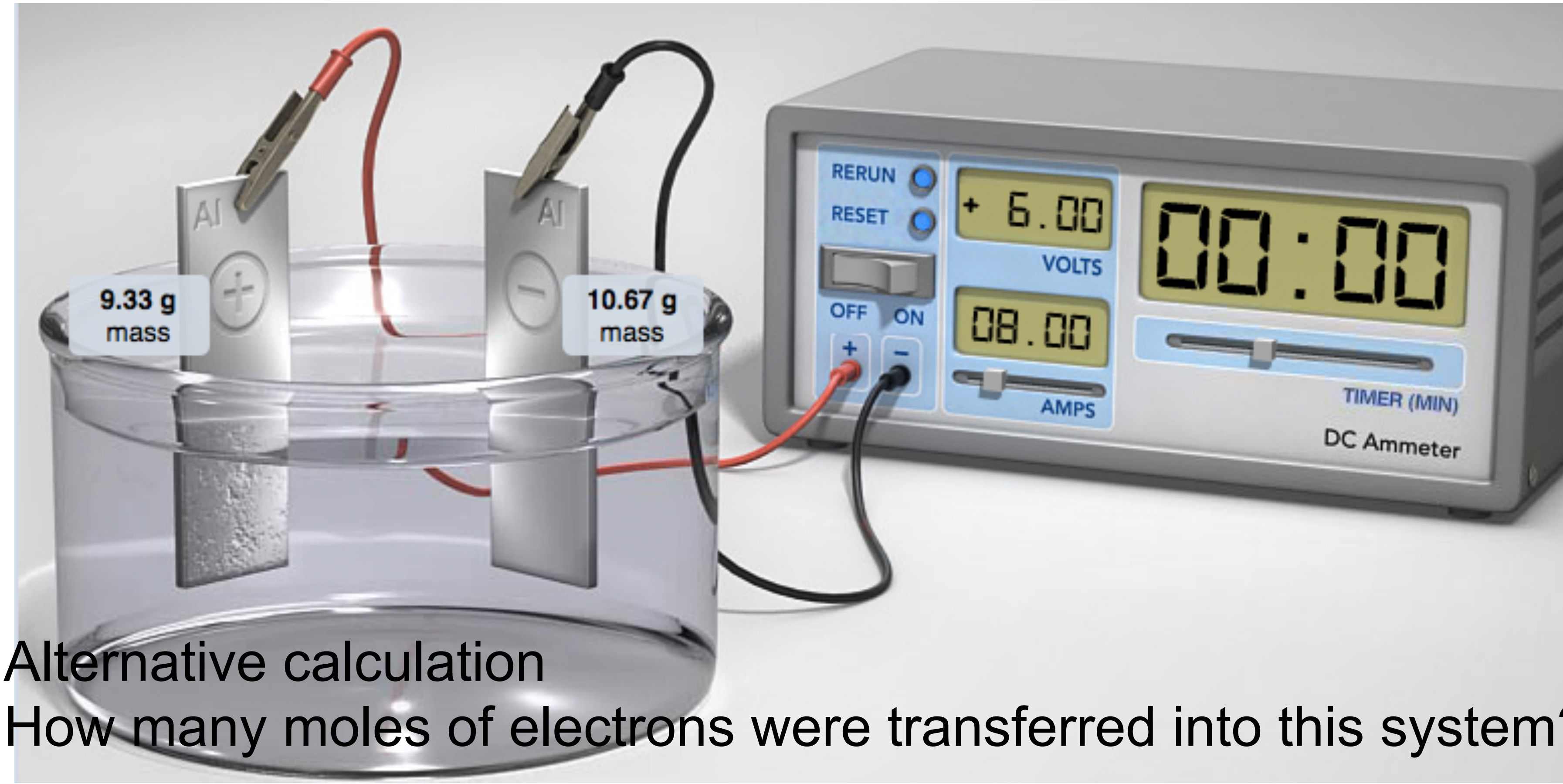
$$C = (8.00\,Amps)(15.00\,\mathrm{min})\left(\frac{60.0\,s}{1\,\mathrm{min}}\right)\left(\frac{1\,C}{1\,Amp\cdot s}\right) = 7.20\,x\,10^3\,C$$

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

$$7.20\,x\,10^3\,C\left(\frac{1\,mol\,e^-}{96,500\,C}\right) = 0.0746\,mol\,e^-$$

POGIL Activity 86 Al-Al Electrolytic Cell

8.00 amps at 15.00 minute.

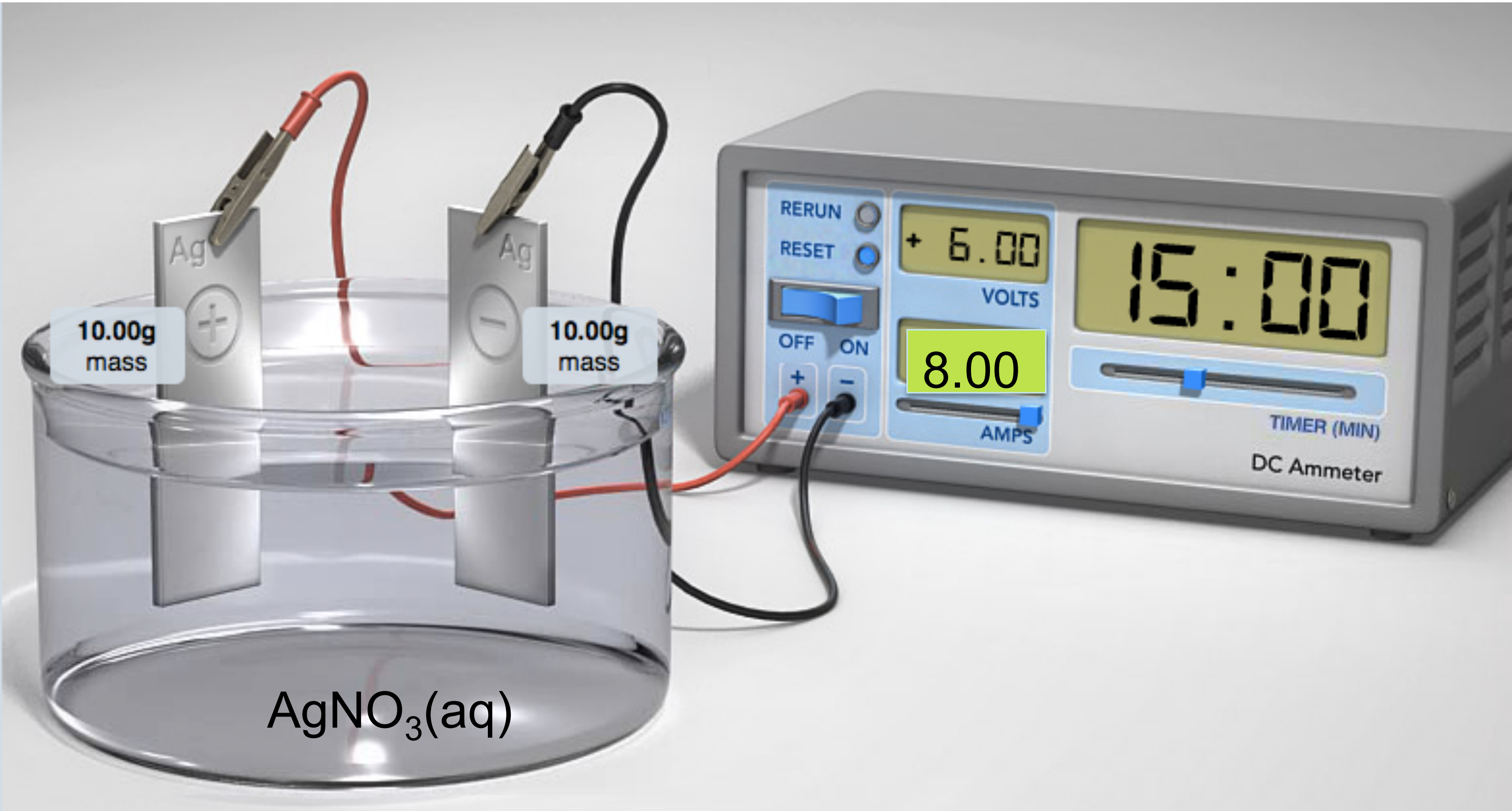


Alternative calculation

How many moles of electrons were transferred into this system?

$$0.67\,g\,Al\ x\left(\frac{1\,mole\,Al}{26.98\,g}\right)\left(\frac{3\ mol\,e^-}{1\,mole\,Al}\right)=\left(0.0744\,mol\,e^-\right)$$

If 8.00 amps flow through a Ag^{+}(*aq*) solution for 15.0 minutes during an AgAg electrolysis experiment, how many moles of e- are tansferred & what is the mass



Compare Electrolysis Experiments
8.00 Amps for 15.0 minutes
Complete the Table

Half-reaction	Current (A)	Time (s)	Moles e-	Moles Metal	Mass (g)
$Ag^{+} + e^{-} \rightarrow Ag$	8.00	900.0	0.0746	0.0746	
$Cu^{2+} + 2e^{-} \rightarrow Cu$	8.00	900.0	0.0746	0.0373	
$Zn^{2+} + 2e^{-} \rightarrow Zn$	8.00	900.0			
$Al^{3+} + 3e^{-} \rightarrow Al$	8.00	900.0			

Compare Electrolysis Experiments
Complete the Table

Half-reaction	Current (A)	Time (s)	Moles e-	Moles Metal	Mass (g)
$Ag^{+} + e^{-} \rightarrow Ag$	8.00	900.0	0.0746	0.0746	8.04 g Ag
$Cu^{2+} + 2e^{-} \rightarrow Cu$	8.00	900.0	0.0746	0.0373	2.37 g Cu
$Zn^{2+} + 2e^{-} \rightarrow Zn$	8.00	900.0	0.0746	0.0373	2.46 g Zn
$Al^{3+} + 3e^{-} \rightarrow Al$	8.00	900.0	0.0746	0.0248	0.669 g Al