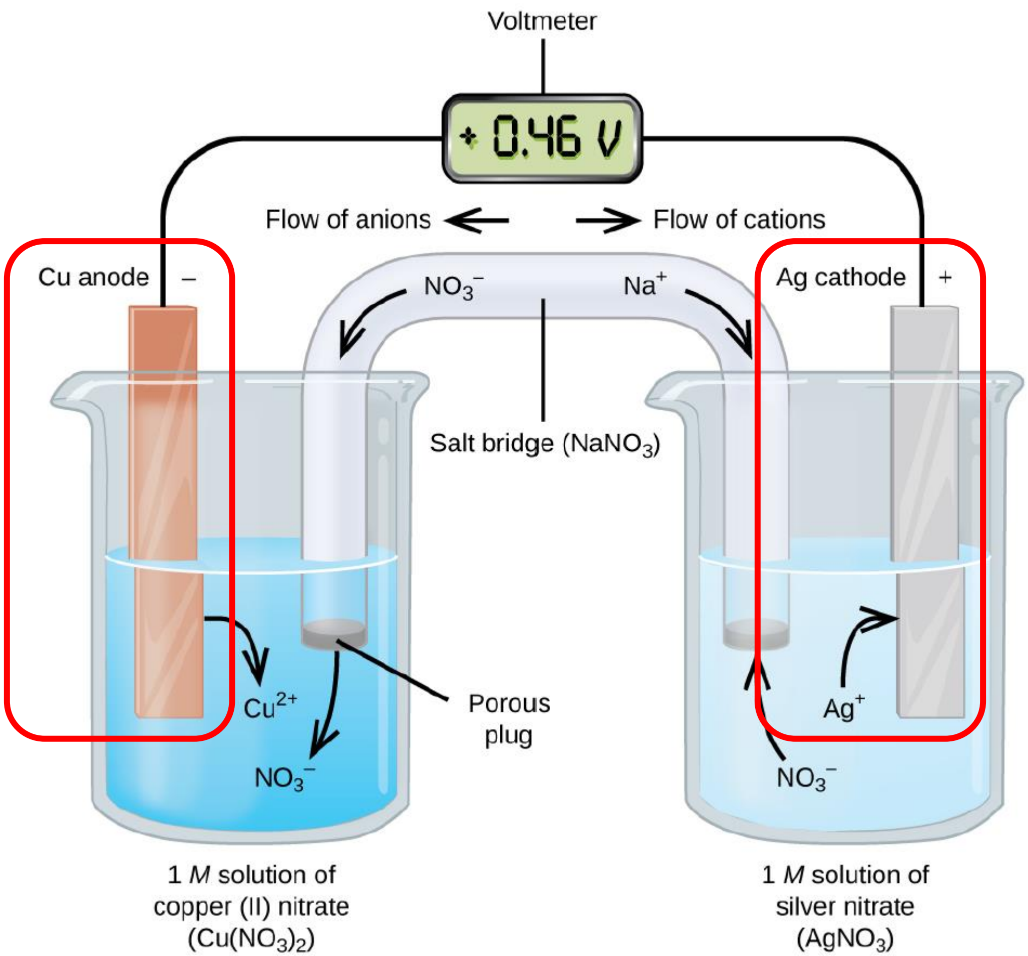
Chapter 17 Part 3

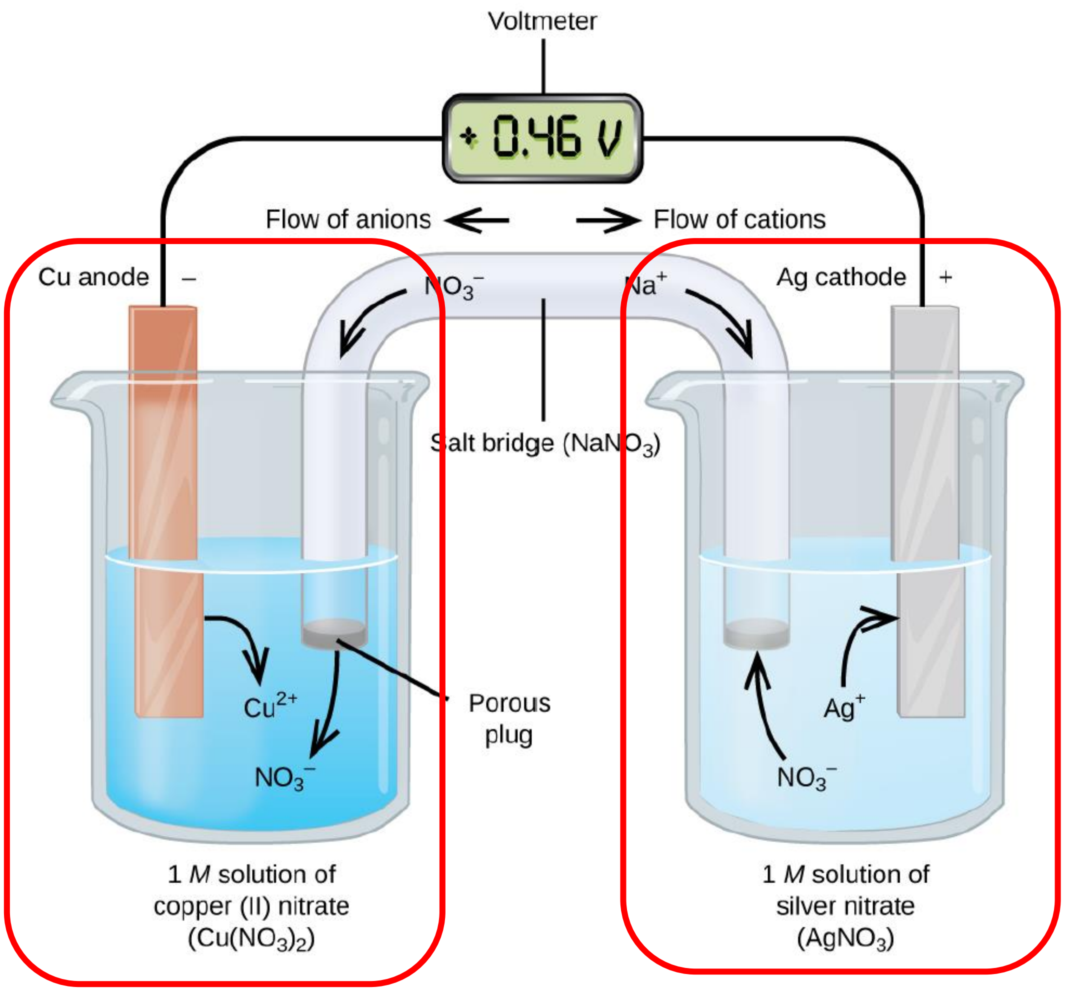
Dr. Turner

Electrodes



Electrodes are the strips of metal (M) in electrochemical cells

Half-Cell



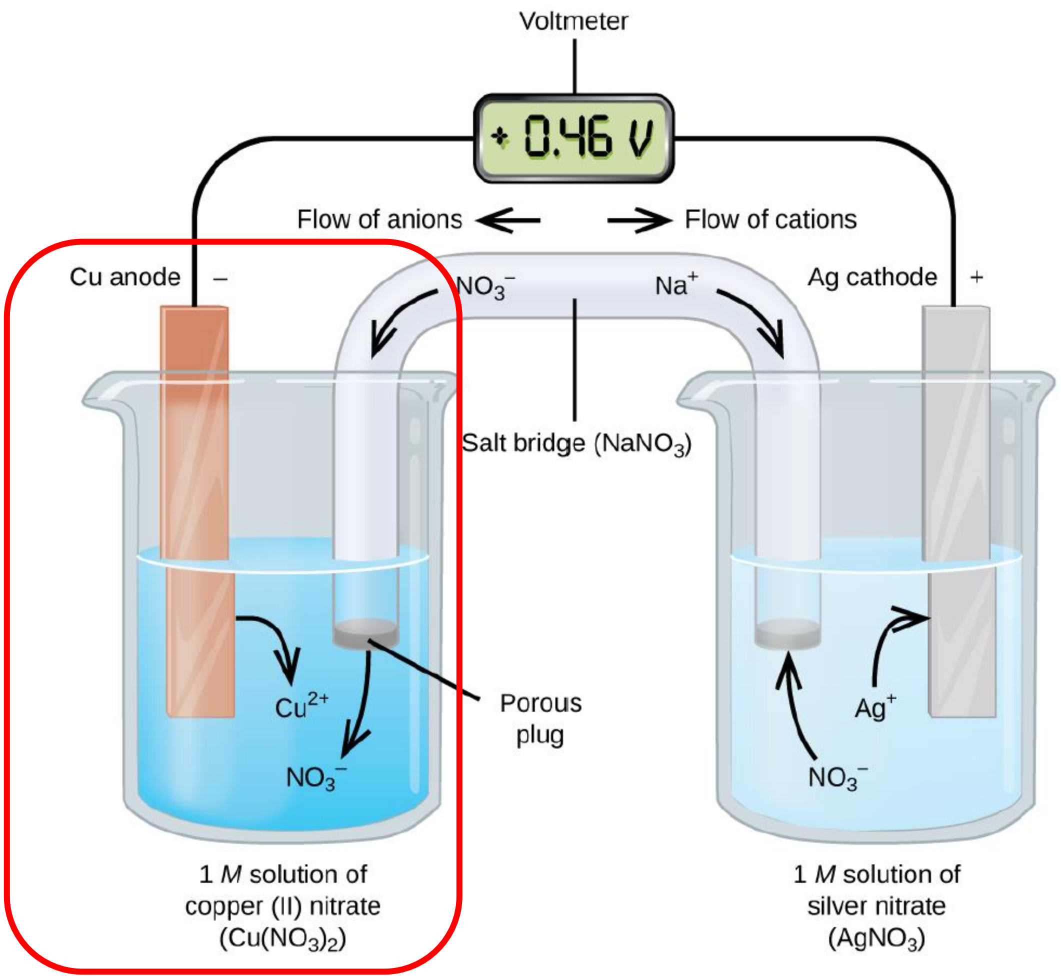
Half-cells are electrodes immersed in a solution containing ions of the same metal (M^{n+})

Anode

- The anode is the electrode where oxidation occurs
- Metal atoms, M, on the surface of the anode will loose n electrons and enter the solution as the ion M^{n+}

  $M(s) \rightarrow M^{n+}(aq) + ne^{-}$
- The anode is the half-cell on the left side of the electrochemical cell
- The half reaction at the circled anode is

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

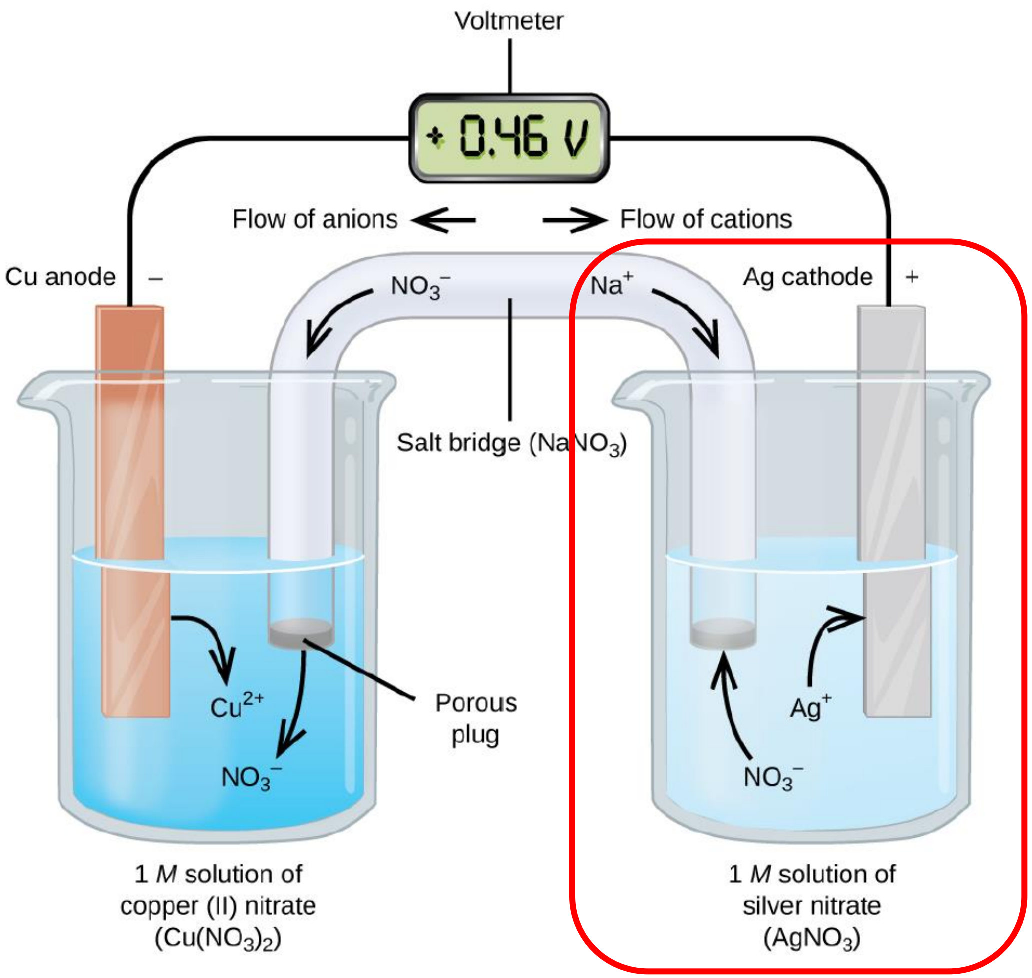


Cathode

- The cathode is the electrode where reduction occurs
- Metal ions in solution, M^{n+} collide with the cathode, gain n electrons, and are converted to the metal atom.

  $M^{n+}(aq) + ne^{-} \rightarrow M(s)$
- The cathode is the half-cell on the right side of the electrochemical cell
- The half reaction at the circled cathode is

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



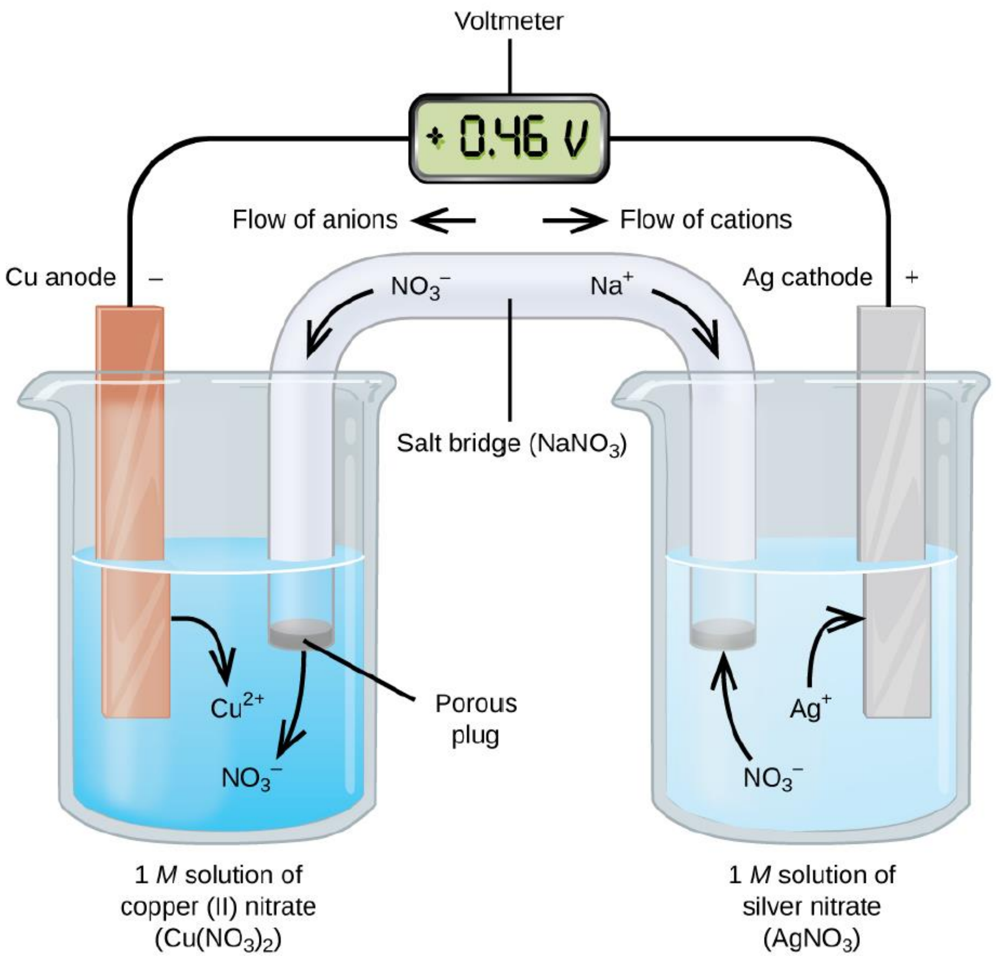
Electron Flow

- The combination of the two half reactions results in the equation

  $Cu(s) + 2\,Ag^{+}(aq)$

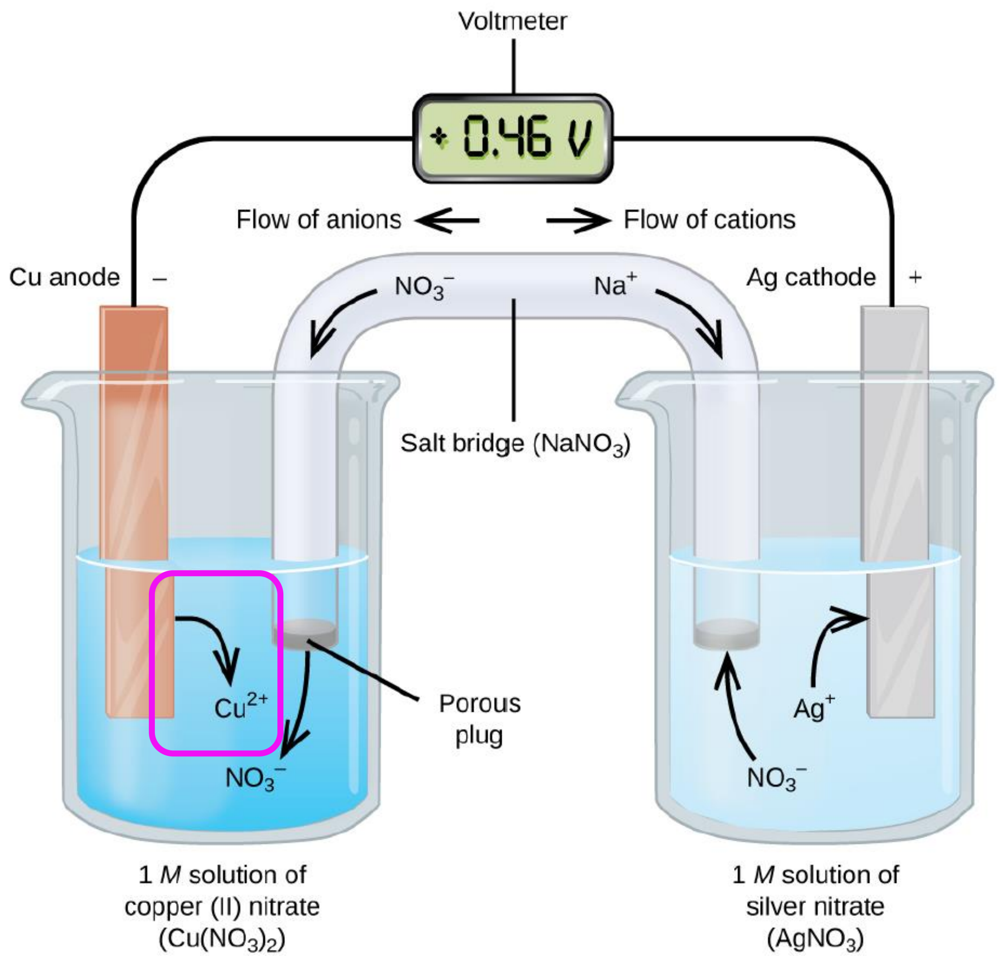
  $\rightarrow Cu^{2+}(aq) + 2\,Ag(s)$

- Thus, the electrons from the solid copper in the left cell must be transferred to the silver ions in solution in the right cell



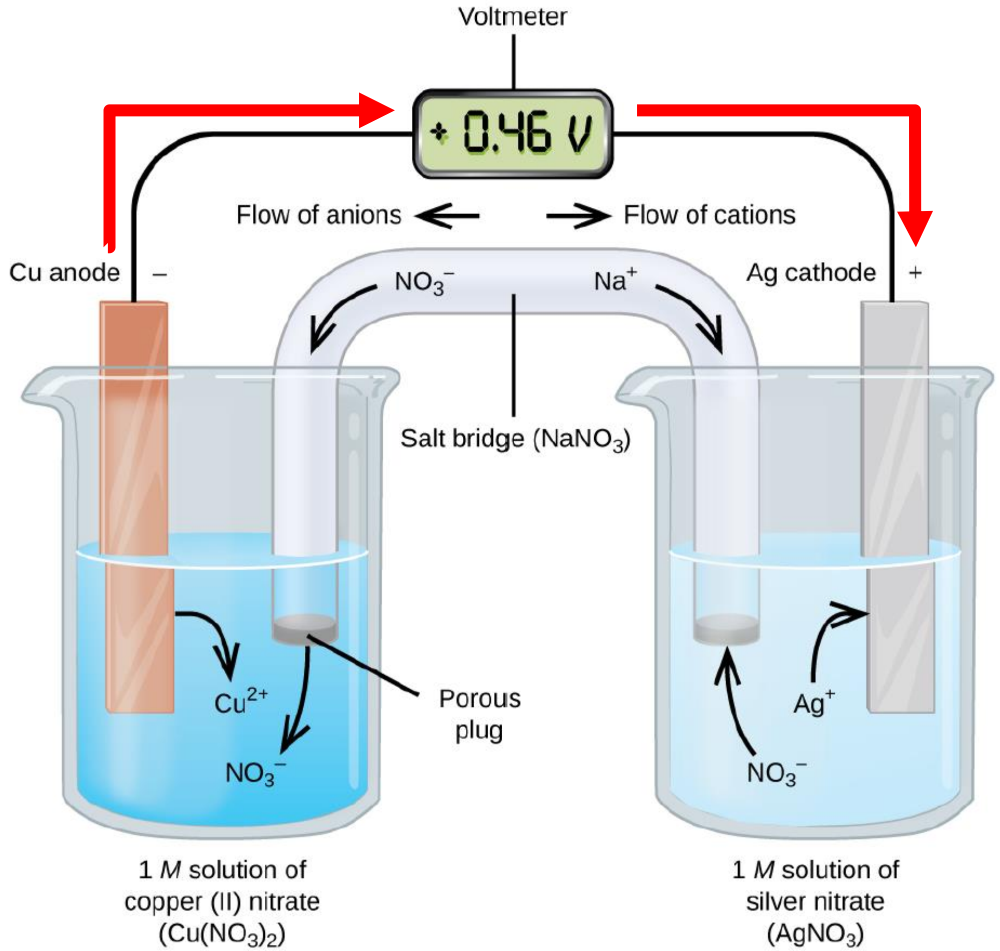
Electron Flow

1. The Cu atoms at the anode release electrons and enter the $Cu(NO_3)_2(aq)$ as Cu^{2+} ions



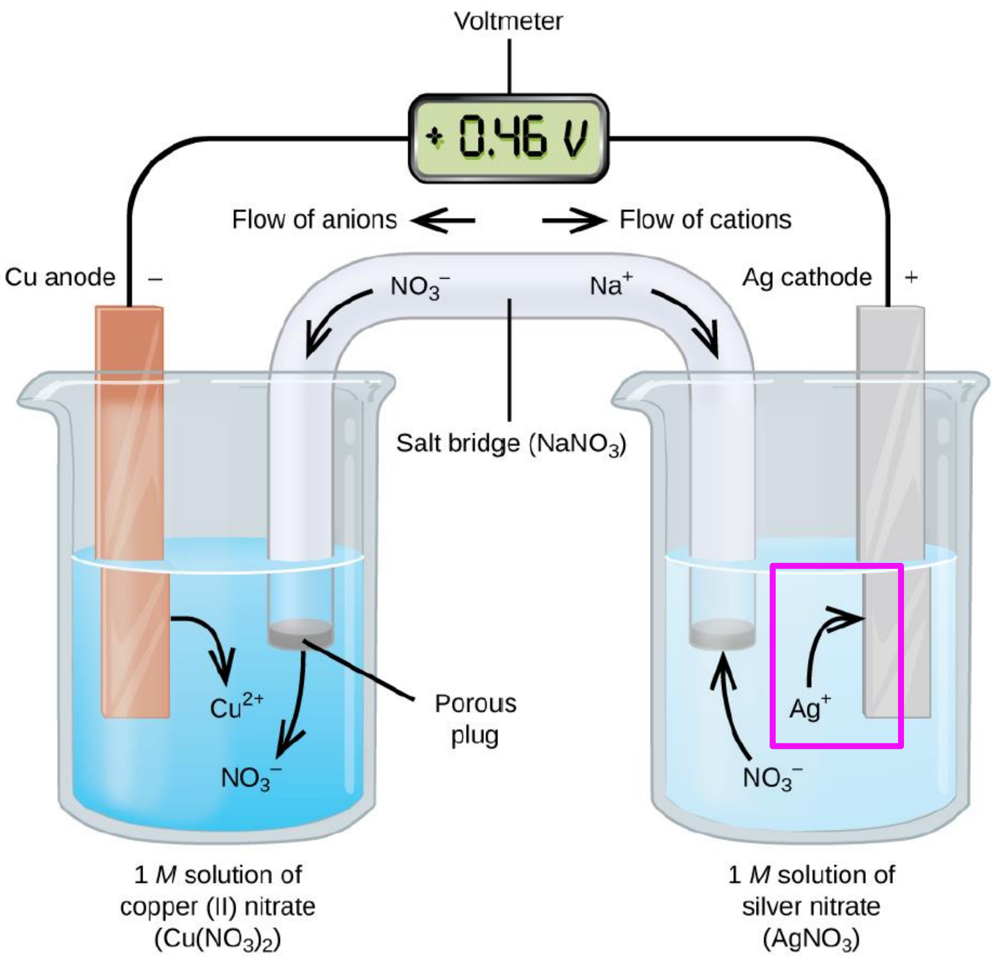
Electron Flow

1. The Cu atoms at the anode release electrons and enter the $Cu(NO_3)_2(aq)$ as Cu^{2+} ions
2. Electrons lost by the Cu atoms pass through the wires and the voltmeter to the Ag cathode



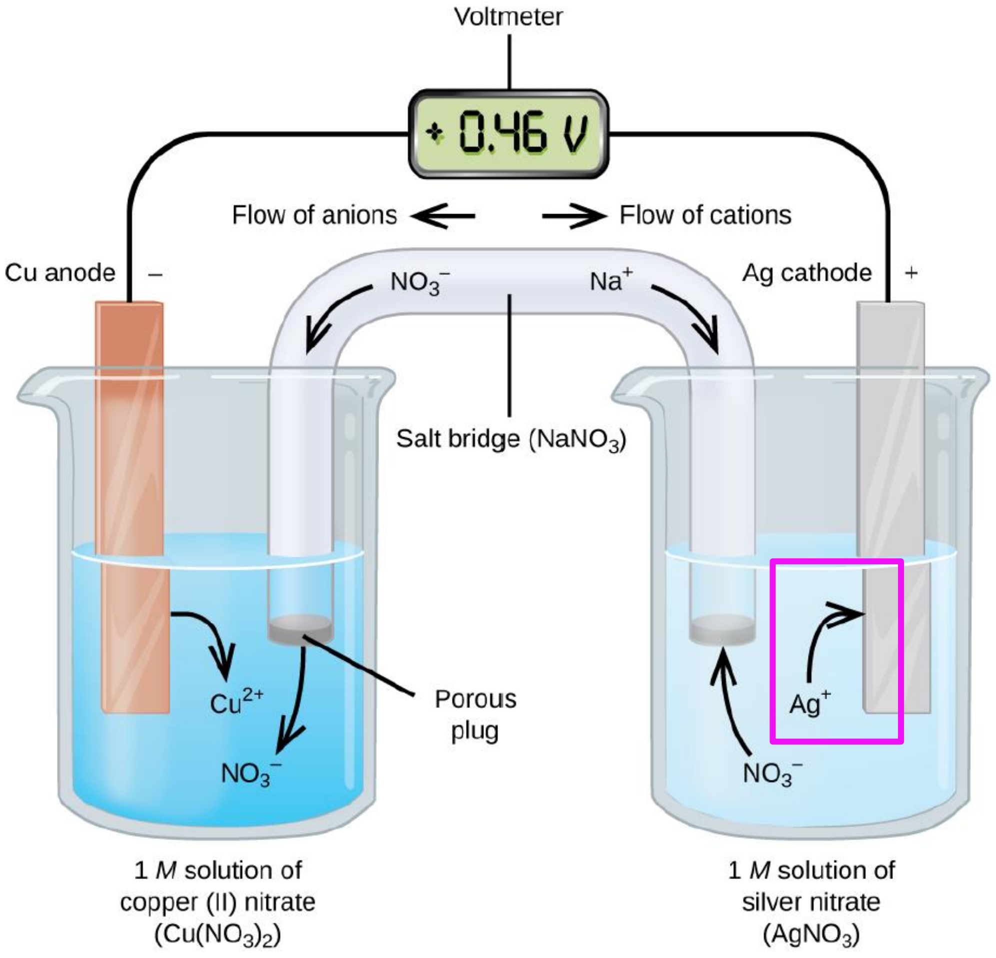
Electron Flow

1. The Cu atoms at the anode release electrons and enter the $Cu(NO_3)_2(aq)$ as Cu^{2+} ions
2. Electrons lost by the Cu atoms pass through the wires and the voltmeter to the Ag cathode
3. The Ag^+ ions from the $AgNO_3(aq)$ gain these electrons and produce solid Ag.



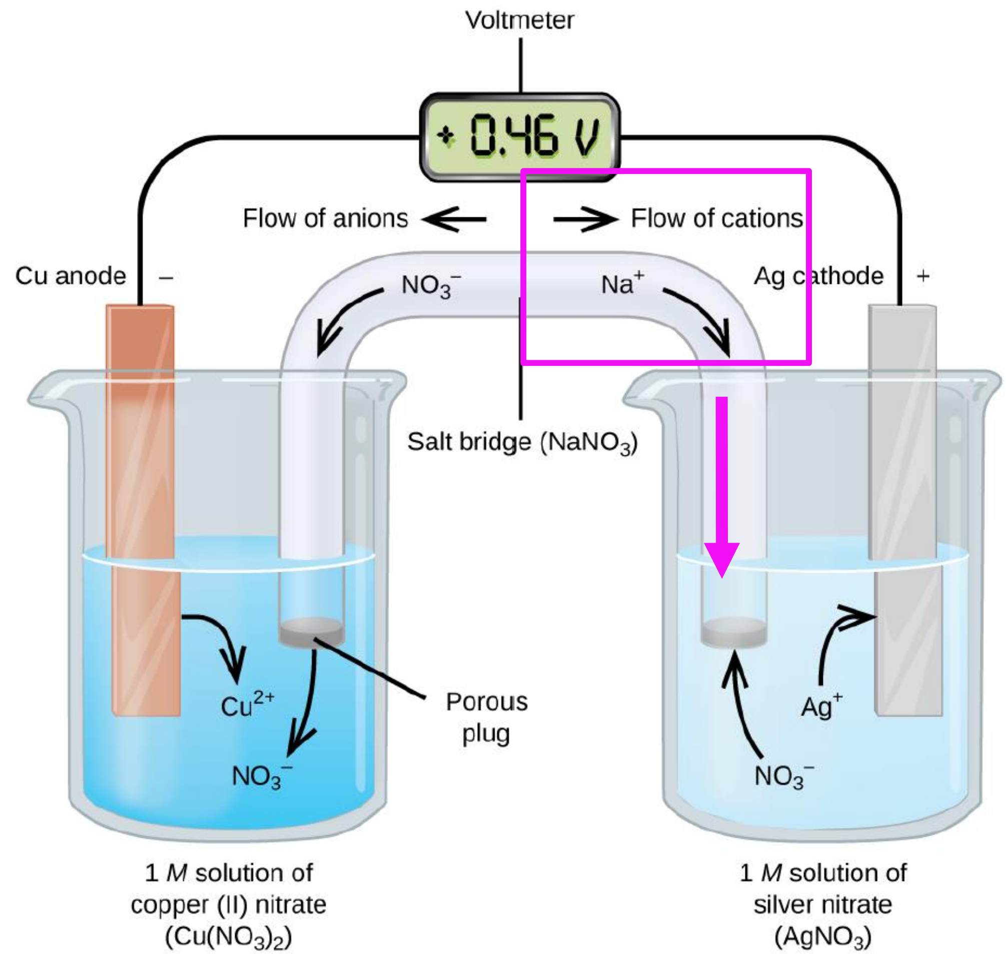
Electron Flow

4. The NO_3^- anions from the salt bridge migrate into the Cu half cell to neutralize the positive charge of the excess Cu^{2+} ions



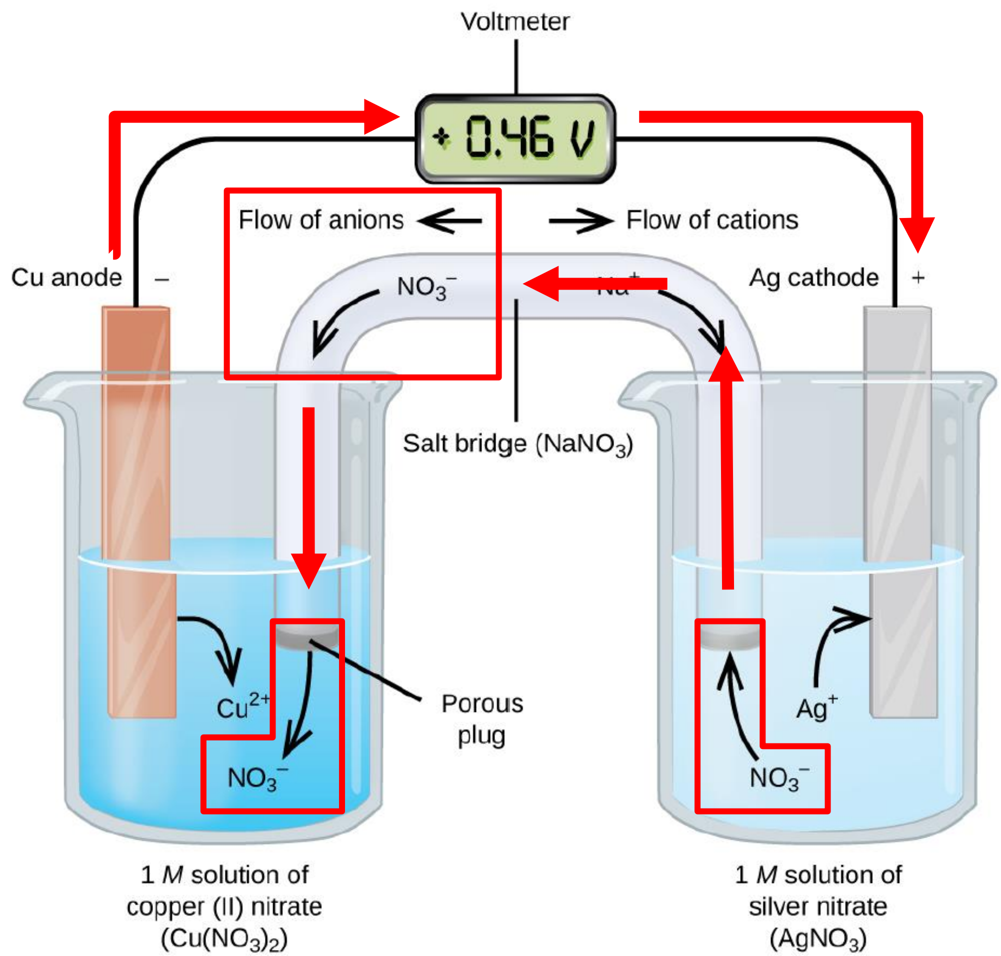
Electron Flow

4. The NO_3^- anions from the salt bridge migrate into the Cu half cell to neutralize the positive charge of the excess Cu^{2+} ions
5. Simultaneously, Na^+ cations migrate into the Ag half cell to neutralize the negative charge of the excess NO_3^- ions.



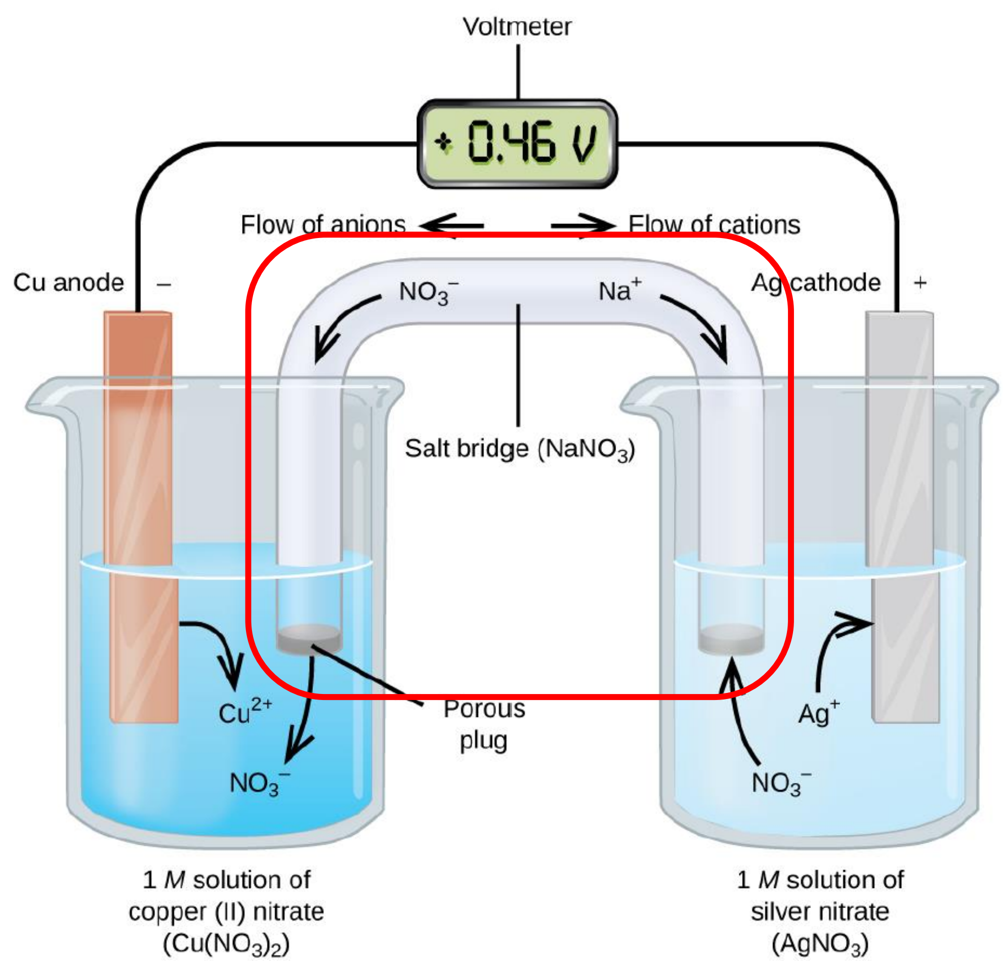
Electric Circuit

- The electric circuit is the path of the electrons through the cell



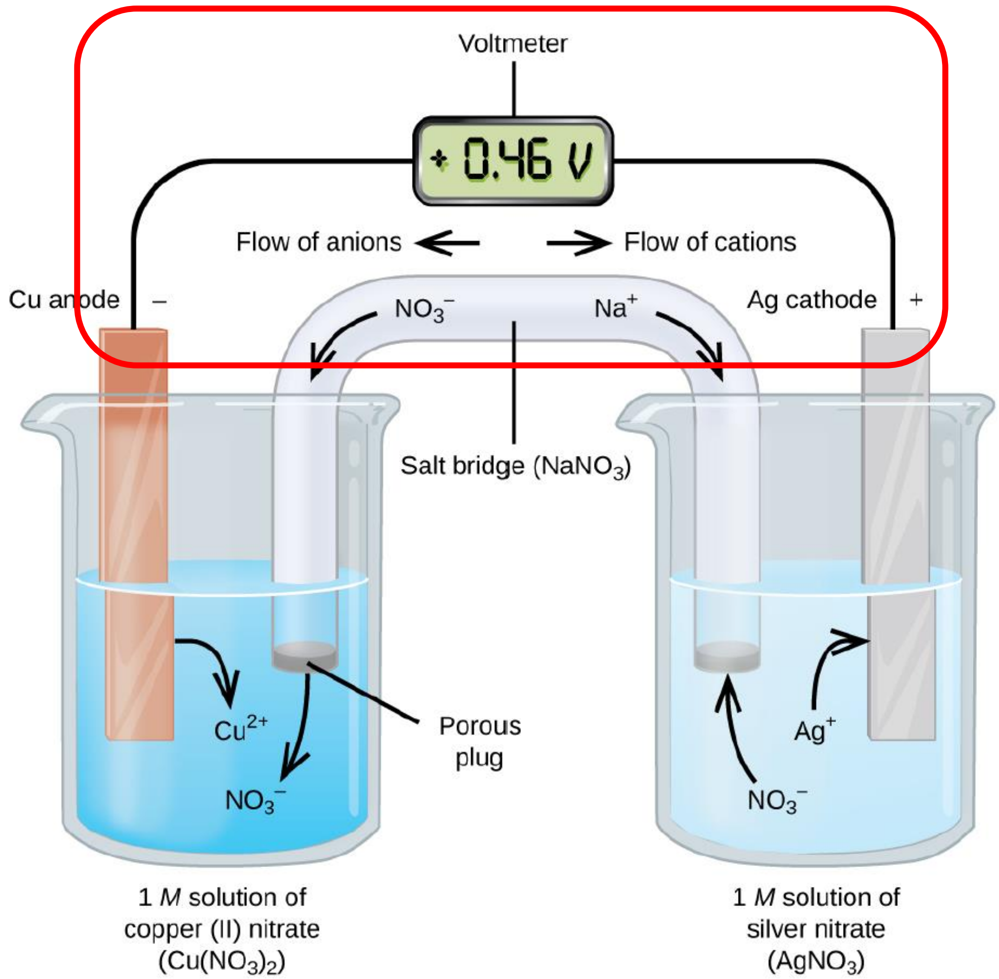
Salt Bridge

- Since the charge is carried through solution by the migration of ions, a wire cannot be used for this connection. Thus, a U-shaped tube called a salt bridge is used.



Voltmeter

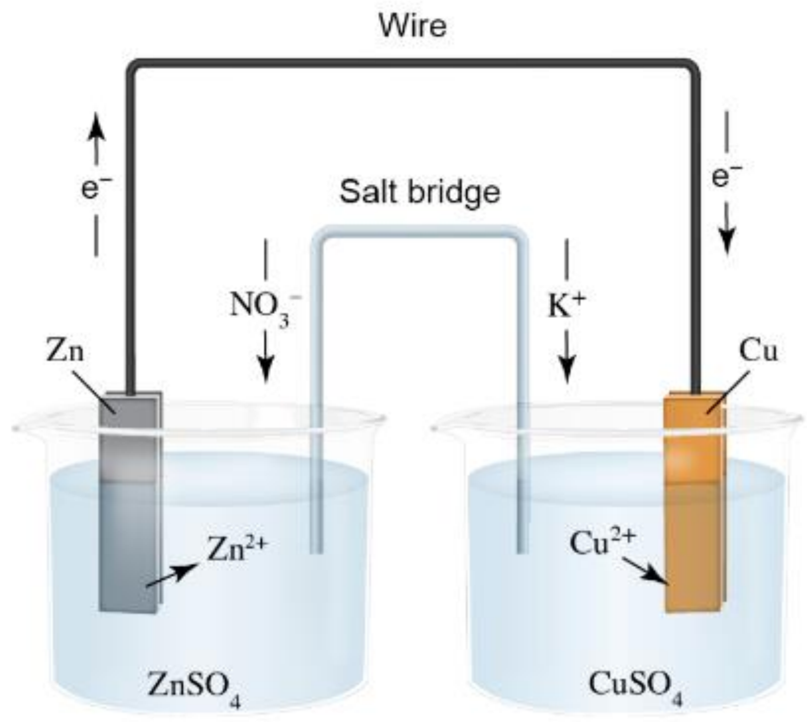
- The voltmeter measures the electric force trying to move the charge.
- It reads 0.46 V because of the inherent differences in the nature of the two materials used to make the half cell



Identify the cathode.

Identify the cathode.

A. solid zinc
B. zinc sulfate solution
C. solid copper
D. copper sulfate solution



Galvanic Cells

- Galvanic cells (also called voltaic cells) are electrochemical cells in which spontaneous redox reactions produce electrical energy
- A redox reaction is spontaneous when it's cell potential (E_{cell}) is a positive value
- The copper and silver cell from the previous slides is spontaneous because it has a cell potential of +0.46 V

Cell Notation

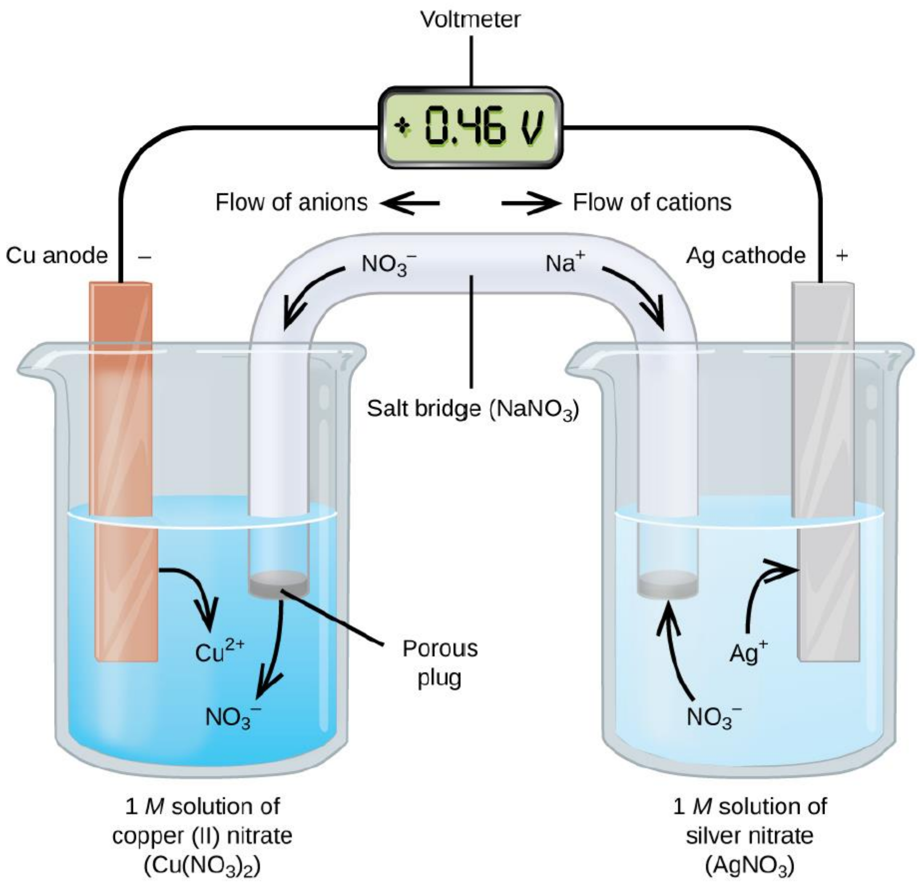
- Cell notation is a shorthand notation used to describe galvanic cells

Anode|Anode sol'n||Cathode sol'n|Cathode

$$Cu(s)|Cu^{2+}(aq)||Ag^{+}(aq)|Ag(s)$$

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

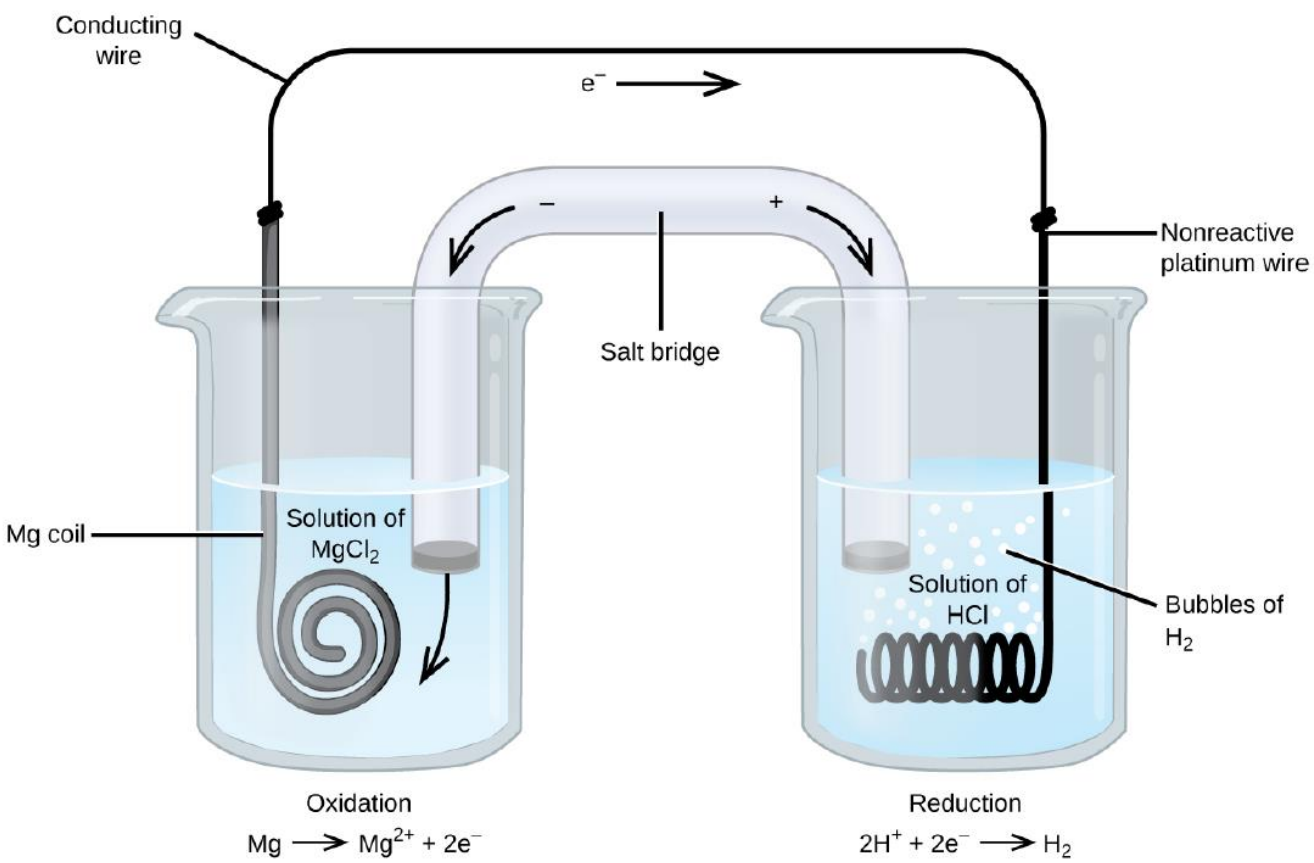
- Since both copper and silver participate in the reaction, they are active electrodes



Cell Notation

Aluminum metal displaces zinc(II) ion from aqueous solution. Write oxidation and reduction half-reactions and an overall equation for this redox reaction. Write a cell-diagram (cell notation) for a voltaic cell in which this reaction occurs.

Inert Electrode



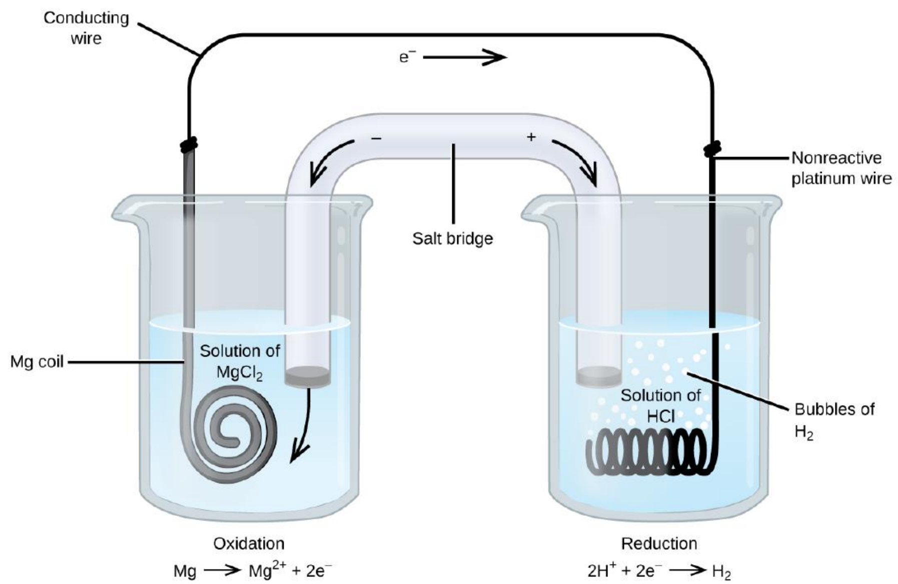
Inert Electrode

- Inert electrodes, like the platinum electrode on the right half-cell, do not participate in bonding and are present so that the current can flow through the cell.

$$Mg(s)|Mg^{2+}(aq)||H^{+}(aq)|H_2(g)|Pt(s)$$

$$Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

- Since platinum doesn't participate in the reaction, it is an inactive active electrode.



Standard Cell Potentials

- E°_{cell} is the standard cell potential for an electrochemical cell
- E°_{cell} is an intensive property as does not change with the amount of metals present in the cell.

$$E^{\circ}_{cell} = E^{\circ}_{red}(\text{reduction}) + E^{\circ}_{ox}(\text{oxidation})$$

- E°_{red}(reduction) is the reduction potential of the reduction reaction
- E°_{ox}(oxidation) is the oxidation potential of the oxidation reaction

Reduction Potentials

For the reaction: $Cu(s) + 2\ Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2\ Ag(s)$

- The reduction half reaction is:

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

- A table of reduction potentials reads:

$$Ag^{+}(aq) + e^{-} \rightarrow Ag(s) \quad E^{\circ} = +0.799\ v$$

- Thus, the reduction potential of the reduction half reaction is +0.799 v

Reduction Potentials

For the reaction: $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

- The oxidation half reaction is:

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

- A table of reduction potentials reads:

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s) \qquad E^{\circ} = +0.337\ v$$

- Thus, the reduction potential of the oxidation half reaction is $+0.337$ v
- Accordingly, the oxidation potential of the oxidation half reaction is $-\ 0.337$ v

$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-} \qquad E^{\circ} = -\ 0.337\ v$$

Standard Cell Potentials

For the reaction: $Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$

- Thus the standard cell potential (E°_{cell}) for the electrochemical cell is

$$E^{\circ}_{cell} = E^{\circ}_{red}(\text{reduction}) + E^{\circ}_{ox}(\text{oxidation})$$

$$E^{\circ}_{cell} = 0.799\ \text{v} - 0.337\ \text{v} = 0.462\ \text{v}$$

Standard Cell Potentials

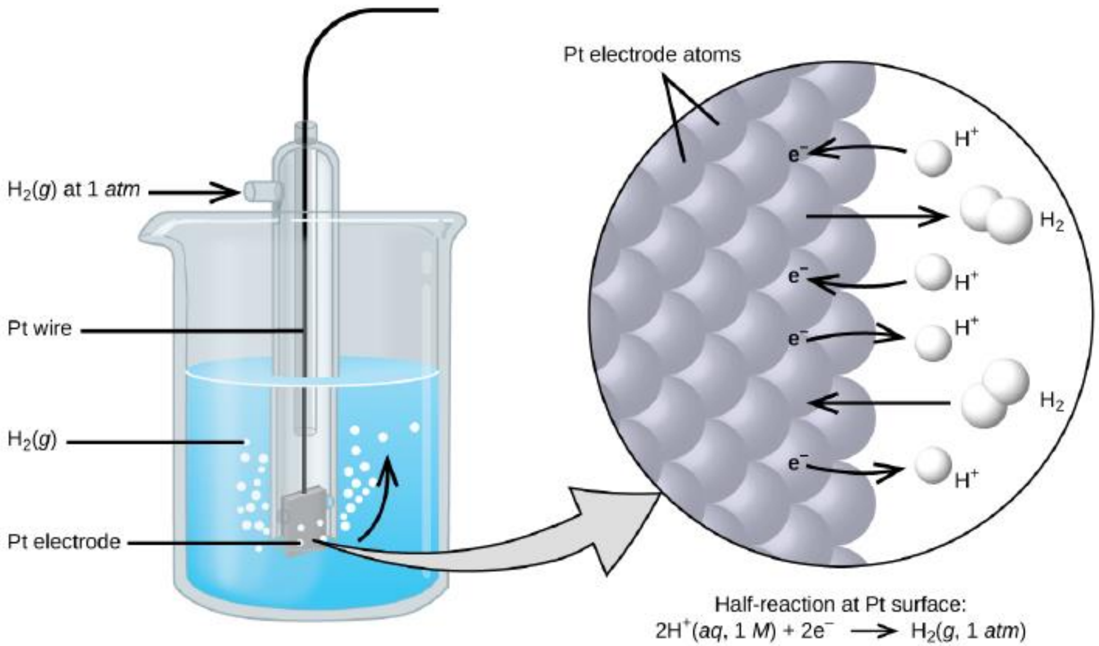
The reaction in a voltaic cell is $Zn(s) + Cl_2(g) \rightarrow ZnCl_2(aq)$. What is E°_{cell} of this voltaic cell? $E^\circ_{Zn^{2+}/Zn} = -0.763$ v. $E^\circ_{Cl_2/Cl^-} = 1.358$ v.

Standard Hydrogen Electrode

- The standard hydrogen electrode (SHE) consists of 1 atm of hydrogen gas bubbled through 1 M HCl solution.
- Platinum, which is chemically inert, is used as the electrode

$$Pt(s)|H_2(g)|H^+(aq)||$$

- This reaction is used as a reference for other reduction reactions

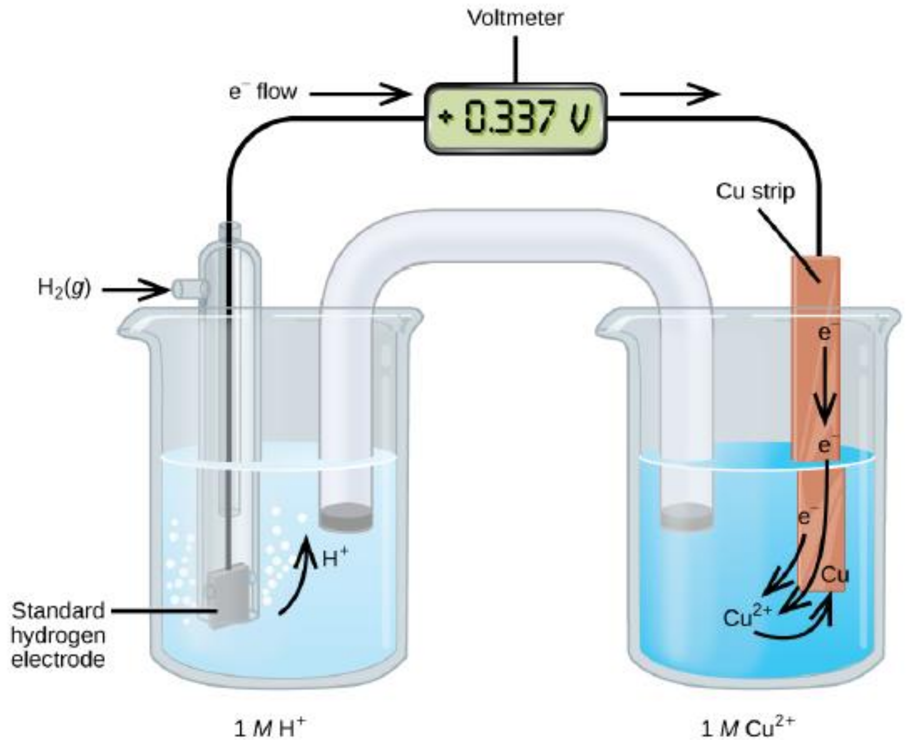


Using the Standard Hydrogen Electrode

- The cell notation of this reaction is

$$H_2(g) + Cu^{2+}(aq) \rightarrow 2H^+(aq) + Cu(s)$$

$$Pt(s)|H_2(g)|H^+(aq)||Cu^{2+}(aq)|Cu(s)$$



Cell Notation

Describe the flow of electrons in the cell.

$$Mg(s)|Mg^{2+}(1\ M)\|Pt(s)|H^{+}(1\ M),\ H_2(1\ atm)$$

A. from the Mg electrode to the Pt electrode
B. from the Mg^{2+} solution to the H^{+} solution
C. from the Pt electrode to the Mg electrode
D. from the Pt electrode to the Mg^{2+} solution

Cell Notation

Identify the component of the voltaic cell symbolized by the double vertical lines.

$$Mg(s)|Mg^{2+}(1\ M)\,\|\,Pt(s)|H^{+}(1\ M),\ H_2(1\ atm)$$

A. anode
B. cathode
C. electrode
D. salt bridge