

2021

Chapter Electrochemistry

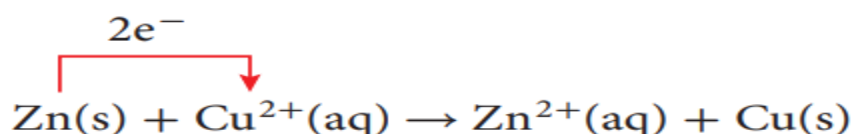
SECTION 1 VOLTAIC CELL

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Redox in Electrochemistry

- ✓ **Electrochemistry** is the study of the redox processes by which chemical energy is converted to electrical energy and vice versa.
- ✓ All redox reactions involve a transfer of electrons from the species that is oxidized to the species that is reduced.
- ✓ In the simple redox reaction in which zinc atoms are oxidized to form zinc (Zn^{2+}) ions. The two electrons donated from each zinc atom are accepted by a copper (Cu^{2+}) ion, which becomes an atom of copper metal.
- ✓ The following net ionic equation illustrates the electron transfer that occurs.

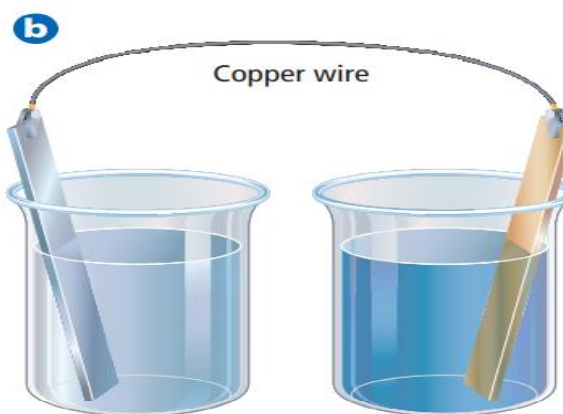


Half-reactions

- ✓ Two half-reactions make up this redox process:
- ✓ $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$ (oxidation half-reaction: electrons lost)
- ✓ $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$ (reduction half-reaction: electrons gained)
- ✓ we separated the oxidation half-reaction from the reduction half-reaction as I the figure.



In a, zinc metal is immersed in 1M zinc sulfate solution, and copper metal in 1M copper sulfate.



In b, a wire joining the zinc and copper strips provides a pathway for the flow of electrons, but the pathway is not complete.

✓ **Two problems prevent a redox reaction from occurring.**

✓ **First**, there is no way for zinc atoms to transfer electrons to copper(II) ions.

✓ **This problem can be solved by**

connecting the zinc and copper strips with a metal wire, as shown in Figure (b) .

✓ The wire can serve as a pathway for electrons to flow from the zinc strip to the copper strip.

✓ **The second problem** is that when the metal strips are placed in their solutions, oxidation begins at the zinc strip and reduction begins at the copper strip . but these reactions cannot continue.

✓ The reason is that as zinc is oxidized, positive zinc ions build up around the zinc electrode.

✓ As copper in the copper sulfate solution is reduced, negative sulfate ions build up around the copper electrode.

✓ The buildup of charges stops any further reaction.

✓ **To solve this problem**, a salt bridge must be added to the system.

✓ **A salt bridge** is a pathway to allow the passage of ions from one side to another.

✓ A salt bridge consists of a tube containing a conducting solution of a soluble salt, such as KCl, held in place by an agar gel or other permeable plug.

✓ Ions can move through the plug, but the solutions in the two beakers cannot mix.

✓ When the connecting metal wire and the salt bridge are in place, the spontaneous redox reaction begins.

✓ Electrons flow through the wire from the oxidation half-reaction to the reduction half-reaction.

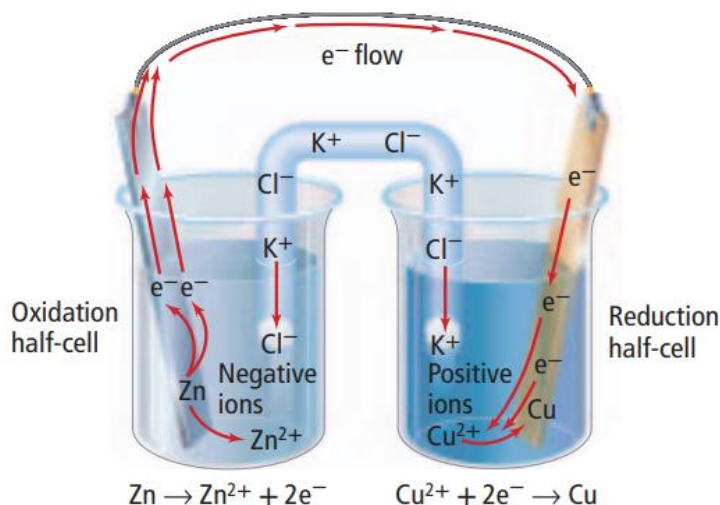
✓ while positive and negative ions move through the salt bridge.

✓ **An electric current** is a flow of charged particles .

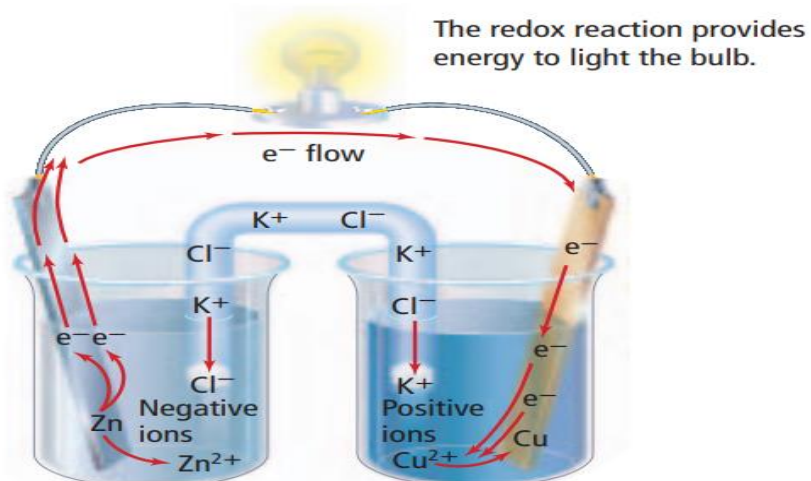
✓ The flow of electrons through the wire and the flow of ions through the salt bridge make up the electric current.

✓ The energy of the flowing electrons can be used to light a bulb, as shown in Figure .

The salt bridge and the wire provide an unbroken pathway for electrical charge to flow.



- ✓ The addition of a salt bridge completes the pathway.
- ✓ Negative ions move through the salt bridge to the zinc side.
- ✓ Positive ions move through the bridge to the copper side.



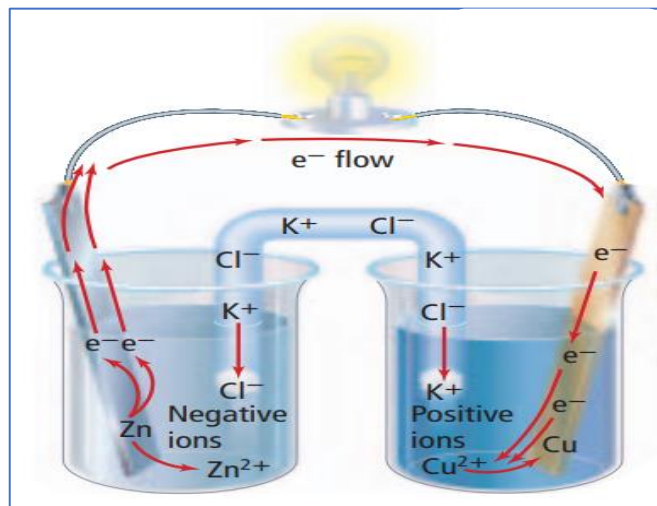
- ✓ **An electrochemical cell** is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction.
- ✓ **A voltaic cell** is a type of electrochemical cell that converts chemical energy to electrical energy by a spontaneous redox reaction.
- ✓ The voltaic cell, also shown in Figure , is named for Alessandro Volta , the Italian physicist who is credited with its invention in 1800.
- ✓ This replica of one of Alessandro Volta's first cells.
- ✓ consists of discs of zinc and copper arranged in alternating layers and separated by cloth or cardboard soaked in an acidic solution.
- ✓ Electric current increased with the number of metal discs used.



Chemistry of Voltaic Cells

- ✓ An electrochemical cell consists of two parts called half-cells, in which the separate oxidation and reduction reactions take place.
- ✓ Each half-cell contains an electrode and a solution containing ions.
- ✓ **An electrode** is an electrically conductive material, usually a metallic strip or graphite, that conducts electrons into and out of the solution in the half cell.

- ✓ the beaker with the zinc electrode is where the oxidation half of the redox reaction takes place.
- ✓ The beaker with the copper electrode is where the reduction half of the reaction takes place.
- ✓ **A half-cell reaction** : the reaction that takes place in each half cell .
- ✓ **The anode** is the electrode where oxidation takes place .
- ✓ **The cathode** is the electrode where reduction takes place .



Voltaic cells and energy

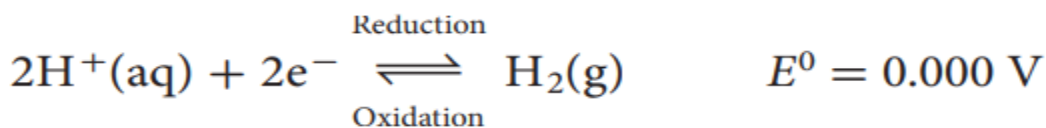
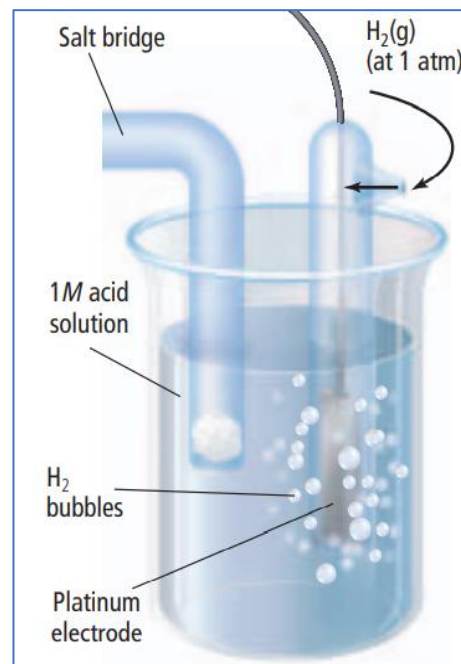
- ✓ **Electric potential energy** is a measure of the amount of current that can be generated from a voltaic cell to do work.
- ✓ Electric charge can flow between two points only when a difference in electric potential energy exists between the two points.
- ✓ In an electrochemical cell, these two points are the two electrodes. Electrons generated at the anode, the site of oxidation, are thought to be pushed or driven toward the cathode by the electromotive force (EMF).
- ✓ **the cell potential** is force that is due to the difference in electric potential energy between the two electrodes .
- ✓ **A volt** is a unit used to measure cell potential.
- ✓ **The electric potential difference of a voltaic cell** is an indication of the energy that is available to move electrons from the anode to the cathode.
- ✓ **In redox terms**, the voltage of a cell is determined by comparing the difference in the tendency of the two electrode materials to accept electrons.
- ✓ The greater the difference, the greater the potential energy difference between the two electrodes and the larger the voltage of the cell will be.
- ✓ In the zinc-copper cell, under standard conditions, copper(II) ions at the cathode accept electrons more readily than the zinc ions at the anode.
- ✓ Thus, the redox reaction occurs spontaneously only when the electrons flow from the zinc to the copper.

Calculating Electrochemical Cell Potentials

- ✓ **Reduction** is gaining electrons .
- ✓ **Reduction potential** is the tendency of a substance to gain electrons .
- ✓ The reduction potential of an electrode cannot be determined directly because the reduction half-reaction must be coupled with an oxidation half-reaction.
- ✓ When two half-reactions are coupled, the voltage generated corresponds to the difference in potential between the half-reactions.
- ✓ The electrical potential difference between two points is expressed in volts (V).

The standard hydrogen electrode

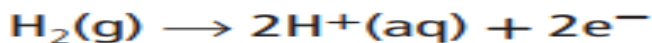
- ✓ The standard hydrogen electrode used to measure the reduction potential of all electrodes .
- ✓ The standard hydrogen electrode consists of a small sheet of platinum immersed in a hydrochloric acid (HCl) solution that has a hydrogen-ion concentration of 1M.
- ✓ Hydrogen gas (H₂), at a pressure of 1 atm is bubbled in and the temperature is maintained at 25°C.
- ✓ The potential, also called the standard reduction potential (E⁰), of this standard hydrogen electrode is defined as 0.000 V.
- ✓ This electrode can act as an oxidation half-reaction or a reduction half reaction, depending on the half-cell to which it is connected.
- ✓ The two possible reactions at the hydrogen electrode are the following.



Reduction:



Oxidation:



Half-cell potentials

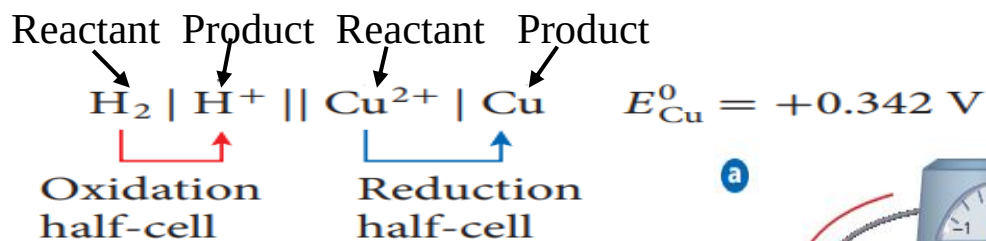
- ✓ Over the years, chemists have measured and recorded the standard reduction potentials of many different half-cells.
- ✓ Table lists some common half-cell reactions in order of increasing reduction potential.
- ✓ The values in the table were obtained by measuring the potential when each half-cell was connected to a standard hydrogen half-cell.
- ✓ All of the half-reactions in Table are written as reductions.
- ✓ In any voltaic cell, which always contains two half-reactions, **the half-reaction with the lower reduction potential will be an oxidation reaction (anode).**
- ✓ **The half-reaction that is more positive will proceed as a reduction (Cathode) .**
- ✓ **The half-reaction that is more negative will proceed as an oxidation (Anode).**
- ✓ The electrode being measured must also be under standard conditions, that is, immersed in a 1M solution of its ions at 25°C and 1 atm.
- ✓ The superscript zero in the notation E^0 is indicating “measured under standard conditions.”

Standard Reduction Potentials

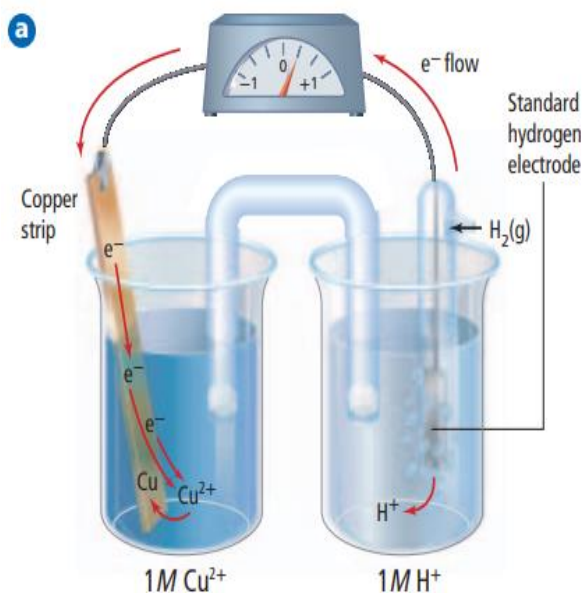
Half-Reaction	E^0 (V)	Half-Reaction	E^0 (V)
$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.0401	$\text{Cu}^{2+} + \text{e}^- \rightarrow \text{Cu}^+$	+0.153
$\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$	-2.868	$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.3419
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$	+0.401
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.372	$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.5355
$\text{Be}^{2+} + 2\text{e}^- \rightarrow \text{Be}$	-1.847	$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.771
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.662	$\text{NO}_3^- + 2\text{H}^+ + \text{e}^- \rightarrow \text{NO}_2 + \text{H}_2\text{O}$	+0.775
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.185	$\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow 2\text{Hg}$	+0.7973
$\text{Cr}^{2+} + 2\text{e}^- \rightarrow \text{Cr}$	-0.913	$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.7996
$2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$	-0.8277	$\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	+0.851
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.7618	$2\text{Hg}_2^{2+} + 2\text{e}^- \rightarrow \text{Hg}_2^{2+}$	+0.920
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.744	$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightarrow \text{NO} + 2\text{H}_2\text{O}$	+0.957
$\text{S} + 2\text{e}^- \rightarrow \text{S}^{2-}$	-0.47627	$\text{Br}_2(\text{l}) + 2\text{e}^- \rightarrow 2\text{Br}^-$	+1.066
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.447	$\text{Pt}^{2+} + 2\text{e}^- \rightarrow \text{Pt}$	+1.18
$\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$	-0.4030	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.229
$\text{PbI}_2 + 2\text{e}^- \rightarrow \text{Pb} + 2\text{I}^-$	-0.365	$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	+1.35827
$\text{PbSO}_4 + 2\text{e}^- \rightarrow \text{Pb} + \text{SO}_4^{2-}$	-0.3588	$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	+1.498
$\text{Co}^{2+} + 2\text{e}^- \rightarrow \text{Co}$	-0.28	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.507
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.257	$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	+1.692
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.1375	$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.776
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.1262	$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	+1.92
$\text{Fe}^{3+} + 3\text{e}^- \rightarrow \text{Fe}$	-0.037	$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	+2.010
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.0000	$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	+2.866

Determining electrochemical cell potentials

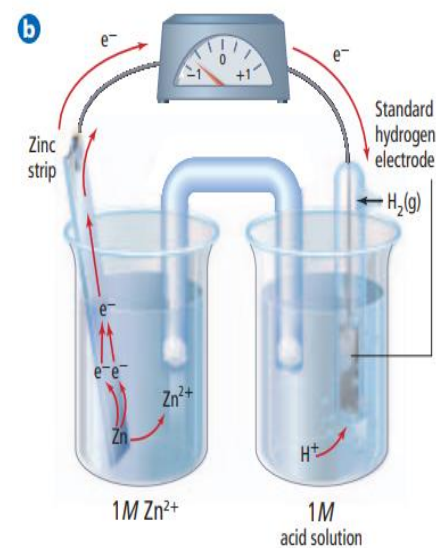
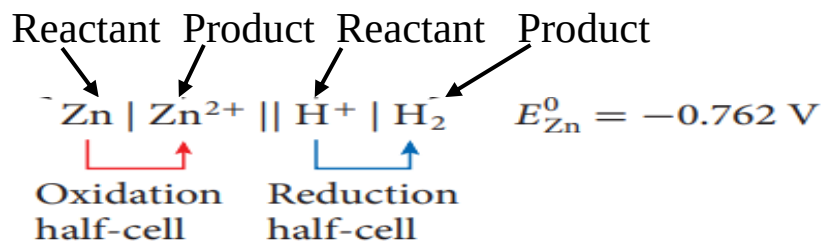
- ✓ The first step is to determine the standard reduction potential for the copper half-cell (E^0_{Cu}).
- ✓ When the copper electrode is attached to a standard hydrogen electrode, as in Figure.
- ✓ Electrons flow from the hydrogen electrode to the copper electrode, and copper ions are reduced to copper metal.
- ✓ The E^0 , measured by a voltmeter, is +0.342 V. The positive voltage indicates that Cu^{2+} ions at the copper electrode accept electrons more readily than do H^+ ions at the standard hydrogen electrode.
- ✓ Therefore, oxidation takes place at the hydrogen electrode and reduction takes place at the copper electrode.
- ✓ The oxidation and reduction half-cell reactions and the overall reaction are
- ✓ $\text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$ (oxidation half-cell reaction)
- ✓ $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$ (reduction half-cell reaction)
- ✓ $\text{H}_2(\text{g}) + \text{Cu}^{2+}(\text{aq}) \rightarrow 2\text{H}^+(\text{aq}) + \text{Cu}(\text{s})$ (overall redox reaction)
- ✓ This reaction can be written in a form called cell notation.



- ✓ The two participants in the oxidation reaction are written first and in the order they appear in the oxidation half-reaction-reactant | product.
- ✓ They are followed by a double vertical line (||) representing the wire and salt bridge connecting the half-cells.
- ✓ Then, the two participants in the reduction reaction are written in the same reactant | product order.
- ✓ **Note that** for positive values of E^0 , it is customary to place a plus sign before the voltage.



- ✓ The next step is to determine the standard reduction potential for the zinc half-cell (E^0_{Zn}).
- ✓ When the zinc electrode is measured against the standard hydrogen electrode under standard conditions, as in Figure .
- ✓ Electrons flow from the zinc electrode to the hydrogen electrode.
- ✓ The E^0 of the zinc half-cell, measured by a voltmeter, is -0.762 V .
- ✓ This means that the H^+ ions at the hydrogen electrode accept electrons more readily than do the zinc ions.
- ✓ Thus, the hydrogen ions have a higher reduction potential than the zinc ions.
- ✓ The hydrogen electrode is assigned a zero potential, so the reduction potential of the zinc electrode must have a negative value.
- ✓ The two half-cell reactions and the overall reaction are written as follows.
- ✓ $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ (oxidation half-cell reaction)
- ✓ $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$ (reduction half-cell reaction)
- ✓ $\text{Zn(s)} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g})$ (overall redox cell reaction)
- ✓ This reaction can be written in a form called cell notation.

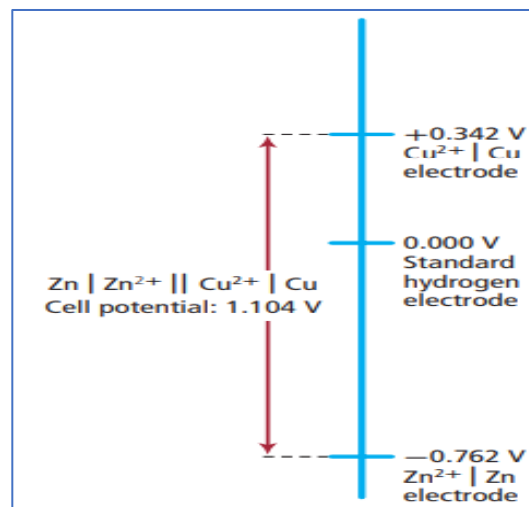


- ✓ The final step in calculating electrochemical cell potential is to combine the copper and zinc half-cells as a voltaic cell.
- ✓ This means calculating the voltaic cell's standard potential using the following formula.

$$E^0_{\text{cell}} = E^0_{\text{reduction}} - E^0_{\text{oxidation}}$$

- ✓ Because reduction occurs at the copper electrode and oxidation occurs at the zinc electrode, the E^0 values are substituted as follows.
- ✓ $E^0_{\text{cell}} = E^0_{\text{Cu}^{2+}|\text{Cu}} - E^0_{\text{Zn}^{2+}|\text{Zn}}$
- ✓ $= +0.342 \text{ V} - (-0.762 \text{ V}) = +1.104 \text{ V}$

- ✓ **Notice that** the negative sign in the formula automatically changes the sign of the oxidation half-reaction, so you do not reverse the sign of the standard reduction potentials, when they are used for the oxidation half-reaction.
- ✓ The graph in Figure shows how the zinc half-cell with the lower reduction potential and the copper half-cell with the higher reduction potential are related.



✓ Solving Problems :

1- Calculate a Cell Potential The following reduction half-reactions represent the half-cells of a voltaic cell.



- ✓ Determine the overall cell reaction and the standard cell potential. Describe the cell using cell notation.

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2- For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation.



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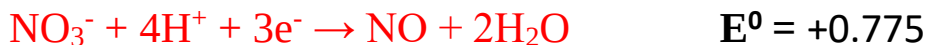
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4. Challenge Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.



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Using Standard Reduction Potentials

- 1. To calculate the standard potential (voltage) of voltaic cells.
 - 2. To determine if a proposed reaction under standard conditions will be spontaneous.
 - ✓ How can standard reduction potentials indicate spontaneity?
 - ✓ Electrons in a voltaic cell always flow from the half-cell with the lower standard reduction potential (anode) to the half-cell with the higher reduction potential (Cathode), giving a positive cell voltage.
 - ✓ If the calculated potential is positive, the reaction is spontaneous.
 - ✓ If the value is negative, the reaction is not spontaneous.
 - ✓ Solving Problem :
 - ✓ Calculate the cell potential to determine if each of the following balanced redox reactions
5. $\text{Sn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + \text{Cu(s)}$
- | | |
|--|-----------------|
| $\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$ | $E^0 = -0.1375$ |
| $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ | $E^0 = +0.3419$ |

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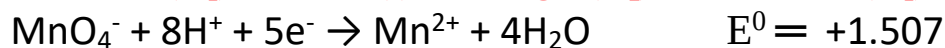
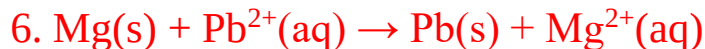
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9. Challenge write the equation and determine the cell voltage (E^0) for the following cell. Is the reaction spontaneous?



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Revision:

1- Describe the conditions under which a redox reaction causes an electric current to flow through a wire.

2- Identify the components of a voltaic cell. Explain the role of each component in the operation of the cell.

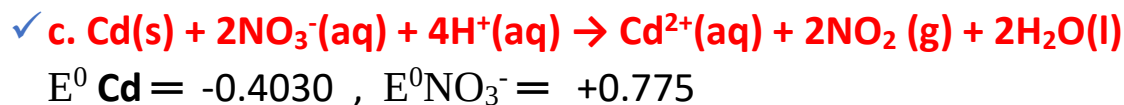
3- Write the balanced equation for the spontaneous cell reaction that occurs in a cell with these reduction half-reactions.





4-Determine the standard potential for electrochemical cells in which each equation represents the overall cell reaction.

✓ Identify the reactions as spontaneous or nonspontaneous as written.



5-Describe the process that releases electrons in a zinc-copper voltaic cell.

6-What is the function of a salt bridge in a voltaic cell?

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7-What information do you need in order to determine the standard voltage of a voltaic cell?

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8-In a voltaic cell represented by $\text{Al} \mid \text{Al}^{3+} \parallel \text{Cu}^{2+} \mid \text{Cu}$, what is oxidized and what is reduced as the cell delivers current?

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9-Under what conditions are standard reduction potentials measured?

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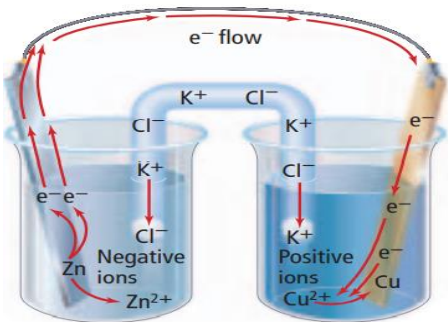
10- In Figure , identify the metal that is being oxidized.
Identify the cathode.

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11- A salt bridge is filled with KNO_3 . Explain why it is necessary that the potassium ions move through the salt bridge to the cathode.

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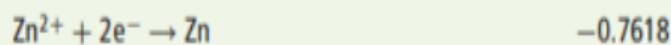
12- A reducing agent is the substance being oxidized and an oxidizing agent is the substance being reduced.

✓ Use Table to select an oxidizing agent that will convert Au to Au^{3+} but will not convert Co^{2+} to Co^{3+} .

$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	+1.35827
$\text{Au}^{3+} + 3\text{e}^- \rightarrow \text{Au}$	+1.498
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.507
$\text{Au}^+ + \text{e}^- \rightarrow \text{Au}$	+1.692
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$	+1.776
$\text{Co}^{3+} + \text{e}^- \rightarrow \text{Co}^{2+}$	+1.92
$\text{S}_2\text{O}_8^{2-} + 2\text{e}^- \rightarrow 2\text{SO}_4^{2-}$	+2.010

13- Using Table , write the standard cell notation for each cell in which each of the following half-cells is connected to the standard hydrogen electrode.

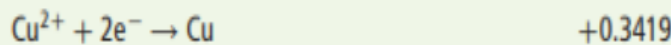
a. $\text{Zn} \mid \text{Zn}^{2+}$



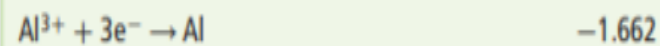
b. $\text{Hg} \mid \text{Hg}^{2+}$



c. $\text{Cu} \mid \text{Cu}^{2+}$



d. $\text{Al} \mid \text{Al}^{3+}$



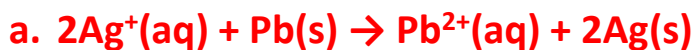
14- Write the balanced chemical equation for the standard cell notations listed below.

a. $\text{I}^- \mid \text{I}_2 \parallel \text{Fe}^{3+} \mid \text{Fe}^{2+}$

b. $\text{Sn} \mid \text{Sn}^{2+} \parallel \text{Ag}^+ \mid \text{Ag}$

c. $\text{Zn} \mid \text{Zn}^{2+} \parallel \text{Cd}^{2+} \mid \text{Cd}$

15- Calculate the cell potential for the following voltaic cells.



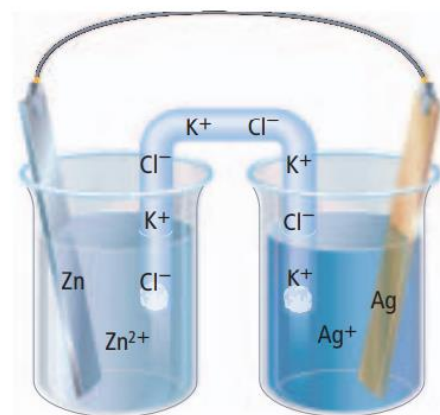
$\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$	+0.5355
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.7996
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.257
$\text{Sn}^{2+} + 2\text{e}^- \rightarrow \text{Sn}$	-0.1375
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.1262
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.185



16- Figure illustrates a voltaic cell consisting of a strip of zinc in a 1.0M solution of zinc nitrate and a strip of silver in a 1.0M solution of silver nitrate.

✓ Use the diagram and Table to answer these questions.

- Identify the anode.
- Identify the cathode.
- Where does oxidation occur?
- Where does reduction occur?
- In which direction is the current flowing through the connecting wire?



- In which direction are positive ions flowing through the salt bridge?
-

g. What is the cell potential at 25°C and 1 atm?

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$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.7618
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.7996

17- Determine whether each redox reaction is spontaneous or nonspontaneous.



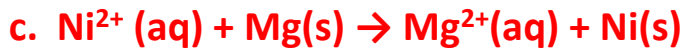
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$\text{Br}_2(\text{l}) + 2\text{e}^{-} \rightarrow 2\text{Br}^{-}$	+1.066
$\text{Mn}^{2+} + 2\text{e}^{-} \rightarrow \text{Mn}$	-1.185



.....

$\text{Fe}^{3+} + \text{e}^{-} \rightarrow \text{Fe}^{2+}$	+0.771
$\text{Sn}^{2+} + 2\text{e}^{-} \rightarrow \text{Sn}$	-0.1375



.....

$\text{Ni}^{2+} + 2\text{e}^{-} \rightarrow \text{Ni}$	-0.257
$\text{Mg}^{2+} + 2\text{e}^{-} \rightarrow \text{Mg}$	-2.372



.....

$\text{Pb}^{2+} + 2\text{e}^{-} \rightarrow \text{Pb}$	-0.1262
$\text{Cu}^{2+} + \text{e}^{-} \rightarrow \text{Cu}^{+}$	+0.153

18- Why do electrons flow from one electrode to the other in a voltaic cell?

.....

19- Use the table below to answer Questions 1 to 4.

1. Which metal ion is most easily reduced?

- A. Mg^{2+} B. Hg^{2+} C. Ag^+ D. Al^{3+}

2- On the basis of the standard reduction potentials shown above, which standard cell notation correctly represents its voltaic cell?

- A. $\text{Ag} \mid \text{Ag}^+ \parallel \text{Al}^{3+} \mid \text{Al}$
 B. $\text{Mg} \mid \text{Mg}^{2+} \parallel \text{H}^+ \mid \text{H}_2$
 C. $\text{H}_2 \mid \text{H}^+ \parallel \text{Pb}^{2+} \mid \text{Pb}$
 D. $\text{Pb} \mid \text{Pb}^{2+} \parallel \text{Al}^{3+} \mid \text{Al}$

Selected Standard Reduction Potentials at 25°C, 1 atm, and 1M Ion Concentration	
Half-Reaction	E° (V)
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.372
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.662
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.1262
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.7996
$\text{Hg}^{2+} + 2\text{e}^- \rightarrow \text{Hg}$	0.851

3. A voltaic cell consists of a magnesium bar dipping into a 1M Mg^{2+} solution and a silver bar dipping into a 1M Ag^+ solution. What is the standard potential of this cell?

- A. 1.572 V B. 3.172 V C. 0.773 V D. 3.971 V

4. Assuming standard conditions, which cell will produce a potential of 2.513 V?

- A. $\text{Al} \mid \text{Al}^{3+} \parallel \text{Hg}^{2+} \mid \text{Hg}$ C. $\text{Mg} \mid \text{Mg}^{2+} \parallel \text{Al}^{3+} \mid \text{Al}$
 B. $\text{H}_2 \mid \text{H}^+ \parallel \text{Hg}^{2+} \mid \text{Hg}$ D. $\text{Pb} \mid \text{Pb}^{2+} \parallel \text{Ag}^+ \mid \text{Ag}$

20- Use the table below to answer Question

Based on the standard reduction potentials given above, if a silver electrode and a chromium electrode are connected in a voltaic cell, which electrode will undergo oxidation, and which will undergo reduction?

Standard Reduction Potentials at 25°C, 1 atm, and 1M Solution	
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	0.7996
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.744

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