Electrochemistry

E\_\_\_\_\_ translates the \_\_\_\_\_ energy of a reduction–oxidation reaction into \_\_\_\_\_ energy.

* In a \_\_\_\_\_ reaction, if the substance being oxidized and the substance being reduced are in \_\_\_\_\_ with one another, a transfer of energy as \_\_\_\_\_ accompanies the transfer of \_\_\_\_\_.
* When the substances being oxidized and reduced are \_\_\_\_\_ from one another, the \_\_\_\_\_ \_\_\_\_\_ is accompanied by a transfer of \_\_\_\_\_ energy instead of heat energy.

\_\_\_\_\_ Types of Electrochemical Cells

* \_\_\_\_\_ cells: Chemical reactions generate \_\_\_\_\_ energy; \_\_\_\_\_
* \_\_\_\_\_ cells: Electrical energy decomposes \_\_\_\_\_ compounds; \_\_\_\_\_

Electrochemical Cells

1. When the substances involved in oxidation and reduction half–reactions are \_\_\_\_\_ \_\_\_\_\_, it is called an \_\_\_\_\_ cell.

2. Each half reaction occurs on the surface of an electrically conductive solid called an \_\_\_\_\_.

3. Each electrode is immersed in an \_\_\_\_\_ solution containing \_\_\_\_\_ needed for the \_\_\_\_\_ –reaction.

4. The electrodes are connected by a wire so that \_\_\_\_\_ can move \_\_\_\_\_ the \_\_\_\_\_ half–reaction \_\_\_ the \_\_\_\_\_ half–reaction.

5. The solutions are connected by a \_\_\_\_\_ \_\_\_\_\_ or porous membrane so \_\_\_\_\_ can move between solutions.

6. The chemical potential energy can be harnessed as the substances undergoing oxidation push \_\_\_\_\_ through the wire to the substances undergoing reduction.

Components of Electrochemical Cells

\_\_\_\_\_

* The electrode in the electrochemical cell where \_\_\_\_\_ occurs.
* Anions are attracted to the anode (negatively charged in this case) that is \_\_\_\_\_ electrons in the half reaction.
* i.e. Zn0*(s)* 🡪 Zn+2(aq) + 2 e-

\_\_\_\_\_

* The electrode in the electrochemical cell where \_\_\_\_\_ occurs.
* Cations are attracted to the cathode (positively charged in this case) that is \_\_\_\_\_ electrons in the half reaction.
* i.e. Cu+2*(aq)* + 2 e- 🡪 Cu0*(s)*

\_\_\_\_\_ \_\_\_\_\_

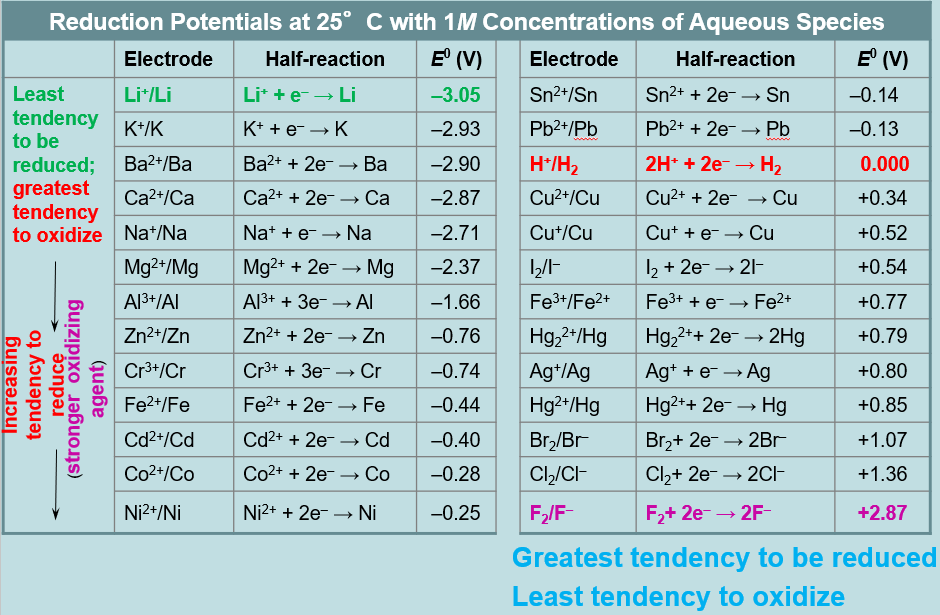
* The salt bridge (porous membrane) allows \_\_\_\_\_ of \_\_\_\_\_ into those solutions to maintain electric \_\_\_\_\_ without mixing of solutions.
* Different kinds of electrochemical cells (electrolytic vs. voltaic) use different kinds of salt bridges or solutions.
* Contains \_\_\_\_\_ solution. Allows ions to flow between half-cells. Prevents \_\_\_\_\_ of charge on the electrodes. Completes the electrical \_\_\_\_\_ by allowing ions to flow.

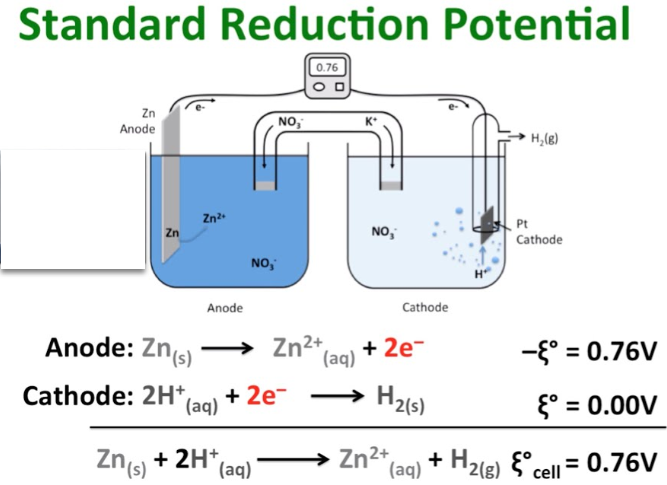
STANDARD \_\_\_\_\_ POTENTIAL (SRP)

* Since one half-reaction does not occur without the other half-reaction, it is impossible to measure electrode potentials of the separate half-reactions directly.
* The \_\_\_\_\_ cell potential is the difference between the reduction potential of the cathode and the reduction potential of the anode (reverse reaction to oxidation):

E0cell = E0\_\_\_\_\_ + E0\_\_\_\_\_ .

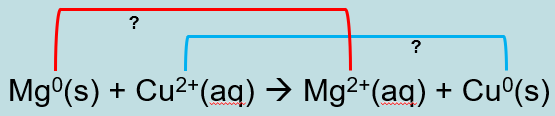
* The term “\_\_\_\_\_” comes from using the hydrogen electrode as the \_\_\_\_\_ half–reaction: [2\_\_\_+ + 2 *e*– 🡪 \_\_\_2]. This electrode is defined as having an electrode potential of exactly \_\_\_ V.
* To obtain standard \_\_\_\_\_ potentials, the magnitude is the same, but the sign is \_\_\_\_\_. [SRP = -3.05; SOP = +3.05]
* E0cell = E0\_\_\_\_\_ + E0\_\_\_\_\_ is used to determine the standard cell \_\_\_\_\_ for any electrochemical cell.





CALCULATING POTENTIALS IN HALF REACTIONS

Consider the reaction between magnesium metal & copper ions:



* In this reaction magnesium metal (Mg) is being \_\_\_\_\_ to magnesium ion, while copper (Cu) in aqueous solution is being \_\_\_\_\_ to copper metal (solid).
* Looking at the standard reduction potentials table for each element we find the following:

Mg2+ (aq) + 2 e- 🡪 Mg (s) E0 = \_\_\_\_\_volts (SRP)\*  
Cu2+ (aq) + 2 e- 🡪 Cu (s) E0 = \_\_\_\_\_ volts (SRP)

*\*notice magnesium is NOT reduced but oxidized*

Change the SRP to match the actual reaction. In the above reaction, magnesium is oxidized and copper is reduced. Therefore:

*Oxid ½ reaction*: \_\_\_(s) 🡪 \_\_\_2+(aq) + 2 e- E0 = \_\_\_\_\_ volts\*

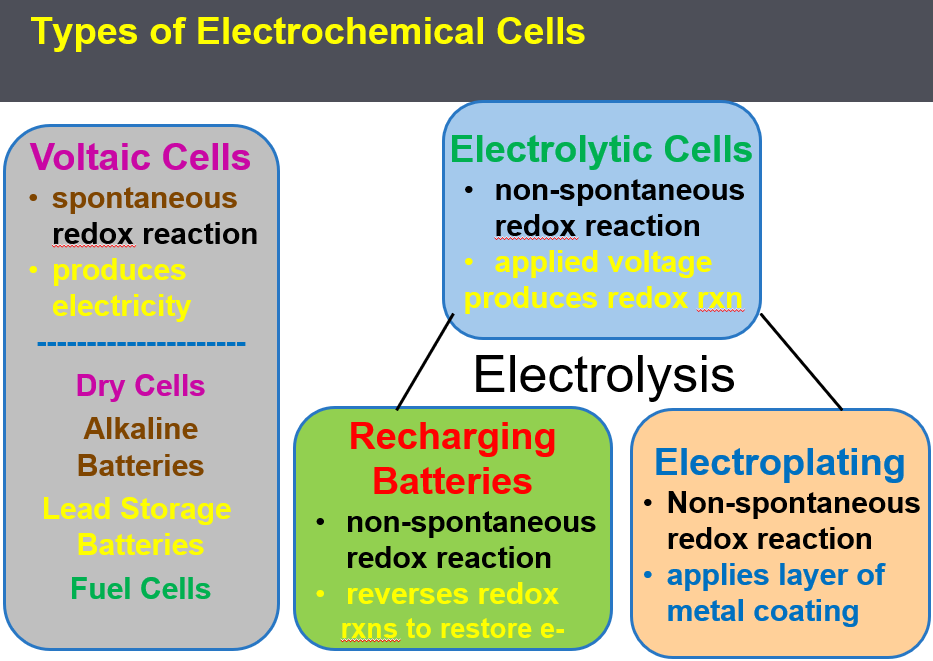
*Red ½ reaction*: \_\_\_2+(aq) + 2 e- 🡪 \_\_\_(s) E0 = \_\_\_\_\_ volts

net potential for the overall reaction: \_\_\_\_\_ V

*\*notice the sign for the oxidation ½ reaction was reversed*

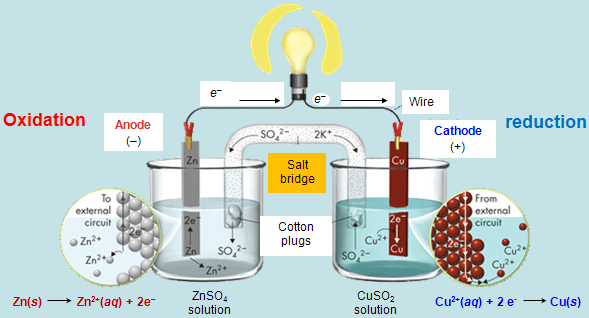
This represents the \_\_\_\_\_ voltage produced in this electrochemical cell.

Types of Electrochemical Cells

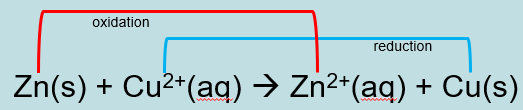


Voltaic Cells

* Also called \_\_\_\_\_ Cells. Electrons are \_\_\_\_\_ produced at the \_\_\_\_\_ (\_\_\_\_\_) and migrate to the \_\_\_\_\_ so that \_\_\_\_\_ can occur. Each electrode is placed in an aqueous solution matching the electrode. The half reactions are kept separate in “cells,” connected by a salt bridge.
* Redox reactions in voltaic cells [\_\_\_\_\_ *cells,* \_\_\_\_\_ *storage &* \_\_\_\_\_ *batteries,* \_\_\_\_\_ \_\_\_\_\_ *as in cars*] take place in aqueous solutions (*in this case ZnSO4(aq) & CuSO4(aq)*) to allow ion migration.
* The electrons are “pumped” from the \_\_\_\_\_ towards the \_\_\_\_\_, supplying \_\_\_\_\_ to some resistance (*bulb, motor, cell, door bell, keypad, etc*.).



* Electrons are produced at the \_\_\_\_\_ so it is designated with a \_\_\_\_\_ charge in voltaic cells. The electrons migrate and are attracted to the cathode, therefore, the cathode is often denoted with a positive sign.
* To maintain electrical \_\_\_\_\_, \_\_\_\_\_ must migrate: cations towards the cathode and anions towards the anode.
* The \_\_\_\_\_ bridge completes the \_\_\_\_\_ so that current (\_\_\_\_\_) can flow. If the circuit is “broken,” the half reactions will not occur.
* The \_\_\_\_\_ are the source of potential in the voltaic cell.
  + As the electrons \_\_\_\_\_ the anode (Zn), the Zn metal disintegrates.
  + As the electrons \_\_\_\_\_ at the cathode (Cu), the Cu metal increases in mass at the cathode.
* Once the redox reaction reaches \_\_\_\_\_, the voltage is \_\_\_\_\_.

Calculate the \_\_\_\_\_ of the Zinc-Copper   


Zn2+(aq) + 2 e- 🡪 Zn(s) E0 = - 0.76 volt (SRP)  
Cu2+(aq) + 2 e- 🡪 Cu(s) E0 = +0.34 volts (SRP)

*Oxid ½ reaction*: \_\_\_(s) 🡪 \_\_\_2+(aq) + 2 e- E0 = \_\_\_\_\_ volts\*

*Red ½ reaction*: \_\_\_2+(aq) + 2 e- 🡪 \_\_\_(s) E0 = \_\_\_\_\_ volts

net potential for the overall reaction: \_\_\_\_\_ V

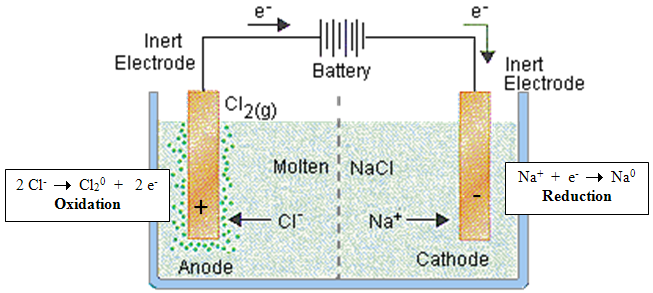
*\*notice the sign for the oxidation ½ reaction was* \_\_\_\_\_

Voltaic Cells

* A \_\_\_\_\_ cell is a voltaic cell in which the electrolyte is a paste. Dry cells have a \_\_\_\_\_ shelf life and a voltage drop when used.
* The \_\_\_\_\_ battery is an improved dry cell. This change eliminates the \_\_\_\_\_ of ammonia gas and maintains the zinc electrode, which corrodes more slowly under basic, or alkaline, conditions.
* A \_\_\_\_\_ Storage Battery is a group of voltaic cells connected together. A 12-V \_\_\_\_\_ battery consists of six voltaic cells connected together. The electrolyte for both half-cells is \_\_\_\_\_ acid. Using the same \_\_\_\_\_ for both half-cells allows the cell to operate without a salt bridge or porous separator.
* \_\_\_\_\_ Cells are voltaic cells in which \_\_\_\_\_ energy is \_\_\_\_\_ obtained. Fuel cells do not have to be recharged.
* Fuel cells consist of an [anode](https://en.wikipedia.org/wiki/Anode), a [cathode](https://en.wikipedia.org/wiki/Cathode), and an [electrolyte](https://en.wikipedia.org/wiki/Electrolyte) that allows positively charged hydrogen ions (or protons) to move between the sides of the fuel cell. Fuel cells are 2-3 times more efficient than combustion engines, producing water and heat.
* Scientists use the known standard oxidation-reduction potentials to predict \_\_\_\_\_: *E*0cell value. E0cell = E0\_\_\_\_\_ + E0\_\_\_\_\_

Electrolytic Cells

* \_\_\_\_\_ uses an electric current (“electro”) to \_\_\_\_\_ (“lysis”) a compound into its elements, including many important industrial reactions such as the electrolysis of brine (*NaCl into sodium and chorine gas*), electrolysis of aluminum oxide to \_\_\_\_\_ and oxygen gas, and the electrolysis of \_\_\_\_\_ into hydrogen and oxygen gases.



* \_\_\_\_\_ -spontaneous: The electrons are being \_\_\_\_\_ in a direction opposite of a spontaneous reaction using an \_\_\_\_\_ DC \_\_\_\_\_.
* The \_\_\_\_\_ (\_\_\_\_\_) is designated \_\_\_\_\_ [*cations (positive sodium ions above) are attracted to it*]. The power source forces reduction at the cathode: Na+ + e- 🡪 Na0, forming sodium metal.
* The \_\_\_\_\_ (oxidation) is \_\_\_\_\_ [anions (*negative chlorine ions above*) are attracted to it]: 2 Cl- 🡪 Cl20 + 2 e-, forming chlorine gas

Use the SRP reference table to determine the potential required to electrolyze sodium chloride into its elements.

Oxid ½ reaction: 2 Cl- 🡪 Cl20 + 2 e- E0 = \_\_\_\_\_ volts

Red ½ reaction: 2(Na+ + e- 🡪 Na0) E0 = \_\_\_\_\_ volts

E0cell = E0oxid + E0red = (\_\_\_\_\_ V) + (\_\_\_\_\_ V) = \_\_\_\_\_ V

Non-Spontaneous Redox Reaction

Electrolytic Cells - \_\_\_\_\_

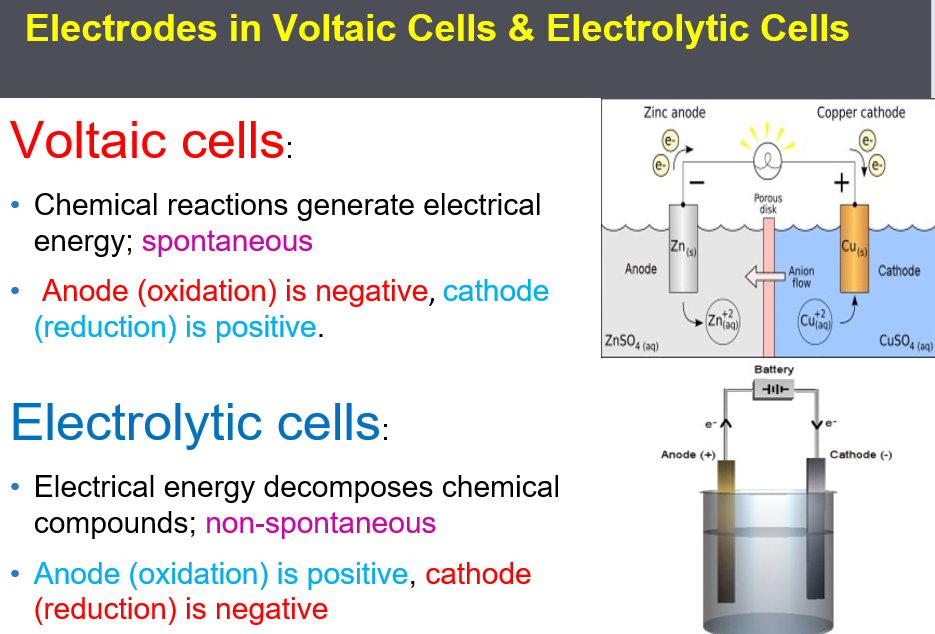
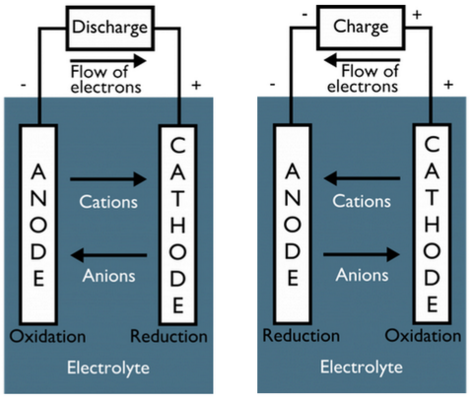
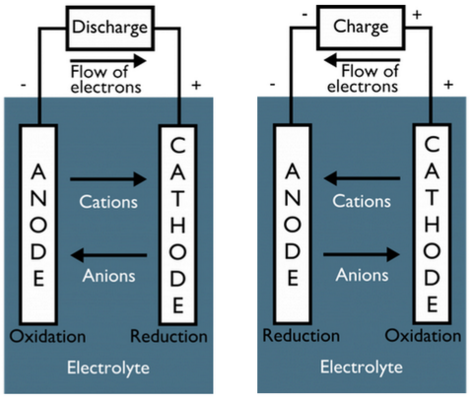
* \_\_\_\_\_ -Spontaneous Redox Reaction requiring a \_\_\_\_\_ source.
* Electroplating is the \_\_\_\_\_ of a metallic \_\_\_\_\_ [*nickel in this case*] on an object at the \_\_\_\_\_ by exposing it to a salt solution [*e.g. nickel chloride in this case*] of the metal to be deposited and then, putting a negative charge on the object, so positively charged metal ions [*Ni+2 ions in this case*] in the salt solution are reduced to the metallic form at the anode.

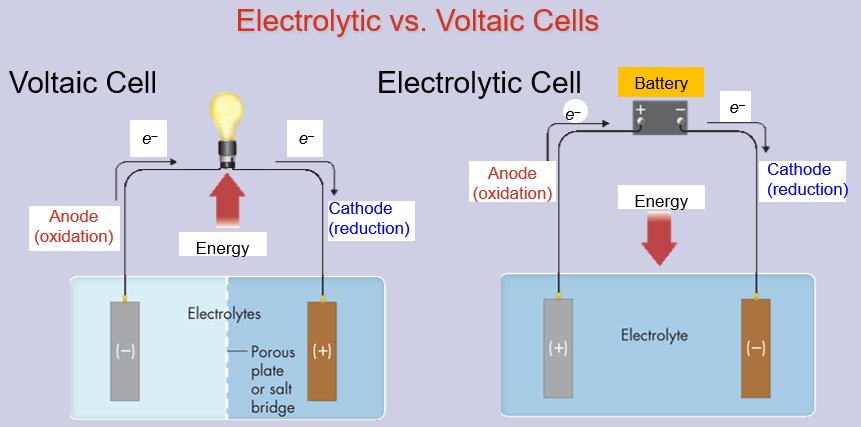
The Process of Electroplating

* \_\_\_\_\_ (\_\_\_\_\_) [+ terminal] … Cu0(s) 🡪 Cu2+ + 2 e-
  + The anode is made of the metal \_\_\_\_\_ for the plating
  + The anode metal gets \_\_\_\_\_ \_\_\_\_\_
* \_\_\_\_\_ (\_\_\_\_\_) [- terminal] … Cu2+ + 2 e- 🡪 Cu0(s)
  + The cathode metal gets \_\_\_\_\_
* Electroplating is used to prevent \_\_\_\_\_, increase wear resistance, enhance the \_\_\_\_\_ of an object, and to build up thickness on undersized parts.

\_\_\_\_\_ Batteries

* When a battery (\_\_\_\_\_ cell) is discharged [image below on the left], \_\_\_\_\_ proceeds at the \_\_\_\_\_ electrode (\_\_\_\_\_), and \_\_\_\_\_ occurs at the \_\_\_\_\_ electrode \_\_\_\_\_.
* To \_\_\_\_\_ a battery (electrolysis), the direction of electric \_\_\_\_\_ flow is \_\_\_\_\_ so the opposite reactions take place. [image below on the right]
* Reduction proceeds at the former anode (negative electrode … now the cathode), and oxidation takes place at the former cathode (positive electrode … now the anode).





* In both voltaic and electrolytic cells, (\_\_\_\_\_ flow (\_\_\_\_\_ the anode (\_\_\_\_\_ the cathode in the external circuit.
* In a voltaic cell, the flow of electrons is the result of a (\_\_\_\_\_ redox reaction, whereas in an electrolytic cell, electrons are caused to flow by an (\_\_\_\_\_ power source, such as a battery.
* In a voltaic cell, the (\_\_\_\_\_ is the negative electrode and the (\_\_\_\_\_ is the positive electrode. In an electrolytic cell, the signs are (\_\_\_\_\_.