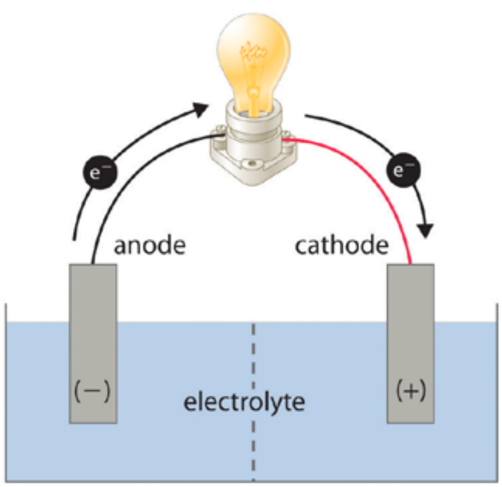
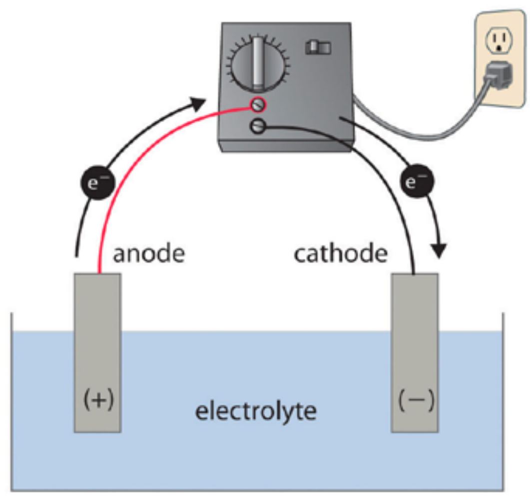
ELECTROCHEMICAL CELLS

Electrochemical Cells: *A Comparison*				
Galvanic (voltaic) cells	spontaneous oxidation-reduction reaction	Is separated into 2 half-cells	Electrodes made from metals (inert Pt or C if ion to ion or gas)	Battery – its cell potential drives the reaction and thus the e^-
Electrolytic cells	non-spontaneous oxidation-reduction reaction	Usually occurs in a single container	Usually inert electrodes	Battery charger – requires an external energy source to drive the reaction and e^-



GALVANIC CELL

Energy released by spontaneous redox reaction is converted to electrical energy.

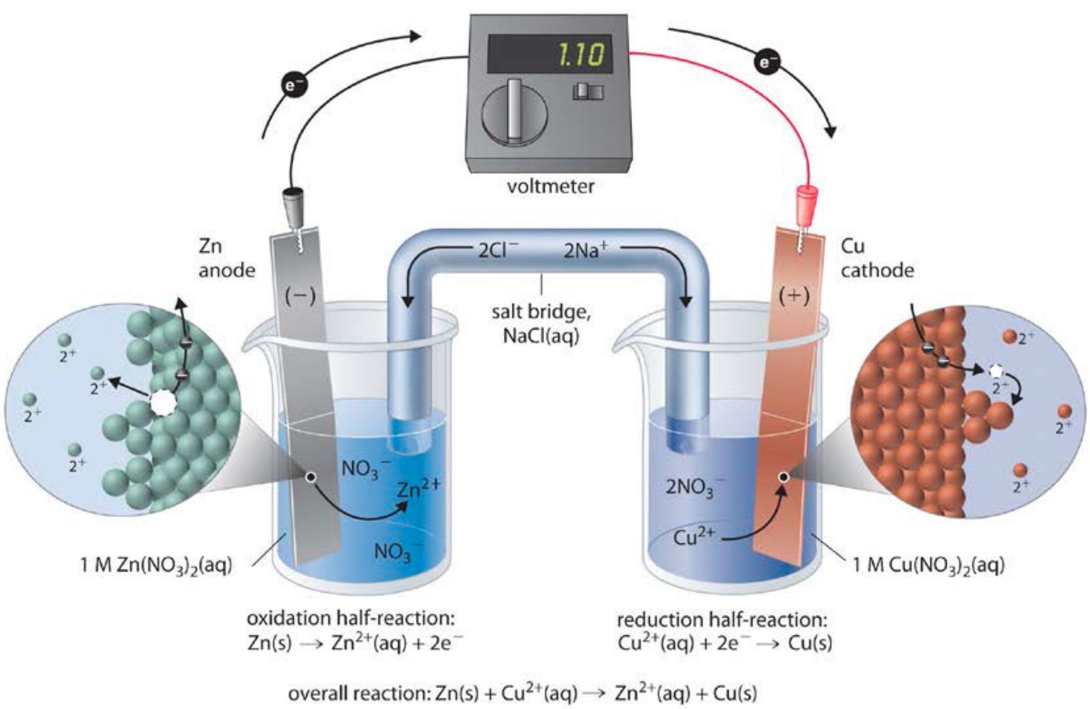


ELECTROLYTIC CELL

Electrical energy is used to drive nonspontaneous redox reaction.

The Galvanic Cell: *What is what and what to know?*	
ANODE: AN OX	<u>Anode</u> – the electrode where ***oxidation*** occurs. Over time the mass of the anode may decrease as the metal is oxidized into ions.
CATHODE: RED CAT	<u>Cathode</u> – the electrode where ***reduction*** occurs. Over time the mass of the cathode may increase as the metal ions in the solution are reduced and plated onto it.
FAT CAT	Electron Flow – **F**rom **A**node **T**o **CAT**hode
Ca+hode	Cathode is + galvanic cells
Salt Bridge	<u>Salt Bridge</u> – provides ions to balance the charge in each cell; contains a neutral salt that is very soluble (avoids precipitation issues). The salt cations flow into the cathode and the salt anions flow into the anode.

The Galvanic Cell: ***How it Works!***



The above picture shows the oxidation-reduction reaction between Zn and Cu. The diagram of the cell clearly shows: the anode and cathode; the half-reaction that occurs at each electrode, the direction of electron flow, the direction of ion movement from the salt bridge, the overall reaction, as well as the $E°_{cell}$ for the reaction.

Calculating Standard Cell Potential ($E°_{cell}$)	
The difference in electrical potential between the two half-reactions is measured with a voltmeter. The difference between the cell potentials of the two half-reactions determines the overall cell potential for the reaction.	
1	Look at a table of Standard Reduction Potentials (or the reduction potentials that are provided). Write both ***reduction*** reactions from the table with their voltages.
2	THE MORE POSITIVE REDUCTION POTENTIAL IS REDUCED. *THE LEAST POSITIVE IS OXIDIZED*.
3	Reverse the equation that will be oxidized; be sure to change the sign of the voltage [this is now $E°_{oxidation}$]
4	Balance and add the two half-reactions together.
5	Now add the two cell potentials together. $E°_{cell} = E°_{oxidation} + E°_{reduction}$
	° indicates *standard conditions* (1 atm, 25°C; 1 *M*)

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ $\quad E°_{reduction} = -\,0.76$ V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E°_{reduction} = +\,0.34$ V

More positive is reduced (+ 0.34 V > – 0.76 V) – Cu^{2+} (+0.34) is reduced and Zn oxidized

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $\quad E°_{oxidation} = +\,0.76$ V
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E°_{reduction} = +\,0.34$ V

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ $\quad E°_{cell} = +\,1.10$ V