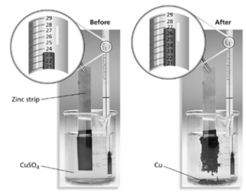
Electrochemistry

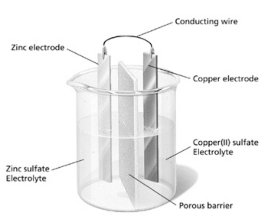
Electrochemical Cells

- Redox reactions involve the ____________ of electrons
- If the two substances are in contact with each other, ____________ will accompany the electron transfer



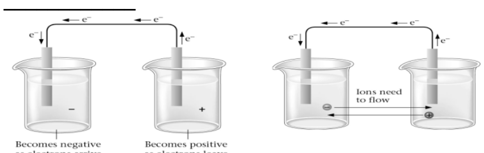
Electrochemical Cells

- But, if we separate the substances that are oxidized & reduced, you will get ____________ energy instead of heat



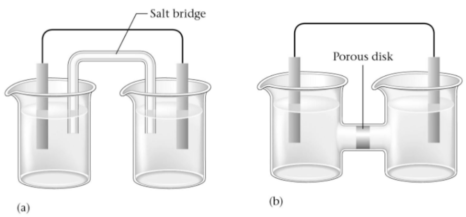
Electrochemical Cells

- If just separated & connected by a wire, ____________ current will flow
- The current stops flowing because of a ____________ in each compartment. We need to complete the ____________



Salt Bridge

- This can be fixed with a ___________ or ___________ connecting the two solutions



Galvanic cell

- _________(_________) cell – a device in which chemical energy is converted into electrical energy
  - ________– compartment in which **oxidation** occurs
  - ________– compartment in which **reduction** occurs

Galvanic Cells

- Cell notation
- anode electrode | anode solution || cathode solution | cathode electrode
- When no solid metal is present in one of the cells, you will need an inert conductor (Pt) & place in ( ) in a cell diagram
  - The double line represents the ________

Cell Potential

- _________ (E_{cell})– the pull or driving force on the electrons
- Also called the _________ (emf)
- Unit is the _________
- 1 volt – 1J/Coulomb
- Voltmeter – measures cell potential

Standard Reduction Potentials

STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			E°(V)
$F_2(g) + 2e^-$	→	$2F^-$	2.87
$Co^{3+} + e^-$	→	Co^{2+}	1.82
$Au^{3+} + 3e^-$	→	$Au(s)$	1.50
$Cl_2(g) + 2e^-$	→	$2Cl^-$	1.36
$O_2(g) + 4H^+ + 4e^-$	→	$2H_2O(l)$	1.23
$Br_2(l) + 2e^-$	→	$2Br^-$	1.07
$2Hg^{2+} + 2e^-$	→	Hg_2^{2+}	0.92
$Hg^{2+} + 2e^-$	→	$Hg(l)$	0.85
$Ag^+ + e^-$	→	$Ag(s)$	0.80
$Hg_2^{2+} + 2e^-$	→	$2Hg(l)$	0.79
$Fe^{3+} + e^-$	→	Fe^{2+}	0.77
$I_2(s) + 2e^-$	→	$2I^-$	0.53
$Cu^+ + e^-$	→	$Cu(s)$	0.52
$Cu^{2+} + 2e^-$	→	$Cu(s)$	0.34
$Cu^{2+} + e^-$	→	Cu^+	0.15
$Sn^{4+} + 2e^-$	→	Sn^{2+}	0.15
$S(s) + 2H^+ + 2e^-$	→	$H_2S(g)$	0.14
$2H^+ + 2e^-$	→	$H_2(g)$	0.00
$Pb^{2+} + 2e^-$	→	$Pb(s)$	-0.13
$Sn^{2+} + 2e^-$	→	$Sn(s)$	-0.14
$Ni^{2+} + 2e^-$	→	$Ni(s)$	-0.25
$Co^{2+} + 2e^-$	→	$Co(s)$	-0.28
$Tl^+ + e^-$	→	$Tl(s)$	-0.34
$Cd^{2+} + 2e^-$	→	$Cd(s)$	-0.40
$Cr^{3+} + e^-$	→	Cr^{2+}	-0.41
$Fe^{2+} + 2e^-$	→	$Fe(s)$	-0.44
$Cr^{3+} + 3e^-$	→	$Cr(s)$	-0.74
$Zn^{2+} + 2e^-$	→	$Zn(s)$	-0.76
$Mn^{2+} + 2e^-$	→	$Mn(s)$	-1.18
$Al^{3+} + 3e^-$	→	$Al(s)$	-1.66
$Be^{2+} + 2e^-$	→	$Be(s)$	-1.70
$Mg^{2+} + 2e^-$	→	$Mg(s)$	-2.37
$Na^+ + e^-$	→	$Na(s)$	-2.71
$Ca^{2+} + 2e^-$	→	$Ca(s)$	-2.87
$Sr^{2+} + 2e^-$	→	$Sr(s)$	-2.89
$Ba^{2+} + 2e^-$	→	$Ba(s)$	-2.90
$Rb^+ + e^-$	→	$Rb(s)$	-2.92
$K^+ + e^-$	→	$K(s)$	-2.92
$Cs^+ + e^-$	→	$Cs(s)$	-2.92
$Li^+ + e^-$	→	$Li(s)$	-3.05

Standard Conditions
1M
1atm
25°C

Standard Reduction Potentials

- Consider the following reaction:
- Al^{+3} (aq) + Mg (s) → Al (s) + Mg^{+2} (aq)
- Write the balanced equation
- Calculate E_{cell}
- Write the cell diagram

Example

- What would the electrical potential be for the following reaction?
- PbO_2 + Na → Pb^{+2} + Na^+

Spontaneous Direction

- A cell will always run spontaneously in the direction that produces a _________ cell potential

Example

- Fe^{+3} + Mg → Mg^{+2} + Fe^{+2}
- Calculate voltage

Example

- Is H_2 capable of reducing Ag^+? Explain why or why not.

Example

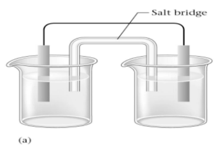
- Is Fe^{+2} capable of reducing Cr^{+3}?

Example

- Given
- Fe^{+2} + 2e- → Fe   E = -0.44v
- MnO_4^- + 5e- + 8H^+ → Mn^{+2} + 4H_20 E = 1.51v
- Write the balanced cell reaction
- What is the E cell for this reaction?

Example

- Draw a sketch of the cell. Label the anode, cathode, & the direction of the electron flow



- Write the cell diagram for the reaction
- Fe | Fe $^{+2}$ || MnO_4^-, Mn^{+2} | Pt