Review

I. Re: Balancing Redox Reactions.
A. Every redox reaction requires a substance to be…
1.
2.
3.
B. The part of the reaction that focuses is known as the

C. Balancing using the half-reaction method (Use if you are given ions)
1. Identify what is being reduced and oxidized.
2. Break the reaction down into two half reactions.
a. If a species is not part of the reaction, leave it out
3. Balance all atoms.
4. Write in the number of electrons gained or lost as a product or reactant.
5. Multiply each half-reaction until the electrons are equal for both.
6. Recombine the half-reactions and cancel out any repeating species (there should be no electrons left).
7. Example: $Fe^{3+}(aq) + Cu(s) \longrightarrow Cu^{2+}(aq) + Fe(s)$

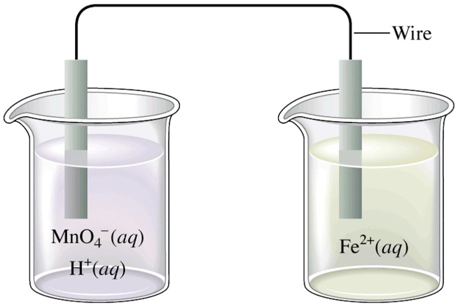
D. For acidic solutions:
1. Complete steps C1-3.
2. Balance the O's with H_2O
3. Balance the H's with H^+
4. Complete steps C4-5.
5. Example: Balance the following in acidic medium: $Al(s) + MnO_4^-(aq) \longrightarrow Al^{3+}(aq) + Mn^{2+}(aq)$

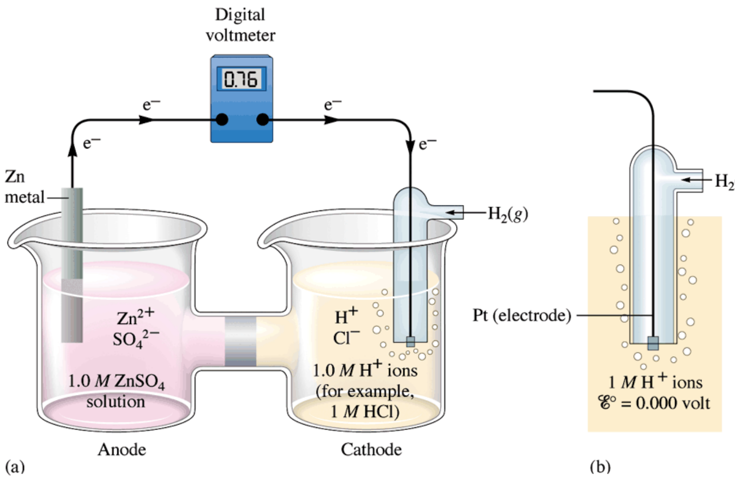
E. For basic solutions:
1. Complete steps D1-3.
2. For each H^+, add an OH^- to both sides.
3. Combine H^+ and OH^- to form H_2O
4. Cancel out H_2Os
5. Complete step D4
6. Ex: Balance the following in a basic medium: $Mg(s) + OCl^-(aq) \longrightarrow Mg(OH)_2(s) + Cl^-(aq)$

Galvanic Cells (Batteries!)

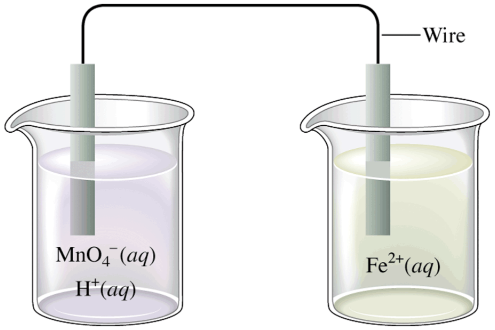
I. An apparatus in which (electricity).
II. Components of a galvanic cell.
A. Need a to occur.
B.
1. A must be present
2. If reactants were placed in the same container,

3. Usually, If the half reaction



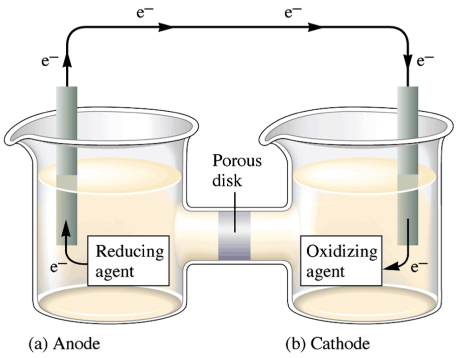


4. Problem with the cell below:



a. The electrons will

b. We ( ).



C. The cell

1. Can be

2.

3.

4. Will for a long time.

D. Other definitions:

1. Oxidation occurs at the <u>anode</u>. Negative charge

2. Reduction occurs at the <u>cathode</u>. Positive charge.

E. Draw a galvanic cell for the following reaction:

$Cu^{2+} + Zn(s) \longrightarrow Zn^{2+}(aq) + Cu(s)$

Show the direction of electron and ion flow.

1. Electrons flow

F. What if one or both of the reactions do not involve metal electrodes? Sometimes a cell requires a gas like H_2.

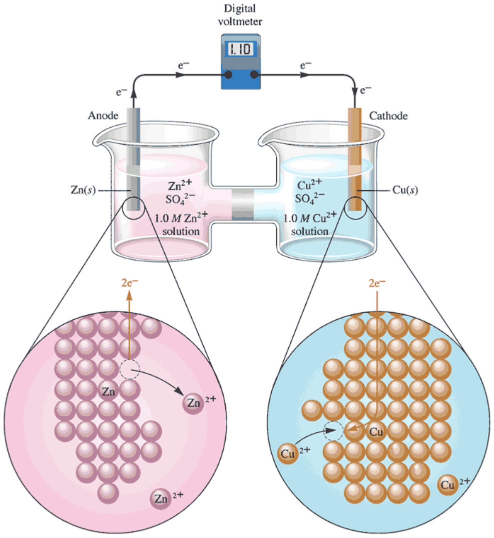
III. What happens to the electrodes over time?

A.

B.

IV. Line notation: How to identify cells:

- A. A
- B. A
- C. Within each half-cell the
- D. The line notation for the
- E.
- F. Example:



Standard Electron Reduction Potentials

I. The most important table (p. 481) in this chapter!!!

A. Hints:

1. All of the
2.
3. You can use this table to help you balance redox reactions!
4. All of these potentials are measured in standard state:
5.
6. You can use it to

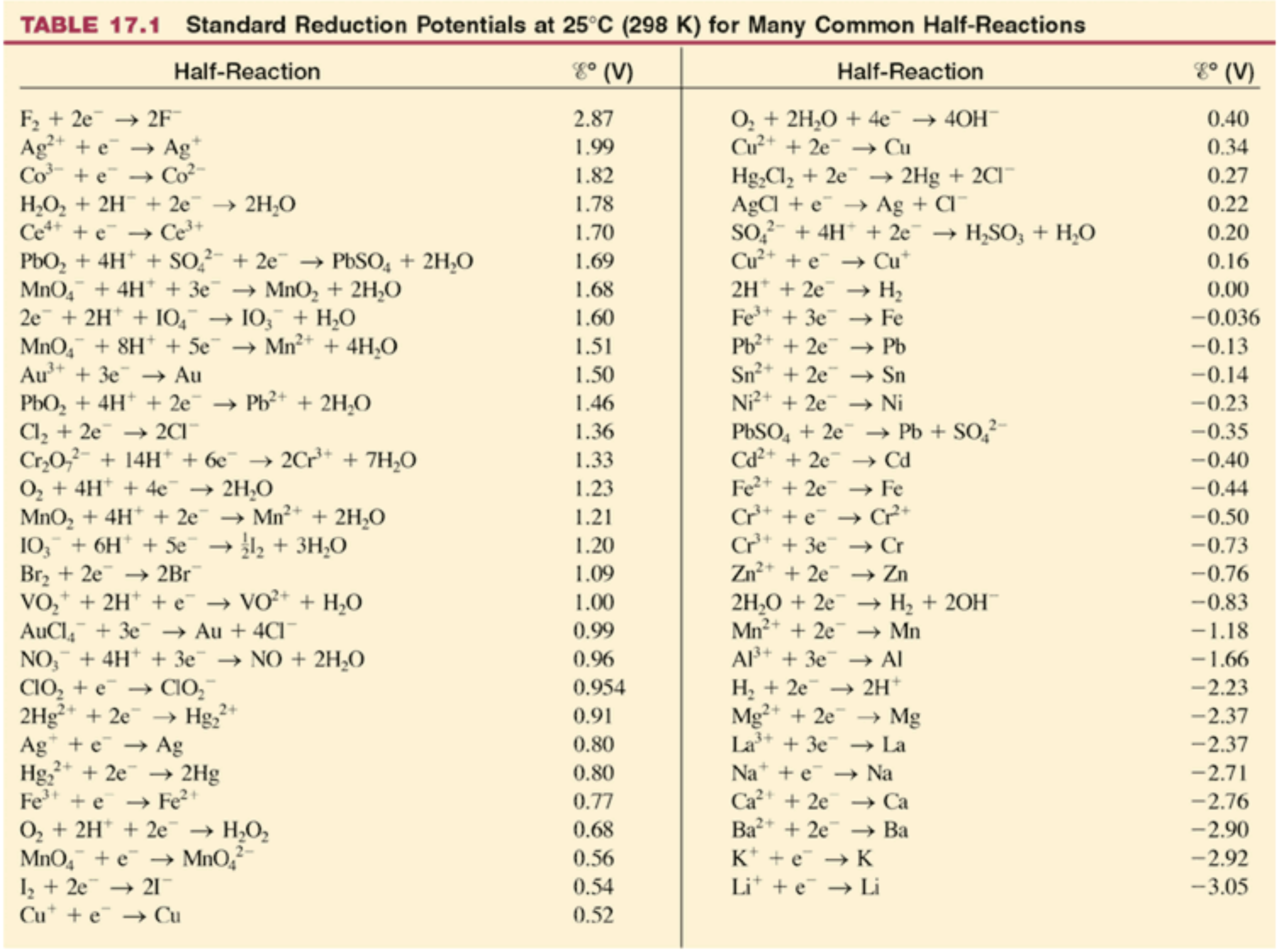


TABLE 17.1 Standard Reduction Potentials at 25°C (298 K) for Many Common Half-Reactions

Half-Reaction	$\mathscr{E}^\circ$ (V)	Half-Reaction	$\mathscr{E}^\circ$ (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	0.40
$Ag^{2+} + e^- \rightarrow Ag^+$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3-} + e^- \rightarrow Co^{2-}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.27
$H_2O_2 + 2H^- + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$2e^- + 2H^+ + IO_4^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	−0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	−0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	−0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	−0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	−0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	−0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	−1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	−1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^+$	−2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	−2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	−2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.80	$Na^+ + e^- \rightarrow Na$	−2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	−2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	−2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	−2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	−3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		

II. Calculating potentials:
A.
B.
C. If you need to multiply a half reaction to balance the electrons,

D. If the
E. Examples:
1. $PbO_2 + Na \rightarrow Pb^{2+} + Na^{+}$
2. $Fe^{3+} + Mg \rightarrow Mg^{2+} + Fe^{2+}$
III. Reduction strength
A.
B. Examples
1. Is $H_2(g)$ capable of reducing $Ag^{+}(aq)$?
2. Is $H_2(g)$ capable of reducing $Ni^{2+}(aq)$?
3. Is $Fe^{2+}(aq)$ capable of reducing $VO_2^{+}(aq)$?
4. Is $Fe^{2+}(aq)$ capable of reducing $Cr^{3+}(aq)$?
5. Rank the following from the strongest oxidizing agent to the weakest oxidizing agent: Ce^{4+}, Fe^{2+}, Fe^{3+}, Mg^{2+}, Mg, Ni^{2+}, Sn

Cell Potential and Free Energy

I. Definitions:
A. Volt (V):
B. Coulomb (C):
C. Faraday (F):
II.
A.
1. n=
2. F=
B. In standard conditions,
C. Ex: What is the free energy change involved in following reaction: $Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$?

D. Ex: Calculate the free energy change for the following reaction: $Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq)$

E. This is why a positive voltage yields a spontaneous reaction.

Cell Potential and Concentration

I. Remember our assumption that everything was at 1.0M
A.
B. These two are related by
II. Free energy (ΔG) is related to cell potential ($\overset{\ominus}{E}$)
A.

1. $\overset{\ominus}{E}{}^{\circ}$ =
2. Q =
3. R= ; T= ; F=96,500C/mol
B. Since R, T and F are all constants, you can simplify the equation
C. Simplified eqn:

1.
2.

III. Ex: What is the cell potential of the following cell with the following concentrations? $VO_2^+ + Zn \rightarrow Zn^{2+} + VO^{2+}$ $[VO_2^+] = 2.0M$; $[H^+] = 0.50M$; $[Zn^{2+}] = 0.10M$; $[VO^{2+}] = 1.0 \times 10^{-2}M$

A. First:
B. Second,
C.

IV. Ex: What is the concentration of Zn^{2+} when the E is 1.20 V given the concentration of Cu^{2+}? $Cu^{2+} + Zn \rightarrow Zn^{2+} + Cu$

A.
B.

V. Concentration Cells:

A. Can you have a current with the same electrodes and a difference in concentration?
B. YES!!! Example:

Cell Potential and Equilibrium

I. At equilibrium:

A.
B.
C. Therefore,

D. Rearranging it, you get:

E. With it, given a galvanic cell.
F.

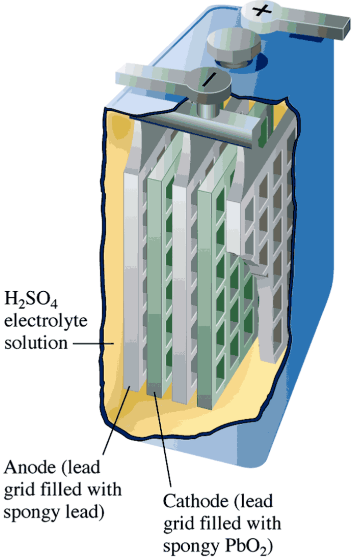
The Faraday

I. What we have learned is the constant (conversion unit).

II. It can be a measured unit.

A. Example: 3.0 faradays passed through the following electrolytic cells. How many m plated?

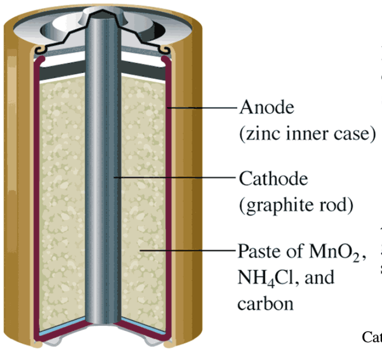
1. $AgNO_3$
2. $FeSO_4$
3. $FeCl_3$



Applications

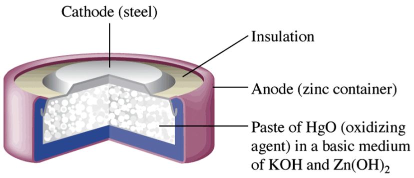
I. A battery is

A. All batteries need
B. Car battery:
  1. battery acid is sulfuric acid.
  2. spongy for
C. Alkaline battery
  1. Cathode is a graphite rod.
  2.
  3. Anode/Cathode are solids.



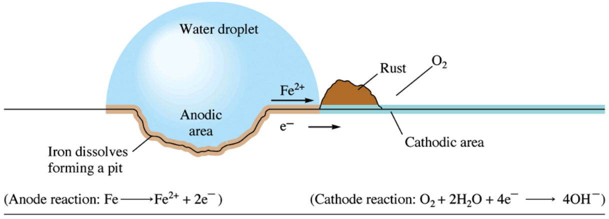
II. Corrosion: When iron rusts…

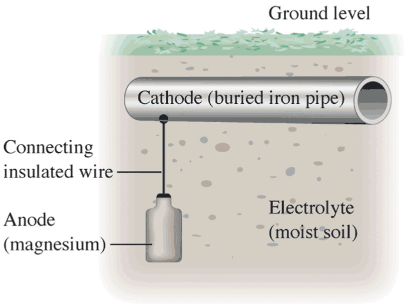
A. Rusting happens when a mini-galvanic cell is formed.
B. The



C. How do you stop this?
  1. Do not expose iron to water (because it forms a solution).
  2. Paint it! Prevents a solution forming with iron.

D. Another way to do this is to use a
  1. Use a





Electrolysis

I. Electrolysis:

II. Electrolytic cell:
  A. This is created by
  B.
  C. The reverse reaction takes place and it is

  D. These reactions

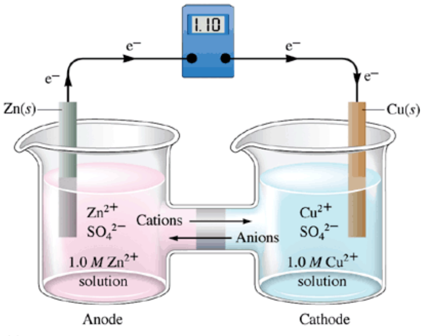
III. Electrolysis problems
  A. First some definitions:
    1. Ampere (A):
    2. Remember the Faraday:
  B. These problems are like stoichiometry problems ( ).
  C. Below: using electrolysis to split water ( ).
  D. Ex.: *Calculate the amount of time required to produce 1000 g of magnesium metal by electrolysis of molten* $MgCl_2$ using a current of 50A.

  E. Ex: *A* Cr^{3+} *(aq) solution is electrolyzed, using a current of 7.60A. What mass of Cr (s) is plated out after* 2.00 days?

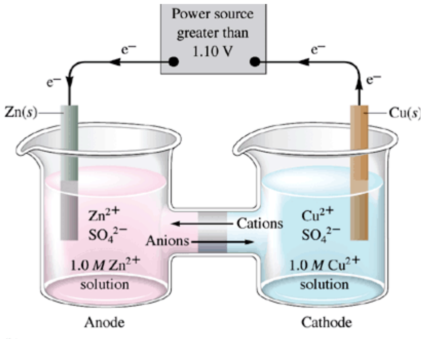
    1. Known as
    2. Can also be used to produce pure metals.
  F. *What amperage is required to plate out 0.250 mol Cr form a* Cr^{3+} *solution in a period of 8.00 hours?*

IV. Determining electrolytic cells.
  A. Reactions can be
  B. You can determine what reaction will occur by

  C. Examples: *What reaction will take place at the cathode and the anode when each of the following is electrolyzed?*
    1. *1.0M KF solution*
    2. *1.0M* $CuCl_2$ *solution*
    3. *1.0M* H_2O_2 *solution containing 1.0M HCl*



(a)



(b)