	V
$Ag^+ \rightarrow Ag$	+0.80
$Cu^{2+} \rightarrow Cu$	+0.34
$Zn^{2+} \rightarrow Zn$	-0.76
$Al^{3+} \rightarrow Al$	-1.66
$Mg^{2+} \rightarrow Mg$	-2.36

12. Which species react with Cu^{2+}?

13. Which species react with Zn°?

14. Which element loves e's the most? Least?

15. $NiCl_2 + H_2 \rightarrow Ni + 2\,HCl$ E° = -0.28 V

a. Product favored or not?

b. Is reduction potential for Ni^{2+} positive?

19.6 $E°_{cell}$ and ΔG°, E° and K

	ΔG	$E°_{cell}$	K
Product Favored	neg	pos	large
Reactant Favored	pos	neg	small
**Equilibrium	0	0	1

ΔG° and $E°_{cell}$ have opposite signs, but are related

- both provide measurements for the favorability or unfavorability of a reaction
- obviously $E°_{cell}$ is more limited, to redox reactions
- K is also related, since it too relates to how favorable or unfavorable a reaction is
- ΔG° = "free energy" available to do be released and do work
- $E°_{cell}$ also reflects the amount of energy that is released to do work when a favorable redox transfer occurs
  - The "free energy" in a cell is really the free energy to do the work of moving electrons and to the work that flowing electricity can do

$$\Delta G° = -nFE°_{cell} \qquad \boxed{\Delta G° = -96.5nE°_{cell}}$$

n = number of electrons transferred in the balanced equation (now coefficients matter!!)

- crucial that you have a correctly balanced redox reaction, and can count how many electrons transfer

F = Faraday's constant **= 96.5** to get ΔG in kJ/mol

$$\boxed{\Delta G° = -nFE°_{cell} \qquad\qquad \Delta G° = -96.5nE°_{cell}}$$

Units $F = \frac{96{,}500C}{\text{mole } e^-}$ $\quad V = \frac{J}{C}$ so $C = \frac{J}{V}$

C = coulomb,
unit of electricity, amount of charge

Substituting $F = \frac{96{,}500J}{\text{mole } e^- \bullet V}$ $\quad F = \frac{96.5\text{ kJ}}{\text{mole } e^- \bullet V}$

Thus when "n" is moles of electrons, and $E°_{cell}$ is in volts, the units cancel and only kJ are left.

<u>Electrochemistry-Related Units/Terms: For interest, not for test</u>

C = Coulomb = <u>quantity</u> <u>of</u> electrical <u>charge</u> = $6.24 \bullet 10^{18}$ electrons
- 1 mole of electrons = 96,500 C

A = amp = <u>rate of charge flow</u> per time = Coulombs/second

V = volt = <u>electrical power/force/strength</u>; difference in electrical potential energy = J/C
- Force for moving electrons and charge
- Not all Coulombs of charge have the same energy/power/force/ability to do work
- Just like dropping a brick from one cm has less force than dropping it from two meters high

F = Faraday = charge per chemical amount (the mole) = $\frac{96{,}500C}{\text{mole } e^-} = \frac{96.5\text{ kJ}}{\text{mole } e^- \bullet V}$

Watt = amount of energy

1. Balance the reaction, and find ΔG° given the reduction potentials shown

$$Cu + \underset{+0.77V}{Fe^{3+}} \rightarrow \underset{+0.34V}{Cu^{2+}} + Fe^{2+}$$

2. $Zn + Cr^{3+} \rightarrow Zn^{2+} + Cr \qquad \Delta G° = -11.6\text{ kJ/mol}$

a. Balance the reaction, and calculate $E°_{cell}$.

b. If the reduction potential for Zn^{2+} is -0.76V, what is the reduction potential for Cr^{3+}?

Cell Voltage and K Likewise voltage and K linked!!

- The more favorable and positive $E°_{cell}$, the larger and more favorable is K
- Again, "n" = number of electrons transferred, so you need balanced reaction
- Caution: K values often work out to be enormous (calculator problems)

$$\boxed{\log K = nE°_{cell} / (0.0592)}$$

3. Calculate K, given reduction potentials.

Ag^+ + Fe^{2+} → $Ag°$ + Fe^{3+}
+0.80V +0.77V

19.7 The Effect of Concentration on Cell Potential: Voltages when Concentrations are not all Equal to Standard 1.0 M

1. E° assumes 1.0 M concentrations for any soluble species (and 1.0 atm pressure for any gas)
   - rarely actually true!
2. For any real reaction, concentrations change as the reaction procedes
   - As the concentrations change, the voltage drops
   - Actual voltage continues to drop until the battery is dead = 0V = equilibrium
   - At equilibrium, $E_{actual} = 0$ V

$$\boxed{\text{Nernst Equation:} \quad \mathbf{E_{actual} = E° - \frac{.0592}{n} \log Q}}$$

n = number of e's transferred (need balanced equation, coefficients)
Q = ratio of actual concentrations (K format, but using actual concentrations)

- Recall: solids, liquids don't appear in K or Q, only aqueous solutes or gases

Problems

1. Calculate actual voltage for $Mg/Mg^{2+}(0.10M)//Cu^{2+}(0.001M)/Cu$
given the following reduction potentials:

$Mg^{2+} = -2.37V$ Logic Steps
$Cu^{2+} = +0.34V$

$$\boxed{\text{Nernst Equation:} \qquad \mathbf{E_{actual} = E° - \frac{.0592}{\mathit{n}} \log Q}}$$

2. Calculate actual voltage for Cu/Cu^{2+}(1.0M)//Ag^+(0.032M)/Ag

$E_{red}°$	
Ag^+	+0.80V
Cu^{2+}	+0.34V

3. $2\ Ag^+(aq) + Zn(s) \rightarrow 2\ Ag(s) + Zn^{2+}(aq)$

$E_{red}°$ +0.80 -0.76

If a cell with $[Ag^+] = 0.20$ M has $E_{actual} = 1.63V$, what is $[Zn^{2+}]$

Cell Potential and Equilibrium

At equilibrium:

a. $E_{actual} = 0$ V

b. Q=K

So, at equilibrium $0 = E° - \frac{.0592}{n} \log K$

At equilibrium: $\boxed{E° = \frac{.0592}{n} \log K}$ Finding E° given K

Rearranged: $\boxed{\log K = nE°_{cell} / (0.0592)}$ Finding K given $E°_{cell}$

"Concentration cells": anode and cathode use the same things, but with ions at different concentrations

- at equilibrium, the concentrations would be equal, so the voltage drive is to equalize

Example: $H_2 + 2\ H^+ \rightleftharpoons H_2 + 2\ H^+$ $E° = 0$

So $E_{actual} = (-0.592/n)\log Q$

This kind of voltage is key to pH meters, neurons (19.8)

- pH meter: dip meter with known $[H^+]$ into a solution, measured voltage reflects solution $[H^+]$
- neurons: The H^+ concentration differs inside and outside cell membranes. This creates a voltage which is the key for nerve sensation

19.9 Common Batteries

A. Primary ("Dry Cell"): Nonrechargeable

- run till concentration achieves equilibrium = dead = toss

1. Alkaline batteries

$$H_2O + \underset{\text{oxidized}}{\underset{\text{anode}}{Zn(s)}} + \underset{\text{reduced}}{\underset{\text{cathode}}{MnO_2(s)}} \rightarrow ZnO(aq) + 2\ MnO(OH)(s) \qquad E° = 1.54V$$

- reduction occurs at a graphite electrode
- this is common when an electrode doesn't involve a redox
- flashlights, radio, toys, Jasperse insulin pump, Jasperse blood testers, tooth brush, etc.

2. Mercury battery

$$Zn + HgO \rightarrow ZnO + Hg \qquad E° = 1.35\ V$$

- Less power than alkaline batteries, but mercury batteries are physically smaller
- used in <u>small</u> things (calculators, watches, cameras,…)
- mercury is poisonous, so battery disposal an environmental issue

B Secondary Batteries ("nicad" and "car") = <u>Rechargeable</u>

1. Lead-acid (car battery)

$$\underset{\text{Anode}}{Pb°(s)} + \underset{\text{cathode}}{\overset{+4}{PbO_2(s)}} + 2\ H_2SO_4 \rightleftharpoons 2\ PbSO_4(s) + 2\ H_20 \quad + \text{energy} \qquad E° = 2.0V$$

↑

CAREFUL!

Leakage caused corrosion!

- The $PbSO_4(s)$ coats electrodes, so reaction can be reversed when "recharged"
- Each cell is 2.0V: six alternating cathode/anodes in series sums to 12V
- Energy during a recharge drives it in the reverse direction, to the left
- Side products during recharge
  - Ox: $6\ H_2O \rightarrow O_2 + 4\ H_3O^+ + 4e^-$
  - Red: $4\ H_2O + 4e^- \rightarrow 2\ H_2 + 4\ OH^-$

- Side Both H_2 and O_2 are produced during the recharge. These are a perfect recipe for an explosion. Why no sparks or cigarette lighting is allowed around a car battery

2. NiCad E°=1.3

- electric shavers, dustbusters, video camcorders, rechargeable power toothbrush, any rechargeable cordless appliances

$$\underset{\text{Anode}}{\overset{0}{Cd(s)}} + 2\ \underset{\text{cathode}}{\overset{+3}{NiO(OH)(s)}} + 2\ H_20 \rightleftharpoons \overset{+2}{Cd(OH)_2} + 2\ \overset{+2}{Ni(OH)_2(s)} + \text{energy} \qquad E°=1.3$$

- again, solid products stay on electrodes, so the reaction can reverse upon treatment with electrical energy

3. Lithium batteries $Li(s) + CoO_2(s) \rightleftharpoons LiCoO_2(s)$ + energy $\boxed{E° = 3.4V}$
(in polymer
with carbon graphite)

Pros:
- Big voltage → good for fueling energy eaters, like laptops, cameras, cell phones
- Light weight

Cons
- More expensive

19.10 Fuel Cell: continuous feed of reactants from outside to electrodes (interest, not test)

$2\ H_2 + O_2 \rightarrow 2\ H_2O$ E° = 0.9V
Anode cathode

- $H_2 + O_2$ light, so good fuels, high energy efficiency
- Spaceships: 500 pounds of fuel enough energy for 11 days
- Dream: come up with some way to use solar/wind energy to produce H_2 from water, then use the H_2/O_2 fuel cell to get energy and regenerate H_2O, pollution free
- Fuel cells future for cars??

19.11 Electrolysis: Using outside electrolysis to force unfavorable redox reactions to proceed to product side
- key route to elements not found in nature: metals, H_2, Cl_2,…

A. Electrolysis of Molten Salts ("molten" = melted, pure liquid salts in absence of solvent, super hot!!)

Ex. Energy + $2\ Na^+Cl^- \rightarrow 2\ Na + Cl_2$ E° = -4.07V

Cathode
e- acceptor

Anode
e- source

(Brown, Gillespie overheads)

- products must be kept separate so can't react

B. Electrolysis of salts in Water: Can only process ions that are more reactive than water
- At each electrode, the most reactive candidate reacts
  - In water, water competes at both the cathode (reduction) and anode (oxidation)
- Reduction/Cathode
  - If a cation is harder to reduce than water itself, water will just get reduced instead
  - If you want to reduce something that is harder to reduce than water, you need to do it as a molten salt rather than in water
  - Only cations with reduction potentials more positive than -0.83 V (in basic water) or -0.41 V (in neutral water) can be reduced in water.
  - Water's reduction potential is pH dependent (since hydroxide concentration factors)

Water reduction: $2\ H_2O + 2\ e^- \rightarrow H_2 + 2\ OH^-$ E° = -0.83 (when hydroxide is 1M)
$2H_2O + 2e \rightarrow H_2 + 2OH^-\ (10^{-7}\ M)$ E = -0.41 (when hydroxide is 10^{-7} M)

Problem: Which of the following metal cations could be converted into elemental metal by electrolysis in water? For which metal cations would you need to use molten salt if you wanted to reduce them?

$Hg^{2+} \rightarrow Hg$	0.9V
$Cu^{2+} \rightarrow Cu$	0.2V
$Co^{2+} \rightarrow Co$	-0.3V
$2\ H_2O \rightarrow H_2 + 2\ OH^-$	-0.4V
$Mn^{2+} \rightarrow Mn$	-1.2V
$Mg^{2+} \rightarrow Mg$	-2.4V

- Easily reduced cations (Zn^{2+}, Ni^{2+}, Cr^{3+}, Sn^{2+}, etc.) can be reduced to elemental form in water.
- Cations of Active metals can't (K^+, Mg^{2+}, Na^+,…). If they are to be reduced to elemental form, they must be reduced as molten salts.

Oxidation/Anode

- If a reduced species is harder to oxidize than water itself, water will just get oxidized instead
- If you want to oxidize something that is harder to oxidize than water, you need to do it as a molten salt rather than in water
- Only reduced species with oxidation potentials more positive than -1.23 V (in acidic water) or -0.82 (in neutral water) can be oxidized in water.

Water Oxidation:	$2\ H_2O \rightarrow O_2 + 4\ H^+ + 4e-$	$E° = -1.23$ ($[H^+] = 1.0$ M)
	$2\ H_2O \rightarrow O_2 + 4\ H^+ + 4e-$	$E = -0.82$ ($[H^+] = 1 \times 10^{-7}$ M)

Problem: Which of the following oxidations could be conducted by electrolysis in water? And which processes would require molten salts??

$2F^- \rightarrow F_2 + 2e$	-2.87
$2Cl^- \rightarrow Cl_2 + 2e$	-1.36
$2Br^- \rightarrow Br_2 + 2e$	-1.08
$2H_2O \rightarrow O_2 + 4\ H^+\ (10^{-7}\ M) + 4e$	-0.82
$2I^- \rightarrow I_2 + 2e$	-0.535
$Cr^{2+} \rightarrow Cr^{3+} + e$	+0.41
$Al \rightarrow Al^{3+} + 3e$	+1.66

$Cl_2 + 2e^- \rightarrow 2Cl^-$	E° = 1.36	$Sn^{2+} + 2e^- \rightarrow Sn$	E° = –0.14 V
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	E° = 1.23 V	$2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$	E° = –0.83 V
$I_2 + 2e^- \rightarrow 2I^-$	E° = 0.54 V	$Mn^{2+} + 2e^- \rightarrow Mn$	E° = –1.18 V

1. Given the eduction potentials, what is the product at the <u>anode</u> and at the <u>cathode</u> when a current is passed through an aqueous solution of $SnCl_2$? (Hint: remember which chemicals and ions are really in the solution and subject to the electrolysis.)

<u>Anode</u>	<u>Cathode</u>
a. Sn	a. Sn
b. Cl_2	b. Cl_2
c. O_2	c. O_2
d. H_2	d. H_2
e. none of the above	e. none of the above

2. Given the eduction potentials, what is the product at the <u>anode</u> and at the <u>cathode</u> when a current is passed through an aqueous solution of MnI_2? (Hint: remember which chemicals and ions are really in the solution and subject to the electrolysis.)

<u>Anode</u>	<u>Cathode</u>
a. Mn	a. Mn
b. I_2	b. I_2
c. O_2	c. O_2
d. H_2	d. H_2
e. none of the above	e. none of the above

<u>Electroplating</u>: metal cation → elemental metal (reduction at cathode)

- metal forms on surface of cathode
- many metals are "plated" on outside of things in their way
- "Silverware" for a long time involved plating a coating of silver over something else
- Art objects, etc.
- Materials that are otherwise subject to rust, corrosions are often electroplated with a coating that is resistant to air, rain, and acid.

<u>Some Famous Electrolyses</u> (trivia):

1. NaCl in H_2O → NaOH (anode) and HCl and O_2 (cathode) NaOH, HCl production

2. NaCl (molten) → Na metal (cathode) + Cl_2 (anode) Cl_2 production

19.12 Electrolysis Calculations

- **1 mol electrons = 96,500 C (Coulombs)**

current, time, and moles of electrons are related

A (amp) = C/sec

A Derivation and 3 Permutations of an equation:

$$\boxed{\textbf{Moles electrons} = \frac{current(A)\cdot \sec}{96{,}500}} \qquad \textbf{Sec} = \frac{(moles)96{,}500}{A} \qquad \textbf{A} = \frac{(moles)96{,}500}{\sec}$$

Finding moles, given current and time

Finding time, given moles and current

Finding current, given moles and time

Qualitative Relationship (and vice versa):

Amps + time → **moles of electrons** → **moles of substance redoxed** → **grams of substance**

Keys:

1. Grams of substance and moles of substance are interconverted by molecular weight
2. Be sure to factors how many moles of electron are involved per moles of chemical formula

1. How many grams of Al (27g/mol) is produced in 1.0 hour by electrolysis of $AlCl_3$ at 10.0A current?

2. At 3.2A, how long will it take to make 10g of Zn (65.4 g/mol) from $ZnBr_2$?

3. What current in amps is required to make 10 grams of Cl_2 (71 g/mol) from $AlCl_3$ in one hour?

19.13 Corrosion

Corrosion involves a product-favored oxidation of a metal exposed to environment (O_2, H^+, H_2O,…)

The metal being oxidized always functions as the oxidation half

Molecular oxygen is reduced to water in the presence of acid as the reduction half

As for any favorable redox reaction, the sum of the two half reactions must give positive E

Thus, the favorability of the oxygen reduction half is critical in determining which metals can and cannot be oxidized

The reduction of oxygen in the presence of acid is a rather favorable reduction half reaction

$O_2 + 4H^+ \rightarrow 2\ H_2O$ $E°_{red} = +1.23V$ very good under standard conditions!!

Obviously acid that is "standard conditions" 1.0 M is rare

The more acidic the water environment, the more favorable oxygen reduction is and the more metals can be corroded

Under 1.0 M acid conditions (pH = 1), any metal that has an oxidation potential better than -1.23 V can be oxidized in air

Under neutral pH = 7 conditions, any metal that has an oxidation potential better than -0.82V can be oxidized in air

Most metals are included, especially under acidic conditions!!

Why most metals are not found in their elemental form in nature, but rather as ions

Exception: gold!!

Metals usually end as metal oxides or sometimes metal hydroxides

Ag tarnish
Cu "greening"
Fe rusting
Rust: $2\ Fe + O_2 + 2\ H_2O \rightarrow 2\ Fe(OH)_2 \rightarrow \rightarrow \rightarrow Fe_2O_3$ red-brown rust

Practical notes:

1. Corrosion often speeded by H^+ and/or ionic salts that acidity water
2. Gold has always been valued because unlike other oxidizable metals, it retains it's elemental form and it's lustrous golden elemental surface appearance.
3. Most metals get coated with a film of hard metal oxide, which ends up protecting the interior or the metal.
   - The interior stays elemental metal, but is protected by sheath of hard metal oxide from exposure to air.
   - Sometimes it takes chemical activation to clear the oxide film and enable the elemental metal inside to be exposed for chemical reactions.
4. Why does iron have such a special rusting problem?
   - Iron is bad because iron oxide (rust) forms flakes that break off.
   - As a result, the interior iron is **not** protected and is continuously exposed for further corrosion.

Prevention

1. Coat iron surface with something that resists corrosion and protects.

Development of improved and more resistant sealants has been a major priority of auto-industry

2. "Galvanized iron"-Iron materials are electroplated with Zn, which is more easily oxidized than iron but oxidizes to give a hard, protecti[e $Zn(OH)_2$ coating.

Chapter 19 Electrochemistry Math Summary

Relating Standard Cell Potential to Standard Half Cell Potentials

$E^o_{cell} = E^o_{oxidation} + E^o_{reduction}$ (standard conditions assume 1.0 M concentrations)

Relating Half Cell Potentials when Written in Opposite Directions

$E^o_{ox} = -E^o_{red}$ for half reactions written in opposite directions

Relating Standard Cell Potentials to ΔG

$\Delta G^o = -nFE^\circ_{cell}$ (to give answer in kJ, use F = 96.485)

F = 96,500 C/mol
n=number of electrons transferred

Relating Actual Cell Potential to Standard Cell Potential when Concentrations aren't 1.0-M

$E_{cell} = E^o_{cell} - [0.0592/n] \log Q$ (Q = ratio of actual concentrations)

Relating Standard Cell Potential to Equilibrium Constant

$\log K = nE^o/0.0592$

Relating Actual Cell Potential to Actual Concentrations in Concentration Cells

$E_{cell} = -[0.0592/n] \log Q$ for concentration cells, where anode and cathode differ only in concentration, but otherwise have same ions

Relating # of Moles of Electrons Transferred as a Function of Time and Current in Electrolysis

1 mol e^- = 96,500 C

moles of electrons = [current (A)•time (sec)]/96,500 for electrolysis, moles, current, and time are related.

rearranged: time (sec)=(moles of electrons)(96500)/current (in A)

Note: 3600 sec/hour

so time (hours)=(moles of electrons)(26.8)/current (in A)

Electrochemistry-Related Units

C = Coulomb = quantity of electrical charge = $6.24 \cdot 10^{18}$ electrons

- 1 mole of electrons = 96,500 C

A = amp = rate of charge flow per time = C/sec
V = volt = electrical power/force/strength = J/C

$$F = \text{Faraday} = \frac{96{,}500\text{C}}{\text{mole e}^-} = \frac{96.5\text{ kJ}}{\text{mole e}^- \bullet \text{V}}$$

Assigning Oxidation Numbers (See Section 5.4)

This is a more complete set of rules than your textbook. It always works.

Use these rules in order.
The sum of all oxidation numbers of all elements = charge on substance.

		Oxidation Number:	Examples:
1.	Atoms in their elemental state	= 0	Fe, H_2, O_2
2.	Monatomic ions	= charge	F^{1-}, Na^{1+}, Fe^{3+}

IN COMPOUNDS

3.	Group 1A	= +1	NaCl, KNO_3
4.	Group 2A	= +2	MgO
5.	Fluorine	= -1	HF, ClF
6.	Hydrogen	= +1	H_2O
7.	Oxygen	= -2	SO_2, $HClO_4$
8.	Group 7A (Halogen family)	= -1	HCl
9.	Group 6A (Oxygen family)	= -2	PbS_2
The sum of all oxidation numbers of all elements = charge on substance.			

Key: For anything else, (or for a group 7A or group 6A in the presence of higher priority atoms), set it's oxidation number = "x", and solve for "x" such that the ox. #'s = actual charge.

Find Ox #'s for

1. H_2OC C:	2. PCl_3 P:
3. HSO_4^- S:	4. $KMnO_4$ Mn:
5. $Mg_3(PO_4)_2$ P:	6. $HClO_2$ Cl:

Balancing Redox: Simple Cases where all Reactants and Products are Provides

1. Identify oxidation numbers for redox actors
2. Set coefficients for them so that the **#e's released = #e's accepted**
    - focus completely on the atoms whose oxidation numbers change
3. Then balance any redox spectators
4. Check at the end to make sure:
    - Charges balance
    - Atoms balance

Note: Test problems will give you all of the species involved. Some OWL problems will be harder and will not include all of the chemicals

Some Harder OWL-Level Redox-Balancing Problems: When some necessary chemicals are ommitted

a. Sometimes H_2O, $OH^{\ominus}$, $H^{\oplus}$ are omitted, and need to be added in order to balance oxygens and hydrogens
b. In knowing how to do this, it is helpful to distinguish acid versus base conditions
c. Under acid conditions, it's appropriate to have $H^{\oplus}$ but not $OH^{\ominus}$
d. Under base conditions, it's appropriate to have $OH^{\ominus}$ but not $H^{\oplus}$

Acid Conditions	**Base Conditions**
1. Identify oxidation numbers for redox actors	**1.** Identify oxidation numbers for redox actors
2. Set coefficients for them so that the **#e's released = #e's accepted**	**2.** Set coefficients for them so that the **#e's released = #e's accepted**
• focus completely on the atoms whose oxidation numbers change Add H_2O's to balance oxygen	• focus completely on the atoms whose oxidation numbers change Add $OH^{\ominus}$'s to balance charge
3. Add H_2O's as needed to balance oxygens	3. Add $OH^{\ominus}$'s as needed to balance charge
4. Add H^+'s as needed to balance hydrogens and charge	4. Add H_2O's to balance hydrogens
5. Check at the end to make sure:	5. Check at the end to make sure:
a. Charges balance	a. Charges balance
b. Atoms balance	b. Atoms balance

Standard Reduction (Electrode) Potentials at 25° C (OWL)

Half-cell reaction	**E_o (volts)**
$F_2 + 2e \rightarrow 2F^-$	2.87
$Ce^{4+} + e \rightarrow Ce^{3+}$	1.61
$MnO_4^- + 8\,H^+ + 5e \rightarrow Mn^{2+} + 4H_2O$	1.51
$Cl_2 + 2e \rightarrow 2Cl$-	1.36
$Cr_2O_7^{2-} + 14\,H^+ + 6e \rightarrow 2Cr^{3+} + 7H_2O$	1.33
$O_2 + 4H^+ + 4e \rightarrow 2H_2O$	1.229
$Br_2 + 2e \rightarrow 2Br^-$	1.08
NO_3- $+ 4H^+ + 3e \rightarrow NO + 2H_2O$	0.96
$2Hg^{2+} + 2e \rightarrow Hg_2^{2+}$	0.920
$Hg^{2+} + 2e \rightarrow Hg$	0.855
$O_2 + 4\,H^+\ (10^{-7}\,M) + 4e \rightarrow 2H_2O$	0.82
$Ag^+ + e \rightarrow Ag$	0.799
$Hg_2^{2+} + 2e \rightarrow 2Hg$	0.789
$Fe^{3+} + e \rightarrow Fe^{2+}$	0.771
$I_2 + 2e \rightarrow 2I^-$	0.535
$Fe(CN)_6^{3-} + e \rightarrow Fe(CN)_4^{4-}$	0.48
$Cu^{2+} + 2e \rightarrow Cu$	0.337
$Cu^{2+} + e \rightarrow Cu^+$	0.153
$S + 2H^+ + 2e \rightarrow H_2S$	0.14
$2H^+ + 2e \rightarrow H_2$	0.0000
$Pb^{2+} + 2e \rightarrow Pb$	-0.126
$Sn^{2+} + 2e \rightarrow Sn$	-0.14
$Ni^{2+} + 2e \rightarrow Ni$	-0.25
$Co^{2+} + 2e \rightarrow Co$	-0.28
$Cd^{2+} + 2e \rightarrow Cd$	-0.403
$Cr^{3+} + e \rightarrow Cr^{2+}$	-0.41
$2H_2O + 2e \rightarrow H_2 + 2OH^-\ (10^{-7}\,M)$	-0.41
$Fe^{2+} + 2e \rightarrow Fe$	-0.44
$Cr^{3+} + 3e \rightarrow Cr$	-0.74
$Zn^{2+} + 2e \rightarrow Zn$	-0.763
$2H_2O + 2e \rightarrow H_2 + 2OH^-$	-0.83
$Mn^{2+} + 2e \rightarrow Mn$	-1.18
$Al^{3+} + 3e \rightarrow Al$	-1.66
$Mg^{2+} + 2e \rightarrow Mg$	-2.37
$Na^+ + e \rightarrow Na$	-2.714
$K^+ + e \rightarrow K$	-2.925
$Li^+ + e \rightarrow Li$	-3.045