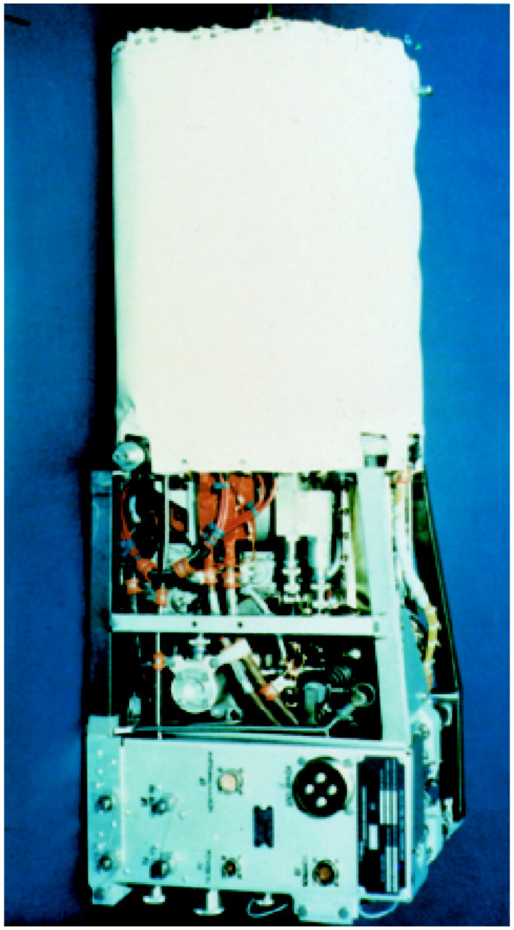
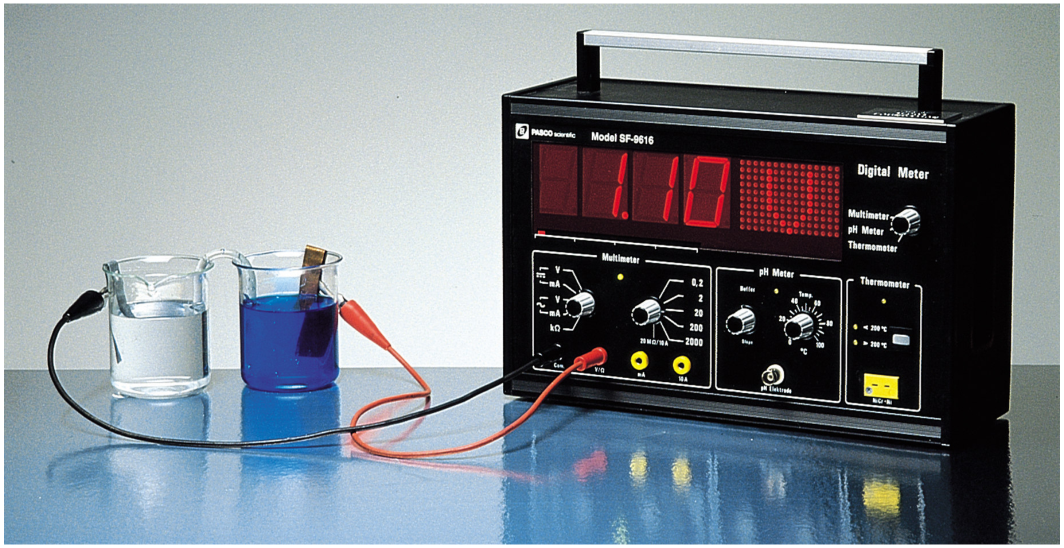




Electrochemistry

Chapter 19

Electrochemical processes are oxidation-reduction reactions in which:

- the energy released by a spontaneous reaction is converted to electricity or
- electrical energy is used to cause a nonspontaneous reaction to occur

$$2\overset{0}{\text{Mg}}\,(s) + \overset{0}{\text{O}_2}\,(g) \longrightarrow 2\overset{2+}{\text{Mg}}\overset{2-}{\text{O}}\,(s)$$

$2\text{Mg} \longrightarrow 2\text{Mg}^{2+} + 4e^-$ ***Oxidation*** half-reaction (lose e^-)

$\text{O}_2 + 4e^- \longrightarrow 2\text{O}^{2-}$ ***Reduction*** half-reaction (gain e^-)

Oxidation number

The charge the atom would have in a molecule (or an ionic compound) if electrons were completely transferred.

1. Free elements (uncombined state) have an oxidation number of zero.

$$\mathrm{Na, Be, K, Pb, H_2, O_2, P_4} = 0$$

2. In monatomic ions, the oxidation number is equal to the charge on the ion.

$$\mathrm{Li^+, Li} = +1;\ \mathrm{Fe^{3+}, Fe} = +3;\ \mathrm{O^{2-}, O} = -2$$

3. The oxidation number of oxygen *is* **usually** –2. In H_2O_2 and O_2^{2-} it is –1.

4. The oxidation number of hydrogen is +1 *except* when it is bonded to metals in binary compounds. In these cases, its oxidation number is –1.

5. Group IA metals are +1, IIA metals are +2 and fluorine is always –1.

6. The sum of the oxidation numbers of all the atoms in a molecule or ion is equal to the charge on the molecule or ion.

Identify the oxidation numbers of all the atoms in HCO_3^- ?

HCO_3^-

O = –2 H = +1

$3 \times (-2) + 1 + ? = -1$

C = +4

Balancing Redox Equations

The oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ in acid solution?

1. Write the unbalanced equation for the reaction ion ionic form.

$$Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

2. Separate the equation into two half-reactions.

Oxidation: $\overset{+2}{Fe^{2+}} \longrightarrow \overset{+3}{Fe^{3+}}$

Reduction: $\overset{+6}{Cr_2O_7^{2-}} \longrightarrow \overset{+3}{Cr^{3+}}$

3. Balance the atoms other than O and H in each half-reaction.

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$$

Balancing Redox Equations

4. For reactions in acid, add H_2O to balance O atoms and H^+ to balance H atoms.

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

$$14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

5. Add electrons to one side of each half-reaction to balance the charges on the half-reaction.

$$Fe^{2+} \longrightarrow Fe^{3+} + \boxed{1e^-}$$

$$\boxed{6e^-} + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

6. If necessary, equalize the number of electrons in the two half-reactions by multiplying the half-reactions by appropriate coefficients.

$$6Fe^{2+} \longrightarrow 6Fe^{3+} + 6e^-$$

$$6e^- + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

Balancing Redox Equations

7. Add the two half-reactions together and balance the final equation by inspection. **The number of electrons on both sides must cancel.**

Oxidation: $6Fe^{2+} \longrightarrow 6Fe^{3+} + \cancel{6e^-}$

Reduction: $\cancel{6e^-} + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$

$$\overline{14H^+ + Cr_2O_7^{2-} + 6Fe^{2+} \longrightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O}$$

8. Verify that the number of atoms and the charges are balanced.

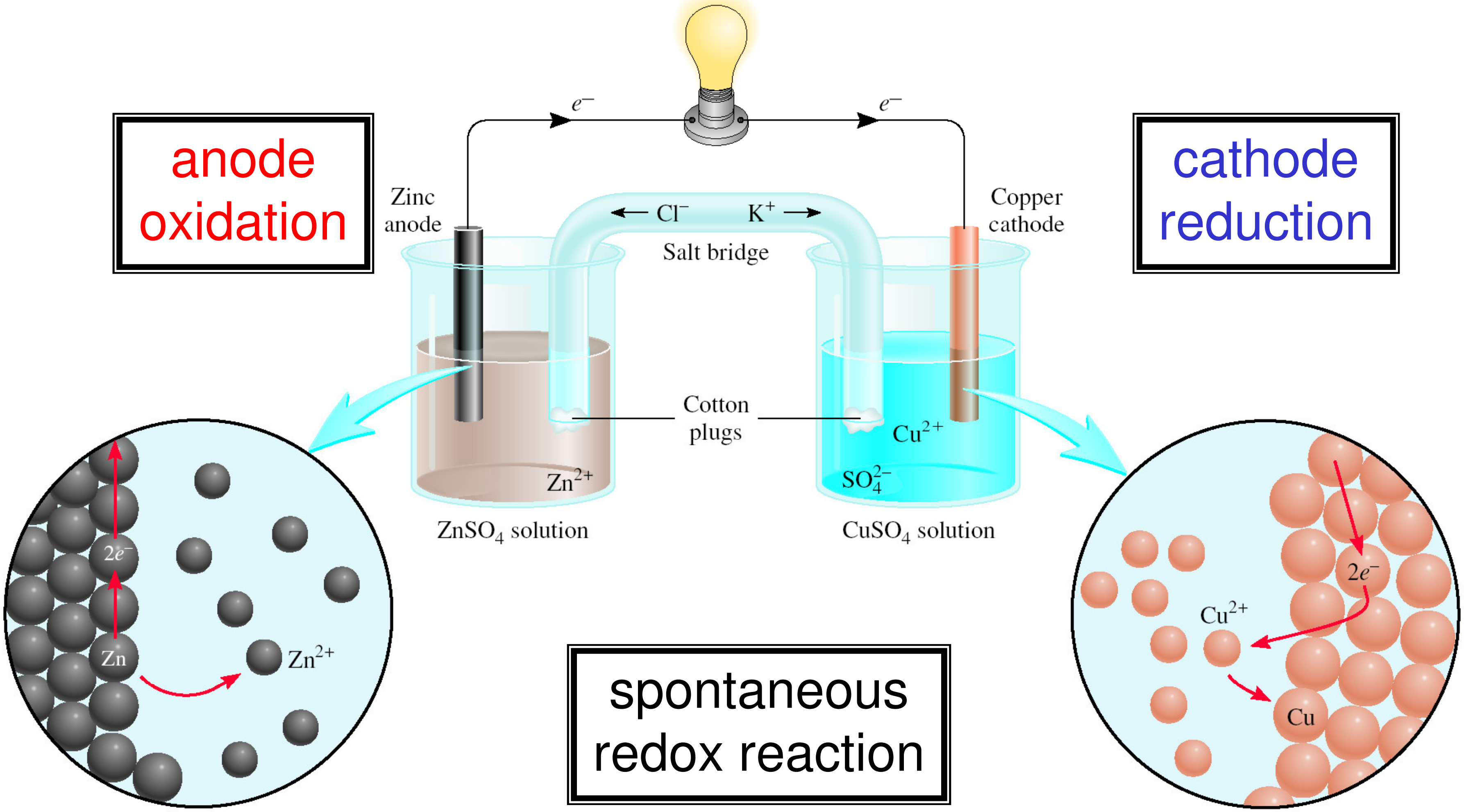
$$14\text{x}1 - 2 + 6 \text{ x } 2 = 24 = 6 \text{ x } 3 + 2 \text{ x } 3$$

9. For reactions in basic solutions, add OH^- to **both sides** of the equation for every H^+ that appears in the final equation.

Galvanic Cells

anode
oxidation

cathode
reduction



spontaneous
redox reaction

Zn is oxidized
to Zn^{2+} at anode.

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

Net reaction

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

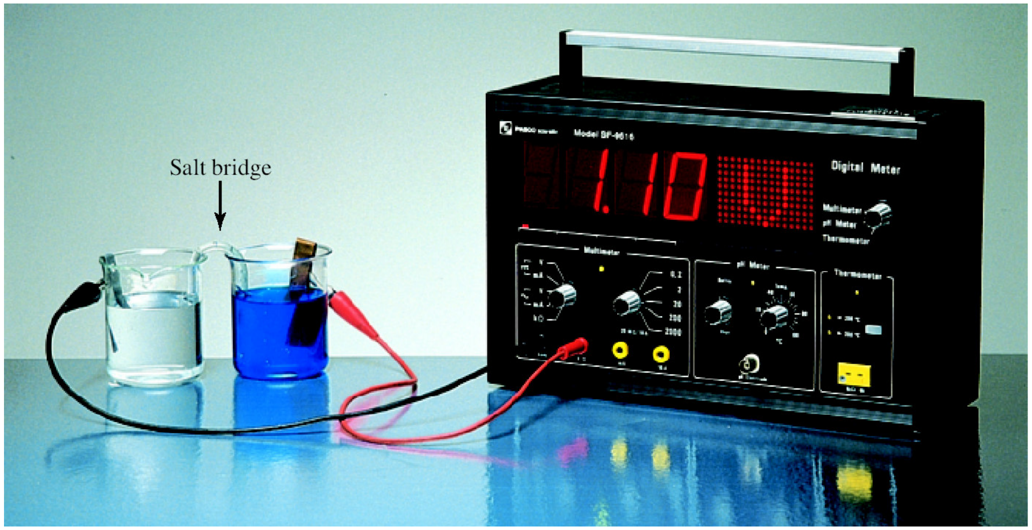
Cu^{2+} is reduced
to Cu at cathode.

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

Galvanic Cells

The difference in electrical potential between the anode and cathode is called:

- ***cell voltage***
- ***electromotive force (emf)***
- ***cell potential***



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

$$[\mathrm{Cu}^{2+}] = 1\ M \text{ and } [\mathrm{Zn}^{2+}] = 1\ M$$

Cell Diagram

phase boundary

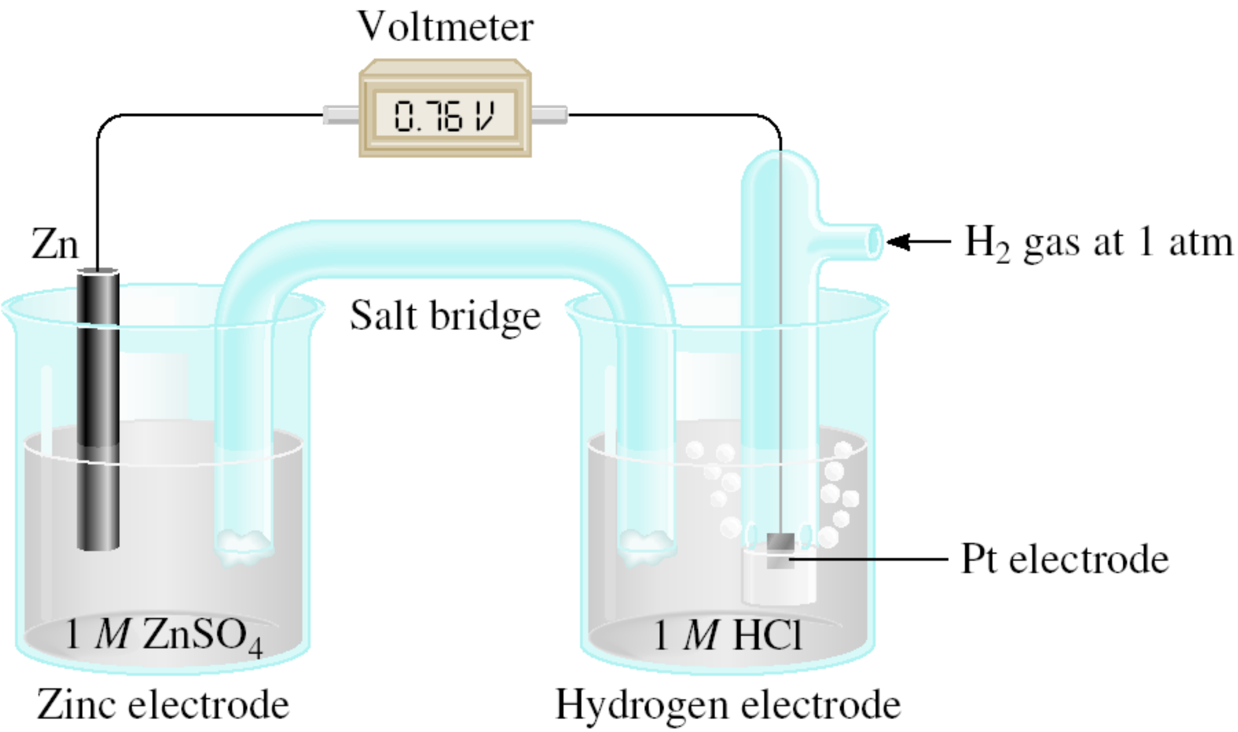
$$\mathrm{Zn}\,(s)\,|\,\mathrm{Zn}^{2+}\,(1\ M)\,||\,\mathrm{Cu}^{2+}\,(1\ M)\,|\,\mathrm{Cu}\,(s)$$

anode

salt bridge

cathode

Standard Reduction Potentials



$\mathrm{Zn}\ (s)\ |\ \mathrm{Zn}^{2+}\ (1\ M)\ ||\ \mathrm{H}^{+}\ (1\ M)\ |\ \mathrm{H}_2\ (1\ \mathrm{atm})\ |\ \mathrm{Pt}\ (s)$

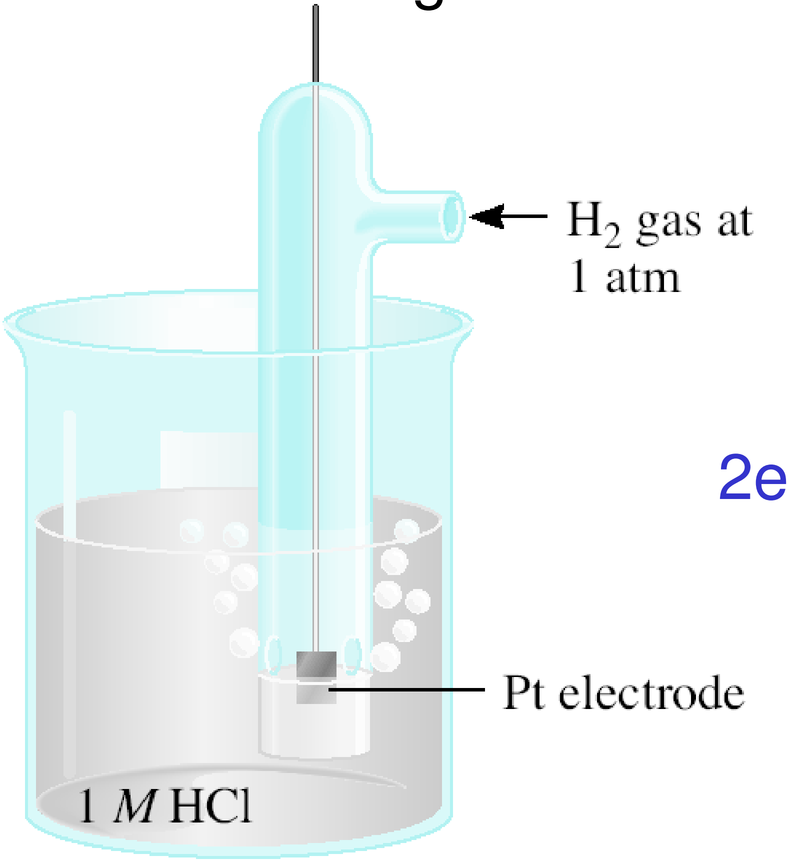
Anode (oxidation): $\mathrm{Zn}\ (s) \longrightarrow \mathrm{Zn}^{2+}\ (1\ M) + 2e^-$

Cathode (reduction): $2e^- + 2\mathrm{H}^+\ (1\ M) \longrightarrow \mathrm{H}_2\ (1\ atm)$

$\mathrm{Zn}\ (s) + 2\mathrm{H}^+\ (1\ M) \longrightarrow \mathrm{Zn}^{2+} + \mathrm{H}_2\ (1\ atm)$

Standard Reduction Potentials

Standard reduction potential (E^0) is the voltage associated with a **reduction reaction** at an electrode when all solutes are 1 *M* and all gases are at 1 atm.



Standard hydrogen electrode (SHE)

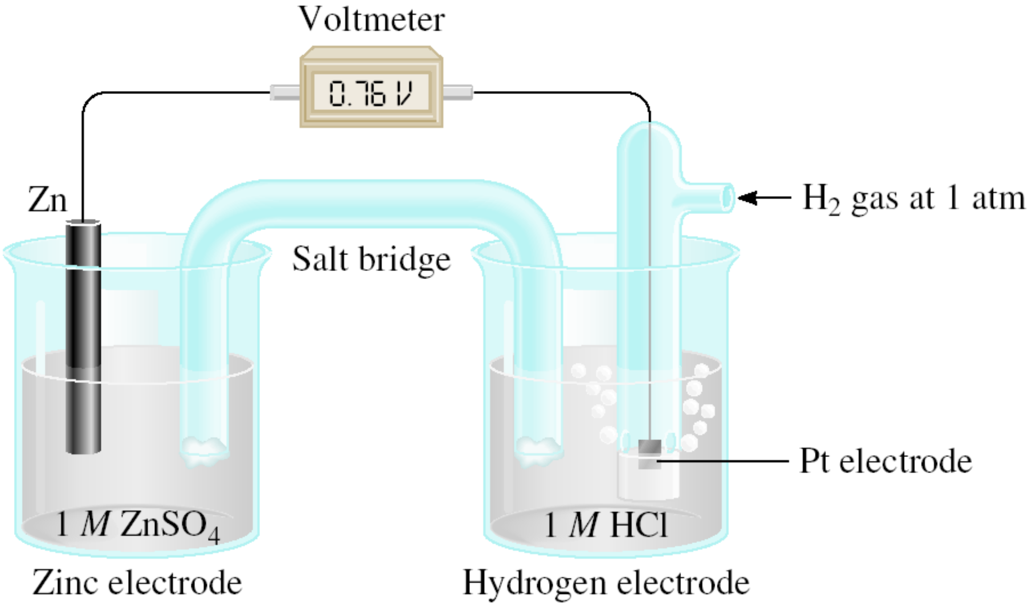
Reduction Reaction

$$2e^- + 2H^+ (1\ M) \longrightarrow H_2\ (1\ atm)$$

$$E^0 = 0\ V$$

Standard Reduction Potentials

$E^0_{cell} = 0.76\ \text{V}$



Standard emf (E^0_{cell})

$$\boxed{E^0_{cell} = E^0_{cathode} - E^0_{anode}}$$

$\text{Zn}\ (s)\ |\ \text{Zn}^{2+}\ (1\ M)\ ||\ \text{H}^+\ (1\ M)\ |\ \text{H}_2\ (1\ \text{atm})\ |\ \text{Pt}\ (s)$

$$E^0_{cell} = E^0_{H^+/H_2} - E^0_{Zn^{2+}/Zn}$$

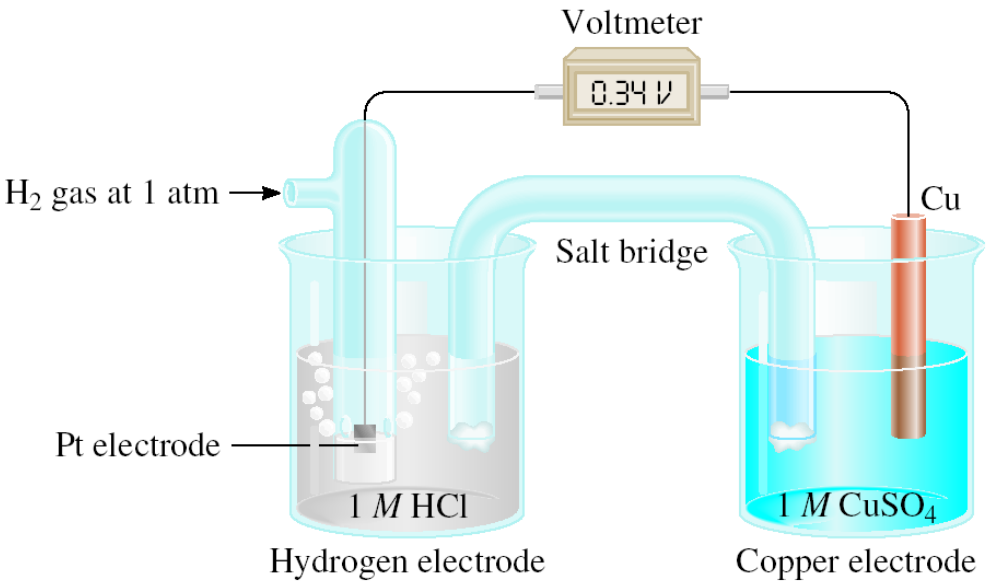
$$0.76\ \text{V} = 0 - E^0_{Zn^{2+}/Zn}$$

$$E^0_{Zn^{2+}/Zn} = -0.76\ \text{V}$$

$\text{Zn}^{2+}\ (1\ M) + 2e^- \longrightarrow \text{Zn} \qquad E^0 = -0.76\ \text{V}$

Standard Reduction Potentials

$E^0_{cell} = 0.34\ V$



$E^0_{cell} = E^0_{cathode} - E^0_{anode}$

$E^0_{cell} = E^0_{Cu^{2+}/Cu} - E^0_{H^+/H_2}$

$0.34 = E^0_{Cu^{2+}/Cu} - 0$

$E^0_{Cu^{2+}/Cu} = 0.34\ V$

Pt (s) | H_2 (1 atm) | H^+ (1 M) || Cu^{2+} (1 M) | Cu (s)

Anode (oxidation): $H_2\ (1\ atm) \longrightarrow 2H^+\ (1\ M) + 2e^-$

Cathode (reduction): $2e^- + Cu^{2+}\ (1\ M) \longrightarrow Cu\ (s)$

$H_2\ (1\ atm) + Cu^{2+}\ (1\ M) \longrightarrow Cu\ (s) + 2H^+\ (1\ M)$

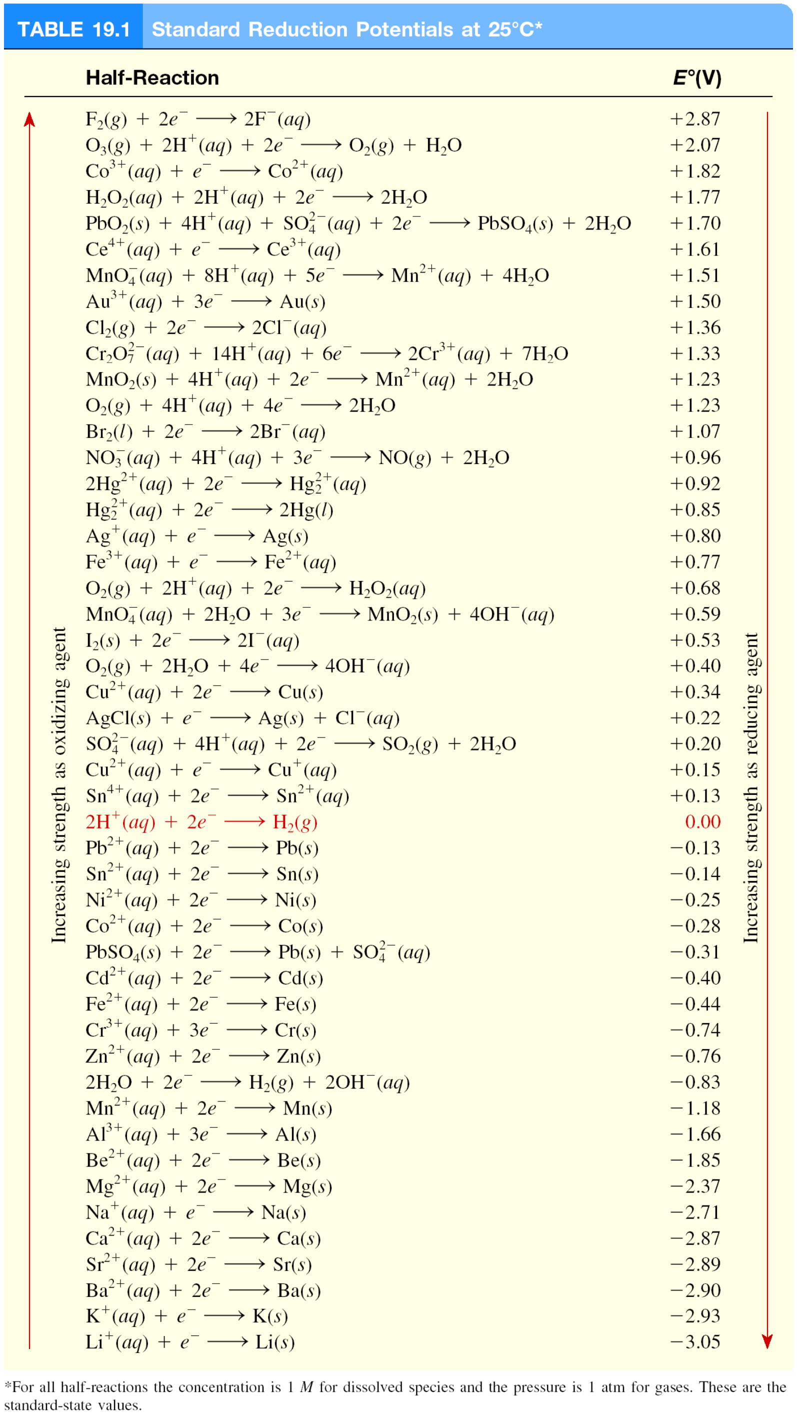


TABLE 19.1 Standard Reduction Potentials at 25°C*

Half-Reaction	$E°$(V)
$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$	+2.87
$O_3(g) + 2H^+(aq) + 2e^- \longrightarrow O_2(g) + H_2O$	+2.07
$Co^{3+}(aq) + e^- \longrightarrow Co^{2+}(aq)$	+1.82
$H_2O_2(aq) + 2H^+(aq) + 2e^- \longrightarrow 2H_2O$	+1.77
$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O$	+1.70
$Ce^{4+}(aq) + e^- \longrightarrow Ce^{3+}(aq)$	+1.61
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O$	+1.51
$Au^{3+}(aq) + 3e^- \longrightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O$	+1.33
$MnO_2(s) + 4H^+(aq) + 2e^- \longrightarrow Mn^{2+}(aq) + 2H_2O$	+1.23
$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O$	+1.23
$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq)$	+1.07
$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O$	+0.96
$2Hg^{2+}(aq) + 2e^- \longrightarrow Hg_2^{2+}(aq)$	+0.92
$Hg_2^{2+}(aq) + 2e^- \longrightarrow 2Hg(l)$	+0.85
$Ag^+(aq) + e^- \longrightarrow Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O_2(aq)$	+0.68
$MnO_4^-(aq) + 2H_2O + 3e^- \longrightarrow MnO_2(s) + 4OH^-(aq)$	+0.59
$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$	+0.53
$O_2(g) + 2H_2O + 4e^- \longrightarrow 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$	+0.34
$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$	+0.22
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow SO_2(g) + 2H_2O$	+0.20
$Cu^{2+}(aq) + e^- \longrightarrow Cu^+(aq)$	+0.15
$Sn^{4+}(aq) + 2e^- \longrightarrow Sn^{2+}(aq)$	+0.13
$2H^+(aq) + 2e^- \longrightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \longrightarrow Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \longrightarrow Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$	−0.28
$PbSO_4(s) + 2e^- \longrightarrow Pb(s) + SO_4^{2-}(aq)$	−0.31
$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$	−0.40
$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$	−0.44
$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$	−0.74
$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$	−0.76
$2H_2O + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$	−0.83
$Mn^{2+}(aq) + 2e^- \longrightarrow Mn(s)$	−1.18
$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$	−1.66
$Be^{2+}(aq) + 2e^- \longrightarrow Be(s)$	−1.85
$Mg^{2+}(aq) + 2e^- \longrightarrow Mg(s)$	−2.37
$Na^+(aq) + e^- \longrightarrow Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \longrightarrow Ca(s)$	−2.87
$Sr^{2+}(aq) + 2e^- \longrightarrow Sr(s)$	−2.89
$Ba^{2+}(aq) + 2e^- \longrightarrow Ba(s)$	−2.90
$K^+(aq) + e^- \longrightarrow K(s)$	−2.93
$Li^+(aq) + e^- \longrightarrow Li(s)$	−3.05

*For all half-reactions the concentration is 1 *M* for dissolved species and the pressure is 1 atm for gases. These are the standard-state values.

- E^0 is for the reaction as written
- The more positive E^0 the greater the tendency for the substance to be reduced
- The half-cell reactions are reversible
- The sign of E^0 changes when the reaction is reversed
- Changing the stoichiometric coefficients of a half-cell reaction ***does not*** change the value of E^0

What is the standard emf of an electrochemical cell made of a Cd electrode in a 1.0 M $Cd(NO_3)_2$ solution and a Cr electrode in a 1.0 M $Cr(NO_3)_3$ solution?

$Cd^{2+}(aq) + 2e^- \longrightarrow Cd(s)$ $E^0 = -0.40\ V$ Cd is the stronger oxidizer

$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$ $E^0 = -0.74\ V$ Cd will oxidize Cr

Anode (oxidation): $Cr(s) \longrightarrow Cr^{3+}(1\ M) + \boxed{3e^-}$ x 2

Cathode (reduction): $\boxed{2e^-} + Cd^{2+}(1\ M) \longrightarrow Cd(s)$ x 3

$$2Cr(s) + 3Cd^{2+}(1\ M) \longrightarrow 3Cd(s) + 2Cr^{3+}(1\ M)$$

$$E^0_{cell} = E^0_{cathode} - E^0_{anode}$$

$$E^0_{cell} = -0.40 - (-0.74)$$

$$E^0_{cell} = 0.34\ V$$

Spontaneity of Redox Reactions

$\Delta G = -nFE_{cell}$ n = number of moles of electrons in reaction

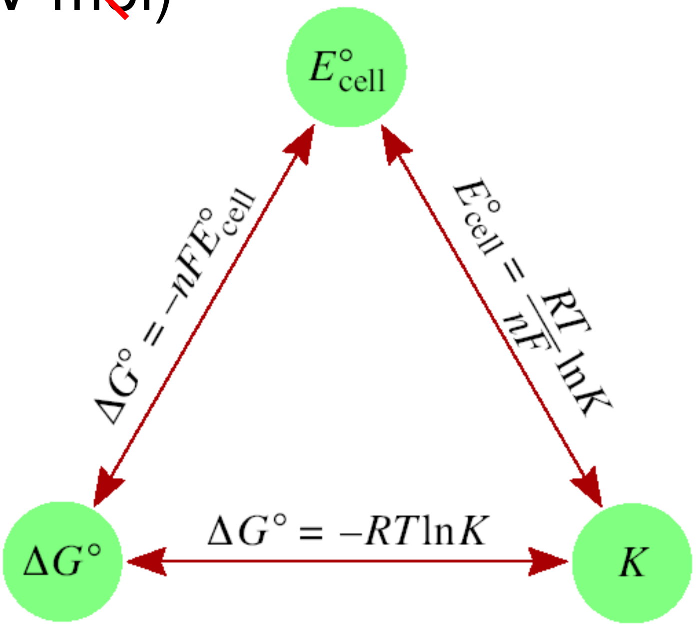
$\Delta G^0 = -nFE^0_{cell}$ $F = 96{,}500 \dfrac{\text{J}}{\text{V} \cdot \text{mol}} = 96{,}500$ C/mol

$\Delta G^0 = -RT \ln K = -nFE^0_{cell}$

$$E^0_{cell} = \frac{RT}{nF} \ln K = \frac{(8.314\ \cancel{\text{J}}/\cancel{\text{K}}\bullet\cancel{\text{mol}})(298\ \cancel{\text{K}})}{n\,(96{,}500\ \cancel{\text{J}}/\text{V}\bullet\cancel{\text{mol}})} \ln K$$

$$E^0_{cell} = \frac{0.0257\ \text{V}}{n} \boxed{\ln} K$$

$$E^0_{cell} = \frac{0.0592\ \text{V}}{n} \boxed{\log} K$$



Spontaneity of Redox Reactions

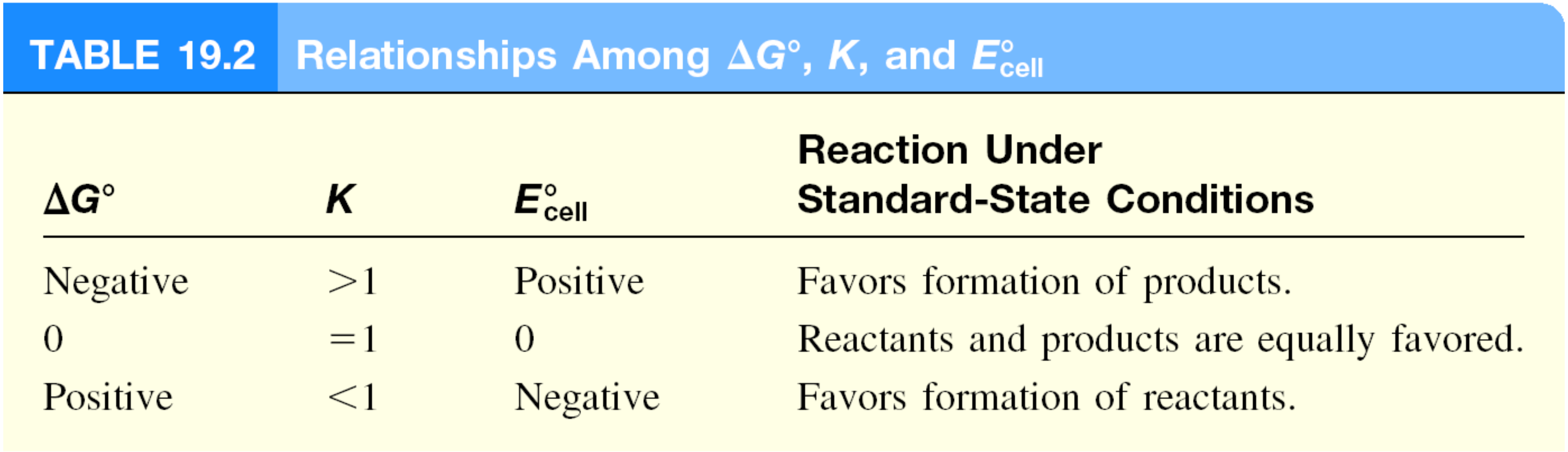


TABLE 19.2 **Relationships Among $\Delta G°$, K, and $E°_{cell}$**

$\Delta G°$	K	$E°_{cell}$	Reaction Under Standard-State Conditions
Negative	>1	Positive	Favors formation of products.
0	$=1$	0	Reactants and products are equally favored.
Positive	<1	Negative	Favors formation of reactants.

$$\Delta G^0 = -RT \ln K = -nFE^0_{cell}$$

What is the equilibrium constant for the following reaction at 25^0C? $Fe^{2+}(aq) + 2Ag(s) \rightleftharpoons Fe(s) + 2Ag^{+}(aq)$

$$E^0_{\text{cell}} = \frac{0.0257\ \text{V}}{n} \ln K$$

Oxidation: $2Ag \longrightarrow 2Ag^{+} + \boxed{2e^{-}}$

Reduction: $\boxed{2e^{-}} + Fe^{2+} \longrightarrow Fe$

$n = 2$

$$E^0 = E^0_{Fe^{2+}/Fe} - E^0_{Ag^{+}/Ag}$$

$$E^0 = -0.44 - (0.80)$$

$$E^0 = -1.24\ \text{V}$$

$$K = e^{\left[\frac{E^0_{cell} \times n}{0.0257\ \text{V}}\right]} = e^{\left[\frac{-1.24\ \text{V} \times 2}{0.0257\ \text{V}}\right]}$$

$$K = 1.23 \times 10^{-42}$$

The Effect of Concentration on Cell Emf

$\Delta G = \Delta G^0 + RT \ln Q \qquad \Delta G = -nFE \qquad \Delta G^0 = -nFE^0$

$$-nFE = -nFE^0 + RT \ln Q$$

Nernst equation

$$E = E^0 - \frac{RT}{nF} \ln Q$$

At 298 K

$$E = E^0 - \frac{0.0257\ \mathrm{V}}{n} \ln Q \qquad E = E^0 - \frac{0.0592\ \mathrm{V}}{n} \log Q$$

Will the following reaction occur spontaneously at 25^0C if $[Fe^{2+}] = 0.60\ M$ and $[Cd^{2+}] = 0.010\ M$?

$$Fe^{2+}\ (aq) + Cd\ (s) \rightleftarrows Fe\ (s) + Cd^{2+}\ (aq)$$

Oxidation: $Cd \longrightarrow Cd^{2+} + \enclose{circle}{2e^-}$

Reduction: $\enclose{circle}{2e^-} + Fe^{2+} \longrightarrow 2Fe$

$n = 2$

$$E^0 = E^0_{Fe^{2+}/Fe} - E^0_{Cd^{2+}/Cd}$$

$$E^0 = -0.44 - (-0.40)$$

$$E^0 = -0.04\ \text{V}$$

$$E = E^0 - \frac{0.0257\ \text{V}}{n} \ln Q$$

$$E = -0.04\ \text{V} - \frac{0.0257\ \text{V}}{2} \ln \frac{0.010}{0.60}$$

$$E = 0.013$$

$E > 0$ Spontaneous

Concentration Cells

Galvanic cell from two half-cells composed of the *same* material but differing in ion concentrations.

$$\mathrm{Zn}(s)\,|\,\mathrm{Zn}^{2+}(0.10\,M)\,||\,\mathrm{Zn}^{2+}(1.0\,M)\,|\,\mathrm{Zn}(s)$$

$$\begin{array}{lrcl}
\text{Oxidation:} & \mathrm{Zn}(s) & \longrightarrow & \mathrm{Zn}^{2+}(0.10\,M) + 2e^- \\
\text{Reduction:} & \mathrm{Zn}^{2+}(1.0\,M) + 2e^- & \longrightarrow & \mathrm{Zn}(s) \\
\hline
\text{Overall:} & Zn^{2+}(1.0\,M) & \longrightarrow & \mathrm{Zn}^{2+}(0.10\,M)
\end{array}$$

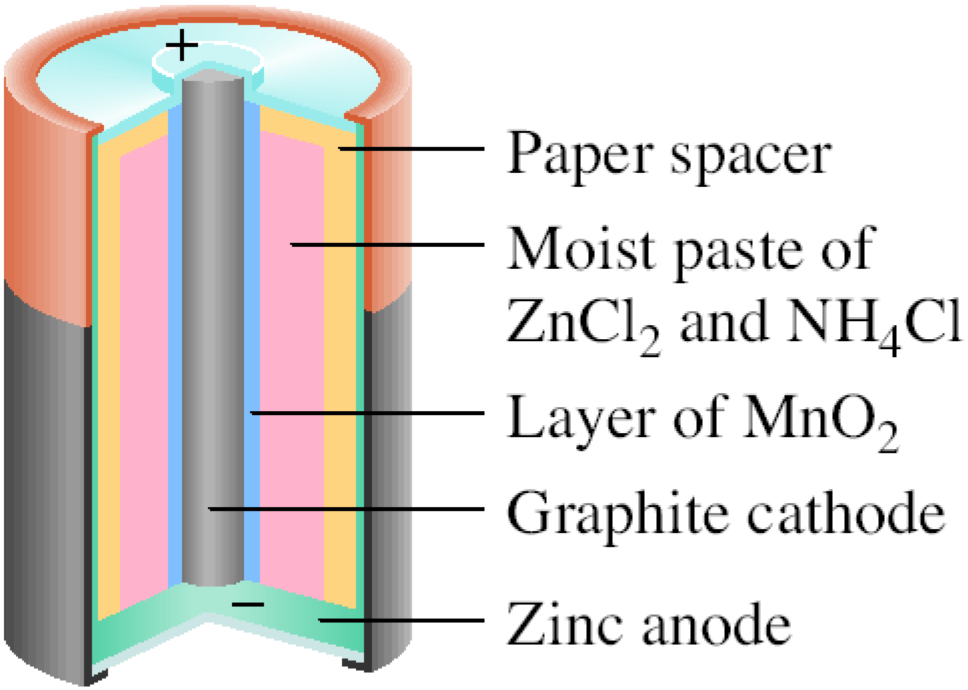
$$E = E^\circ - \frac{0.0257\ \mathrm{V}}{2}\ln\frac{[\mathrm{Zn}^{2+}]_{\mathrm{dil}}}{[\mathrm{Zn}^{2+}]_{\mathrm{conc}}}$$

$$E = 0 - \frac{0.0257\ \mathrm{V}}{2}\ln\frac{0.10}{1.0} = 0.0296\ \mathrm{V}$$

Batteries

Dry cell

Leclanché cell



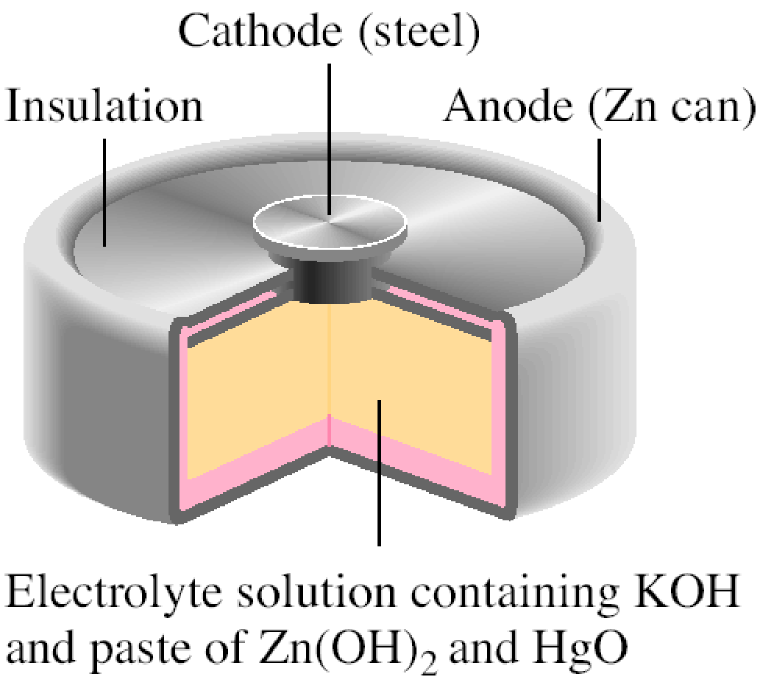
Anode: $Zn\ (s) \longrightarrow Zn^{2+}\ (aq) + 2e^-$

Cathode: $2NH_4^+\ (aq) + 2MnO_2\ (s) + 2e^- \longrightarrow Mn_2O_3\ (s) + 2NH_3\ (aq) + H_2O\ (l)$

$$Zn\ (s) + 2NH_4\ (aq) + 2MnO_2\ (s) \longrightarrow Zn^{2+}\ (aq) + 2NH_3\ (aq) + H_2O\ (l) + Mn_2O_3\ (s)$$

Batteries

Mercury Battery



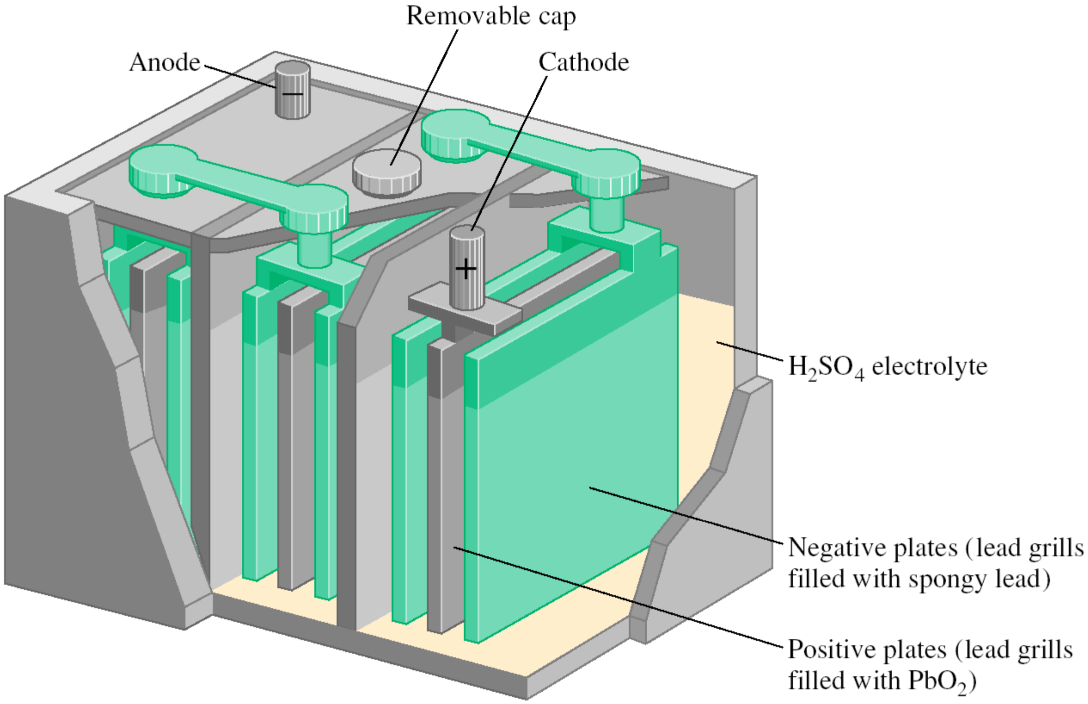
Anode: $Zn(Hg) + 2OH^- (aq) \longrightarrow ZnO (s) + H_2O (l) + 2e^-$

Cathode: $HgO (s) + H_2O (l) + 2e^- \longrightarrow Hg (l) + 2OH^- (aq)$

$Zn(Hg) + HgO (s) \longrightarrow ZnO (s) + Hg (l)$

Batteries

Lead storage battery

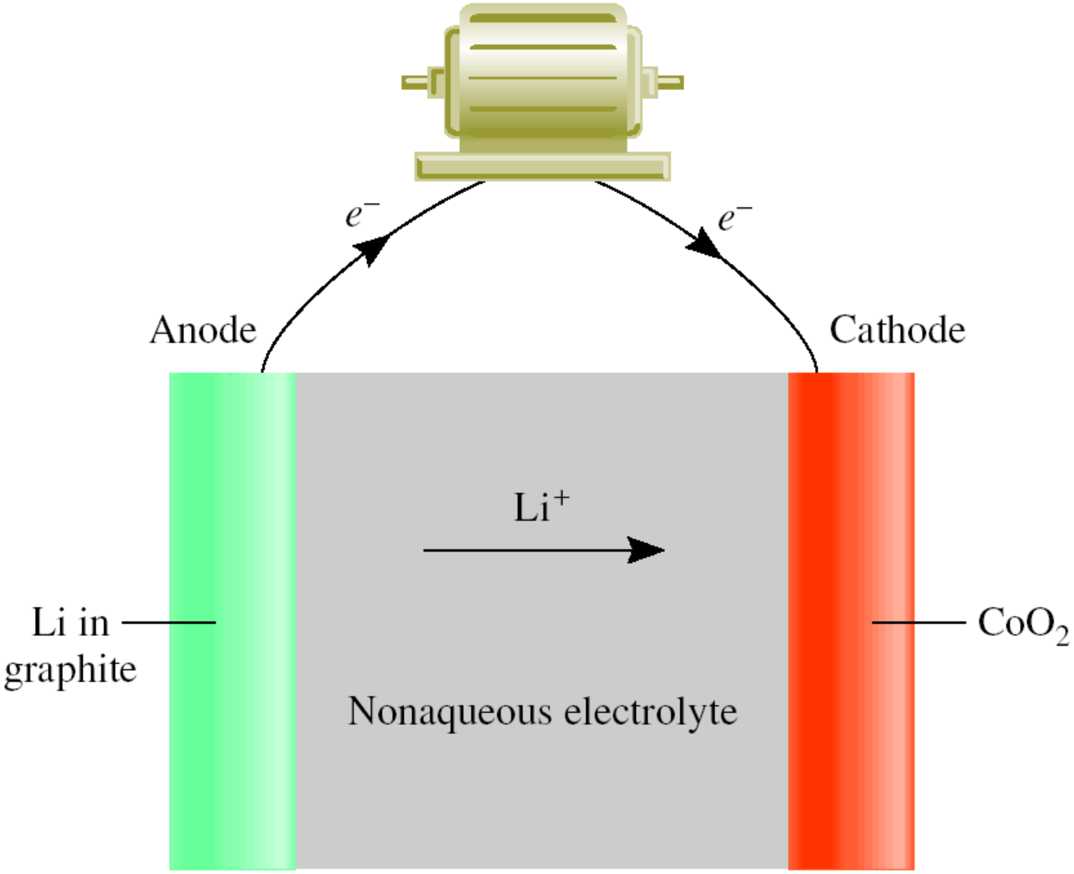


Anode: $Pb\ (s) + SO_4^{2-}\ (aq) \longrightarrow PbSO_4\ (s) + 2e^-$

Cathode: $PbO_2\ (s) + 4H^+\ (aq) + SO_4^{2-}\ (aq) + 2e^- \longrightarrow PbSO_4\ (s) + 2H_2O\ (l)$

$Pb\ (s) + PbO_2\ (s) + 4H^+\ (aq) + 2SO_4^{2-}\ (aq) \longrightarrow 2PbSO_4\ (s) + 2H_2O\ (l)$

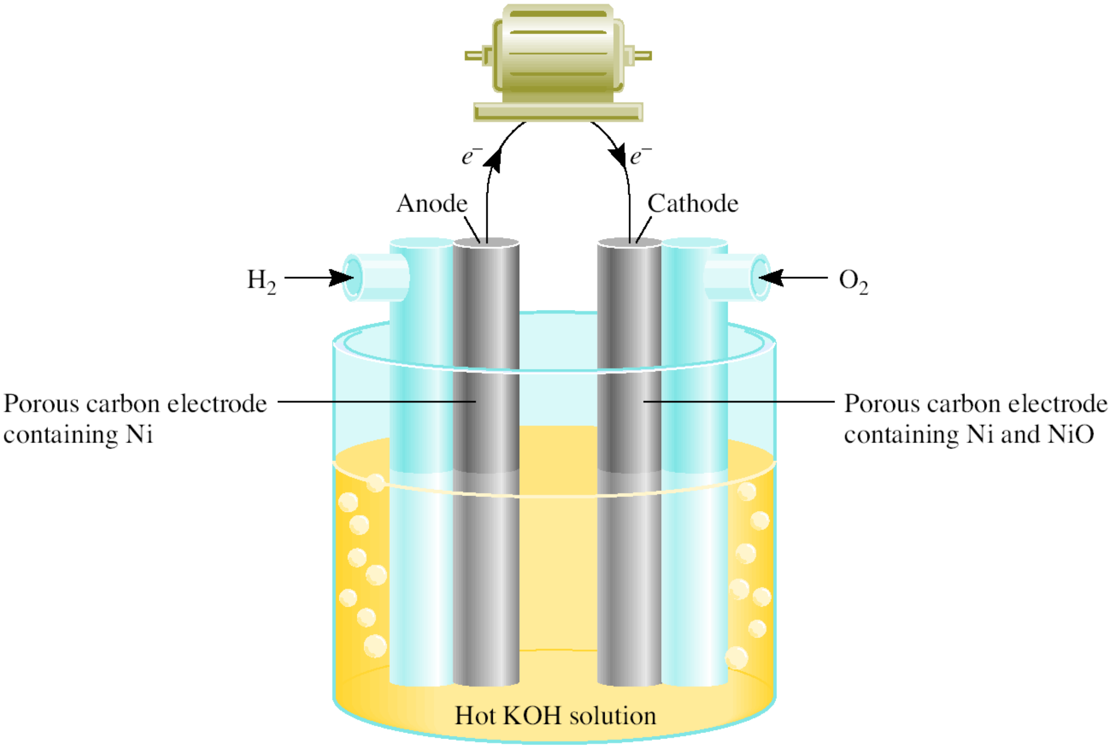
Batteries



$$Li \longrightarrow Li^+ + e^- \qquad\qquad Li^+ + CoO_2 + e^- \longrightarrow LiCoO_2$$

Solid State Lithium Battery

Batteries



A ***fuel cell*** is an electrochemical cell that requires a continuous supply of reactants to keep functioning

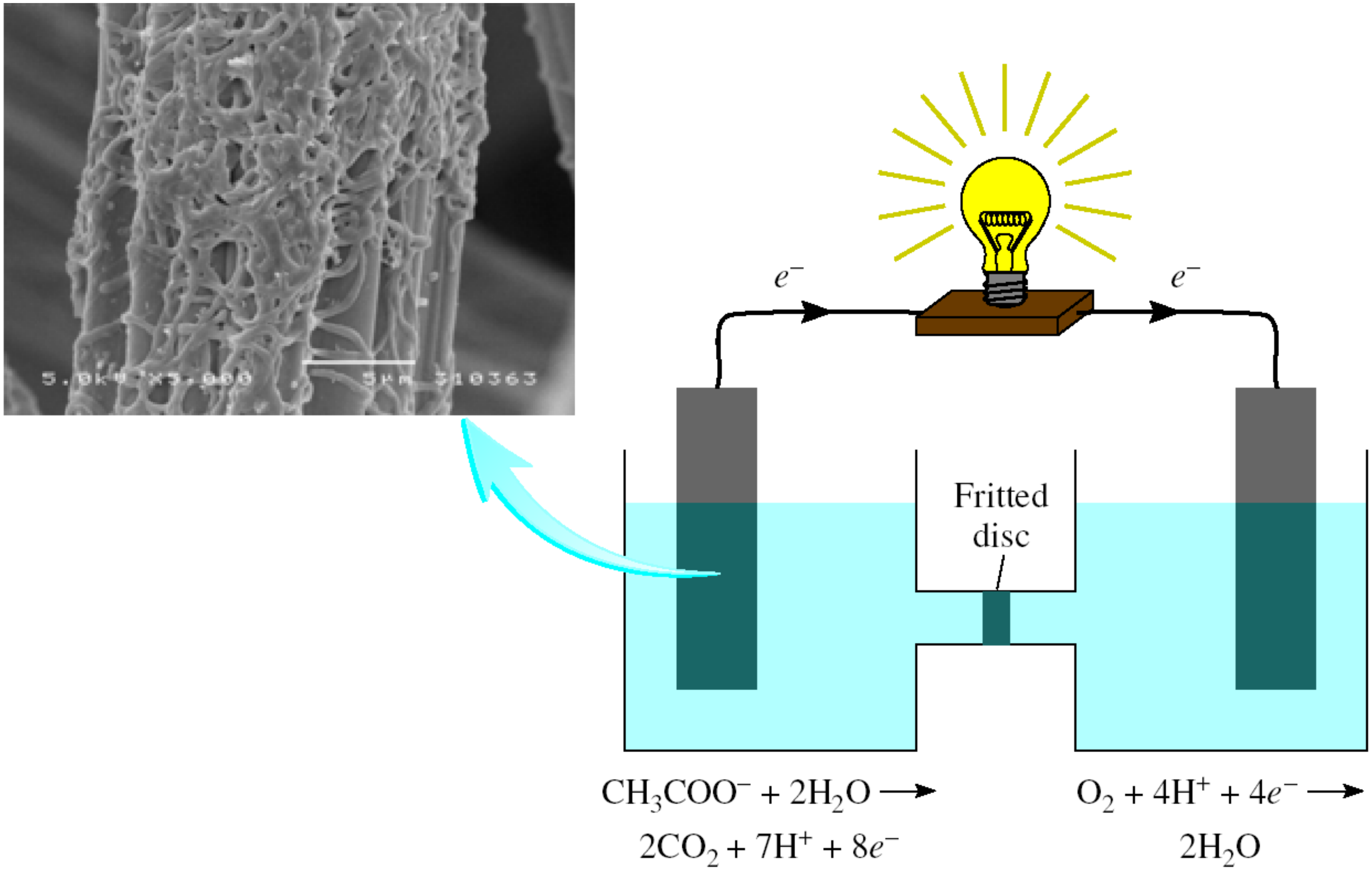
Anode: $2H_2\ (g) + 4OH^-\ (aq) \longrightarrow 4H_2O\ (l) + 4e^-$

Cathode: $O_2\ (g) + 2H_2O\ (l) + 4e^- \longrightarrow 4OH^-\ (aq)$

$2H_2\ (g) + O_2\ (g) \longrightarrow 2H_2O\ (l)$

Chemistry In Action: Bacteria Power

$$CH_3COO^- + 2O_2 + H^+ \longrightarrow 2CO_2 + 2H_2O$$

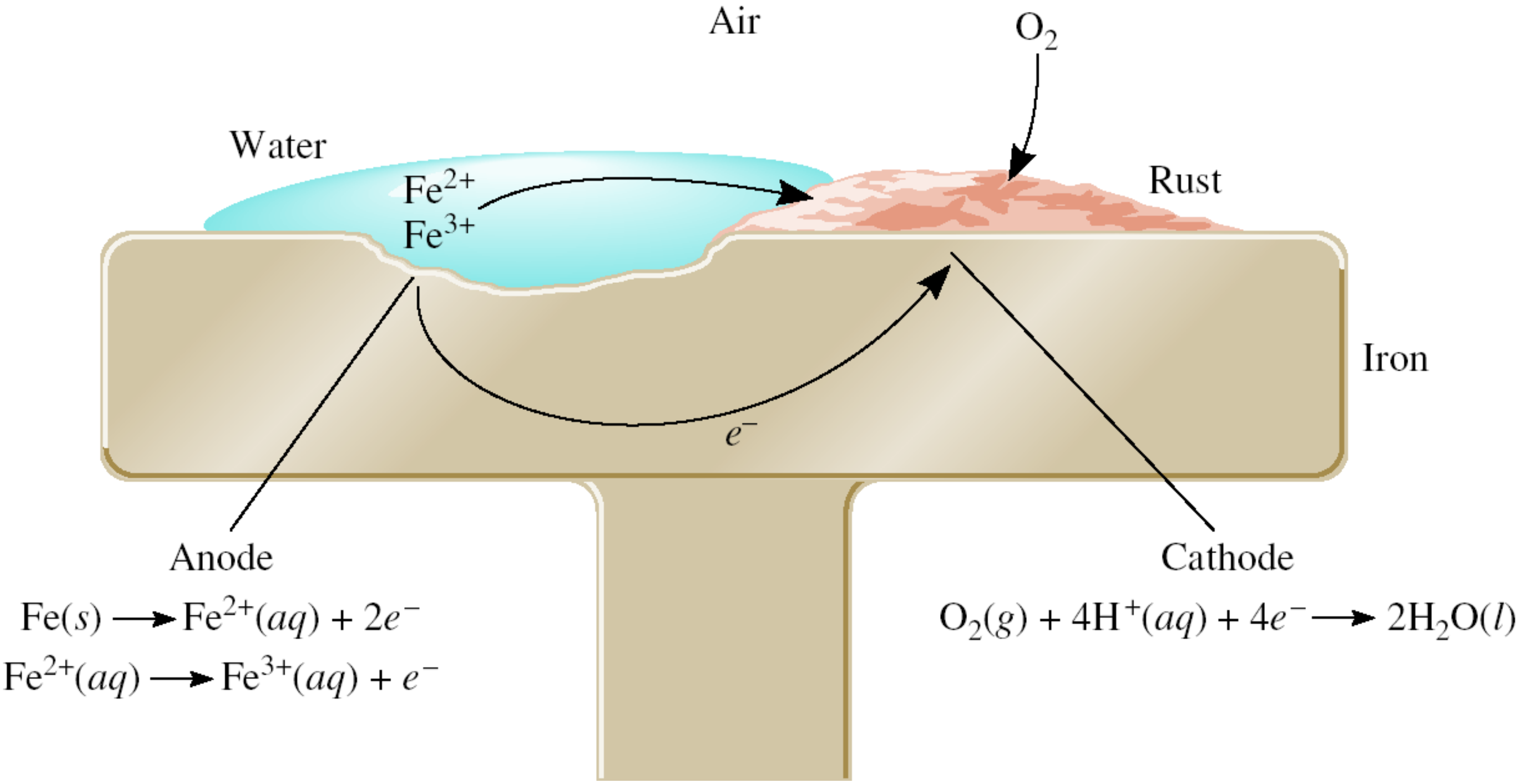


$$CH_3COO^- + 2H_2O \longrightarrow 2CO_2 + 7H^+ + 8e^-$$

$$O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$$

Corrosion

Corrosion is the term usually applied to *the deterioration of metals by an electrochemical process*.

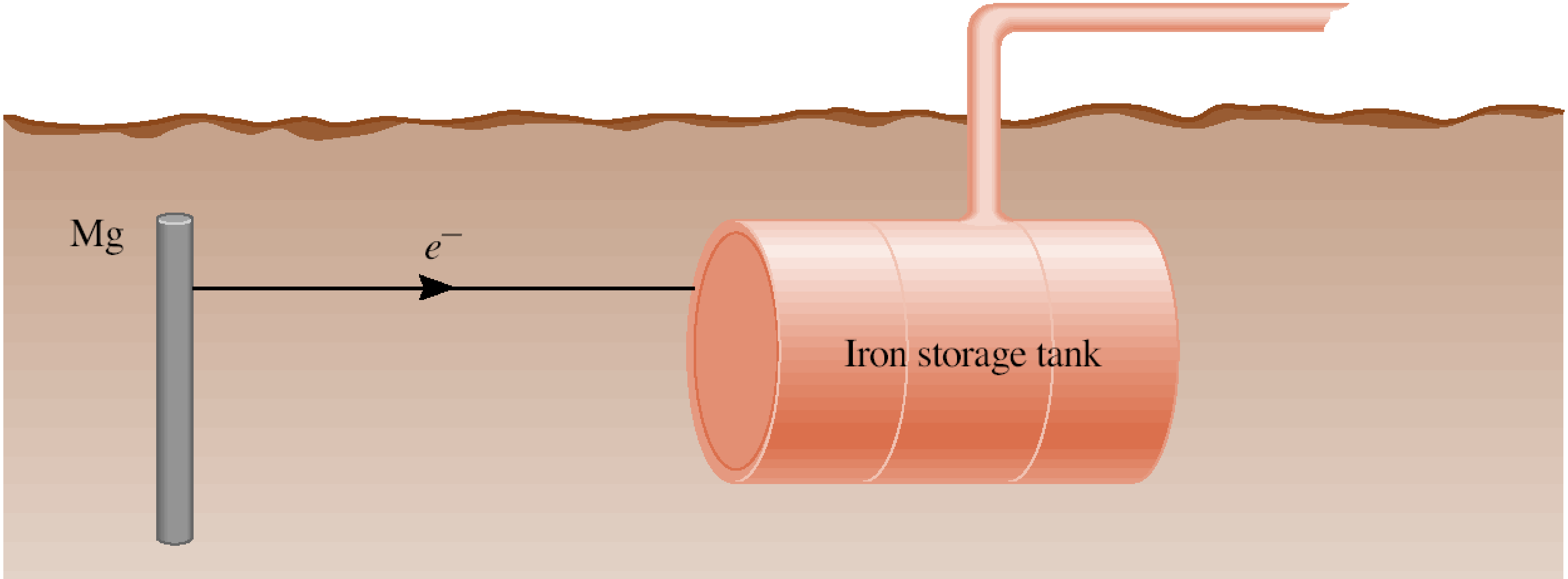


$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$

$Fe^{2+}(aq) \longrightarrow Fe^{3+}(aq) + e^-$

$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$

Cathodic Protection of an Iron Storage Tank

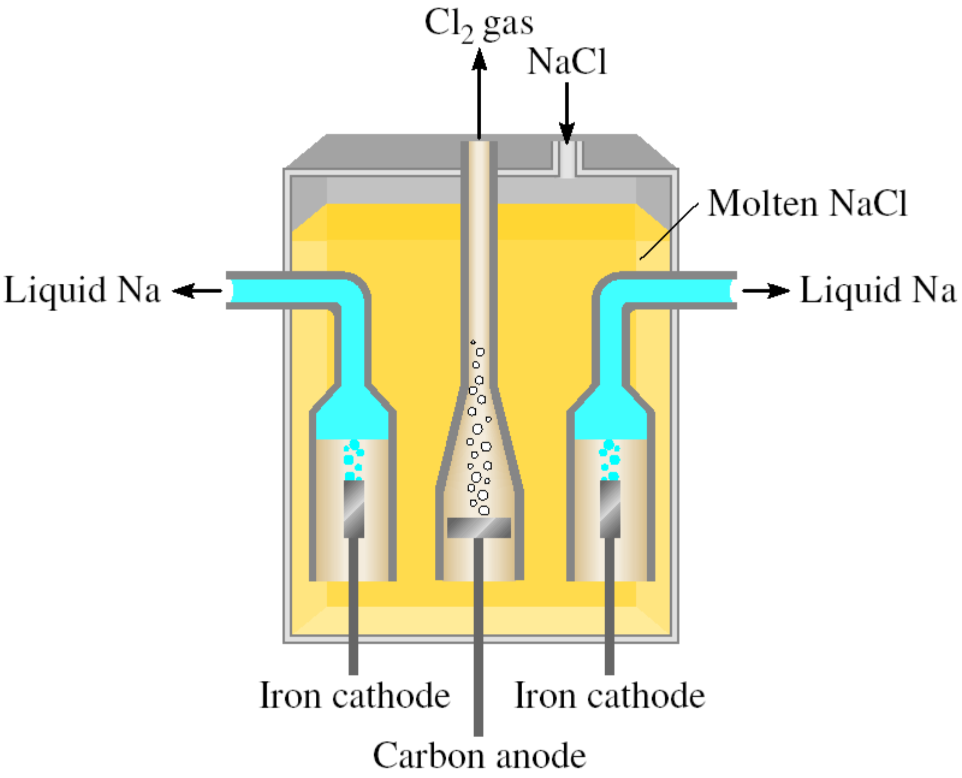


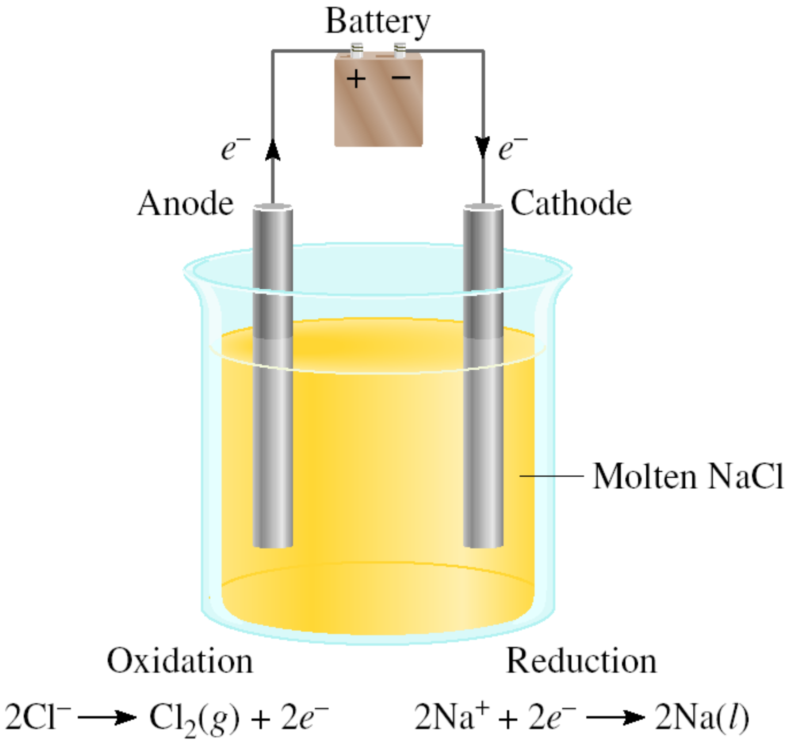
Oxidation: $Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$

Reduction: $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$

Electrolysis is the process in which electrical energy is used to cause a **nonspontaneous** chemical reaction to occur.

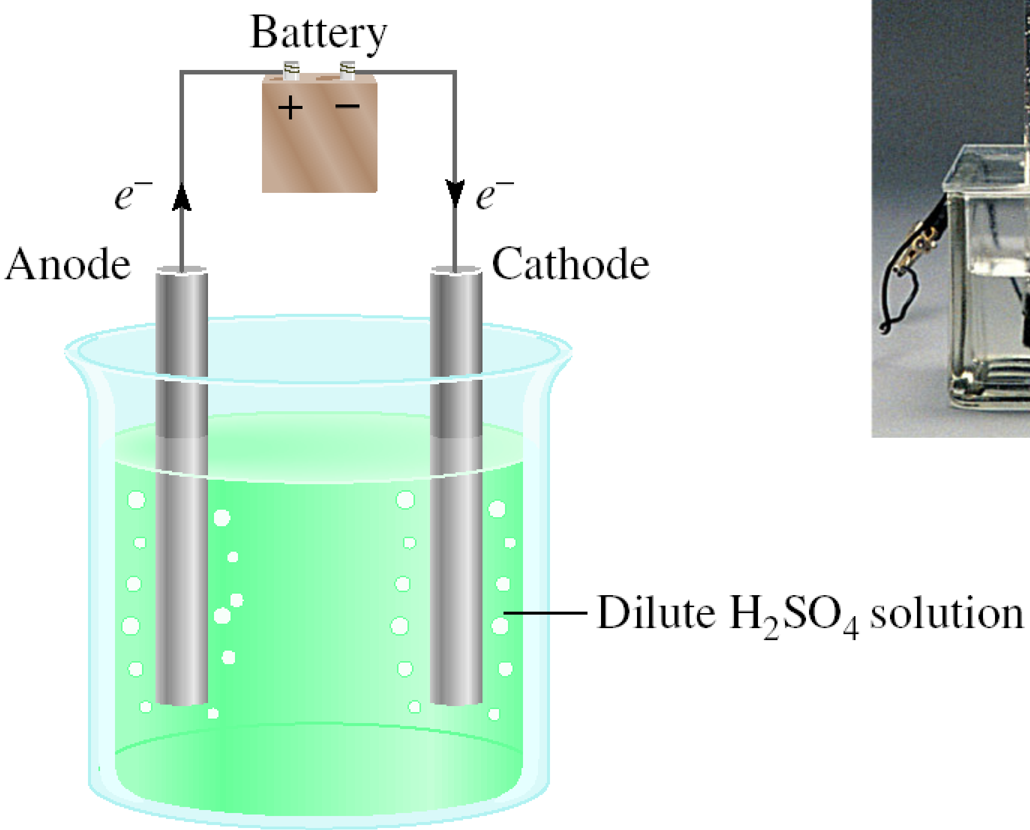
Electrolysis of molten NaCl

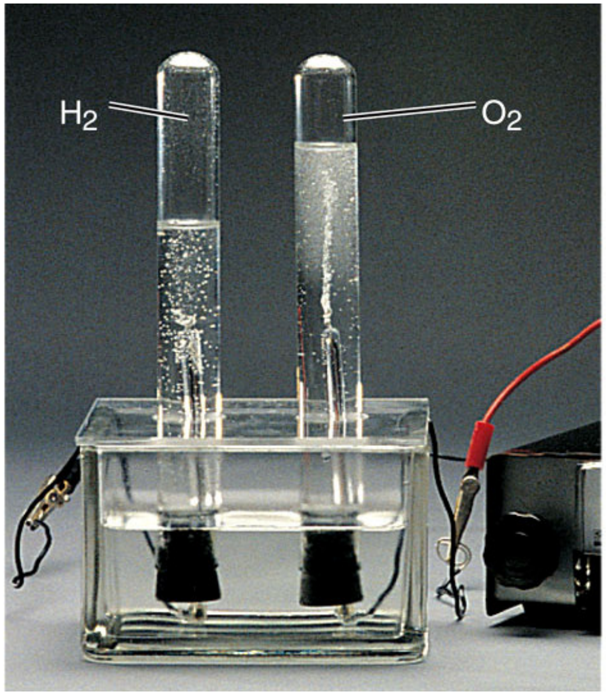




$2Cl^- \longrightarrow Cl_2(g) + 2e^-$ $2Na^+ + 2e^- \longrightarrow 2Na(l)$

Electrolysis of Water





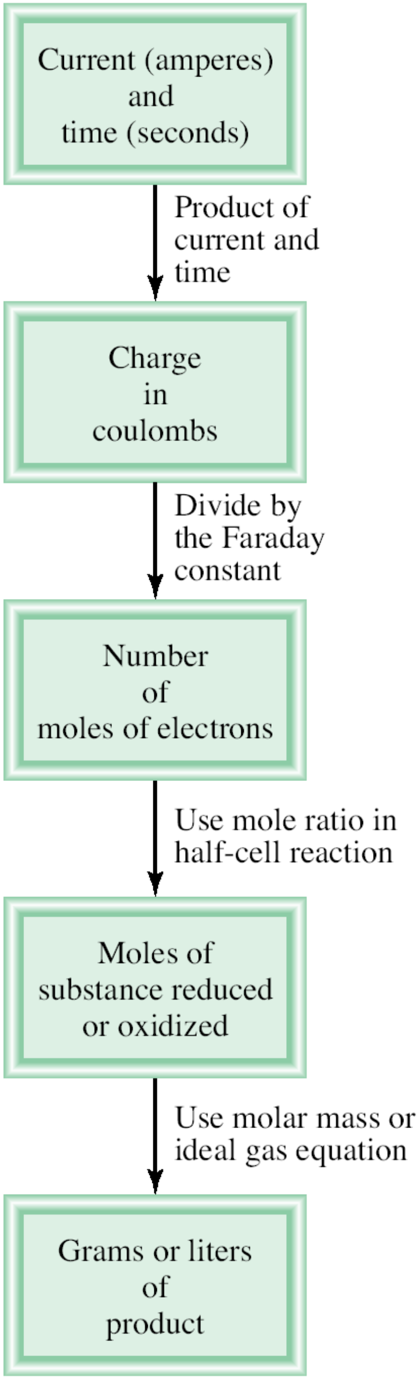
Oxidation

$2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-$

Reduction

$4H^+(aq) + 4e^- \longrightarrow 2H_2(g)$

Electrolysis and Mass Changes



Current (amperes) and time (seconds)

Product of current and time

Charge in coulombs

Divide by the Faraday constant

Number of moles of electrons

Use mole ratio in half-cell reaction

Moles of substance reduced or oxidized

Use molar mass or ideal gas equation

Grams or liters of product

charge (C) = current (A) x time (s)

$$\boxed{1 \text{ mol } e^{-} = 96{,}500 \text{ C}}$$

How much Ca will be produced in an electrolytic cell of molten $CaCl_2$ if a current of 0.452 A is passed through the cell for 1.5 hours?

Anode: $2Cl^- (l) \longrightarrow Cl_2 (g) + 2e^-$

Cathode: $Ca^{2+} (l) + \boxed{2e^-} \longrightarrow Ca (s)$

$Ca^{2+} (l) + 2Cl^- (l) \longrightarrow Ca (s) + Cl_2 (g)$

2 mole e^- = 1 mole Ca

$$\text{mol Ca} = 0.452 \frac{\cancel{\text{C}}}{\cancel{\text{s}}} \times 1.5\ \cancel{\text{hr}} \times 3600 \frac{\cancel{\text{s}}}{\cancel{\text{hr}}} \times \frac{1\ \cancel{\text{mol e}^-}}{96{,}500\ \cancel{\text{C}}} \times \frac{1\ \text{mol Ca}}{2\ \cancel{\text{mol e}^-}}$$

$$= 0.0126\ \text{mol Ca}$$

$$= 0.50\ \text{g Ca}$$

Chemistry In Action: Dental Filling Discomfort

Hg_2^{2+}/Ag_2Hg_3 0.85 V

Sn^{2+}/Ag_3Sn -0.05 V

Sn^{2+}/Ag_3Sn -0.05 V