Chapter 18: Electrochemistry

Oxidation States

An oxidation-reduction reaction, or redox reaction, is one in which electrons are transferred.

$2Na + Cl_2 \rightarrow 2NaCl$

Each sodium atom is losing one electron to form Na^+

$Na \rightarrow Na^+ + 1e^-$ This loss of electrons is called oxidation.

Each chlorine atom is gaining 1 electron to form Cl^-

$Cl_2 + 2e^- \rightarrow 2Cl^-$ This gain of electrons is called reduction (charge reduced from 0 to -1)

These equations are called half-reactions
In a redox reaction the number of electrons lost must equal the number of electrons gained.

For ionic compounds, it is obvious that electrons are lost or gained. For covalent compounds, it is necessary to devise rules for counting electrons to determine if electrons are lost or gained, pg. 721. From a Lewis structure, oxidation numbers can be determined by assuming all shared electrons are on the more electronegative atom. (do H_2CO_3)

Rules for determinig oxidation states

1. Oxidation states for atoms in uncombined elements is 0.
2. The oxidation state of a monatomic ion is equal to the charge of the ion.
3. In compounds, F is always -1. Other halogens are -1 unless combined with a more electronegative element.
4. In compounds, H is +1 except for metal hydrides where it is -1.
5. In compounds, oxygen is always -2 except for peroxides (O_2^{-2}) and with F
6. The sum of the oxidation states muse equal zero in a neutral compound, or the charge of the ion for polyatomic ions.

$C_2H_2 + O_2 \rightarrow CO_2 + H_2O$ C: $-1 \rightarrow +4$ (oxidation) O: $0 \rightarrow -2$ (reduction)

do $KMnO_4$, $FeSO_4$, CH_2O

Balancing Oxidation-Reduction Equations

$\overset{0}{H_2} + 1/2\ \overset{0}{O_2} \rightarrow \overset{+1}{H_2}\overset{-2}{O}$

H_2 is oxidized by O_2, O_2 is reduced by H_2

$\overset{0}{Mg} + 2\overset{+1}{H}\overset{-1}{Cl} \rightarrow \overset{0}{Mg}\overset{-1}{Cl_2} + \overset{0}{H_2}$

Mg is oxidized by HCl, HCl is reduced by Mg

Ion-electron method:

- assign oxidation states and determine what is being oxidized and what is being reduced
- write 2 skeleton half reactions
- Balance each half-reaction seperately
  - balance each element except H and O
  - balance O by adding H_2O as needed
  - balance H by adding H^+
  - add electrons to make charge balance
- multiply one or both half reactions by a small integer so the number of electrons lost is equal to the number of electrons gained.
- add the two half reactions, cancelling out the electrons and any ions or molecules which appear on both sides of the equation.

$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$

$Cu + 2NO_3^- \rightarrow Cu(NO_3)_2 + 2\ e^-$

$HNO_3 \rightarrow NO + H_2O$

$HNO_3 \rightarrow NO + 2H_2O$

$HNO_3 + 3H^+ \rightarrow NO + 2H_2O$

$HNO_3 + 3H^+ + 3e^- \rightarrow NO + 2H_2O$

$3Cu + 6NO_3^- \rightarrow 3Cu(NO_3)_2 + 6\ e^-$

$2HNO_3 + 6H^+ + 6e^- \rightarrow 2NO + 4H_2O$

$3Cu + 6NO_3^- + 2HNO_3 + 6H^+ \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

do $Cr_2O_7^{-2} + C_2H_5OH \rightarrow Cr^{+3} + CO_2 + H_2O$

In basic solutions, balance as above and add OH^- to convert H^+ to H_2O

$$Ag + CN^- + O_2 \rightarrow Ag(CN)_2^-$$

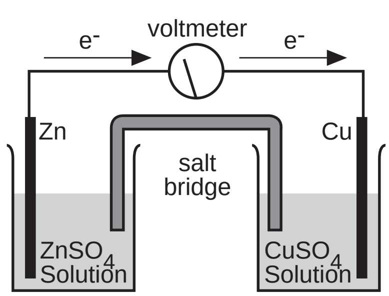
$$Ag + 2CN^- \rightarrow Ag(CN)_2^- + e^-$$
$$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$$

$$4Ag + 8CN^- + O_2 + 4H^+ \rightarrow 4Ag(CN)_2^- + 2H_2O$$
$$4Ag + 8CN^- + O_2 + 4H^+ + 4OH^- \rightarrow 4Ag(CN)_2^- + 2H_2O + 4OH^-$$
$$4Ag + 8CN^- + O_2 + 2H_2O \rightarrow 4Ag(CN)_2^- + 4OH^-$$

Voltaic Cells

$$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$$
Zn is oxidized, Cu^{2+} is reduced

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ oxidation half cell
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ reduction half cell



anode: electrode at which oxidation occurs
cathode: electrode at which reduction occurs

The two half reactions must be physically separated into **half cells**, otherwise Cu^{2+} will react directly with the Zn electrode. A salt bridge is a tube containing a gel of an inert salt such as KCl or NH_4NO_3. Ions are able to move through the bridge and complete the circuit through the cell.

The voltage difference between the electrodes is called the **electromotive force** or **emf** (E), measured in volts.

Cell diagram: anode to cathode, | to separate phases

$$Zn_{(s)} | Zn^{2+} \text{ (aq, 1 M)} | \text{KCl (sat'd)} | Cu^{2+} \text{ (aq., 1 M)} | Cu_{(s)}$$

Electrodes may be active or inert (Pt, Au, C). Gas phase reactions and reactions such as $Fe^{+2} \rightarrow Fe^{+3} + e^-$ may occur at an inert electrode.

Voltages of Voltaic Cells

Standard Electrode Potentials

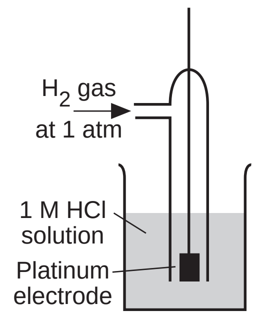
$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$ oxidation half cell
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ reduction half cell

We can treat each half cell as if it has its own electrical potential, and so the cell EMF is equal to the sum of the half-cell EMFs. However, EMF has to be measured relative to a reference. We define the Standard Hydrogen Electrode (SHE)

$H_{2(g)} \rightarrow 2H^+_{(aq)} + 2e^-$

the potential for the SHE is defined to be exactly zero, or

$E°_{H_2/H^+} = 0$ V



Standard reduction potential: reduction potential of a half cell when all solutes are at 1 M and all gases are at 1 ATM

Standard reduction potentials are measured relative to the SHE and are listed in table19.1 on pg. 785. Relative to SHE, the potential for the reaction

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$

is +0.76, because the EMF of the cell

$Zn_{(s)}|Zn^{2+}$ (aq, 1 M)|KCl (sat'd)|H^+ (aq., 1 M) |$H_{2(s)}$

is 0.76 volts.

$$
\begin{array}{ll}
Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^- & E°_{Zn/Zn^{2+}} = 0.76 \text{ V} \\
2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)} & E°_{H^+/H_2} = 0 \text{ V} \\
\hline
Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)} & E°_{cell} = 0.76 \text{ V}
\end{array}
$$

The standard cell potential is the sum of the standard oxidation potential and the standard reduction potential.

$E°_{cell} = E°_{ox} + E°_{red}$

$E°_{ox} = -E°_{red}$

By convention, only standard reduction potentials are listed.

Under standard conditions, a reaction is spontaneous in the forward direction if the cell emf is positive, and spontaneous in the reverse direction if the cell emf is negative.

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$	$E°_{Zn/Zn^{2+}} = 0.76$ V
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$	$E°_{Cu^{2+}/Cu} = 0.34$ V
$Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$	$E°_{cell} = 1.10$ V

For Standard Reduction Potentials

- E° applies only to reaction as read in the forward (reduction) direction
- For oxidation potentials, E°E°$_{ox}$ = -E°$_{red}$
- The more positive the SRP, the greater the tendency for the substance to be reduced, i.e., greater strength as an oxidizing agent. F_2 is the strongest oxidizing agent and Li is the weakest oxidizing agent.
- Oxidative and reductive strengths are relative. Strong oxidizing agents are weak reducing agents, and vice versa. Any given substance will act as an oxidizing agent when reacted with a substance above it in the SRP table, and will act as a reducing agent when reacted with a substance below it in the SRP table.
- Diagonal Rule: any species on the left side of a given half-cell reaction will react with a species on the right side of a half-cell located above it on the SRP table.

$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$	$E°_{Zn/Zn^{2+}} = -0.76$ V
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$	$E°_{Cu^{2+}/Cu} = +0.34$ V

Zn reacts with Cu^{2+} but Cu does not react with Zn^{2+}

Conversely, any species on the right side of a given half-cell reaction will react with a species on the left side of a half-cell located below it on the SRP table.

Arrange in order of increasing oxidative strength: O_3, Cl_2, MnO_4^-, $Cr_2O_7^{2-}$

Answer: $Cr_2O_7^{2-}$, Cl_2, MnO_4^-, O_3

What will happen if Fe is added to a solution of 1 M $Zn(NO_3)_2$ and 1 M $Co(NO_3)_2$

Answer:

$Zn^{2+}_{(aq)} + 2e^- \rightarrow Zn_{(s)}$	E° = -0.76 V
$Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$	E° = -0.44 V
$Co^{2+}_{(aq)} + 2e^- \rightarrow Co_{(s)}$	E° = -0.28 V

By the diagonal rule, $Fe_{(s)}$ will react with $Co^{2+}_{(aq)}$ but not with $Zn^{2+}_{(aq)}$

The reaction is $Fe_{(s)} + Co^{2+}_{(aq)} \rightarrow Fe^{2+}_{(aq)} + Co_{(s)}$

A galvanic cell is made up of an Al electrode in a 1 M solution of $Al(NO_3)_3$ and a Cu electrode in a 1 M solution of $Cu(NO_3)_2$. What is the cell emf, which electrode is the anode, which electrode is the cathode, what is the oxidizing agent, what is the reducing agent?

Answer: $Al^{3+}_{(aq)} + 3e^- \rightarrow Al_{(s)}$ $E°_{Al^{3+}/Al} = -1.66$ V
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$ $E°_{Cu^{2+}/Cu} = +0.34$ V

By the diagonal rule, $Cu^{2+}_{(aq)}$ will react with $Al_{(s)}$

$2(Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-)$ $E°_{Al/Al^{3+}} = +1.66$ V
$3(Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)})$ $E°_{Cu^{2+}/Cu} = +0.34$ V

$2Al_{(s)} + 3Cu^{2+}_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3Cu_{(s)}$ $E° = 2.00$ V

Al is the anode, Cu is the cathode, Cu^{2+} is the oxidizing agent, Al is the reducing agent.

Cell Potentials, G, and K

Gibbs Free Energy G = H - TS
G = H - TS = w_{max}
at standard conditions, G° = H° - TS°
G = G° + RT lnQ
at equilibrium, G = 0 and Q = K so
G° = -RT lnK

In a galvanic cell, electrical energy = volts x coulombs = joules
Total charge (coulombs) = nF where F, the Faraday, is 1 mole of electrons = 96,500 coulombs.

1 F = 96,500 coulombs/mol
1 F = 96,500 J/V mol

$w_{max} = -nFE_{cell}$
$G = -nFE_{cell}$

$2Al_{(s)} + 3Cu^{2+}_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3Cu_{(s)}$ $E° = 2.00$ V
$G = -nFE_{cell}$ = (-6)(96,500 J/V)(+2.00 V) = 1,160 kJ

$G° = -nFE°_{cell}$ = -RT ln K

$$E°_{cell} = \frac{RT}{nF} \ln K$$

at 298K,

$$E°_{cell} = \frac{.0591}{n} \log K$$

Dependence of Voltage on Concentration: the Nernst Equation

$$G = G° + RT \ln Q$$
$$-nFE_{cell} = -nFE°_{cell} + RT \ln Q$$
$$E_{cell} = E°_{cell} - \frac{RT}{nF} \ln Q$$

at 298 K

$$E_{cell} = E°_{cell} - \frac{.0591}{n} \log Q \qquad \text{this is the Nernst equation}$$

What is the cell potential for a cell made up of Zn, 0.500 M $Zn(NO_3)_2$, Cu, and 1.500 M $Cu(NO_3)_2$?

$$Zn^{2+}{}_{(aq)} + 2e^- \rightarrow Zn_{(s)} \qquad E°_{Zn/Zn^{2+}} = -0.76 \text{ V}$$
$$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \qquad E°_{Cu^{2+}/Cu} = +0.34 \text{ V}$$
$$Zn_{(s)} + Cu^{2+}{}_{(aq)} \rightarrow Zn^{2+}{}_{(aq)} + Cu_{(s)} \qquad E°_{cell} = 1.10 \text{ V}$$

$$E_{cell} = E°_{cell} - \frac{.0591}{n} \log Q = 1.10 - \frac{.0591}{2} \log \frac{[Zn^{2+}]}{[Cu^{2+}]} = 1.10 - .0295 \log \frac{0.500}{1.500} = 1.11 \text{ V}$$

As LeChatelier's principle predicts, increasing [reactant] and decreasing [product] drives reaction harder to the right.

Concentration Cells

$$Cu^{2+}{}_{(aq)} + 2e^- \rightarrow Cu_{(s)} \qquad E°_{Cu^{2+}/Cu} = +0.34 \text{ V}$$

For a half-cell with $[Cu^{2+}] = 0.0100$ M, $E_{cell} = E°_{cell} - \frac{.0591}{n} \log Q$

$$= 0.34 - \frac{.0591}{2} \log \frac{1}{[Cu^{2+}]} = 0.34 - .0295 \log \frac{1}{0.0100} = 0.28 \text{ V}$$

For a cell made up of one 1 M half cell and one 0.0100 M half cell, the cell emf would be

$$Cu^{2+}{}_{(aq,\ 1\ M)} + 2e^- \rightarrow Cu_{(s)} \qquad E°_{Cu^{2+}/Cu} = +0.34 \text{ V}$$
$$Cu_{(s)} + 2e^- \rightarrow Cu^{2+}{}_{(aq,\ 0.01\ M)} \qquad E°_{Cu^{2+}/Cu} = -0.28 \text{ V}$$
$$Cu^{2+}{}_{(aq,\ 1\ M)} \rightarrow Cu^{2+}{}_{(aq,\ 0.01\ M)} \qquad E°_{Cu^{2+}/Cu} = +0.06 \text{ V}$$

The system would tend to move toward an intermediate concentration. As the concentrations change, the half-cell potentials would converge on a common value and the cell potential would drop to zero.

Applications of Voltaic Cells in Chemistry

Ion-selective electrodes

Reference electrode: Calomel $Hg_2Cl_{2(s)} + 2e^- \rightarrow 2Hg_{(l)} + 2Cl^-_{(sat'd\ KCl)}$

Determination of K

$RT \ln K = nFE$ measure E, determine K

Commercial Voltaic Cells

Battery: a galvanic cell (or series of cells) which are designed to be self-contained.

Lead Storage Battery: Six cells in series. Each cell has a lead anode, 38% H_2SO_4 (sulfuric acid), and a PbO_2/metal cathode.

$$Pb_{(s)} + SO_4^{-2}{}_{(aq)} \rightarrow PbSO_{4(s)} + 2e^-$$

$$PbO_{2(s)} + 4H^+_{(aq)} + SO_4^{-2}{}_{(aq)} + 2e^- \rightarrow PbSO_{4(s)} + 2H_2O_{(l)}$$

$$Pb_{(s)} + PbO_{2(s)} + 4H^+_{(aq)} + 2SO_4^{-2}{}_{(aq)} \rightarrow 2PbSO_{4(s)} + 2H_2O_{(l)}$$

The battery is rechargeable because both reactions are reversible.
The battery can be checked by testing the density of the electrolyte; H_2SO_4 is replaced by water as the battery discharges.

Dry Cell: Zn can is anode, in contact with moist paste of MnO_2, NH_4Cl, and $ZnCl_2$ in starch. A carbon rod cathode is in contact with the metal cap on top of the cell.

$$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$$

$$2NH_4^+{}_{(aq)} + MnO_{2(s)} + 2e^- \rightarrow Mn_2O_{3(s)} + 2NH_{3(aq)} + H_2O_{(l)}$$

$$Zn_{(s)} + 2NH_4^+{}_{(aq)} + MnO_{2(s)} \rightarrow Zn^{2+}_{(aq)} + Mn_2O_{3(s)} + 2NH_{3(aq)} + H_2O_{(l)}$$

Alkaline Battery: NH_4Cl is replaced with KOH or NaOH. Zn anode corrodes less under alkaline conditions and cell voltage is more constant because net reaction contains all condensed-phase species.

$$Zn_{(s)} + 2OH^-_{(aq)} \rightarrow ZnO_{(s)} + H_2O_{(l)} + 2e^-$$

$$2MnO_{2(s)} + H_2O_{(l)} + 2e^- \rightarrow Mn_2O_{3(s)} + 2OH^-_{(aq)}$$

$$Zn_{(s)} + MnO_{2(s)} \rightarrow ZnO_{(s)} + Mn_2O_{3(s)}$$

Mercury Battery: Zn/Hg amalgam anode in contact with a paste containing KOH, $Zn(OH)_2$, and HgO, in contact with a steel cathode.

$$Zn(Hg)_{(s)} + 2OH^-_{(aq)} \rightarrow ZnO_{(s)} + H_2O_{(l)} + 2e^-$$

$$HgO_{(s)} + H_2O_{(l)} + 2e^- \rightarrow Hg_{(l)} + 2OH^-_{(aq)}$$

$$Zn(Hg)_{(s)} + HgO_{(s)} \rightarrow ZnO_{(s)} + Hg_{(l)}$$

Since all reactants and products are in condensed phases, the cell voltage remains constant until the reactants are spent.

Ni-Cad Battery:

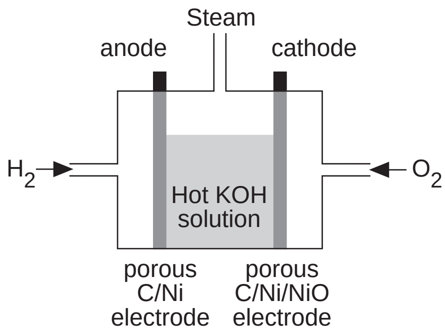
$$Cd_{(s)} + 2OH^-_{(aq)} \rightarrow Cd(OH)_{2(s)} + 2e^-$$

$$NiO_{2(s)} + 2H_2O_{(l)} + 2e^- \rightarrow Ni(OH)_{2(s)} + 2OH^-_{(aq)}$$

$$Cd_{(s)} + NiO_{2(s)} + 2H_2O_{(l)} \rightarrow Cd(OH)_{2(s)} + Ni(OH)_{2(s)}$$

The battery is rechargeable because both reactions are reversible.

Fuel Cell: A galvanic cell in which the reactants are continuously supplied.



Ni and NiO are catalysts. E° = 1.23 V

steam generated by cell is condensed to water for use as drinking water

Electrolysis

Electrolysis: process in which electrical energy is used to make a non-spontaneous chemical reaction occur.

Electrolysis of molten NaCl

Anode (oxidation)	$2Cl^- \rightarrow Cl_{2(g)} + 2e^-$	E° = -1.36 V
Cathode (reduction)	$2Na^+ + 2e^- \rightarrow 2Na_{(l)}$	E° = -2.71 V
overall	$2Na^+ + 2Cl^- \rightarrow 2Na_{(l)} + Cl_{2(g)}$	E° = -4.07 V

Electrolysis of water (in acid solution)

Anode (oxidation)	$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$	E° = -1.23 V
Cathode (reduction)	$2H^+ + 2e^- \rightarrow H_{2(g)}$	E° = 0 V
overall	$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$	E° = -1.23 V

Electrolysis of aqueous NaCl

Anode (oxidation)	$2Cl^- \rightarrow Cl_{2(g)} + 2e^-$	E° = -1.36 V
or	$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$	E° = -1.23 V
Cathode (reduction)	$2Na^+ + 2e^- \rightarrow 2Na_{(l)}$	E° = -2.71 V
or	$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$	E° = -0.83 V
or	$2H^+ + 2e^- \rightarrow H_{2(g)}$	E° = 0 V

overall $2H_2O_{(l)} + 2Cl^- \rightarrow H_{2(g)} + Cl_{2(g)} + 2OH^-_{(aq)}$ E°= -4.07 V

Overvoltage: difference between calculated potential and actual potential needed to cause the reaction.

Quantitative Aspects of Electrolysis

A current of 0.452 A is passed for 1.50 hrs through a cell containing molten $CaCl_2$. What masses ofch product is formed?

$Ca^{+2} + 2e^- \rightarrow Ca_{(s)}$

$2Cl^- \rightarrow Cl_{2(g)} + 2e^-$

0.452 A x 1.50 hr x 3600 s/hr = 2,440 coul

$$2{,}440 \text{ coul} \times \frac{1 \text{ F}}{96500 \text{ coul}} = 0.0253 \text{ F} \times \frac{1 \text{ mol Ca or } Cl_2}{2\text{F}} = 0.0126 \text{ mol}$$

0.0126 mol Ca x 40.08 g/mol = 0.507 g Ca
0.0126 mol Cl_2 x 70.90 g/mol = 0.897 g Cl_2

Energy consumed E = Q x V What is the energy used above if the cell voltage was 5.0 V?

E = Q x V = (2440 coul)(5.0 V) = 12,200 J (1 V = 1 J/coul)

1 kw-hr = 3.60×10^6 J so 12,200 J x 1 kw-hr/3.60×10^6 J = 0.0034 kw-hr

Electrolysis in Industry

Aluminum - the Hall process. Al_2O_3 is mixed with Na_3AlF_6 (cryolite) to lower the mp to 980°. Carbon electrodes are used at 4.2 V to form Al at the cathode and O_2 at the anode, which reacts with the electrodes to form CO_2.

Sodium - NaCl is mixed with $CaCl_2$ to lower the mp and electrolyzed to form Na metal and Cl_2 gas

Fluorine - mixture of HF and KF is electrolyzed to form H_2 and F_2

Chlorine - mostly produced by electrolysis of brine

$2NaCl + 2H_2O \rightarrow Cl_2 + H_2 + NaOH$

Electrorefining of Cu - Cu is dissolved at the anode and pure Cu is deposited at the cathode

Corrosion

Corrosion: deterioration of metal by an electrochemical process.

$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$$

$$\underline{O_{2(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(l)}}$$

$$2Fe_{(s)} + O_{2(s)} + 4H^+_{(aq)} \rightarrow 2Fe^{2+}_{(aq)} + 2H_2O_{(l)} \qquad E° = 1.67 \text{ V}$$

H^+ is supplied by reaction of CO_2 with water to form H_2CO_3

Fe^{2+} is further oxidized

$$4Fe^{2+}_{(aq)} + O_{2(s)} + (4 + 2x)H_2O_{(l)} \rightarrow 2Fe_2O_3{\cdot}xH_2O(s) + 8H^+_{(aq)}$$

iron (III) oxide is rust.

Some metals, such as Al, oxidize to form a hard, non-porous coating which is then resistant to additional oxidation.

Cathodic Protection: connecting the metal to be protected to a more electropositive metal (such as Fe to Zn or Mg), which makes the protected metal the cathode in the electrochemical cell.