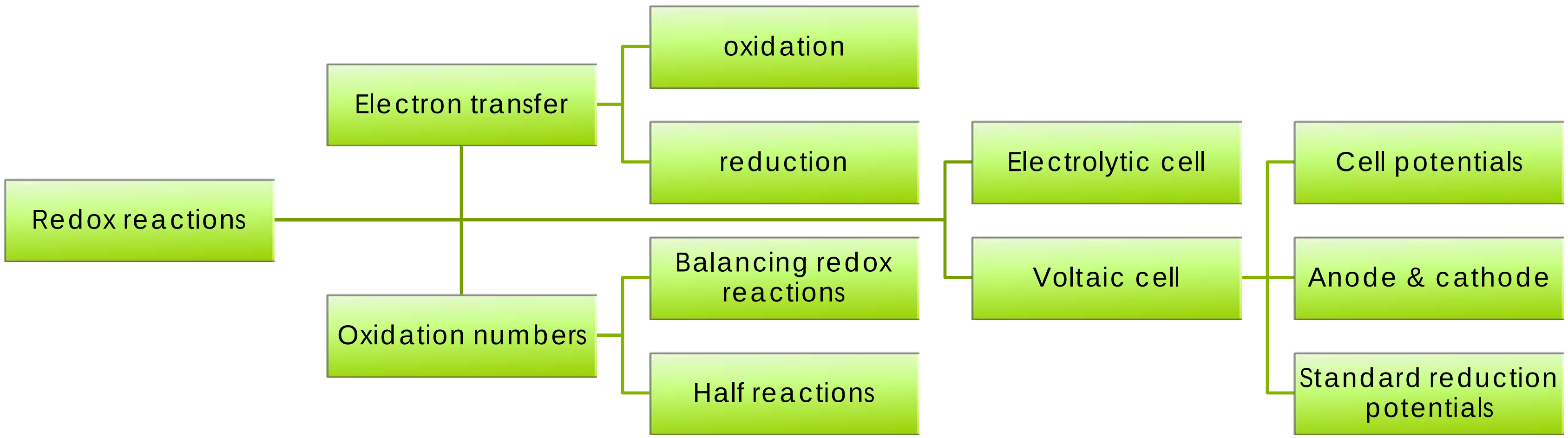
Redox

Electrochemical Cells

Redox unit advance organizer



Assigning Oxidation Numbers

- Elemental samples of matter have an oxidation number of 0.
  - Examples:

Assigning Oxidation Numbers

- For a monotomic ion, the charge of the ion is the oxidation number.
  - Examples:

Assigning Oxidation Numbers

- Hydrogen in a compound usually has an oxidation number of +1.
  - Example:

Assigning Oxidation Numbers

- Hydrogen in a compound usually has an oxidation number of +1.
- Exception:
  - When in a metal hydride, H has an oxidation number of -1.
    - Ex.

Assigning Oxidation Numbers

- Fluorine in a compound has an oxidation number of -1.
  - Ex

Assigning Oxidation Numbers

- Oxygen in a compound usually has an oxidation number of -2.
  - Ex.

Assigning Oxidation Numbers

- Oxygen in a compound usually has an oxidation number of -2.
- Exception: In peroxides, oxygen has an oxidation number of -1.
  - Ex.

Assigning Oxidation Numbers

- Exception: In compounds with fluorine, oxygen has an oxidation number of +2.
  - Ex.

Assigning Oxidation Numbers

- In a neutral compound, the sum of the oxidation numbers equals zero.
- In a polyatomic ion, the sum of the oxidation numbers equal the overall charge of the ion.

Oxidation

- Oxidation = loss of electrons
- In oxidation, the oxidation number increases
- Oxidation half reactions:
  - $Na \rightarrow Na^{+} + e^{-}$
  - $Mg \rightarrow Mg^{2+} + 2e^{-}$

Reduction

- Reduction = gain of electrons
- In reduction, the oxidation number decreases
- Reduction half reactions:
  - $Fe^{3+} + 3e^- \rightarrow Fe$
  - $F_2 + 2e^- \rightarrow 2F^-$

Redox

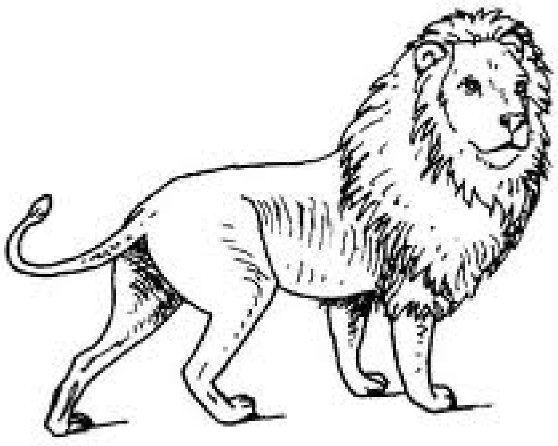
- Oxidation can't occur without reduction!
- The number of electrons lost must equal the number of electrons gained.
- In a redox reaction, oxidation numbers change.

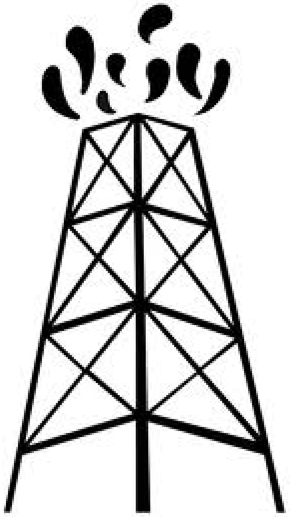
Are these redox reactions?

- $CuO + H_2 \rightarrow Cu + H_2O$
- $H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + H_2O$
- $MgCO_3 \rightarrow MgO + CO_2$
- $I_2O_5 + CO \rightarrow I_2 + CO_2$
- $P_4 + S_8 \rightarrow P_2S_5$

Mnemonics

- LEO the lion says GER
- OIL RIG





What is oxidized? What is reduced?

- $2Na + Br_2 \rightarrow 2\ NaBr$
- $N_2 + 3\ H_2 \rightarrow 2\ NH_3$
- $Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

More about redox reactions

- Oxidizing agent = the substance that gets reduced
- Reducing agent = the substance that gets oxidized

Identify the oxidizing agent. Identify the reducing agent.

- $2Na + Br_2 \rightarrow 2\ NaBr$

- $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

- $Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

Electrochemical cells

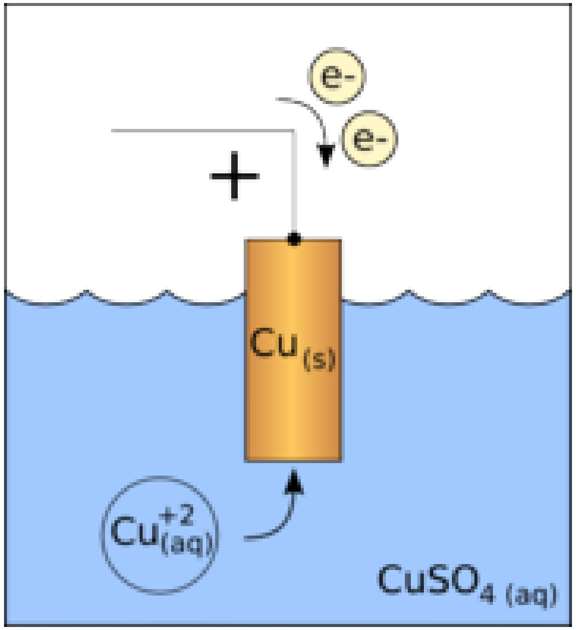
- Voltaic cells
  - Aka galvanic cells
  - Set up to release electrical energy
  - batteries

Electrochemical cells

- Electrolytic cells
  - Consume electrical energy to drive a chemical reaction
  - electroplating

Galvanic cells

- Electrode = metal strip
- Half-cell = electrode immersed in a solution of its ions
- Two half cells are needed for an electrochemical cell



Completing the circuit

- Connect the electrodes with an external wire
- Connect the half-cells with a "salt bridge" to complete the circuit
  - The salt bridge allows the movement of ions, but keeps the solutions physically separated

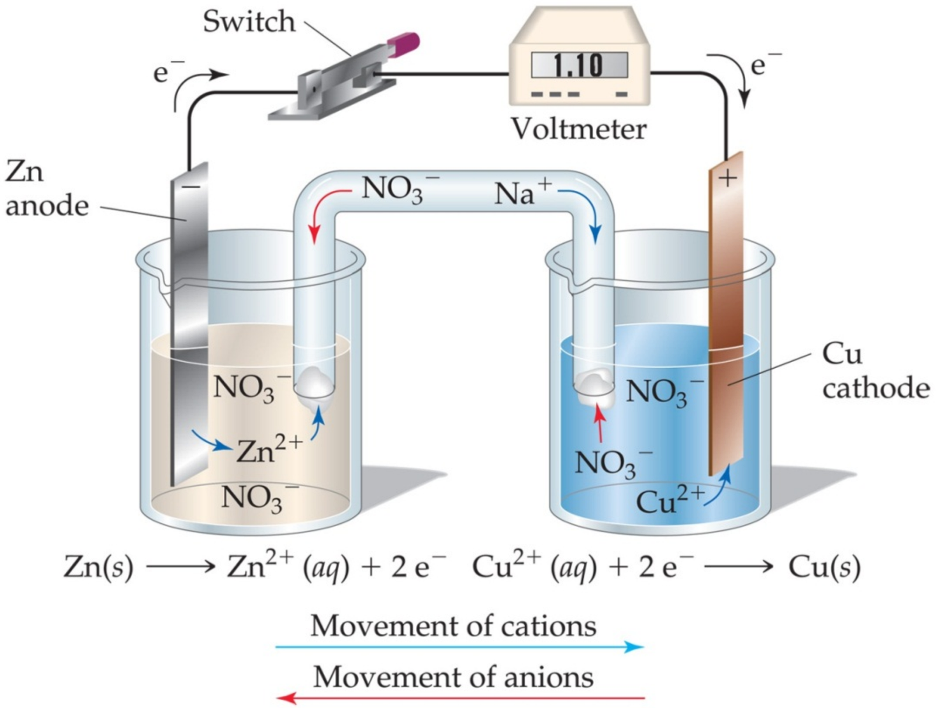
Parts of the galvanic cell

- Oxidation occurs at the anode
  - The anode is the negative electrode
- Reduction occurs at the cathode
  - The cathode is the positive electrode

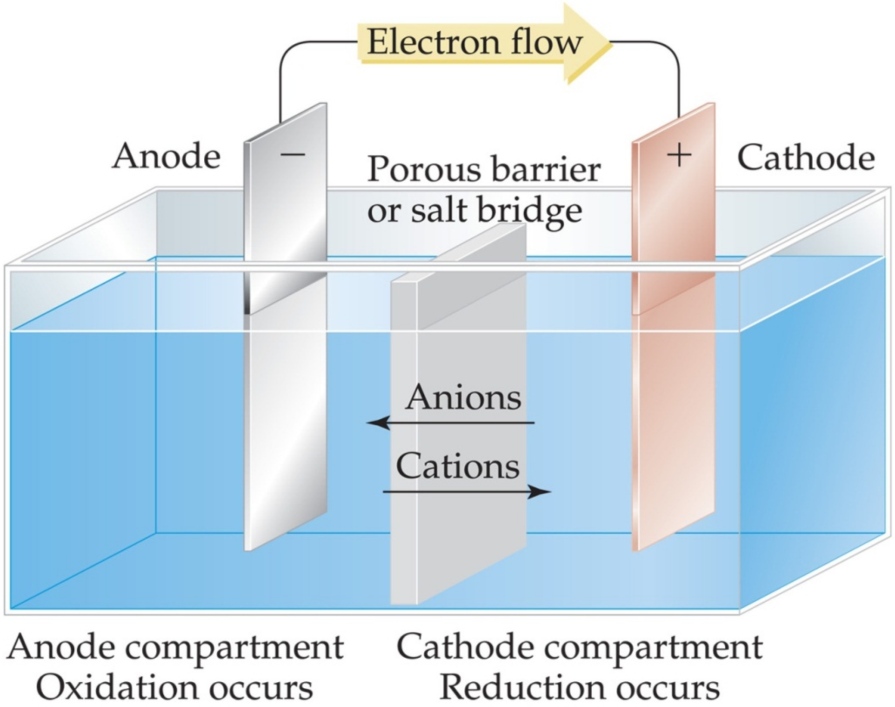
Parts of the galvanic cell

- Electrons flow from the anode to the cathode
- Mnemonic:
  - RED CAT
  - AN OX

A galvanic cell



A visual summary



As the galvanic cell runs…

- The anode will get less massive, as cations are generated
  - $A \rightarrow A^+ + e^-$
  - The concentration of metal cations will increase in the anode compartment

As the galvanic cell runs…

- The cathode will get more massive, as cations are reduced to form the metal
  - $C^+ + e^- \rightarrow C$
  - The concentration of metal cations in the cathode will decrease

A virtual model

- Go to http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/galvan5.swf

Standard Reduction Potentials

- A relative measure of tendency to undergo reduction
  - More positive Eº value (Higher up): greater tendency to reduce
  - Less positive Eº value (lower down): greater tendency to oxidize

TABLE 20.1 Standard Reduction Potentials in Water at 25°C

Potential (V)	Reduction Half-Reaction
+2.87	$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$
+1.51	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$
+1.36	$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$
+1.33	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
+1.23	$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$
+1.06	$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq)$
+0.96	$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$
+0.80	$Ag^+(aq) + e^- \longrightarrow Ag(s)$
+0.77	$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$
+0.68	$O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O_2(aq)$
+0.59	$MnO_4^-(aq) + 2H_2O(l) + 3e^- \longrightarrow MnO_2(s) + 4OH^-(aq)$
+0.54	$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$
+0.40	$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$
+0.34	$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$
0 [defined]	$2H^+(aq) + 2e^- \longrightarrow H_2(g)$
−0.28	$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$
−0.44	$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$
−0.76	$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$
−0.83	$2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$
−1.66	$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$
−2.71	$Na^+(aq) + e^- \longrightarrow Na(s)$
−3.05	$Li^+(aq) + e^- \longrightarrow Li(s)$

Calculating E_{cell}

- Find the half-reactions on the Table of Standard Reduction Potentials
- The reaction that is higher on the table:
  - Greater tendency to reduce, so keep as written
  - This reaction will occur at the cathode
- The reaction that is lower on the table:
  - Lesser tendency to reduce, so flip it to be an oxidation half reaction, change sign of E^o
  - This reaction will occur at the anode
- Add the E^o values to find E_{cell}

Calculate E_{cell}

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