AP Chemistry Prep Session
Saturday, December 12, 2009

John I. Gelder
Department of Chemistry
Oklahoma State University

John.gelder@okstate.edu
http://intro.chem.okstate.edu

Oxidation-Reduction

Electron transfer reactions

Recognizing Oxidation reduction reactions

- Elemental form of a substance reacts or is formed;
- Change in the number of oxygen atoms around an element;
- An increase or decrease in oxidation state;

Some example reactions

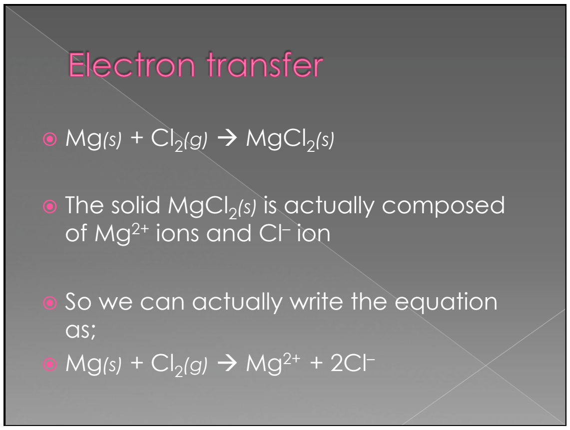
- $S(s) + O_2(g) \rightarrow SO_2\,(g)$
- $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$
- $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
- $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

Electron transfer

- $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

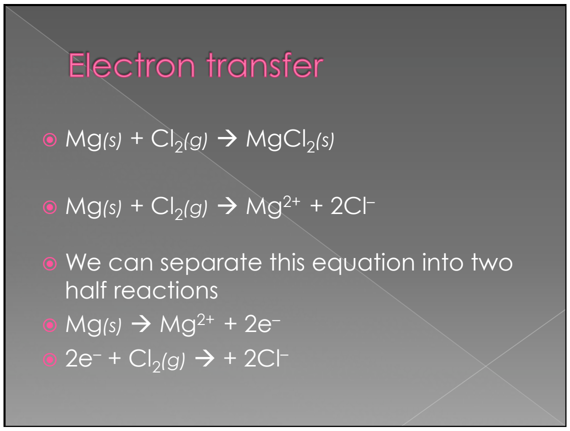
Electron transfer

- $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
- The solid $MgCl_2(s)$ is actually composed of Mg^{2+} ions and Cl^- ions



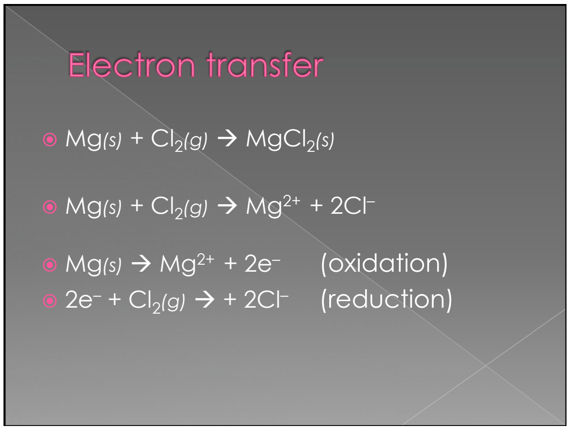
Electron transfer

- $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
- The solid $MgCl_2(s)$ is actually composed of Mg^{2+} ions and Cl^- ion
- So we can actually write the equation as;
- $Mg(s) + Cl_2(g) \rightarrow Mg^{2+} + 2Cl^-$



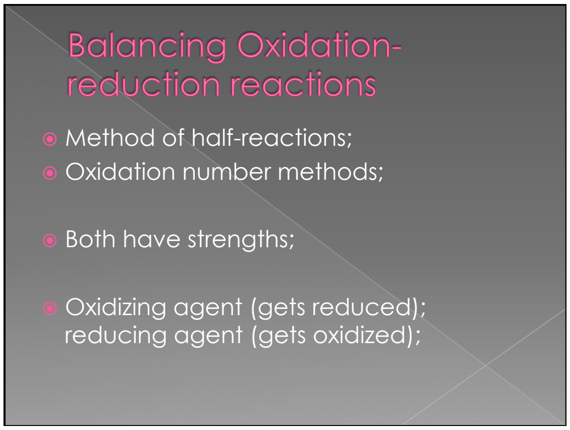
Electron transfer

- $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
- $Mg(s) + Cl_2(g) \rightarrow Mg^{2+} + 2Cl^-$
- We can separate this equation into two half reactions
- $Mg(s) \rightarrow Mg^{2+} + 2e^-$
- $2e^- + Cl_2(g) \rightarrow + 2Cl^-$



Electron transfer

- $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
- $Mg(s) + Cl_2(g) \rightarrow Mg^{2+} + 2Cl^-$
- $Mg(s) \rightarrow Mg^{2+} + 2e^-$ (oxidation)
- $2e^- + Cl_2(g) \rightarrow + 2Cl^-$ (reduction)

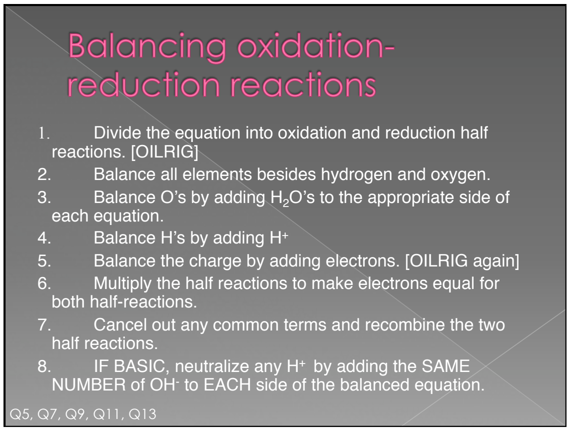


Balancing Oxidation-reduction reactions

- Method of half-reactions;
- Oxidation number methods;

- Both have strengths;

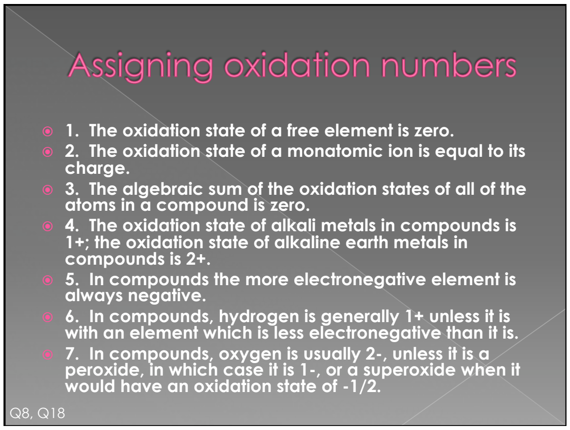
- Oxidizing agent (gets reduced); reducing agent (gets oxidized);



Balancing oxidation-reduction reactions

1. Divide the equation into oxidation and reduction half reactions. [OILRIG]
2. Balance all elements besides hydrogen and oxygen.
3. Balance O's by adding H_2O's to the appropriate side of each equation.
4. Balance H's by adding H^+
5. Balance the charge by adding electrons. [OILRIG again]
6. Multiply the half reactions to make electrons equal for both half-reactions.
7. Cancel out any common terms and recombine the two half reactions.
8. IF BASIC, neutralize any H^+ by adding the SAME NUMBER of OH^- to EACH side of the balanced equation.

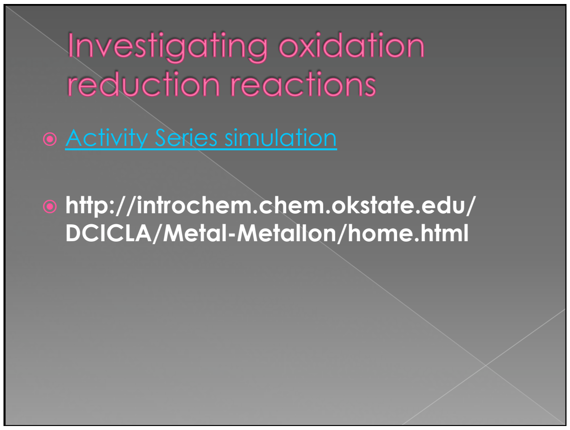
Q5, Q7, Q9, Q11, Q13



Assigning oxidation numbers

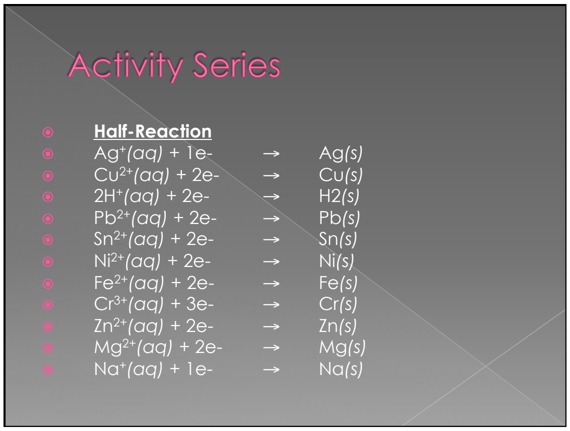
- **1. The oxidation state of a free element is zero.**
- **2. The oxidation state of a monatomic ion is equal to its charge.**
- **3. The algebraic sum of the oxidation states of all of the atoms in a compound is zero.**
- **4. The oxidation state of alkali metals in compounds is 1+; the oxidation state of alkaline earth metals in compounds is 2+.**
- **5. In compounds the more electronegative element is always negative.**
- **6. In compounds, hydrogen is generally 1+ unless it is with an element which is less electronegative than it is.**
- **7. In compounds, oxygen is usually 2-, unless it is a peroxide, in which case it is 1-, or a superoxide when it would have an oxidation state of -1/2.**

Q8, Q18



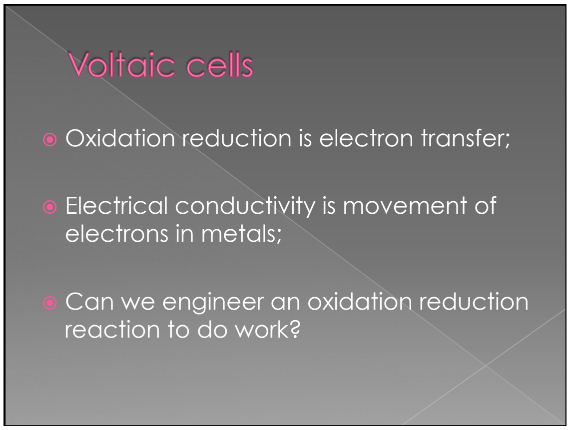
Investigating oxidation reduction reactions

- Activity Series simulation
- **http://introchem.chem.okstate.edu/DCICLA/Metal-MetalIon/home.html**



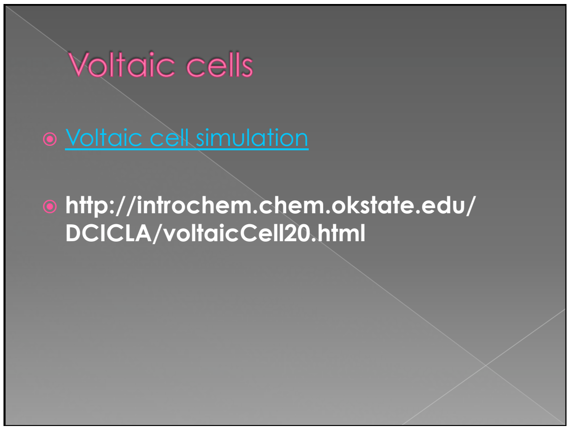
Activity Series

Half-Reaction		
$Ag^{+}(aq)$ + 1e-	→	$Ag(s)$
$Cu^{2+}(aq)$ + 2e-	→	$Cu(s)$
$2H^{+}(aq)$ + 2e-	→	$H2(s)$
$Pb^{2+}(aq)$ + 2e-	→	$Pb(s)$
$Sn^{2+}(aq)$ + 2e-	→	$Sn(s)$
$Ni^{2+}(aq)$ + 2e-	→	$Ni(s)$
$Fe^{2+}(aq)$ + 2e-	→	$Fe(s)$
$Cr^{3+}(aq)$ + 3e-	→	$Cr(s)$
$Zn^{2+}(aq)$ + 2e-	→	$Zn(s)$
$Mg^{2+}(aq)$ + 2e-	→	$Mg(s)$
$Na^{+}(aq)$ + 1e-	→	$Na(s)$



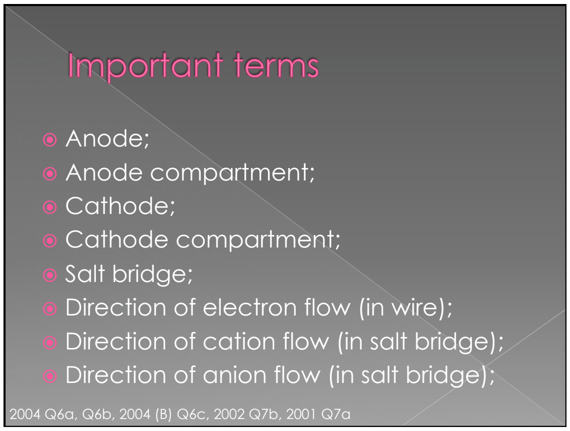
Voltaic cells

- Oxidation reduction is electron transfer;
- Electrical conductivity is movement of electrons in metals;
- Can we engineer an oxidation reduction reaction to do work?



Voltaic cells

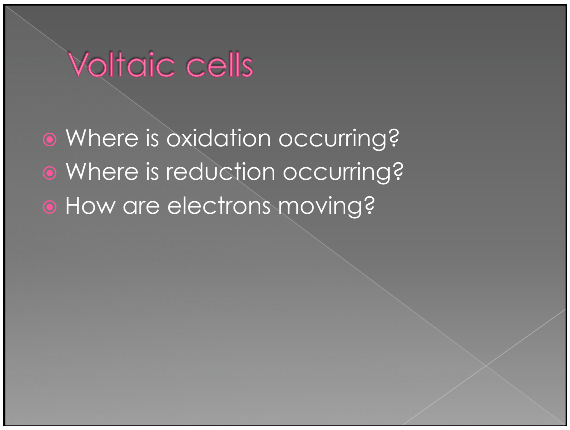
- Voltaic cell simulation
- **http://introchem.chem.okstate.edu/DCICLA/voltaicCell20.html**



Important terms

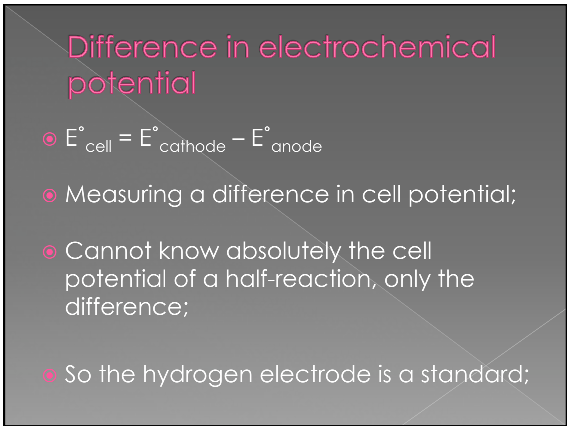
- Anode;
- Anode compartment;
- Cathode;
- Cathode compartment;
- Salt bridge;
- Direction of electron flow (in wire);
- Direction of cation flow (in salt bridge);
- Direction of anion flow (in salt bridge);

2004 Q6a, Q6b, 2004 (B) Q6c, 2002 Q7b, 2001 Q7a



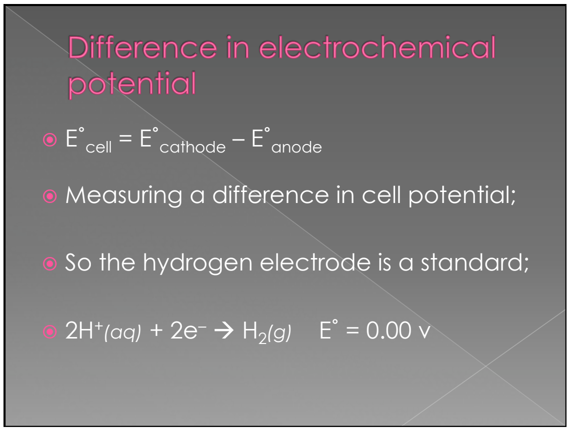
Voltaic cells

- Where is oxidation occurring?
- Where is reduction occurring?
- How are electrons moving?



Difference in electrochemical potential

- $E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$
- Measuring a difference in cell potential;
- Cannot know absolutely the cell potential of a half-reaction, only the difference;
- So the hydrogen electrode is a standard;



Difference in electrochemical potential

- $E^{\circ}_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$
- Measuring a difference in cell potential;
- So the hydrogen electrode is a standard;
- $2H^+(aq) + 2e^- \rightarrow H_2(g)$ $E^{\circ} = 0.00$ v

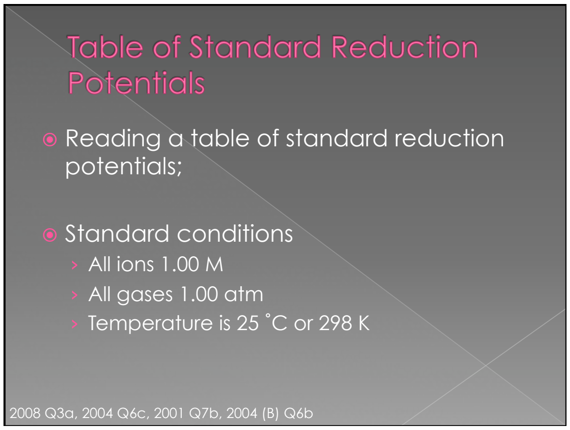
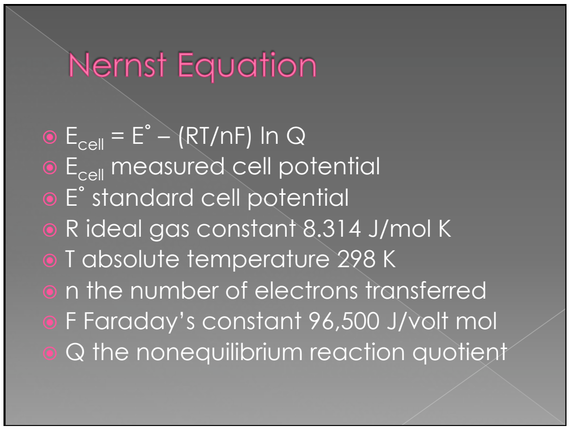


Table of Standard Reduction Potentials

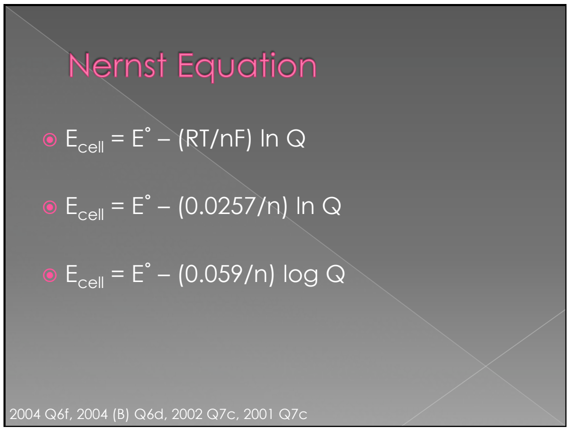
- Reading a table of standard reduction potentials;
- Standard conditions
  - All ions 1.00 M
  - All gases 1.00 atm
  - Temperature is 25 ˚C or 298 K

2008 Q3a, 2004 Q6c, 2001 Q7b, 2004 (B) Q6b



Nernst Equation

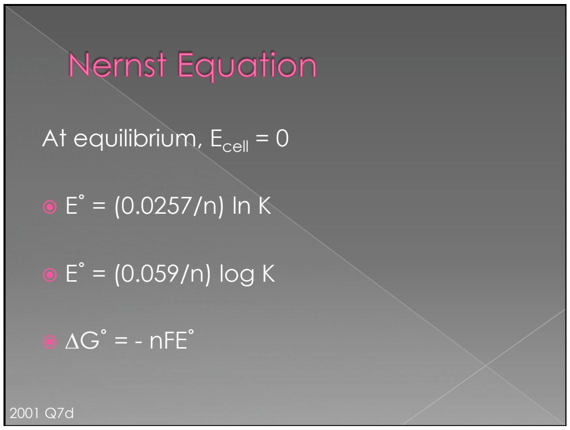
- $E_{cell} = E^{\circ} – (RT/nF) \ln Q$
- E_{cell} measured cell potential
- E° standard cell potential
- R ideal gas constant 8.314 J/mol K
- T absolute temperature 298 K
- n the number of electrons transferred
- F Faraday's constant 96,500 J/volt mol
- Q the nonequilibrium reaction quotient



Nernst Equation

- $E_{cell} = E^{\circ} – (RT/nF) \ln Q$
- $E_{cell} = E^{\circ} – (0.0257/n) \ln Q$
- $E_{cell} = E^{\circ} – (0.059/n) \log Q$

2004 Q6f, 2004 (B) Q6d, 2002 Q7c, 2001 Q7c



Nernst Equation

At equilibrium, $E_{cell} = 0$

- $E^{\circ} = (0.0257/n) \ln K$
- $E^{\circ} = (0.059/n) \log K$
- $\Delta G^{\circ} = - nFE^{\circ}$

2001 Q7d