CHEM 116

Redox Reactions and Electrochemistry

November 30, 2006
Prof. Sevian

Note: Start with the lecture slides from Tuesday that we didn't finish

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Electrochemistry overview

- What you need to know (mostly boils down to vocabulary)
  - How to recognize an electron transfer (redox) reaction
  - Oxidation vs. reduction
    - Loss of electrons is oxidation, gain of electrons is reduction (LEO the lion says GER)
    - Reduction is moving (oxidation number) down on a number line, oxidation is up
  - Agent means "that which causes"
- What you need to be able to do
  - Assign oxidation numbers to elements to figure out which **species** contain elements that are changing
  - Balance redox rxns in neutral, acid, and base solutions
  - Diagram an electrochemical cell and indicate direction of flow of negative charge through the system
  - Read and write electrochemical cell notation
  - Use reduction potentials to predict cell voltage under standard conditions
  - Use the Nernst equation to predict cell voltage at non-standard conditions

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Agenda

- Review oxidation states (from first semester course)
- Balancing redox reactions
  - Inspection method
  - SOHe method
- Redox reactions that occur spontaneously
  - Predicting whether a reaction will be spontaneous
  - How to separate the halves of a reaction to generate electricity
  - How much electricity?

The **final exam** is scheduled for Monday, December 18, 11:30AM-2:30PM. Location TBA.

Redox Reactions and Electrochemistry

Part I: Foundational

- Assigning oxidation numbers
- Balancing redox equations

Part II: Applied

- Diagramming electrochemical cells
- Facility with cell notation

Part III: Making mathematical predictions of properties

- Calculating cell potential at standard conditions
- Calculating cell potential at non-standard conditions
- Relating the amount of a substance oxidized or reduced to the electrical current and the time during which current is allowed to flow

Some vocabulary that applies to redox reactions

- Species = a reactant that contains an element whose oxidation number changes
  - On the reactant side, there must be one species that gets oxidized (oxidation number of an element in it goes up) <u>and</u> one species that gets reduced (oxidation number of an element in it goes down)
- Oxidation half-reaction: the reaction showing the species that gets oxidized, its product, and how many electrons get produced
- Reduction half-reaction: the reaction showing the species that gets reduced, its product, and how many electrons react in order for this to happen
- Species that gets oxidized is called the reducing agent because it causes the reduction half-reaction to occur
- Species that gets reduced is called the oxidizing agent because it causes the oxidation half-reaction to occur

Electrochemical cells

- Electrochemical, voltaic, galvanic = all the same thing
- Divide the half-reactions into separate cells (locations) so that electrons generated at the oxidation half-cell are forced to travel through a wire to get to the reduction half-cell
- http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/electroChem/voltaicCellEMF.html

- What's the role of the salt bridge?
- Why do reactions occur in only one direction?
- How can you predict the voltage?

Key ideas so far about electrochem

- You must review ions and ionic charges
- You must be able to assign oxidation numbers to elements in a compound or ion
- Redox reactions are when oxidation numbers change during a reaction
  - One element (part or all of a "species") has its oxidation number increase (this is oxidation)
  - Another element (part or all of a "species") has its oxidation number decrease (this is reduction)
  - Both processes must occur, because redox is electron transfer
- Study the vocabulary

Parts of an electrochemical cell

- Anode = oxidation half-cell
- Cathode = reduction half-cell
- Wire
- Salt bridge

Predicting voltage under standard conditions

- Voltage = cell potential (E^o_{cell}) = related to ΔH
- Table of reference voltages, all referenced to the hydrogen half-cell
- All reference voltages are reduction potentials, so if you need an oxidation potential you just take the opposite
- See table of standard reduction potentials
- Examples

Examples

Non-standard conditions: Nernst equation

- Concentrations of solutions not at standard 1.0 M

$$E_{cell} = E^o_{cell} - \frac{0.0257}{n}\ln Q$$

$$= E^o_{cell} - \frac{0.0592}{n}\log Q$$

Reaction quotient

Least common multiple of electrons exchanged in the redox reaction

- Temperature not at standard 25 ºC

$$E_{cell} = E^o_{cell} - \frac{RT}{n\Im}\ln Q$$

Electrolysis: electrochemical cell forced to run backwards (against its will)

- Note: sections 20.7 and 20.8 in the text book are fascinating reading, but they are not required reading
- Electroplating: The longer you run the cell, the more metal electroplate builds up
- Current means how many electrons pass by per second
- If you know how much metal electroplate you want to make, then stoichiometry tells you how many electrons are required
  - Cu^{2+} (*aq*) + 2 e^- → Cu (*s*)
  - Au^{3+} (*aq*) + 3 e^- → Au (*s*)
  - Ag^+ (*aq*) + 1 e^- → Ag (*s*)
- If you know what current is being applied (e.g., 0.800 amperes = 0.800 Coulombs/second), then you can figure out how much time you must run the cell for in order to build the amount of electroplated metal that you want

Electrolysis example

Similar to Sample Exercise 20.14, pp. 885-886

Calculate the mass of O_2 produced in the electrolysis of water, using a current of 0.445 A for a period of 45 minutes.