CHEM 116 Electrochemical Cells

Lecture 22 Prof. Sevian



Today's agenda

• Big picture of electrochemistry

- Redox reactions and oxidation numbers (last lecture)
- Charge flow in electrochemical cells and diagramming a cell
- Using the mathematical model to predict current and voltage under standard and non-standard conditions

Announcements

- Exam 3 is this Thursday, Dec 4
- Final exam has been scheduled for Tuesday, Dec 16, 3:00pm in S-1-006

SOHe⁻ method of balancing redox equations

Separate the redox reaction into reduction and oxidation halves, and then for each half:

- Species in the species that contains the element whose oxidation number is changing, balance that element
- Oxygen balance the oxygens by adding H₂O on whichever side of the equation is missing oxygen
- Hydrogen balance the hydrogens by:
 - If acid solution, add H+ to one side
 - If basic solution, add OH⁻ to one side and H₂O to the other side, so H₂O is added to side needing more H atoms
- Electrons balance the charge on both sides by adding electrons (e⁻) to whichever side requires negative charge

Now add the two half reactions together in such a way that the electrons cancel (least common multiple)

Practice using SOHer method

1. $\operatorname{Cu}(s) + \operatorname{NO}_{3}(aq) \rightarrow \operatorname{Cu}^{2+}(aq) + \operatorname{NO}_{2}(g)$ in acid

Practice using SOHe⁻ method

2. $\operatorname{Cd}(s) + \operatorname{NiO}_2(s) + \operatorname{H}_2\operatorname{O}(l) \rightarrow \operatorname{Cd}(\operatorname{OH})_2(s) + \operatorname{Ni}(\operatorname{OH})_2(s)$ in acid

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

Review of 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) and 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 mean 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
- <u>http://www.chem.iastate.edu/group/Greenbowe/sections/proje</u> ctfolder/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?



Parts of an electrochemical cell and cell notation

- Anode is the site of oxidation in the oxidation half-cell
- · Cathode is the site of reduction in the reduction half-cell
- Wire
- Salt bridge

Note: Standard cell notation is not explained in the text book, but it is on the final exam so you need to learn it.

Standard cell notation:
Anode | Anode ion || Cathode ion | Cathode
(oxidation) (salt (reduction)

bridge)

Standard cell notation for the example that we did in class
Zn (s) | Zn²⁺ (aq, 1M) || Cu²⁺ (aq, 1M) | Cu (s)

oxidation is $Zn \rightarrow Zn^{2+} + 2e^{-}$

Ways to remember cell notation:

reduction is $Cu^{2+} + 2e^{-} \rightarrow Cu$

- 1) it shows the direction that electrons flow through the wire (from oxidation half-cell to reduction half-cell)
- 2) oxidation (anode) to reduction (cathode) is alphabetical order

Predicting voltage under standard conditions

- Voltage = cell potential (E^{o}_{cell}) = related to ΔG , which is related to ΔH , T, and ΔS
- 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 of using standard reduction potentials to predict Eocell

Non-standard conditions: Nernst equation

Concentrations of solutions not at standard 1.0 M

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

= $E_{cell}^{o} - \frac{0.0592}{n} \log Q$
• Temperature not at standard 25 °C
 $E_{cell} = E_{cell}^{o} - \frac{RT}{n} \ln Q$
Least common multiple of electrons exchanged in the redox reaction

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

redox reaction

where \mathcal{F} is the Faraday constant: \mathcal{F} = 96,500 C/mol

http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/electroChem /voltaicCellEMF.html set up Cu|Cu²⁺(0.001M)||Ag⁺(2.0M)|Ag cell

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. However, section 20.9 is required reading.
- <u>Electroplating</u>: 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
 - $\operatorname{Cu}^{2+}(aq) + 2 e^{-} \rightarrow \operatorname{Cu}(s)$
 - $\operatorname{Au}^{3+}(aq) + 3 e^{-} \rightarrow \operatorname{Au}(s)$
 - $\operatorname{Ag}^+(aq) + 1 e^- \rightarrow \operatorname{Ag}(s)$
- 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.