CHM 158 COURSE SYLLABUS
Lecture portion only – See separate lab syllabus
(Subject to revision)
Oakland University College of Arts and Sciences
Summer term 2005

I. Course Overview

A. Basic Information

Department: Chemistry
Course Prefix: CHM 158, Section 001 (CRN 31367)
Course Title: General Chemistry II
Credit Hours: 5

Class meets Mondays, Tuesdays, Wednesdays and Thursdays, 12:00 PM – 1:35 PM in Room 168 Science and Engineering Building (SEB) with laboratory sessions on Tuesdays and Thursday from 8:00 AM – 11:00 AM in Room 230 Hannah Hall of Science. No separate recitation sessions are scheduled during summer terms.

B. Instructors’ Information

Name: Professor Dagmar Cronn, Ph.D.
Office: Room 285, Science and Engineering Building (SEB)
Mailbox: 264 SEB under name of Cronn.
Phones: Office: (248) 370-4064  Home: (248) 693-7845
E-mail Address: cronn@oakland.edu
Office FAX number: (248) 370-2321 (address to Cronn to assure delivery)
Mailing Address: Chemistry Department, 285 SEB
Oakland University, Rochester, MI 48309-4477

Office Hours: Before and after class or by appointment. Feel free to leave me a message at either phone number or on my e-mail. Please do not call me at home after 9:00 PM.

C. Catalog Course Description

Integrated lecture-laboratory. Chemical reactions, kinetics, equilibrium, acid-base chemistry, thermodynamics and electrochemistry. Offered fall and winter [and summer].
Prerequisite: CHM 144 and 147 or 157.

Course Learning Outcomes:
By the end of this course you will be able to:
• Apply rate laws to problems of chemical kinetics
• Solve equilibrium problems
• Calculate acid-base equilibria
• Determine solubility and complex-ion equilibria
• Understand three laws of thermodynamics
• Describe electrochemical cells chemically and quantitatively

You will also be able to (a) recognize, understand, and be able to discuss basic chemical principals which are encountered in daily life, (b) become familiar with how chemists approach problems and systematically work toward solutions, and (c) become familiar with applying chemical principals toward problems encountered in other disciplines.


Optional material: Study Guide (available in book stores)

Grading
The lecture portion of this course will account for 80% of your final grade and the lab 20%. All students must register for the laboratory section. The exception is noted below. Performance on the lab experiments requires all of the skills of critical thinking including formulating questions, acquiring data, analyzing, reaching conclusions and noting any assumptions used.

Homework will be assigned but not collected or graded. Success on exams is related to completing homework.

Course Credit will be assigned as follows: (Note: there are three quizzes, three tests and a final exam.)

• Quizzes: 3 x 50 points each = 150 points (Scantron form 815 required)
• Tests: 3 x 100 points each = 300 points (Scantron form 882 required)
• Final exam: 1 x 200 points = 200 points (2 Scantron forms 882 required)

All exams, quizzes, and the final are multiple-choice questions.
The final exam consists of two parts:
The American Chemical Society standardized exam portions covered in this course
Additional questions to balance coverage of course content

NO make-up quizzes, tests, or final exam will be given.
NO lowest grades will be dropped.
NO extra credit is available.

For the course grade, a numerical grade between 1.0 and 4.0 will be assigned. The University Senate has approved publication of the following conversion for external purposes:
A: 3.6 to 4.0 B: 3.0 to 3.5 C: 2.0 to 2.9 D: 1.0 to 1.9 No credit: 0.0

Grading is on a sliding scale
The quizzes and exams are multiple-choice questions.
No partial credit
Your grades will be posted in the Grades section of the WebCT site including bar graphs of exam results.
Grading in lecture is on a modified curve

\[ \approx 90\% = 4.0 \]
\[ \approx 75\% = 3.0 \]
\[ \approx 60\% = 2.0 \]

1.0 requires 50\% of 4.0 points (typically 45\% of total points.)

Mid-semester evaluations for 100- and 200-level courses will be available through the Banner system on-line during the fourth week of the semester prior to the **last day to withdraw from classes (August 1 this term).** If no evaluation shows next to your name, it indicates that you are making satisfactory progress. You must continue to make satisfactory progress throughout the semester to ensure a passing grade. A “U” for unsatisfactory progress is intended to alert you to problems and to give you an opportunity to take steps to improve your performance.

**NOTE:** An I (incomplete) grade is only possible if a student is doing satisfactory work (2.0 or better). A signed agreement involving the instructor and student is required and must be completed before final grades are submitted. A grade of 2.0 or better is necessary for an S grade (if S/U grading is the chosen option).

**Regrading:** Test regrading requests should be submitted to the instructor. When inquiring about possible regrading, do not make any marks on the answer sheet.

**IF YOU ARE REPEATING THIS COURSE,** you may be able to waive the laboratory portion of the current course. You may choose to use the laboratory scores from the first course if: (1) your average in the laboratory portion of that course was at least 75\% and (2) you were enrolled in that course no more than three years ago. If you wish to apply for a laboratory waiver, you must complete a **General Chemistry Laboratory Waiver form.** These forms are available in the Department of Chemistry office, 260 Science and Engineering Building (SEB). This laboratory waiver form **must** be completed and returned to your current lecture instructor during the first week of the current semester. You must attend the laboratory portion of the course until the waiver is approved.

**Class attendance:** Attendance is strongly encouraged although not mandatory and you are responsible for all material presented in the lectures including changes made in class to the tentative schedule below. Cell phones are required to be turned off during class.

**Study Advice:** DO THE HOMEWORK PROBLEMS! This is the best way for you to learn the material. Keep up with the reading material since it is difficult to catch up once you get behind. A suggested study strategy: First, read the appropriate material, work the in-chapter exercises, study your lecture notes, and do a few problems with your notebook and/or book open for assistance. Then do as many problems as possible under test-taking conditions (closed book, limited time). After correcting yourself, ask yourself: "Why did I get this one and/or that one wrong?" If you can't answer that question, get help from the instructor (at office hours, or by appointment) or others. As perhaps might be obvious (but still worth stating), the worst way to do the problems is by consulting the answers while you are still working out the problems.
**Study Groups:** It is encouraged for students to seek a study partner or participate in study groups to help cement concepts through discussion of material.

**Classroom Courtesy:** The instructor of this course has a strong commitment to the development and maintenance of an instructional climate that supports respect for everyone in the classroom. Your enrollment in this course requires that you will treat your fellow classmates and course instructor with respect. The instructor reserves the right to adjust course grades for disrespectful behavior.

**Cheating:** A student will be referred to the Academic Conduct Committee for review if suspected of cheating. Examples of cheating include copying on exams, changing answers on exams after they are scored (scored answer forms are photocopied before being returned), having another person take an exam, or obtaining exam questions prior to the exam time. Students found guilty of academic misconduct face suspension or permanent dismissal from the university. For further details see the *Undergraduate Catalog*. You will earn a grade of zero for the course if you are caught cheating.

**Course Website:** There is a web site for this course. You will not be required to use the site although you will almost certainly find the site useful. Assignments will be available through handouts but other material, including class notes, will be available only on the web site. You may use the "Discussion" board to post questions and to receive answers.

Instructions for accessing the site are available online at the orientation site: [http://www2.oakland.edu/iti/orientation](http://www2.oakland.edu/iti/orientation)

The web site address (URL) is: [https://webct.oakland.edu](https://webct.oakland.edu)

When you click on "My Website," you will use your OU Grizzly student number for the User name and your six-digit SAIL Pin number as your password. (Your Pin number is most probably your birth date in the mmdyy format, unless you have previously changed your OU Pin number.) Warning: Your User Name and password are case sensitive. If you are unsure about your SAIL User ID and Pin, please contact the Registrar’s Office at registra@oakland.edu or (248) 370-3450.

The course syllabus is posted. The assignments are posted in the assignment area. The class notes are those I use during class. Some students choose to bring a printed copy of the notes to class so they can pay attention without having to make very many handwritten notes.

Note: Because so much of the material in this class requires tables, superscripts and subscripts, etc., some of the documents may be placed on the site using "rich text format." If you have trouble seeing some of the text materials, it may be because the computer or software you are using cannot translate this format. If you have access to both, try switching between Netscape and Explorer. Please let me know if you have trouble reading some files. I may be able to make them more accessible for you.

**Free tutoring** is available in the Academic Skills Center, 100 North Foundation Hall (370-4215).

**Closing the University:** If the University is officially closed on the day a quiz or test is scheduled, the quiz or test will be given during the next scheduled class meeting.
Calculators: Every student should have a calculator for both homework and exams. Arithmetic functions, square roots, logarithms, and exponentials will be most useful. Also, the calculator must be capable of scientific notation (e.g. $6.02 \times 10^{23}$). Calculators may be used (but not shared) for computational purposes during exams. NOTE: Small (hand-held) personal computers may not be used during examinations. If in doubt, ask instructor for clarification. Cell phones must be turned off and may not be on your person during quizzes and exams.

Students with Disabilities that affect their ability to participate fully in class or to meet all course requirements are encouraged to bring this to the attention of the instructor so that appropriate accommodations can be arranged. Further information is available from the Office of Disability Support Services located in 106 Foundation Hall: 370-3266 (voice); 370-3268 (TDD).

Media Resources: There are electronic resources associated with the textbook. The General Chemistry Interactive 5.0 Student CD assists in the visualization and explanation of concepts. The practice questions component may be particularly useful in learning the material through repetition. In addition, there are practice tests which are available to students through the Ebbing/Gammon website.

Lecture Schedule

Subject to revision; revisions will only be stated orally in class and posted on the WebCT site.
<table>
<thead>
<tr>
<th></th>
<th>MONDAY</th>
<th>TUESDAY</th>
<th>WEDNESDAY</th>
<th>THURSDAY</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>June</strong> 2006</td>
<td><strong>26</strong> Introduction &amp; Chapter 14</td>
<td><strong>27</strong> Chapter 14 Rates of Reactions</td>
<td><strong>28</strong> Chapter 14</td>
<td><strong>29</strong> Chapter 15 Chemical Equilibrium</td>
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<tr>
<td><strong>July</strong> 2006</td>
<td><strong>3</strong> Holiday No class</td>
<td><strong>4</strong> Holiday No class</td>
<td><strong>5</strong> QUIZ 1 Chapter 15</td>
<td><strong>6</strong> Chapter 15</td>
</tr>
<tr>
<td></td>
<td><strong>10</strong> TEST 1</td>
<td><strong>11</strong> Chapters 2, 4 &amp; 16</td>
<td><strong>12</strong> Chapters 16/17 Acids &amp; Bases</td>
<td><strong>13</strong> Chapters 16/17</td>
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<tr>
<td></td>
<td><strong>17</strong> Chapter 17 Acid/Base Equilibria</td>
<td><strong>18</strong> QUIZ 2 Chapter 17</td>
<td><strong>19</strong> Chapter 17</td>
<td><strong>20</strong> Chapter 17</td>
</tr>
<tr>
<td></td>
<td><strong>24</strong> TEST 2</td>
<td><strong>25</strong> Chapter 18 Solubility &amp; Complex Equilibria</td>
<td><strong>26</strong> Chapter 18</td>
<td><strong>27</strong> Chapter 18</td>
</tr>
<tr>
<td><strong>Aug</strong> 2006</td>
<td><strong>31</strong> QUIZ 3 Chapters 6 + 1</td>
<td><strong>1</strong>** Chapter 19 Thermodynamics &amp; Equilibria</td>
<td><strong>2</strong> Chapter 19</td>
<td><strong>3</strong> Chapter 20 Electrochemistry</td>
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<tr>
<td></td>
<td><strong>7</strong> TEST 3</td>
<td><strong>8</strong> Chapter 20</td>
<td><strong>9</strong> Chapter 20</td>
<td><strong>10</strong> Review</td>
</tr>
<tr>
<td></td>
<td><strong>15</strong> FINAL EXAM 12:00 – 3:00 PM</td>
<td></td>
<td></td>
<td><strong>14</strong></td>
</tr>
</tbody>
</table>

* Mid-semester grades will be provided prior to Tuesday, July 18

** August 1 is the last day to withdraw from any course.
There will be three lectures over the subject of rates of reaction. The following coverage can be correlated with the various sections of Chapter 14 from the Ebbing and Gammon textbook.

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
<tbody>
<tr>
<td>Monday, June 26</td>
<td>Introduction to the course. Introduction to rates of reaction; factors affecting rates. Sections 14.1 - 14.2</td>
<td>Chapter 14: All Exercises and Concept Checks</td>
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<tr>
<td></td>
<td></td>
<td>Problems: 1, 2, 23, 33, 35, 39</td>
</tr>
<tr>
<td>Tuesday, June 27</td>
<td>Rate dependence on concentration; concentration/time relations for 0, 1st, 2nd orders, half-life. Section 14.3 - Section 14.4</td>
<td>Problems: 7, 9, 25, 27, 31, 43, 45, 47, 49, 51, 53, 55, 57, 59, 61, 63, 93, 115</td>
</tr>
<tr>
<td>Wednesday, June 28</td>
<td>Collision theory; Arrhenius equation, rates laws and reaction mechanisms. Sections 14.5 - 14.9</td>
<td>Problems: 11, 14, 17, 18, 20, 75, 79, 87, 95, 97, 101, 109, 111</td>
</tr>
<tr>
<td>Wednesday, July 5</td>
<td>QUIZ 1</td>
<td>Scantron 815 required</td>
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<tr>
<td></td>
<td></td>
<td>Coverage: Chapter 14</td>
</tr>
</tbody>
</table>

**Learning Objectives**

Upon completion of the chapter, you should be able to:

- Relate the different ways of expressing reaction rates
- Calculate the average reaction rate
- Determine the order of reaction from the rate law
- Determine the rate law from initial rates
- Use the concentration-time equation for a first-order reaction
- Relate the half-life of a reaction to the rate constant
- Use the Arrhenius equation
- Write the overall chemical equation from a mechanism
- Determine the molecularity of an elementary reaction
- Write the rate equation for an elementary reaction
- Determine the rate law from a mechanism
There will be three lectures over the subject of chemical equilibrium covered in Chapter 15 from the Ebbing and Gammon text.

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
<tbody>
<tr>
<td>Thursday, June 29</td>
<td>Introduction to dynamic equilibrium; the equilibrium constant, $K_c$; forms of $K$. Sections 15.1, 15.2, 15.3</td>
<td>Chapter 15: All Exercises and Concept Checks</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Problems: 13, 21, 23, 25, 27, 29, 31</td>
</tr>
<tr>
<td>Wednesday, July 5</td>
<td>Magnitude of $K_c$; reaction quotient, $Q$; calculations with $K_c$; Sections 15.4, 15.5, 15.6</td>
<td>Problems: 33, 35, 37, 39, 47, 51, 53, 55, 57, 59, 61, 65</td>
</tr>
<tr>
<td>Thursday, July 6</td>
<td>LeChatelier’s Principle Section 15.7, 15.8, 15.9</td>
<td>Problems: 67, 69, 71, 73, 75, 79, 81, 83, 85, 91</td>
</tr>
<tr>
<td>Monday, July 10</td>
<td>TEST 1 Rates (Chapter 14): 5 questions Equilibrium (Chapter 15): 15 questions</td>
<td>Need: Scantron form 882 Calculator #2 pencil (See note below)</td>
</tr>
</tbody>
</table>

**NOTE:** The test will be available at 12:00 PM. You will have the entire class period in which to finish, although it is not likely you will need the entire period to complete the questions.

**Learning Objectives**

Upon completion of the chapter, you should be able to:

- Apply stoichiometry to an equilibrium mixture
- Write equilibrium-constant expressions
- Obtain an equilibrium constant from reaction composition
- Use the reaction quotient
- Obtain one equilibrium concentration given the others
- Solve equilibrium problems
- Apply Le Chatelier’s principle
There will be three lectures over the subject of acids and bases covered in Chapter 16 and Sections 17.1 through 17.4 of Chapter 17 from the Ebbing and Gammon text.

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
</table>
| Tuesday, July 11   | Chapters 2 and 4 review of Sections 2.8, 4.1, 4.2, 4.4, 4.7 Ionic compounds; ions in aqueous solution; net ionic reactions; introduction to acids/bases; molarity of aqueous solutions. Bronsted acids/bases*; conjugate pairs, amphoteric, conjugate acid/base strength; strong acids Ch. 16: Sections 1, 2, 4, 6, 7, 8 | Chapter 2: 111, 113, 115; Chapter 4: 27, 29, 35  
Chapter 16: All Exercises. And Concept Checks  
Handout Problems: 1, 2, 3, 4 (next pages) |
| Wednesday, July 12 | Weak acids and weak bases  
Chapter 17: Sections 1, 3 | Chapter 16 Problems: 75, 83, 95  
Chapter 17: All Exercises and Concept Checks  
Chapter 17 Problems: 27, 29, 31, 33, 35, 43, 45, 47, 87, 91, 93, 107  
Handout Problems: 5, 6 |
| Thursday, July 13  | Polyprotic acids; pH of salt solutions (hydrolysis) Ch. 17: Sections 2, 4 | Chapter 17: Problems: 41, 49, 51, 53, 57, 59, 115, 117 |
| Monday, July 17    | Common ion effect Ch. 17: Section 5 | Chapter 17 Problems: 61, 63, 65, 67, 91, 95, 105 |
| Tuesday, July 18   | **QUIZ 2**  
Chapter 16 and Sections 17.1 – 17.4 | Need: Scantron form 815  
Calculator  
#2 pencil |

*If you plan to take organic chemistry courses, you will need to know the definition and uses of Lewis acids and bases. You should cover the material in Section 16.3 of the chapter and work Problems 31, 33, 35 and 37. This material will not be covered in class, nor will it be tested.

**Learning Objectives**

Upon completion of the material in Chapter 16 & Section 1 - 4 of Chapter 17, you should be able to:

- **Identify acid and base species**
- Decide whether reactants or products are favored in acid-base reactions
• Calculate concentrations of H$_3$O$^+$ and OH$^-$ in solutions of a strong acid or base
• Calculate the pH from the hydronium concentration and vice versa
• Determine Ka (or Kb) from the solution pH
• Calculate concentrations of species in a weak acid solution using Ka
• Calculate concentrations of species in a weak base solution using Kb
• Predict whether a salt solution is acidic, basic, or neutral
• Obtain Ka from Kb or Kb from Ka
• Calculate concentrations of species in a salt solution
• Calculate the pH at the equivalence point in the titration of a weak acid by a strong base
CHM 158

Chapters 16 and 17.1 - 17.2
Handout Problems:

1. Write balanced equations showing how the HPO_{4}^{2-} ion of sodium monohydrogen phosphate, Na_{2}HPO_{4}, can be a Brønsted acid or a Brønsted base.

2. Several acids are listed here with their respective equilibrium constants:

- HF(aq) + H_{2}O(l) ⇌ H_{3}O^{+}(aq) + F^{-}(aq) \quad K_{a} = 7.2 \times 10^{-4}.
- HPO_{4}^{2-}(aq) + H_{2}O(l) ⇌ H_{3}O^{+}(aq) + PO_{4}^{3-}(aq) \quad K_{a} = 3.6 \times 10^{-13}.
- CH_{3}CO_{2}H(aq) + H_{2}O(l) ⇌ H_{3}O^{+}(aq) + CH_{3}CO_{2}^{-}(aq) \quad K_{a} = 1.8 \times 10^{-5}.

a. Which is the strongest acid? Which is the weakest?
b. What is the conjugate base of the acid HF?
c. Which acid has the weakest conjugate base?
d. Which acid has the strongest conjugate base?

3. Several bases are listed here with their respective Kb values:

- NH_{3}(aq) + H_{2}O(l) ⇌ NH_{4}^{+}(aq) + OH^{-}(aq) \quad K_{b} = 1.8 \times 10^{-5}.
- C_{5}H_{5}N(aq) + H_{2}O(l) ⇌ C_{5}H_{5}NH^{+}(aq) + OH^{-}(aq) \quad K_{b} = 1.5 \times 10^{-9}.
- N_{2}H_{4}(aq) + H_{2}O(l) ⇌ N_{2}H_{5}^{+}(aq) + OH^{-}(aq) \quad K_{b} = 8.5 \times 10^{-7}.

a. Which is the strongest base? Which is the weakest?
b. What is the conjugate acid of C_{5}H_{5}N?
c. Which base has the strongest conjugate acid? Which has the weakest?

4. State which of the following ions or compounds has the strongest conjugate base and briefly explain your choice:
   a. HSO_{4}^{-}
   b. CH_{3}CO_{2}H
   c. HOCl

5. Ammonium chloride and sodium dihydrogen phosphate, NaH_{2}PO_{4}, are mixed in water. Using Table 17.4 or the tables in your text (See class handout for Table 17.4), write a balanced equation for the acid-base reaction that could, in principle, occur. Does the reaction occur to a significant extent?

6. For each reaction below, predict whether the equilibrium lies predominantly to the left or to the right. Explain your prediction briefly.
   a. NH_{4}^{+}(aq) + Br^{-}(aq) ⇌ NH_{3}(aq) + HBr(aq)
   b. HPO_{4}^{2-}(aq) + CH_{3}CO_{2}^{-}(aq) ⇌ PO_{4}^{3-}(aq) + CH_{3}CO_{2}H(aq)
   c. NH_{2}^{-}(aq) + H_{2}O(l) ⇌ NH_{3}(aq) + OH^{-}(aq)
   d. Fe(H_{2}O)_{6}^{3+}(aq) + HCO_{3}^{-}(aq) ⇌ Fe(H_{2}O)_{5}(OH)^{2-}(aq) + H_{2}CO_{3}(aq)
Answers to handout problems:

1. As an acid: \( \text{HPO}_4^{2-}(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{PO}_4^{3-}(aq) \)
   As a base: \( \text{HPO}_4^{2-}(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{OH}^-(aq) + \text{H}_2\text{PO}_4^-(aq) \)

2. 
   a. HF is the strongest acid, and \( \text{HPO}_4^{2-} \) is the weakest acid;
   b. \( \text{F}^- \);
   c. HF;
   d. \( \text{HPO}_4^{2-} \).

3. 
   a. \( \text{NH}_3 \) is the strongest base, \( \text{C}_5\text{H}_5\text{N} \) is the weakest.
   b. \( \text{C}_5\text{H}_5\text{NH}^+ \).
   c. \( \text{C}_5\text{H}_5\text{N} \) has the strongest conjugate acid; \( \text{NH}_3 \) has the weakest conjugate.

4. Answer is c. HOCl, the weakest acid in this list (Table 17.4), has the strongest conjugate base.

5. \( \text{NH}_4^+(aq) + \text{H}_2\text{PO}_4^-(aq) \leftrightarrow \text{NH}_3(aq) + \text{H}_3\text{PO}_4(aq) \)
   The reaction does not occur to a significant extent. Equilibria favor the weaker acid and base, in this case, the reactants.

6. Each equilibrium favors the weaker acid and base. The equilibrium positions are therefore:
   a. left
   b. left
   c. left
   d. right
1. For each of the following:
   a) Write net ionic equations for all acid/base reactions that take place when 1.0 mol of the
      compound is dissolved in 1.0 L of water; and
   b) Give a rough estimate of the concentration of each species that is present in the resulting
      solution.

   1) NaCN
   2) NH₄I
   3) KNO₃
   4) HNO₂
   5) KHSO₃
   6) Ba(OH)₂
   7) Ni(ClO₄)₂
   8) Na₂HPO₄
   9) H₃PO₄
   10) NH₃
   11) NH₄F

2. For each of the following:
   a) Write a net ionic equation for the additional acid/base reaction that takes place upon
      mixing 1.0 L of each solution (if there is a reaction), and
   b) State whether the solution after mixing is acidic, basic, or neutral.

   1) 1M K₂CO₃ and 1M HBr
   2) 1M LiOH and 1M HNO₃
   3) 1 M NH₄NO₃ and 1M NaOH
   4) 1 M KH₂PO₄ and 1M KOH
   5) 1 M HCN and 1M KOH
   6) 1M HClO₄ and 1 M NaHCO₃
   7) 1 M LiBr and 1 M HI
   8) 1M NaC₂H₅O₂ and 1M HCl
Solutions
Observations: sl = slightly less than; vs = very small

1.

a)  
1) \( \text{CN}^- (aq) + H_2O(l) \rightarrow HCN(aq) + OH^-(aq) \)
2) \( \text{NH}_4^+ (aq) + H_2O(l) \rightarrow NH_3(aq) + H_3O^+(aq) \)
3) No net ionic acid/base reaction takes place.
4) \( \text{HNO}_2(aq) + H_2O(l) \rightarrow NO_2^-(aq) + H_3O^+(aq) \)
5) \( \text{HSO}_3^-(aq) + H_2O(l) \rightarrow SO_3^{2-}(aq) + H_3O^+(aq) \quad K_a(\text{HSO}_3^-) = 6.3 \times 10^{-8} \)
6) No net ionic reaction takes place.
7) \( \text{Ni(H}_2\text{O})_4^{2+}(aq) + H_2O(l) \rightarrow \text{Ni(H}_2\text{O})_3\text{OH}^+(aq) + H_3O^+(aq) \)
8) \( \text{HPO}_4^{2-}(aq) + H_2O(l) \rightarrow H_2\text{PO}_4^-(aq) + OH^-(aq) \quad K_a(\text{HPO}_4^{2-}) = 1.6 \times 10^{-7} \)
9) \( \text{H}_2\text{PO}_4^-(aq) + H_2O(l) \rightarrow \text{PO}_4^{3-}(aq) + H_3O^+(aq) \quad K_a(\text{HPO}_4^{2-}) = 4.8 \times 10^{-13} \)
10) \( \text{F}^-(aq) + H_2O(l) \rightarrow HF(aq) + OH^-(aq) \quad K_b(F^-) = 1.5 \times 10^{-11} \)

1.

b)  
1) \([\text{Na}^+] = 1.0 \, \text{M}; \, [\text{CN}^-] \, \text{sl 1.0 M}; \, [\text{HCN}] = [\text{OH}^-] \, \text{vs}; \, [\text{OH}^-] > [H_3O^+]; \) basic solution
2) \([\text{I}^-] = 1.0 \, \text{M}; \, [\text{NH}_4^+] \, \text{sl 1.0 M}; \, [\text{NH}_3] = [H_3O^+] \, \text{vs}; \, [H_3O^+] > [OH^-]; \) acidic solution
3) \([\text{K}^+] = [\text{NO}_3^-] = 1.0 \, \text{M}; \) neutral solution
4) \([\text{HNO}_2] \, \text{sl 1.0 M}; \, [\text{NO}_2^-] = [H_3O^+] \, \text{vs}; \, [H_3O^+] > [OH^-]; \) acidic solution
5) \([\text{K}^+] = 1.0 \, \text{M}; \, [\text{HSO}_3^-] = \text{sl 1.0 M}; \, [\text{SO}_3^{2-}] = [H_2\text{SO}_3] \, \text{vs}; \, [H_3O^+] > [OH^-]; \) acidic solution
6) \([\text{Ba}^{2+}] = 1.0 \, \text{M}; \, [\text{OH}] = 2.0 \, \text{M}; \, [H_2O] = 5.0 \times 10^{-15}; \) basic solution
7) \([\text{ClO}_4^-] = 2.0 \, \text{M}; \, [\text{Ni(H}_2\text{O})_4^{2+}] \, \text{sl 1.0 M}; \, [\text{Ni(H}_2\text{O})_3\text{OH}^+] = [H_3O^+] \, \text{vs}; \, [H_3O^+] > [OH^-]; \) acidic solution
8) \([\text{Na}^+] = 2.0 \, \text{M}; \, [\text{HPO}_4^{2-}] \, \text{sl 1.0 M}; \, [\text{PO}_4^{3-}] = [H_2\text{PO}_4^-] \, \text{vs}; \, [OH^-] > [H_3O^+]; \) basic solution
   Also very, very small amount of H_3PO_4 due to: \( H_3\text{PO}_4(aq) + H_2O(l) \rightarrow H_3\text{PO}_4 + OH^-(aq) \)
9) \([\text{H}_2\text{PO}_4] \, \text{sl 1.0 M}; \, [\text{H}_2\text{PO}_4^-] \approx [H_3O^+] \, \text{vs}; \, [H_3O^+] > [OH^-]; \) acidic solution
   Also very, very small amounts of HPO_4^{2-} and PO_4^{3-}.
10) \([\text{NH}_3] \, \text{sl 1.0 M}; \, [\text{NH}_4^+] = [\text{OH}^-] \, \text{vs}; \, [OH^-] > [H_3O^+]; \) basic solution
11) \([\text{NH}_4^+] \, \text{sl 1.0 M}; \, [\text{F}^-] \, \text{sl 1.0 M}; \, [\text{NH}_3] = [HF] = [H_3O^+] = [OH^-] \, \text{vs}; \, [H_3O^+] > [OH^-]; \) acidic solution

2.

sl = slightly less than

a)  
1) \([\text{K}^+] = 1.0 \, \text{M (Mix 1 l + 1 l = 2 l, 2 mol K'/2 l = 1.0 M)}\)
   \([\text{Br}^-] = 0.5 \, \text{M (1 mol Br'/2 l = 0.5 M)}\)
   \( \text{CO}_3^{2-} + H_3O^+ \leftrightarrow HCO_3^- + H_2O \quad K >> 1. : [\text{CO}_3^{2-}] \, \text{sl 0.50 M} \)
If all of the carbonate ion initially present is converted to bicarbonate ion, then the following two reactions take place to a small extent.

\[
\text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{CO}_3^{2-} \quad K = K_a(\text{HCO}_3^-)
\]

\[
\text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{OH}^- + \text{H}_2\text{CO}_3 \quad K = K_b(\text{HCO}_3^-)
\]

Since \( K_b(\text{HCO}_3^-) > K_a(\text{HCO}_3^-) \), the solution is basic.

2) \([\text{Li}^+] = [\text{NO}_3^-] = 0.5 \text{ M} \) (1 mol/2 l = 0.5 M)

\[
\text{H}_3\text{O}^+ + \text{OH}^- \leftrightarrow 2 \text{H}_2\text{O} \quad K >> 1 \\ [\text{OH}^-] = [\text{H}_3\text{O}^+] = 10^{-7} \text{ M}
\]

Solution is neutral

3) \([\text{Na}^+] = [\text{NO}_3^-] = 0.5 \text{ M} \) (1 mol/2 l = 0.5 M)

\[
\text{NH}_4^+ + \text{OH}^- \leftrightarrow \text{H}_2\text{O} + \text{NH}_3 \quad K >> 1 \\ \therefore [\text{NH}_4^+] \text{ sl} \ 0.50 \text{ M}
\]

If all of the ammonium ion initially present is converted to ammonia, then the following reaction takes place to a small extent.

\[
\text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{NH}_4^+ + \text{OH}^- \quad K = K_b(\text{NH}_3) \quad \text{Solution is basic.}
\]

4) \([\text{K}^+] = 1.0 \text{ M} \) (Mix 1 l + 1 l = 2 l, 2 mol K+/2 l = 1.0 M) (Potassium ion comes from both 1.0 M solutions.)

\[
\text{H}_2\text{PO}_4^- + \text{OH}^- \leftrightarrow \text{HPO}_4^{2-} + \text{H}_2\text{O} \quad K >> 1 \\ \therefore [\text{HPO}_4^{2-}] \text{ sl} \ 0.50 \text{ M}
\]

If all the \( \text{H}_2\text{PO}_4^- \) initially present is converted to \( \text{HPO}_4^{2-} \), then the following two reactions each take place.

\[
\text{HPO}_4^{2-} + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{PO}_4^- + \text{OH}^- \quad K = K_b(\text{HPO}_4^{2-})
\]

\[
\text{HPO}_4^{2-} + \text{H}_2\text{O} \leftrightarrow \text{PO}_4^{3-} + \text{H}_3\text{O}^+ \quad K = K_a(\text{HPO}_4^{2-})
\]

Since \( K_b(\text{HPO}_4^{2-}) > K_a(\text{HPO}_4^{2-}) \), solution is basic.

5) \([\text{K}^+] = 0.5 \text{ M} \) (Mix 1 l + 1 l = 2 l, 1 mol K+/2 l = 0.5 M)

\[
\text{HCN} + \text{OH}^- \leftrightarrow \text{H}_2\text{O} + \text{CN}^- \quad K >> 1 \\ \therefore [\text{CN}^-] \text{ sl} \ 0.50 \text{ M}
\]

If all the HCN initially present is converted to CN\(^-\), then the following reaction takes place to a small extent.

\[
\text{H}_2\text{O} + \text{CN}^- \leftrightarrow \text{HCN} + \text{OH}^- \quad K = K_b(\text{CN}^-)
\]

Solution is basic.

6) \([\text{Na}^+] = [\text{ClO}_4^-] = 0.50 \text{ M} \)

\[
\text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{CO}_3 + \text{H}_2\text{O} \quad K >> 1 \\ [\text{H}_2\text{CO}_3] \text{ sl} \ 0.50 \text{ M}
\]

If all the bicarbonate initially present is converted to carbonic acid, then the following reaction takes place to a small extent.

\[
\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \leftrightarrow \text{HCO}_3^- + \text{H}_3\text{O}^+ \quad K = K_a(\text{H}_2\text{CO}_3)
\]

Solution is acidic.

7) \([\text{Li}^+] = [\text{I}^-] = 0.5 \text{ M} \)

Also \([\text{H}_2\text{O}^+] = [\text{Br}^-] = 0.5 \text{ M} \)

This is because there is no reaction between hydronium ion and either iodide ion or bromide ion since both halogen ions are the “conjugate bases” of strong acids and have no basic character.

Solution is acidic.

8) \([\text{Na}^+] = [\text{Cl}^-] = 0.50 \text{ M} \)

\[
\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \leftrightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \quad K >> 1 \\ [\text{HC}_2\text{H}_3\text{O}_2] \text{ sl} \ 0.50 \text{ M}
\]

If all of the acetate ion initially present is converted to acetic acid, then the following reaction takes place to a small extent.

\[
\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \leftrightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_3\text{O}^+ \quad K = K_a(\text{HC}_2\text{H}_3\text{O}_2) \quad \text{Solution is acidic.}
\]

Chemistry 158

Acids and Bases (Part II)  Cronn

There will be four more lectures continuing the subject of acids and bases covered in Chapter 17 from the Ebbing and Gammon text.

**Omit Section 16.4**

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
<tbody>
<tr>
<td>Monday, July 17</td>
<td>Chapter 17: Section 5 Common Ion Effect</td>
<td>Chapter 17: All Exercises and Concept Checks</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Problems: 63, 65, 67, 91, 95, 105</td>
</tr>
<tr>
<td>Tuesday, July 18</td>
<td>Chapter 17: Sections 5 (contd.) Common Ion Effect</td>
<td>Chapter 17: See problems from class notes and accompanying class handouts and additional problems 1 – 6 on following pages.</td>
</tr>
<tr>
<td>Wednesday, July 19</td>
<td>Chapter 17: Section 6 Buffers</td>
<td>Chapter 17: 69, 71, 73, 75, 77, 95, 97</td>
</tr>
<tr>
<td>Thursday, July 20</td>
<td>(Review Acid-Base Titrations in Chapter 4, Section 10, pages 161 - 163)</td>
<td>Chapter 17: 79, 81, 83, 85, 99, 113, 115, 117, 119</td>
</tr>
<tr>
<td>Monday, July 24</td>
<td><strong>TEST 2</strong> Chapters 16 and 17</td>
<td>Need: Scantron form 882 Calculator #2 pencil</td>
</tr>
</tbody>
</table>

**Learning Objectives**

Upon completion of the chapter, you should be able to:

- Calculate the common-ion effect on acid ionization
- Calculate the pH of a buffer from given volumes of solution
- Calculate the pH of a buffer when a strong acid or strong base is added
- Calculate the pH of a solution of a strong acid and a strong base
CHM 158  
Chapter 17 Additional Problems

1. For each of the following cases, decide whether the pH is less than 7, equal to 7, or greater than 7.
   a) Equal volumes of 0.10 M acetic acid, CH₃CO₂H, and 0.10 M KOH are mixed.
   b) 25 mL of 0.015-M NH₃ is mixed with 25 mL of 0.015-M HCl.
   c) 150 mL of 0.20-M HNO₃ is mixed with 75 mL of 0.40-M NaOH

2. Does the pH of the solution increase, decrease, or stay the same when you
   a) Add solid ammonium chloride to a dilute aqueous solution of NH₃?
   b) Add solid sodium acetate to a dilute aqueous solution of acetic acid?
   c) Add solid NaCl to a dilute aqueous solution of NaOH?

3. Which of the following combinations would be the best to buffer a solution at a pH of approximately 9?
   a) HCl/NaCl
   b) Ammonia/ammonium nitrate (NH₃/NH₄Cl)
   c) Acetic acid/sodium acetate (CH₃CO₂H/NaCH₃CO₂)

4. Does the pH of the solution increase, decrease, or stay the same when you
   a) Add 10.0 mL of 0.10-M HCl to 25.0 mL of 0.10-M NH₃?
   b) Add 25.0 mL of 0.050-M NaOH to 50.0 mL of 0.050-M acetic acid?

5. You have 0.30 mol of K₂CO₃ available. You could form a buffer by dissolving all the K₂CO₃ in 1.0 L of water and then dissolving
   a) 0.20 mol of KHCO₃
   b) 0.20 HCl
   c) 0.40 mol KHCO₃
   d) a or b
   e) a or b or c

6. You require 36.78 mL of 0.0105 M HCl to reach the equivalence point in the titration of 25.0 mL of aqueous ammonia. What are the concentrations of H₃O⁺, OH⁻, and NH₄⁺ at the equivalence point? What is the pH of the solution at the equivalence point? What was the concentration of NH₃ in the original ammonia solution?

Problem Answers
1. a) Greater than 7  b) Less than 7  c) Equal to 7
2. a) Decrease  b) Increase  c) Stay the same
3. b
4. a) Decrease  b) Increase
5. e
6. [NH₄⁺] = 6.25 x 10⁻³ M; [H₃O⁺] = 1.85 x 10⁻⁶ M; [OH⁻] = 5.41 x 10⁻⁹ M; pH = 5.73
   Concentration of original NH₃ solution: 1.54 x 10⁻² M.
Chemistry 158  Precipitation Reactions  Cronn

There will be three lectures on precipitation reactions covered in Chapter 18 from the Ebbing and Gammon text.

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
<tbody>
<tr>
<td>Tuesday, July 25</td>
<td>Chapter 18: Sections 1, 3</td>
<td>Chapter 18: All Exercises and Concept Checks</td>
</tr>
<tr>
<td></td>
<td>Solubility and $K_{sp}$; precipitation</td>
<td>Problems: 21, 23, 25, 27, 29, 31, 39, 41, 43, 45, 47, 49</td>
</tr>
<tr>
<td>Wednesday, July 26</td>
<td>Chapter 18: Sections 2, 3, 4</td>
<td>Supplemental Problems (See next pages)</td>
</tr>
<tr>
<td></td>
<td>Precipitation calculations; common ion effect; fractional precipitation; effect of pH on solubility; simultaneous equilibria</td>
<td>Problems: 33, 35, 51, 53, 73, 75, 77, 79, 81, 89</td>
</tr>
<tr>
<td>Thursday, July 27</td>
<td>Chapter 18: Section 5, 6 (Skip Section 18.7.)</td>
<td>Problem: 55, 57, 59</td>
</tr>
<tr>
<td>Monday, July 31</td>
<td>Quiz 3</td>
<td>Chapter 18</td>
</tr>
</tbody>
</table>

**Learning Objectives**

Upon completion of the chapter, you should be able to:

- Write solubility product expressions
- Calculate $K_{sp}$ from the solubility, or vice versa
- Calculate the solubility of a slightly soluble salt in a solution of a common ion
- Predict whether precipitation will occur
- Determine the qualitative effect of pH on solubility
- Calculate the concentration of a metal ion in equilibrium with a complex ion
- Predict whether a precipitate will form in the presence of the complex ion
- Calculate the solubility of a slightly soluble ionic compound in a solution of the complex ion
Supplemental Homework Problems for Chapter 18
Simultaneous Equilibria: Solubility in Acids

Calculate the solubility of CuCN in a buffer with pH = 4. Use $K_{sp} = 3.2 \times 10^{-20}$ and $K_a(HCN) = 4.0 \times 10^{-10}$.
Answer: $8.9 \times 10^{-8}$ M.

Precipitation Reactions

Which of the following:
- Na$_2$CO$_3$, PbI$_2$, Cu(OH)$_2$, AgBr, CaCO$_3$, Pb(NO$_3$)$_2$

a) will dissolve in water?
b) will dissolve in acid?
c) will dissolve in NH$_3$?
d) will dissolve in both acid and NH$_3$?

List two ways in which you could get more Ag$_2$CO$_3$ to dissolve in aqueous solution than the amount that will dissolve in pure water.
Ag$_2$CO$_3$ $\leftrightarrow$ 2 Ag$^+$ (aq) + CO$_3$$^-$(aq); $K_{sp} = [Ag^+]^2[CO_3^-]$

Fractional Precipitation

1. Magnesium ions are slowly added to a solution, which has a carbonate ion concentration of 0.005 M.
   a. What is the minimum concentration of magnesium ions at which a ppt will start to form?
   b. If the magnesium ion concentration is adjusted to 0.05 M, what is the carbonate ion concentration in the solution?
   c. When the [Mg$^{2+}$] = 0.05 M, what percent of the carbonate ions remain dissolved in solution?
   d. When the [Mg$^{2+}$] = 0.05 M, what percent of the carbonate ions have been removed from solution?

2. An aqueous solution contains 0.10 M Ca$^{2+}$ and 0.10 M Ba$^{2+}$. An aqueous solution of Na$_2$CO$_3$ is slowly added to this solution.
   a. Which ion will precipitate out of solution first? Why?
   b. What is the calcium concentration when the second ion begins to precipitate?
   c. What percent of the first ion (see your answer to part a) remains in solution when the second first starts to precipitate?

3. A solution is prepared by dissolving 3.3 g of Pb(NO$_3$)$_2$ in 1.0 liter of water (Assume total solution volume is 1 l).
   a. What is the lead ion concentration in this solution?
   b. Calculate the mass of NaCl, which must be added to this solution to remove 99% of the lead ions from solution.

Answers to Fractional Precipitation Problems

1. a. $8 \times 10^{-3}$ M  
   b. $8 \times 10^{-4}$M  
   c. 16%
   d. 84%

2. a. CaCO$_3$ will ppt first because $K_{sp}$(BaCO$_3$) is greater than $K_{sp}$(CaCO$_3$).
   b. 0.047 M  
   c. 47%

3. a. 0.010 M  
   b. 24 g
There will be three lectures on thermodynamics covered in Chapter 19 from Ebbing and Gammon including a review of Chapter 6, Sections 6.3, 4, 7 and 8.

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
</table>
| Monday, July 31 | Enthalpy; thermochemical equations; Hess’s Law; standard enthalpy of formation  
Chapter 6 review Sections 3, 4, 7, & 8  
First and second Laws of Thermodynamics, entropy, standard entropies Chapter 19: Sections 1, 2, 3 | Chapter 6 Problems: 47, 49, 53, 65, 67, 69, 71, 73, 103  
Chapter 19: All Exercises and Concept Checks  
Chapter 19 Problems: 29, 33, 35(a, b), 37, 39, 41(a,b), 43 |
| Tuesday, August 1 | Free energy and spontaneity  
Chapter 19: Sections 1, 2, 3, 4, 5 | Additional Problems 1-3 (See next page). |
| Wednesday, August 2 | Relating free energy to K and T  
Chapter 19: Sections 4, 5, 6, 7 | Chapter 19 Problems: 53, 57, 59, 61, 77, 81 |
| Thursday, August 3 | Chapter 20 |  |
| Monday, August 7 | **Test 3**  
Chapters 18 and 19 | Need: Scantron form 882  
Calculator  
#2 pencil |

**Learning Objectives**

Upon completion of the chapter, you should be able to:

- Calculate the entropy change for a phase transition
- Predict the sign of the entropy change of a reaction
- Calculate $\Delta H^\circ$ and/or $\Delta S^\circ$ for a reaction
- Calculate $\Delta G^\circ$ from $\Delta H^\circ$ and $\Delta S^\circ$
- Calculate $\Delta G^\circ$ from standard free energies of formation
- Interpret the sign of $\Delta G^\circ$
- Write the expression for a thermodynamic equilibrium constant
- Calculate $K$ from the standard free-energy change
- Calculate $\Delta G^\circ$ and $K$ at various temperatures
Additional Problems for Chapter 19:

1. Find the sign of $\Delta S^\circ$ for the reaction
   \[ 2 \text{N}_2\text{O}_5(s) \rightarrow 4 \text{NO}_2(g) + \text{O}_2(g) \]
   The reaction is endothermic and spontaneous at 25°C. Explain the spontaneity of the reaction in terms of enthalpy and entropy changes.

2. The reaction
   \[ \text{CO}_2(g) + \text{H}_2(g) \rightarrow \text{CO}(g) + \text{H}_2\text{O}(g) \]
   is non-spontaneous at room temperature but becomes spontaneous at a much higher temperature. What can you conclude from this about the signs of $\Delta H^\circ$ and $\Delta S^\circ$, assuming that the enthalpy and entropy changes are not greatly affected by the temperature change? Explain your reasoning.

3. Consider the reaction
   \[ \text{CS}_2(g) + 4 \text{H}_2(g) \rightarrow \text{CH}_4(g) + 2 \text{H}_2\text{S(g)} \]
   Calculate $\Delta H^\circ$, $\Delta S^\circ$, and $\Delta G^\circ$ at 25°C for this reaction. Assume $\Delta H^\circ$ and $\Delta S^\circ$ are constant with respect to a change of temperature. Now calculate $\Delta G^\circ$ at 650°C. Compare the two values of $\Delta G^\circ$. Briefly discuss the spontaneity of the reaction at 25°C and 650°C.
There will be three lectures on electrochemistry covered in Chapter 20 from the Ebbing and Gammon text, including a review of Chapter 4, Section 5.

<table>
<thead>
<tr>
<th>Days</th>
<th>Topics</th>
<th>Homework after lecture</th>
</tr>
</thead>
<tbody>
<tr>
<td>Thursday, August 3</td>
<td>Oxidation-Reduction reactions Chapter 4: Section 5 (Review)</td>
<td>Chapter 4: Review Problem 47a,b, 49a,d, 53, 55</td>
</tr>
<tr>
<td></td>
<td>Construction and notation for voltaic cells Chapter 20: Sections 1, 2, 3, 4</td>
<td>Chapter 20: All appropriate Practice Exercises.</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Chapter 20 Problems: 35, 39, 41, 43, 47</td>
</tr>
<tr>
<td>Tuesday, August 8</td>
<td>Standard electrode potentials emf; equilibrium constants from emf; change of emf with concentration Chapter 20: Sections 5, 6, 7</td>
<td>Additional Problems 1 &amp; 2 (See next page.)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Chapter 20 Problems: 53, 55, 57, 61, 65, 73, 77, 81, 19, 99</td>
</tr>
<tr>
<td>Wednesday, August 9</td>
<td>Electrolytic cells Chapter 20: Sections 8, 9, 10, 11</td>
<td>Chapter 20 Problems: 83, 85, 87, 89, 113, 117</td>
</tr>
<tr>
<td>Thursday, August 10</td>
<td>Chapter 20 (completed)</td>
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<tr>
<td></td>
<td>Review for final if time permits</td>
<td></td>
</tr>
<tr>
<td>Tuesday, August 15</td>
<td><strong>Final Exam 12:00 pm – 3:00 pm</strong></td>
<td>Need Scantrons 882 and 812, #2 pencil and calculator</td>
</tr>
<tr>
<td></td>
<td>ACS Standardized exam</td>
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<td>Plus extra questions as needed to accurately reflect the course material covered</td>
<td></td>
</tr>
</tbody>
</table>
Learning Objectives

Upon completion of the chapter, you should be able to:

- Sketch and label a voltaic cell
- Write the cell reaction from the cell notation
- Calculate the quantity of work from a given amount of cell reactant
- Determine the relative strengths of oxidizing and reducing reagents
- Determine the direction of spontaneity from electrode potentials
- Calculate the free-energy change from electrode potentials
- Calculate the cell emf from free-energy change
- **Calculate the equilibrium constant from cell emf**

Additional Problems for Chapter 20

1. Under standard conditions, identify a chemical species that can oxidize Cd to Cd^{2+}(aq), but cannot oxidize Sn to Sn^{2+}(aq).

2. Under standard conditions, identify a chemical species that can reduce Ag^{+}(aq) to Ag, but cannot reduce I_{2}(s) to I^{-}(aq).