

PURDUE UNIVERSITY

CHM II STUDY GUIDE (Chemistry 11600/12400)

This study guide describes the topics to be mastered prior to attempting the examination for the second semester of chemistry at Purdue University. The material covered can be found in current textbooks used for CHM 11500/11600, CHM 12300/12400 at Purdue University.

Suggested Textbooks

Chemistry & Chemical Reactivity, 7th ed by Kotz, Treichel & Townsend; Thomson/Brooks/Cole. ISBN-13: 978-0495390299.

Chemistry: The Molecular Nature of Matter and Change, 5th edition by Silberberg; McGraw-Hill. ISBN-10: 0077216504; ISBN-13: 978-0077216504.

Chemistry, the Central Science by Brown, 11th edition by LeMay, Bursten & Murphy; Prentice-Hall. ISBN-13: 978-0136006176.

Texts can be purchased at Local Book Stores:

- University Book Store, 360 W. State Street, W. Lafayette, IN 47906.
- University Book Store, 720 Northwestern, W. Lafayette, IN 47906.
- Follett's Purdue West Book Store, 1400 W. State Street, W. Lafayette, IN 47906.

IMPORTANT!

1. Read this material thoroughly if you contemplate trying for advanced placement (and extra credit which counts toward graduation).
2. Study all the material listed in the outline.
3. Work many practice problems.
4. When you feel prepared for it, take the sample examination.
5. Come to the actual examination rested and confident.

Test Topics/Preparation for Exam:

The subject matter of General Chemistry II is varied. It usually deals with the following topics: solutions, reduction-oxidation reactions, acids and bases, electrochemistry, chemical kinetics, chemical equilibrium (e.g., gas-phase, acid-base, solubility and complex ion equilibria), and chemical thermodynamics. Many of these topics are related, such as electrochemistry and thermodynamics. Other relationships will become apparent as you study the concepts presented in this outline.

You should study the topics listed in the attached outline prior to attempting the simple examination included with this study guide. These topics are categorized to be consistent with the content studied during the past several years in the various Chemistry II courses offered at Purdue.

In preparing for the examination, it is important to work many problems. The problems should be used to measure your understanding of the concepts and principles involved. In the section dealing with electrochemistry, for example, you must understand the relationship that exists between the positions of various reduction-oxidation half reactions in an activity series and the reactions occurring at the anode and cathode of an electrochemical cell. These relationships will allow you to predict which electrode in the cell will be the cathode, what the reaction would be at the cathode, and the direction of electron flow in the cell.

At the end of this study guide you will find a sample examination over this material. Allow yourself 60 minutes to complete this sample exam. Naturally, it does not cover every topic, or every aspect of a topic. No examination extending over a reasonable time period could do that. However, if you have no difficulty with the sample examination, you should not have difficulty with the actual one. Following the sample exam are the answers to the questions.

A Word of Advice

No one does well on an examination when he/she is excessively fatigued; therefore, you are urged to provide yourself with an adequate rest period before taking the actual examination. If your trip to campus necessitates travel into the late hours of the night or an extremely early departure from your home, you should consider allowing yourself one night of rest in the Lafayette area before you take the examination. Many students who are unsuccessful with the examination tell us that failing to take the above precautions contributed strongly to their inability to complete the exam successfully. Most of these students find that their first year in college was somewhat less rewarding than it might have been because of the time spent retracing material studied in high school. Please consult your advanced credit schedule for the actual time and place of the examination. It is usually given both morning and afternoon.

Listed below are a set of major topics that was compiled from recent Chemistry II courses offered at Purdue. Study the list carefully before attempting the sample examination.

1. **Solutions**

Describe solute, solvent, and solution and the intermolecular forces between particles which determine if a solution will form.

Be able to calculate solution concentration in units of molarity, molality, percent by mass, and mole fraction. Be able to convert between these units and know which do not depend on temperature.

Given the absorbance of a solution find its concentration.

2. **Kinetics**

Describe a reaction rate in terms of a change in concentrations divided by a change in time.

Calculate a rate law from given data.

Know the integrated rate equations and half-life for zero order, first order, and second order reactions. Be able to give the rate constant in appropriate units for each order.

Describe how the energy of molecules effects their ability to react. Know how collision frequency, kinetic energy, and orientation of collision affect the rate of reaction.

Know how activation energy is experimentally determined using the Arrhenius Equation.

Describes how a catalyst increases the rate of a chemical reaction.

3. **Chemical Equilibrium.**

Write an equilibrium constant for a chemical reaction and calculate its value for given concentrations of reactants and products.

Describe chemical equilibrium in terms of forward and reverse reaction rates, and changes in concentration of reactants and products.

Be able to use the reaction quotient Q to calculate if a reaction is at equilibrium and determine which way the reaction must proceed to reach equilibrium.

Describe Le Chatelier's Principle and how it is used.

4. Acids and Bases

Use the following models to describe acids and bases (this includes comparing and contrasting the models): Arrhenius, Bronsted, and Lewis. Be able to identify acid-base reactions.

Know the term K_w and be able to write the dissociation reaction for water. What are the concentrations of H_3O^+ and OH^- at $25^\circ C$ in pure water?

Define pH and be able to calculate pH and pOH for any solution.

Define K_a and/or K_b and be able to calculate it for weak acids/bases.

Describe how a buffer functions, be able to identify buffer systems, and describe the term buffer capacity. Be able to calculate the effect on pH of adding an acid or a base to a buffer using the Henderson-Hasselbach equation.

Know what the pH titration curve for a weak acid titrated with a strong base, and a weak base titrated with a strong acid would look like. How would K_a or K_b be found from this data?

5. Chemical Thermodynamics

Define terms associated with thermodynamics such as system, surroundings, universe, and state functions.

Be able to state the First Law of Thermodynamics in terms of changes in internal energy accompanied by heat flow and work done on or by the system. Be careful to note the sign conventions when stating this law.

Be able to describe two types of work associated with chemical reactions. If discussing pressure-volume work, be able to calculate this quantity at constant opposing pressure.

Use the Second Law of Thermodynamics to describe the spontaneity of a system. Describe spontaneity and entropy. Predict the sign of an entropy change from a chemical reaction based upon the states of the products and reactants.

Calculate standard entropy, enthalpy, and Gibbs energy changes for a chemical reaction.

Know how Gibbs energy is related to entropy and enthalpy, and how Gibbs energy defines spontaneous processes.

Be able to relate Gibbs energy to equilibrium constants and cell potentials.

6. **Redox (Oxidation and reduction) reactions**

Balance an oxidation/reduction reaction and be able to define the oxidation numbers of all atoms. Identify the species being oxidized and reduced, and the oxidizing agent and reducing agent.

7. **Electrochemistry**

Apply your knowledge of redox reactions to understand electrochemical reactions.

Be able to calculate cell potentials using the Nernst equation. If the cell potential is known, then know how to calculate Gibbs energy and equilibrium constant.

Draw an electrochemical cell and identify the anode, cathode, solutions, salt bridge, and species being oxidized and reduced. Identify the direction in which electrons flow.

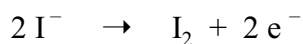
Describe an electrolytic and galvanic cell.

- _____ 6. Calculate the H_3O^+ ion concentration (in M) in a solution with a $\text{pH} = 7.80$.
- (a) $1.6 \times 10^{-8} \text{ M}$ (c) $8.0 \times 10^{-7} \text{ M}$
(b) $2.0 \times 10^{-7} \text{ M}$ (d) 0.89 M
- _____ 7. Ammonium chloride is used as an electrolyte in dry cells. Which of the following statements about a 0.10 M solution of NH_4Cl is correct?
- (a) The solution is basic.
(b) The solution is neutral.
(c) The solution is acidic.
(d) The values for K_a and K_b for the species in solution must be known before a prediction can be made.
- _____ 8. Calculate the $[\text{H}_3\text{O}^+]$ concentration (in M) in a 0.1 M aqueous solution of NH_3 , $[K_b = 1.8 \times 10^{-5}]$
- (a) $7.5 \times 10^{-12} \text{ M}$ (c) $1.8 \times 10^{-6} \text{ M}$
(b) $3.0 \times 10^{-10} \text{ M}$ (d) $1.3 \times 10^{-3} \text{ M}$
- _____ 9. A pH buffer is best described as a solution containing:
- (a) a weak acid.
(b) a strong acid.
(c) a mixture of a weak acid and a strong acid.
(d) a mixture of a weak acid and the salt of a weak acid.
- _____ 10. Calculate the pH of a solution prepared by dissolving 0.50 moles of acetic acid (HOAc , $K_a = 1.8 \times 10^{-5}$) and 0.020 moles of sodium acetate (NaOAc) in enough water produce 1.0 L of solution.
- (a) 2.52 (b) 3.35 (c) 6.14 (d) 7.71
- _____ 11. Calculate the pH of a 0.10 M aqueous solution of sodium acetate ($\text{MM} = 82 \text{ g/mol}$). For acetic acid, $K_a = 1.8 \times 10^{-5}$.
- (a) 5.13 (b) 8.87 (c) 9.37 (d) 10.25

Questions 12-13 refer the *unbalanced* equation shown below.



- _____ 12. How many hydroxide ions are involved in the balanced half-reaction involving HSnO_2^- ?
- (a) 2 (b) 3 (c) 4 (d) 5
- _____ 13. What is the coefficient of H_2O in the final balanced equation?
- (a) 1 (b) 3 (c) 4 (d) 6
- _____ 14. I_2 (MM = 254 g/mol) can be produced by passing an electric current through a solution of KI. Calculate the number of minutes a current of 10.0 A would have flow in order produce 6.0 g of I_2 .



- (a) 3.8 min (c) 7.6 min
(b) 4.6 min (d) 76 min

You may need the following list of Standard Reduction Potentials for questions 15-18.

<u>half-reaction</u>	<u>E°, V</u>
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.38
$\text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}$	-1.03
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{Cr}^{3+} + 3\text{e}^- \rightarrow \text{Cr}$	-0.74
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.23
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.00
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.80
$14\text{H}^+ + \text{Cr}_2\text{O}_7^{2-} + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$	+1.33
$8\text{H}^+ + \text{MnO}_4^- + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.49

- _____ 15. Which one of the following species is the best reducing agent?
- (a) Mn^{2+} (b) MnO_4^- (c) Mg (d) Mg^{2+}

- _____ 16. Calculate the value for the equilibrium constant for the following reaction under standard conditions.



- (a) 1.9×10^{-37} (c) 4.4×10^{18}
 (b) 1.6×10^{16} (d) 1.9×10^{37}
- _____ 17. The potential for the cell shown below is 0.31 V. What is the Mn^{2+} concentration?



- (a) 0.029 M (b) 0.066 M (c) 0.32 M (d) 34 M
- _____ 18. For the cell described in question 25, electrons will flow from _____.
- (a) Mn Zn (b) Zn Mn

- _____ 19. Addition of a strong acid a solution in which Ag^+ , AgCl , $\text{Ag}(\text{NH}_3)^+$, ammonia and Cl^- are at equilibrium will cause:

- (a) more AgCl dissolve.
 (b) some AgCl precipitate from solution.
 (c) more $\text{Ag}(\text{NH}_3)^+$ form.
 (d) the concentrations of Ag^+ , $\text{Ag}(\text{NH}_3)^+$ and Cl^- increase.

- _____ 20. Silver ion, Ag^+ , reacts with thiosulfate ion, $\text{S}_2\text{O}_3^{2-}$, in two steps form $\text{Ag}(\text{S}_2\text{O}_3)^-$ in the first step and $\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$ in the second step. If the stepwise formation constants are $K_{f1} = 6.6 \times 10^8$ and $K_{f2} = 4.4 \times 10^4$, what is the overall formation constant, K_f , of $\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$?

- (a) 6.7×10^{-5} (c) 6.6×10^8
 (b) 1.5×10^4 (d) 2.9×10^{13}

- _____ 21. Which one of the following thermodynamic properties is *not* a state function of a system?

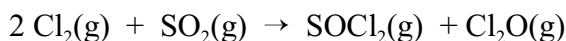
- (a) a transfer of heat
 (b) a change in temperature
 (c) a change in internal energy
 (d) a change in free energy

- _____ 22. Which of the following statements is **TRUE**?
- (a) An exothermic process will always be spontaneous.
 - (b) A process in which the entropy of the system increases will always be spontaneous.
 - (c) An endothermic process can never be spontaneous.
 - (d) A process in which the entropy of the surroundings increases will always be spontaneous.
 - (e) An exothermic process that is accompanied by an increase in the entropy of the system will always be spontaneous.

- _____ 23. Acetic acid, CH_3COOH , has an enthalpy of vaporization equal 52.25 kJ and an entropy of vaporization equal 122 J K^{-1} at its boiling point. Calculate the boiling point (in K) of acetic acid.

- (a) 0.43 K
- (b) 2.3 K
- (c) 100 K
- (d) 428 K
- (e) 563 K

- _____ 24. Use the given data at 298 K calculate ΔG° for the reaction:



Substance	$\text{Cl}_2(\text{g})$	$\text{SO}_2(\text{g})$	$\text{SOCl}_2(\text{g})$	$\text{Cl}_2\text{O}(\text{g})$
ΔH°_f (kJ/mol)	0	-296.8	-212.5	80.3
S° (J/K·mol)	223.0	248.1	309.77	266.1

- (a) 129.3 kJ
 - (b) 133.6 kJ
 - (c) 196.0 kJ
 - (d) 199.8 kJ
- _____ 25. Which of the following results in a decrease in entropy?
- (a) $\text{O}_2(\text{g}), 300 \text{ K} \rightarrow \text{O}_2(\text{g}), 400 \text{ K}$
 - (b) $\text{H}_2\text{O}(\text{s}), 0^\circ\text{C} \rightarrow \text{H}_2\text{O}(\text{l}), 0^\circ\text{C}$
 - (c) $\text{N}_2(\text{g}), 25^\circ\text{C} \rightarrow \text{N}_2(\text{aq}), 25^\circ\text{C}$
 - (d) $\text{NH}_3(\text{l}), -34.5^\circ\text{C} \rightarrow \text{NH}_3(\text{g}), -34.5^\circ\text{C}$

The following information applies for questions 26-27. Consider the following reaction:



It was experimentally shown that when the concentration of A is tripled and the concentration of B is held constant, the reaction rate increases by a factor of nine. When the concentrations of both A and B are doubled, the reaction rate increases by a factor of eight.

- _____ 26. What is the order of the reaction with respect A?
- (a) 0 (b) 1 (c) 2 (d) 3
- _____ 27. What is the order of the reaction with respect B?
- (a) 0 (b) 1 (c) 2 (d) 3

Use the following information answer questions 28 and 29.



<u>experiment</u>	<u>A, M</u>	<u>B, M</u>	<u>initial rate, M min⁻¹</u>
1	0.30	0.30	5.250
2	0.90	0.30	15.75
3	0.30	0.60	21.00

- _____ 28. Calculate the value for the rate constant for the reaction shown above.
- (a) 58.3 (b) 194 (c) 648 (d) 2160
- _____ 29. Calculate the initial rate (in M min⁻¹) for the above reaction if the initial concentration of both A and B is 0.90 M.
- (a) 15.75 M min⁻¹ (c) 94.50 M min⁻¹
 (b) 31.50 M min⁻¹ (d) 141.75 M min⁻¹

***Additional help (problem solving by topic) is available here:

<http://www.chem.purdue.edu/gchelp/>

This is a site with multiple questions on many different topics!

ANSWERS CHEM II SAMPLE TEST

- | | | | |
|-----|---|-----|---|
| 1. | A | 16. | D |
| 2. | B | 17. | B |
| 3. | D | 18. | A |
| 4. | C | 19. | B |
| 5. | C | 20. | D |
| 6. | A | 21. | A |
| 7. | C | 22. | E |
| 8. | A | 23. | D |
| 9. | D | 24. | D |
| 10. | B | 25. | C |
| 11. | B | 26. | C |
| 12. | A | 27. | B |
| 13. | A | 28. | B |
| 14. | C | 29. | D |
| 15. | C | | |