Purdue University
Chemistry 11600 Study Guide
Revised May 2018

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 at Purdue University.

Suggested Textbooks

Chemistry & Chemical Reactivity, 8th ed by Kotz, Treichel & Townsend; Thomson/Brooks/Cole.
ISBN-13: 978-0840048288.

Chemistry: The Molecular Nature of Matter and Change, 8^{th} edition by Silberberg; McGraw-Hill.
ISBN-13: 978-1259916229.

Chemistry, the Central Science by Brown, 14^{th} edition by LeMay, Bursten, et. al.; Prentice-Hall.
ISBN-13: 978-0134414232.

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 Information

1. You may only take the CHM 11600 test out exam **ONCE**.
2. You are NOT ELIGIBLE to take this test out exam if you
    a. Do not have credit for CHM 11500 or equivalent (meaning CHM 11100 and CHM 11200).
    b. Have enrolled in CHM 11600 and remained in the class longer than four weeks.
    c. Have a grade in CHM 11600 that does award you credit in your major.
3. If you have questions please contact the general chemistry office at 494-5250.

Test Topics/Preparation for Exam:

The subject matter of CHM 11600 deals with the following topics: solutions, intermolecular forces, chemical equilibrium, acids/bases/buffers, reduction-oxidation, electrochemistry, and chemical thermodynamics. Many of these topics are related, such as chemical equilibrium and acids/bases/buffers, or electrochemistry and thermodynamics. These topics rest upon a foundation of understanding fundamental concepts covered in a general chemistry I course, or CHM 11500.

You should study the topics listed in the attached outline prior to attempting the simple examination included with this study guide.

At the end of this study guide you will find a sample examination over this material. Naturally, it does not cover every topic, or every aspect of a topic. Following the sample exam are the answers to the questions.

Words of Advice

It is a student's responsibility to meet with his or her advisor to discuss myPurduePlan and to graduate on time with all requirements completed. Do not wait until the last weeks of your undergraduate program to establish credit in CHM 11600 if it is part of your degree requirements.

If you fail the test out your next best option is to take the CLEP exam https://admissions.purdue.edu/transfercredit/clep.php.

Do not request special consideration if you fail the test out exam and you are scheduled to graduate and have a job waiting. It is your responsibility to regularly meet with your advisor, to work within MyPurduePlan, and to complete your degree on time.

Listed below are a set of major topics may be covered on the CHM 11600 testout examination. The examination contains a formula sheet for the exam that gives any formula and/or constant you may need to complete an item on the exam. You do not need to memorize formulas in order to pass the exam.

Solutions

Describe solute, solvent, and solution in terms of the species in solution and their concentrations.

Given the absorbance of a solution find its concentration (laboratory data).

Intermolecular Forces

List and briefly describe the different types of intermolecular forces: dipole-dipole, hydrogen bonding, dispersion forces and ion-dipole.

Identify the intermolecular forces that a given molecule would participate in.

Use Lewis structures to draw how molecules might interact and be able to identify interactions (IMF) from a drawing.

Use Lewis structures, surface area of the molecule, and possible intermolecular interactions to compare the boiling point, vapor pressure, or surface tension of a set of molecules.

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 and associated graphs. Be able to give the rate constant in appropriate units for each order.

Describe how the energy of molecules affects their ability to react and how this can be represented in graphical form. 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.

Describe how a catalyst impacts the rate of a chemical reaction and how this can be represented graphically.

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, and using graphs.

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

Describe and apply Le Chatelier’s Principle to chemical systems.

Acids and Bases

Use the following models to describe acids and bases (this includes comparing and contrasting the models): Arrhenius, Bronsted-Lowry, and Lewis.

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^oC?

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 apply the concept of 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 the following systems looks like and how to extract information from it: strong acid/strong base, weak acid/strong base, and weak base/strong acid.

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 given data.

Apply Hess's Law – meaning find the change in enthalpy for a specific reaction given chemical reactions and change in enthalpy values of other reactions.

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

Relate cell potentials and Gibbs energy.

Redox (Oxidation and reduction) reactions

Balance an oxidation/reduction reaction and be able to define the oxidation numbers of all atoms.

Given a redox reaction, identify the species being oxidized and reduced, and the oxidizing agent and reducing agent.

Identify strongest oxidizing agent or reducing agent from experimental data.

Electrochemistry

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 and 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.

Laboratory

Apply your knowledge of safe laboratory practices

Know how to use a pipet, electronic balance, buret, and volumetric flask.

Know how to construct and use a calibration curve.

Given laboratory data from experiments designed to determine concentrations (titrations or spectroscopy), find the values requested.

Chemistry 11600 TESTOUT PRACTICE EXAM

_______ 1. Consider the reaction below. If the rate of reaction of nitric oxide is 0.066 M/s, what is the rate of consumption of oxygen?

$$2\ NO\ (g) + O_2\ (g) \rightarrow 2\ NO_2\ (g)$$

(a) 0.132 M/s
(b) 0.066 M/S
(c) 0.033 M/s
(d) −0.132 M/s

_______ 2. In 500. ml of a 0.15 M solution of NaCl what is the concentration of sodium ions?

(a) 0.30 M
(b) 0.15 M
(c) 8.8 g
(d) 4.4 g

_______ 3. Calculate the pH of a solution containing 0.15 M NH_3 and 0.35 M NH_4Cl.

$K_a = 5.6 \times 10^{-10}$.

(a) 10.1
(b) 9.62
(c) 8.88
(d) 8.40

_______ 4. What is the concentration (in M) of hydronium ions in a solution at 25°C with pH = 4.282?

(a) 1.92×10^{-10} M
(b) 5.22×10^{-5} M
(c) 5.22×10^{-4} M
(d) 9.718

_______ 5. Which has the highest normal boiling point?

(a) CH_4
(b) C_2H_6
(c) C_3H_8
(d) C_4H_{10}

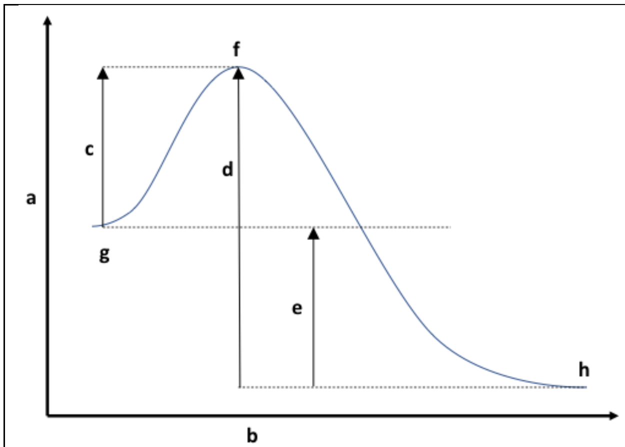
________ 6. Find the pH of a 0.235 M solution of acetic acid, CH_3COOH. The K_a for acetic acid is 1.8×10^{-5}.

(a) 4.23×10^{-6} M
(b) 2.06×10^{-3} M
(c) 2.69
(d) 5.37

________ 7. Which of the following liquids has the lowest vapor pressure at room temperature and 1 atm pressure?

(a) CH_4
(b) CH_3CH_3
(c) CH_3OH

Use the following diagram for questions 8-11



- Reactants
- Transition State
- Reaction Coordinate
- Products
- Enthalpy Change
- Energy
- Activation Energy (forward)
- Activation Energy (reverse)

______ 8. Which identifies the activation energy?

(a) c
(b) d
(c) e
(d) d-c

______ 9. Which identifies the enthalpy change?

(a) c
(b) d
(c) e
(d) d-c

_______ 10. Which identifies the transition state?

(a) g
(b) f
(c) h
(d) b

_______ 11. Which identifies the products?

(a) g
(b) f
(c) h
(d) b

For questions 12. and 13. Use the following reaction and data table. Consider the reaction:
P_4 (*g*) + 6 H_2 (*g*) → 4 PH_3 (*g*)

[P_4] (M)	**[H_2] (M)**	**Initial Rate (M/s)**
0.0110	0.0075	3.20×10^{-4}
0.0110	0.0150	6.40×10^{-4}
0.0220	0.0150	6.39×10^{-4}

_______ 12. What is the rate law for this reaction?

(a) Rate = $k[P_4][H_2]$
(b) Rate = $k[P_4]$
(c) Rate = $k[H_2]$
(d) Rate = $k[P_4][H_2]^6$

_______ 13. What is the value and units of the rate constant, k?

(a) 3.87 $M^{-1}\ s^{-1}$
(b) 0.0291 s^{-1}
(c) 0.0427 s^{-1}
(d) 2.18×10^{13} $M^{-6}\ s^{-1}$

_______ 14. In which substance does bromine have an oxidation number of +1?

(a) Br_2
(b) HBr
(c) HBrO
(d) $HBrO_2$

______ 15. In aqueous solution, Fe^{2+} reacts with permanganate ion (MnO_4^-) to form Fe^{3+} and Mn^{2+}. What is the reducing agent in this reaction?

(a) H_2O
(b) Fe^{2+}
(c) MnO_4^-
(d) Fe^{3+}
(e) Mn^{2+}

______ 16. In which of the following species does sulfur have the same oxidation number as it does in H_2SO_4?

(a) H_2SO_3
(b) $S_2O_3^{2-}$
(c) S^{2-}
(d) S_8
(e) SO_2Cl_2

Use the following for 17-19: The data in the table below were obtained for the reaction:

$2\ ClO_2\ (aq) + 2\ OH^-\ (aq) \rightarrow ClO_3^-\ (aq) + ClO_2^-\ (aq) + H_2O$

Experiment number	$[ClO_2]$ (M)	$[OH^-]$ (M)	Initial Rate (M/s)
1	0.060	0.030	0.0248
2	0.020	0.030	0.00276
3	0.020	0.090	0.00828

______ 17. What is the order of the reaction with respect to ClO_2?

(a) 4
(b) 1
(c) 0
(d) 2
(e) 3

______ 18. What is the order of the reaction with respect to OH^-?

(a) 0
(b) 1
(c) 2
(d) 3
(e) 4

______ 19. What is the overall order of the reaction?

(a) 0
(b) 1
(c) 4
(d) 3
(e) 2

______ 20. What is the magnitude of the rate constant for the reaction?

(a) 115
(b) 1.15×10^4
(c) 713
(d) 4.6
(e) 230

______ 21. For a first-order reaction, a plot of __________ versus __________ is linear.

(a) $\ln [A]_t$, t
(b) $1/[A]_t$, t
(c) In [A]t, 1/t
(d) $[A]_t$, t
(e) t, $1/[A]_t$

______ 22. The thermodynamic quantity that expresses the degree of disorder in a system is __________.

(a) entropy
(b) internal energy
(c) heat flow
(d) enthalpy
(e) bond energy

________ 23. Which one of the following is always positive when a spontaneous process occurs?

(a) ΔH_{univ}

(b) ΔH_{surr}

(c) ΔS_{surr}

(d) ΔS_{univ}

(e) ΔS_{sys}

_______ 24. $N_2(g) + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$

The reaction indicated above is thermodynamically spontaneous at 298 K, but becomes nonspontaneous at higher temperatures. Which of the following is true at 298 K?

(a) ΔG, ΔH, and ΔS are all positive.

(b) ΔG, ΔH, and ΔS are all negative.

(c) ΔG and ΔH are negative, but ΔS is positive.

(d) ΔG and ΔS are negative, but ΔH is positive.

(e) ΔG and ΔH are positive, but ΔS is negative.

_______ 25. When solid NH_4SCN is mixed with solid $Ba(OH)_2$ in a closed container, the temperature drops and a gas is produced. Which of the following indicates the correct signs for ΔG, ΔH, and ΔS for the process?

	ΔG	ΔH	ΔS
(a)	−	−	−
(b)	−	+	−
(c)	−	+	+
(d)	+	−	+
(e)	+	−	−

_______ 26. The energy diagram for the reaction X + Y → Z is shown. The addition of a catalyst to this reaction would cause a change in which of the indicated energy differences?

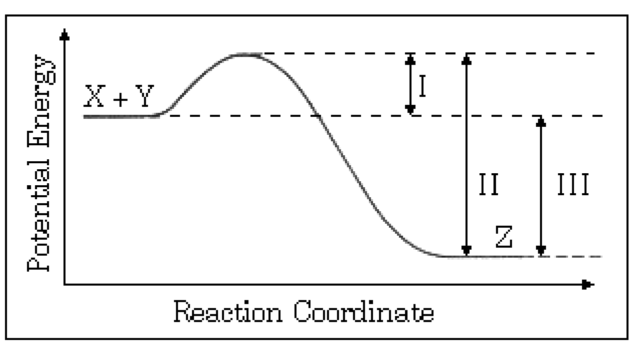
(a) I only

(b) II only

(c) III only

(d) I and II only

(e) I, II, and III



_______ 27. Given the following balanced chemical equation: $3\ C_2H_{2(g)} \rightarrow C_6H_{6(g)}$

What is the standard enthalpy change, ΔH° , for the reaction represented above?
(ΔH°_f of $C_2H_{2(g)}$ is 230 kJ mol^{-1} ; ΔH°_f of $C_6H_{6(g)}$ is 83 kJ mol^{-1})

(a) −607 kJ
(b) −147 kJ
(c) −19 kJ
(d) + 19 kJ
(e) +773 kJ

_______ 28. Calculate the pH of a solution prepared by dissolving 2×10^{-3}moles of HCl in enough water produce 1.0 L of solution.

(a) –2.7
(b) 2.3
(c) 2.7
(d) 3.3

_______ 29. 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.

_______ 30. Calculate the $[H_3O^+]$ concentration (in M) in a 0.1 M aqueous solution of NH_3 $[K_b = 1.8 \times 10^{-5}]$

(a) 7.5×10^{-12} M
(b) 3.0×10^{-10} M
(c) 1.8×10^{-6} M
(d) 1.3×10^{-3} M

______ 31. 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.

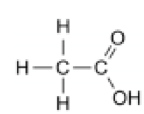
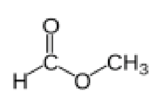
______ 32. Which of the following best describes all the intermolecular forces exhibited by a pure sample of CH_3NH_2?

(a) dispersion only
(b) dipole-dipole and hydrogen bonding
(c) dispersion and hydrogen bonding
(d) dispersion, dipole-dipole, and hydrogen bonding
(e) dispersion and dipole-dipole

______ 33. The pH of a solution prepared by mixing 50.0 mL of 0.125 M KOH and 50.0 mL of 0.125 M HCl is __________.

(a) 0.00
(b) 6.29
(c) 8.11
(d) 5.78
(e) 7.00

______ 34. The boiling point for isomers of $C_2H_4O_2$ are shown in the figure below. The reason acetic acid has a higher boiling point is

Acetic acid		118°C
Methyl acetate		56.9°C

(a) Acetic acid has a larger molar mass.
(b) Methyl acetate has stronger dispersion forces between its molecules.
(c) The –CH_3 group on the methyl acetate interacts with the carbon oxygen double bond differently than in acetic acid.
(d) Acetic acid forms hydrogen bonds while methyl acetate does not.

ANSWERS

1. C
2. B
3. C
4. B
5. D

6. C
7. C
8. A
9. C
10. B

11. C
12. C
13. C
14. C
15. B

16. E
17. D
18. B
19. D
20. E

21. A
22. A
23. D
24. B
25. C

26. D
27. A
28. C
29. C
30. A

31. D
32. D
33. E
34. D