General Chemistry II (CHM 112) Spring 2012

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This course provides a continuation of the study of the fundamental principles and laws of chemistry. Topics include gases, solutions, intermolecular forces, kinetics, equilibrium, acid-base theory, and electrochemistry. Upon completion, you should be able to demonstrate an understanding of chemical concepts as needed to pursue further study in chemistry and related professional fields.

Chemistry and the other sciences are processes of inquiry, not static collections of facts. While I certainly want to facilitate your understanding of chemical concepts and the "language" chemists speak, I also want you to come to see chemistry as a way of learning. A significant portion of this class will be aimed at helping you to apply chemical logic to the investigation and analysis in the world. It is extremely important that you understand the foundations of chemistry and not simply memorize specific problems or topics. You will be able to solve a variety of problems if you THINK and APPLY the principles you have learned. You will also advance your level of problem solving skills that will be applicable to many other aspects of life.

Prerequisite

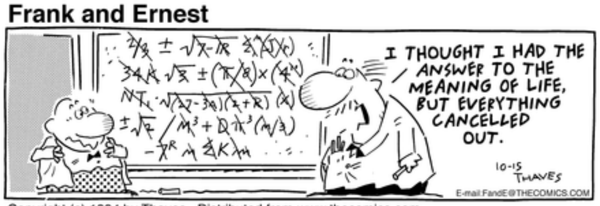
CHM 111, General Chemistry I

Required Materials
Textbook: Chemistry,6th Ed. McMurray and Fay, Prentice Hall, 2012 (custom ed. for La Salle U).

A Scientific Calculator (it must have a key for y^x and log x). Learn how to use it. (ask if you are unsure).

Bring your text and calculator to each lecture period.

Frank and Ernest



Learning Objectives and Homework Problems for CHM 112 (General Chemistry II)

When we are finished discussing a chapter and you have **MASTERED** the concepts and the homework problems you should be able to:

Chapter 9 Gases: Their Properties and Behavior

38, 39, 40, 48, 50, 51, 59, 64, 70, 79, 86, 90, 122

1. Explain how the height of a liquid in a barometer depends on the density of the liquid.
2. Interconvert units of pressure.
3. Describe how the pressure of a gas is measured using a manometer.
4. Use the ideal gas law to calculate pressure, volume, moles of gas, or temperature, given the other three variables.
5. Use the ideal gas law to calculate final pressure, volume, moles of gas, or temperature from initial pressure, volume, moles of gas, and temperature.
6. Perform stoichiometric calculations relating the mass of a reactant to the mass, moles, and volume or pressure of a gaseous product.
7. Use the ideal gas law to determine the molar mass of a gas.
8. Use the ideal gas law to determine the density of a gas.
9. Use Dalton's law to calculate the partial pressure of a gas in a mixture.
10. Explain each of the gas laws using the Kinetic Molecular Theory.
11. Use Graham's law to calculate the relative rates of effusion of two different gases.
12. State conditions under which a gas is expected to behave ideally or nonideally.

Chapter 10 Liquids, Solids, and Phase Changes

30, 31, 32, 46, 52, 54, 57, 58

1. Determine whether a molecule is expected to be polar using its VSEPR geometry and electronegativity.
2. Identify major intermolecular forces present in substances and determine which of two substances will exhibit the stronger intermolecular force.
3. Sketch a phase diagram, labeling the axes and each of the regions, and locate the triple point, critical point, normal melting point, and the normal boiling point. Use the phase diagram to describe physical changes.

Chapter 11 Solutions and Their Properties

34, 39, 48, 54, 55, 64, 66 ,72, 76, 80, 94, 97, 100, 105, 108, 110

1. Explain the rule of thumb "like dissolves like." Analyze the solution process in terms of solute-solute, solute-solvent, and solvent-solvent interactions.

2. Perform calculations using solution density, molarity, mole fraction, weight percent, parts per million, parts per billion, and molality.
3. Use Henry's law to predict concentrations of gases in solution.
4. Use Raoult's law to calculate the vapor pressure over a solution containing a nonvolatile solute and a solution containing two volatile liquids.
5. Employ freezing point depression and boiling point elevation to determine the molar mass of a solute.
6. Explain the origin of osmotic pressure, and use it to determine the molar mass of a solute.

Chapter 12 Chemical Kinetics

38, 39, 42, 45, 49, 50, 52, 59, 76, 78, 79, 81, 82, 85, 86, 88, 106, 117

1. Use concentration versus time data to calculate an average rate of reaction over a period of time.
2. Express the relative rates of consumption of reactants and formation of products using the coefficients of a balanced chemical equation.
3. From data on initial concentrations of reactants and initial rates, determine the order of reaction with respect to each reactant, the overall order of reaction, the rate law, the rate constant, and the initial rate for any other set of initial concentrations.
4. Use integrated first- and second-order rate laws to describe the progress of a reaction.
5. Determine the order of reaction from plots of log concentration *versus* time and reciprocal concentration *versus* time.
6. Relate half-life to the rate constant for first- or second-order reactions.
7. Estimate the half-life of a first-order reaction from a plot of concentration *versus* time.
8. Given a reaction mechanism and an experimental rate law, identify the reaction intermediates, determine the molecularity of each elementary reaction, and determine if the mechanism is consistent with the experimental rate law.
9. Prepare an Arrhenius plot and determine the activation energy from the slope of the line.
10. Use the Arrhenius equation to relate rate constant to temperature and activation energy.
11. Sketch a potential energy profile showing activation energies for the forward and reverse reactions, and how they are affected by the addition of a catalyst.

Chapter 13 Chemical Equilibrium

27, 28, 29, 35, 40, 41, 44, 48, 58, 60, 65, 70, 71, 78, 80, 81, 83, 86, 87, 94, 99, 106, 118

1. Write the equilibrium equation for any balanced chemical equation representing a homogenous or heterogeneous equilibrium.

2. Calculate the equilibrium constant K_c from the equilibrium concentrations of products and reactants.
3. Calculate the equilibrium constant K_p from the equilibrium partial pressures of reactants and products.
4. Interconvert between K_c and K_p using a balanced equation.
5. Determine whether mainly products or mainly reactants exist at equilibrium using the value of K_c or K_p.
6. Determine whether or not a system is at equilibrium for a given mixture of reactants and products. If it is not, determine the direction in which the reaction must go to achieve equilibrium.
7. Calculate the final concentrations of reactants and/or products from K_c and initial concentrations of reactants and/or products.
8. Determine the reaction direction when a system at equilibrium reacts to a stress applied to the system, including changes in concentrations, pressure and volume, or temperature.
9. Describe the effect of adding a catalyst to a system at equilibrium.
10. Describe the relationship between the equilibrium constant and the ratio of the rate constants for the forward and reverse reactions. Solve problems involving this relationship.

Chapter 14 Aqueous Equilibria: Acids and Bases

34, 35, 37, 46, 47, 49, 50, 54, 55, 61, 67, 70, 71, 74, 82, 84, 93, 96

1. Describe an acid or base according to the Arrhenius, Brønsted-Lowry, and Lewis theories.
2. From Lewis structures, determine which chemical species can act as a Brønsted-Lowry acid, a Brønsted-Lowry base, or both.
3. Identify the conjugate acid-base pairs in a chemical equation for a proton-transfer reaction.
4. Determine whether an acid is a stronger or weaker acid than water and whether the conjugate base of the acid is a stronger or weaker base than water, based on the extent of dissociation of an acid in water.
5. Given a chemical equation representing a proton transfer reaction and the relative strengths of each acid and base involved in the reaction, determine whether the reaction is favored to the right or to the left.
6. Calculate H_30^+ concentration from OH^- concentration and vice versa. From these concentrations determine whether the solution is acidic, neutral, or basic.
7. Interconvert pH and $[H_30^+]$. Classify the solution as acidic, neutral, or basic.
8. Determine the pH of a solution from the molar concentration of a strong acid or a strong base.
9. Determine the K_a of the acid from the pH of a weak-acid solution.

10. Given the K_a value and the initial concentration of a weak monoprotic acid, calculate the concentrations of all species at equilibrium, the pH of the solution, and the percent dissociation of the acid.
11. Calculate the concentrations of all species at equilibrium and the pH of the solution from the K_a values and initial concentration of a weak diprotic acid.
12. Calculate the concentrations of all species at equilibrium and the pH of the solution from the K_b value and initial concentration of a weak base.
13. Interconvert K_a and K_b.
14. Classify salt solutions as acidic, neutral, or basic. Calculate the pH of these solutions.
15. Identify which of two substances is more acidic.
16. Identify Lewis acids and Lewis bases in chemical reactions.

Chapter 15 Applications of Aqueous Equilibria

35, 44, 53, 55, 60, 88, 90, 95, 96, 100, 101

1. Predict whether the pH will be equal to, greater than, or less than 7.00 at the equivalence point using the relative strengths of acid and base in a neutralization reaction, p.
2. Write balanced net ionic equations for the four types of neutralization reactions.
3. Describe the effect on pH when the conjugate base of a weak acid is added to a solution of the weak acid, and the conjugate acid of a weak base is added to a solution of the weak base. Calculate the concentrations of all species present at equilibrium.
4. Determine the pH and equilibrium concentrations of all species in a buffer solution using the initial concentration of weak acid (or weak base) and its conjugate base (or weak acid).
5. Calculate the pH of a buffer after the addition of OH^- or H_30^+.
6. Use the Henderson-Hasselbalch equation to calculate the pH of a buffer.
7. From a table of weak acids and their K_a values, select the weak acid/conjugate base pair that would make the best buffer at a given pH.
8. Calculate pH values for a strong acid-strong base titration.
9. Calculate pH values for a weak acid-strong base titration.
10. Select which indicator(s) could be used to detect the equivalence point for a particular a titration curve.
11. Calculate pH values for a weak base-strong acid titration.
12. Write the solubility product expression for a given ionic compound.
13. Calculate the solubility using the K_sp of an ionic compound, and vice versa.
14. Calculate the solubility of an ionic compound in the presence of a common ion.

Chapter 16 Thermodynamics: Entropy, Free Energy, and Equilibrium

30, 32, 37

1. Qualitatively determine whether simple chemical or physical changes are spontaneous.

2. Qualitatively predict whether the sign of ΔS is positive or negative for a chemical or physical change.
3. On the basis of probability, determine which of two states has the higher entropy.
4. Calculate the standard entropy of reaction from the standard molar entropies of products and reactants.
5. Determine whether a reaction is spontaneous by determining the sign of ΔS_{total}.
6. Use the equation $\Delta G = \Delta H - T\Delta S$ to calculate the free energy of reaction and to determine the temperature at which a nonspontaneous reaction becomes spontaneous.
7. Calculate the standard free energy of reaction from standard free energies of formation.
8. Calculate the free energy of reaction for a system having nonstandard pressures and concentrations.
9. From the standard free energy of reaction, calculate the value of the equilibrium constant.

Chapter 17 Electrochemistry

27, 28, 44, 46, 56

1. Sketch a galvanic cell, identifying the anode and cathode half-reactions, the sign of each electrode, and the direction of electron and ion flow.
2. Write balanced chemical equations for reactions occurring in a galvanic cell.
3. Write and interpret shorthand notations for galvanic cells.
4. Interconvert cell potential and free-energy change for a reaction.
5. Use a table of standard reduction potentials to calculate standard cell potentials.
6. Use a table of standard reduction potentials to rank substances in order of increasing oxidizing strength or reducing strength and to determine whether a reaction is spontaneous.
7. Use the Nernst equation to calculate cell potentials for reactions occurring under nonstandard conditions.
8. From a measured cell potential for a reaction involving hydrogen ion and a reference cell potential, calculate the pH of the solution.
9. Calculate equilibrium constants from standard cell potentials and vice versa.
10. Write balanced chemical equations for reactions occurring in common batteries.
11. Describe the reactions that occur when iron rusts.
12. Describe half-cell and overall reactions occurring in electrolytic processes.
13. Perform electrolytic cell calculations interconverting current and time, charge, moles of electrons, and moles (or grams) of product.

<u>Course grade</u>

a) Class Participation: There will be many opportunities for you to participate in class, e.g., by asking questions and solving problems. A portion of your grade will be determined by your ability to regularly and meaningfully contribute to class discussions. Your final grade will be negatively impacted by frequent absences and/or persistent, disruptive talking.

Please note: Cell phones and pagers must be turned off or placed in silent mode at the beginning of class as a courtesy to everyone in the class. Grade penalties of 5-25 points may also be applied for failing to follow this policy.

b) Quizzes: Short in-class quizzes will be given once or twice a week. Quizzes may or may not be announced ahead of time. The question format will include short answer and problems . No make-up quizzes will be available. Missed quizzes will count as a zero.

c) Exams: Although homework problems will help you prepare from the exam, It will generally be the case that the exam questions will test what you have LEARNED, by asking you to apply your knowledge. It is therefore important that you UNDERSTAND what you are doing and that you do not just memorize various problem types. All exams are cumulative!

Three midterm exams will be given on the indicated class dates, and **d) a final comprehensive exam** . Exam questions may take the form of short answer, and/or problem solving formats. There are no make–up exams. Only official documentation confirming legal or medical emergencies will be considered.

Homework: I expect you to read the assigned sections and homework problems from the textbook. Homework problems are **not** collected; it is your responsibility to complete your work. It is important that you do your homework. The more you do, the more you will learn. From time to time **assignments**, which will be submitted, will be handed out. These are meant to check your progress in the course. Chemistry cannot be efficiently learned without working problems, and you will not realize what you do not know until you try to do a problem. Most students do poorly in this course because they neglect to do homework! Answers to the textbook problems are available on my web page. I am happy to help you outside of class. You may consult with me as often as you wish; **I encourage it**. Please consult early rather than waiting just before a test; it will be better for your learning. In order to compete successfully in this course, you will need to spend at least 12–15 hours per week studying chemistry.

There is ample research which proves that students who study together perform considerably better than those who study alone. You are encouraged to find fellow students interested in spending at least a portion of their study time with others. If you know of no other student taking this course, I will be happy to assist your efforts by making a class announcement for other interested person(s) to meet with you after class. All of us rely on our senses for the learning process. However, the extent to which we rely on sight, sound, even touch and smell differs from person to person. You are encouraged to assess your own learning style preference and capitalize on the approach best suited for you.

Supplemental Instruction (SI): This course is a “historically difficult” course and therefore Supplemental Instruction (SI) has been attached to it to help you get the most out of your efforts with this course. With SI, a trained student, who has already successfully taken this course, attends the class again and conducts weekly review sessions for the students in the class. The SI Leader knows the course content and can also share with you successful study techniques for this course. Additionally, the review sessions offer you an opportunity to work