Preparation for the Chem 1C Final

Chem 1C (Kahn, Spring 2006) final covers Chapters 13, 14, 16, 17, 19 (parts of), 20, 21, 22 (parts of). The questions on the exam test your knowledge and understanding of topics covered in my lectures. The Key Concepts that you need to know are outlined the Midterm I and Midterm II guide with the more recent additions highlighted below

Questions and grading.
The final has
1) **four** multiple-choice questions, each 10 points, no partial credit
2) **two** true/false question sets, each 10 points, 2 points for each correct statement
3) **six** problem-solving questions, each 12 points, 50% partial credit
4) **one** question on definitions (glossary), total 18 points with 3 pts each, partial credit of 1 point for a close but not sufficiently complete answer.
Partial credit of 50% will be given for the problem-solving questions for the work that has correctly accomplished 50% or more of the task but has a small error that prevents student from arriving to a correct answer. Some examples of such small errors are i) being off by a factor 1000 (typical unit error), ii) sign error, or iii) trivial math errors with correct formulas. Errors arising from using incorrect formulae or wrong molecular masses do not qualify for partial credit.

Study tips.
As you have probably figured out by now, a significant fraction of my questions are variations of the assigned textbook problems. The syllabus has the list of these problems. You should be able to solve pretty much all these problems independently, without looking up the answers or solution strategies, if you plan to do well in the final. I have emphasized various concentration calculations on the caffeine problems and pop quizzes and you need to know how to do these in order to solve some of the problems on the test.
Key concepts (in addition to those in Midterm I and Midterm II)

1. Distinction between inorganic and organic compounds
2. Structure and nomenclature of alkanes
3. Structural isomerism of alkanes
4. Chemical properties of alkanes, reactions with oxygen and chlorine
5. Polyethylene as a common organic polymer
6. Structure and nomenclature of alkenes
7. Cis-trans isomerism of alkenes
8. Chemical properties of alkenes: addition reactions
9. Structure and nomenclature of alkynes
10. Cyclic alkanes
11. Structure of benzene
12. Halogenated hydrocarbons: structure and nomenclature
13. The concept of nucleophiles and electrophiles
14. Alcohols: structure, nomenclature, chemical properties
15. Aldehydes: structure, nomenclature, chemical properties
16. Ketones: structure, nomenclature, chemical properties
17. Carboxylic acids: structure, nomenclature, chemical properties
18. Esters and amides
19. Amino acids: general structure, structures of glycine and alanine
20. Optical isomerism of amino acids
21. Structure and properties of the peptide bond
22. Proteins as linear polymers of amino acids
23. Polysaccharides as linear or branched polymers of carbohydrates
24. Purines and pyrimidines as components of nucleic acids
25. General structure of the nucleic acid chain
26. General structure of DNA
27. Structure and properties of N₂
28. Structure and properties of O₂
29. Structure and properties of ammonia and hydrazine
30. Industrial synthesis of ammonia and biological nitrogen fixation
31. Structure and properties of nitrogen oxides (N₂O, NO, NO₂)
32. Nitric acid
33. Structure and properties of H₂S
34. Structure and properties of sulfur dioxide and sulfur trioxide
35. Sulfuric acid
**Equations and constants sheet.** If you have a programmable calculator, feel free to stuff it with all the equations and fundamental constants that you desire. However, do not store any textual/factual chemistry information in your calculator. I reserve the right to check your calculator during the exam and finding inappropriate content in your calculator is grounds for failure in the class. For the benefit of students who do not have programmable calculators, the information sheet below will be provided.

**Units:**

- 1 gram = 0.001 kilograms = 1000 mg
- 1 liter = 1000 milliliters = $1 \cdot 10^6 \mu L$
- 1 M (molar) = $1 \cdot 10^6 \mu M = 1 \cdot 10^9 nM$
- 1 atm = 760 torr (mmHg) = 1.01325 bar = 101.325 kPa
- 0 °C = 273.15 K

**Fundamental constants:**

- Universal gas constant: $R = 8.31451 \text{ J K}^{-1} \text{ mol}^{-1} = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$
- Avogadro’s number: $N_A = 6.022 \cdot 10^{23} \text{ mol}^{-1}$
- Faraday’s constant: $F = 96,485 \text{ C mol}^{-1}$
- Planck constant: $h = 6.626 \cdot 10^{-34} \text{ J s}$
- Speed of light: $c = 3 \cdot 10^8 \text{ m s}^{-1}$

**Mathematical Relationships:**

- **Logarithms**
  
  $\ln \left( \frac{A}{B} \right) = -\ln \left( \frac{B}{A} \right)$

  $\ln(A \cdot B) = \ln A + \ln B$

  $\ln(1 - A) \approx -A$ for $A << 1$

**Physical relationships:**

- Energy and mass: $E = m \cdot c^2$
- Energy and frequency: $E = h \cdot \nu$
- Wavelength and frequency: $\lambda = \frac{c}{\nu}$

**Color and Wavelength:**

The wavelength of the absorbed light and is indicated on the color wheel, the observed color of the material is the color opposite to the absorbed color. For example, sample absorbing at 590 nm will appear blue to the human observer. Note that green color can be observed by samples that absorb a little above 400 as well as by samples that absorb a little below 800 nm. The region below 400 nm is the ultraviolet; sample absorbing ultraviolet appears colorless. The region above 800 in the infrared; the sample absorbing infrared appears colorless.
Definitions of concentrations

Mass percent (w/w) \( \left( \frac{\text{grams of solute}}{\text{grams of solution}} \right) \times 100\% \)

Mass percent (w/vol) \( \left( \frac{\text{grams of solute}}{\text{milliliters of solution}} \right) \times 100\% \)

Mole fraction of solute, \( \chi \) \( \frac{\text{moles of solute}}{\text{moles of solute + moles of solvent}} \)

Molarity, M \( \frac{\text{moles of solute}}{\text{liters of solution}} \)

Molality, m \( \frac{\text{moles of solute}}{\text{kg of solvent}} \)

Physical Chemistry:

Ideal gas law \( PV = nRT \)

Free Energy & Enthalpy \( \Delta G = \Delta H - T\Delta S \)

Free Energy & Equil. Const. \( \Delta G^\circ = -RT\ln K_{eq} \)

Free Energy & Potential \( \Delta G^\circ = -nF\varepsilon^\circ \)

Temp. dep. of K \( \ln K = \frac{-\Delta H^0}{RT} + \frac{\Delta S^0}{R} \)

Temp. dep. of \( P_{vap} \) \( \ln P_{vap} = -\frac{\Delta H^0_{vap}}{R} \left( \frac{1}{T} - \frac{1}{T^0} \right) = -\frac{\Delta H^0_{vap}}{R} \left( \frac{1}{T} - \frac{1}{T^0} \right) \)

Clausius-Clapeyron relation. \( \ln \left( \frac{P_{vap}^1}{P_{vap}^2} \right) = \frac{\Delta H^0_{vap}}{R} \left( \frac{1}{T^1} - \frac{1}{T^2} \right) \)

Henry’s law \( P_{gas} = k_H \chi \)

Raoult’s law (nonvolatile) \( P_{solution} = \chi_{\text{solvent}} P^0_{\text{solvent}} \)

Raoult’s law (ideal volatile) \( P_{solution} = \chi_A P^0_A + \chi_B P^0_B \)

Boiling point elevation \( \Delta T = iK_b m_{\text{solute}} \)

Freezing point depression \( \Delta T = iK_f m_{\text{solute}} \)

Boiling point elevation const. \( K_b = \frac{RT^2}{1000[g/kg] \Delta H_{vap}} M_{\text{solvent}} [g/mol] \)

Freezing point depr. const. \( K_f = \frac{RT^2}{1000[g/kg] \Delta H_{fus}} M_{\text{solvent}} [g/mol] \)

Osmotic pressure \( \pi = iMRT \)

Natural decay \( \ln \left( \frac{N_t}{N_o} \right) = -kt \)

Half-life \( \tau = \frac{\ln(2)}{k} = \frac{0.693}{k} \)