

Preparation for the Chem 1C second midterm

Chem 1C (Kahn, Spring 2006) second midterm covers Chapters 16 (starting from 16.10), 17, 20, and 21. The questions on the exam test your knowledge and understanding of topics covered in my lectures.

Questions and grading. The second midterm has **six** multiple-choice questions and **four** problem-solving questions. The six multiple-choice questions, each **5** points, focus on your knowledge of material. The four problem-solving questions, **(8, 8, 8, 6 points)**, require that you apply the learned concepts to solve practical problem and typically include calculations. Partial credit of 50% will be given for the four 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.

Alternative early testing.

If you cannot attend the regular test because of family or friends situation, I will provide an early test on Thursday, May 25 at starting at 4:50 PM in PSB-N 2619. The format of the early test will be different: instead of five multiple-choice questions you will have four essay questions and one multiple choice question. You will still have three problem-solving questions.

Study tips.

See the syllabus for advice and for the list of practice problems. Your individual practice is crucial for successfully doing the problem-solving questions. I have emphasized simple concentration calculations on the caffeine problems and pop quizzes and you need to know how to do these in order to solve the problems on the test. However, the test questions will be more difficult than the simple examples we did on Quiz 4.

Key concepts.

1. Kinetic energy of molecules and temperature
2. The equilibrium between vapor and liquid phase;
3. Vapor pressure, its temperature dependence
4. Heat of vaporization and heat of fusion, relation to molecular structure
5. Heating curves, boiling and melting points
6. Relation between vapor pressure, atmospheric pressure and phase change
7. Superheating and supercooling
8. Phase diagram of water. Triple point, critical point
9. Negative slope of the solid-liquid curve: origin and consequences
10. Solution composition: percent, mole fraction, molarity, molality
11. Thermodynamic analysis of dissolving ionic solids, factors determining solubility
12. Thermodynamic analysis of dissolving nonpolar solutes
13. Factors affecting solubility of gases; Henry's law
14. Effect of nonvolatile solutes on the vapor pressure of water, Raoult's law
15. Ideal and nonideal mixtures of volatile liquids, exo and endothermic mixing
16. Boiling point elevation and freezing point depression by a nonvolatile solute
17. Osmotic pressure due to a nonvolatile solute
18. Living cells in isotonic, hypertonic, and hypotonic solutions
19. Dialysis
20. Dissociation and association of solutes, van't Hoff factor.
21. Nature and properties of colloids, light scattering and the Tyndall effect
22. Forces affecting stability of colloids, coagulation
23. Nature and properties of ferrofluids
24. Transition metals: defining characters
25. Trends in ionization potential and ionic radius
26. Reactions of transition metals with acids, reduction potential
27. Properties of low oxidation state cations of Cr, Mn, and Fe
28. Oxyanions of Cr(VI) and Mn(VII): structure and properties; redox reactions
29. Coordination compounds: general structure and terminology
30. Complex formation constants as a measure of weak/strong ligand
31. Shapes of s, p, and d orbitals, d^2sp^3 and dsp^2 hybridization
32. Nature of bonding in coordination complexes
33. Crystal field model: splitting of d-energy levels in the octahedral field
34. Crystal field model: splitting of d-energy levels in the tetrahedral field
35. Prediction of color and low-spin/high-spin in coordination complexes
36. Isomerism of coordination compounds
37. Transition metal complexes in organic chemistry and biochemistry
38. Structure of the atomic nucleus
39. Factors determining instability of nuclei; thermodynamic and kinetic stability
40. Radioactive decay by the loss of helium nuclei (α -particles)
41. Radioactive decay via electron (β -particle) or positron emission
42. Neutron-induced nuclear fission, nuclear chain reactions
43. Kinetics of radioactive decay, half-life
44. Nuclear fusion; mass defect and energy from nuclear fusion

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 = $1 \cdot 10^6 \mu\text{L}$
1 M (molar)	1000 mM = $1 \cdot 10^6 \mu\text{M}$ = $1 \cdot 10^9 \text{nM}$
1 atm	760 torr (mmHg) = 1.01325 bar = 101.325 kPa
0 °C	273.15 K

Fundamental constants:

Universal gas constant:	$R = 8.31451 \text{ J K}^{-1} \text{ mol}^{-1}$
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:

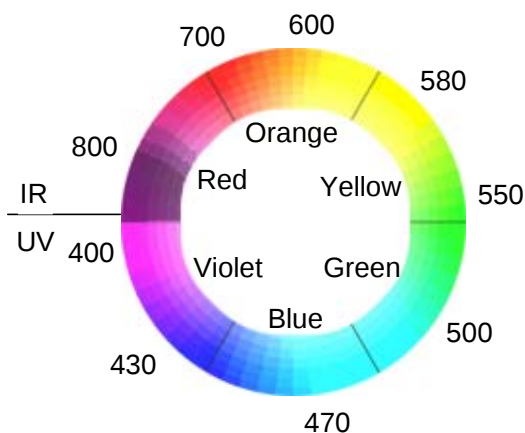
Logarithms	$\ln\left(\frac{A}{B}\right) = -\ln\left(\frac{B}{A}\right)$
	$\ln(A \cdot B) = \ln A + \ln B$
	$\ln(1 - A) \approx -A \quad \text{for } A \ll 1$

Physical relationships:

Energy and mass	$E = m \cdot c^2$
Energy and frequency	$E = h \cdot \nu$
Wavelength and frequency	$\lambda = c/\nu$

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 nm is 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, χ	$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$
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.	$\Delta G^\circ = -RT \cdot \ln K_{eq}$
Free Energy & Potential	$\Delta G^\circ = -n \cdot F \cdot \mathcal{E}^\circ$
Temp. dep. of K	$\ln K = -\frac{\Delta H^0}{RT} + \frac{\Delta S^0}{R}$
Temp. dep. of P_{vap}	$\ln P_{vap} = -\frac{\Delta H_{vap}^0}{R} \cdot \frac{1}{T} + \frac{\Delta S^0}{R}$
Clausius-Clapeyron relation.	$\ln \left(\frac{P_{vap}^{T1}}{P_{vap}^{T2}} \right) = \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right) = -\frac{\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$
Henry's law	$P_{gas} = k_H \cdot \chi$
Raoult's law (nonvolatile)	$P_{solution} = \chi_{solvent} P_{solvent}^0$
Raoult's law (ideal volatile)	$P_{solution} = \chi_A P_A^0 + \chi_B P_B^0$
Boiling point elevation	$\Delta T = i \cdot K_b \cdot m_{solute}$
Freezing point depression	$\Delta T = i \cdot K_f \cdot m_{solute}$
Boiling point elevation const.	$K_b = \frac{RT_b^2 M_{solvent} [g/mol]}{1000 [g/kg] \cdot \Delta H_{vap}}$
Freezing point depr. const.	$K_b = \frac{RT_f^2 M_{solvent} [g/mol]}{1000 [g/kg] \cdot \Delta H_{fus}}$
Osmotic pressure	$\pi = i \cdot M \cdot R \cdot T$
Natural decay	$\ln \left(\frac{N_t}{N_o} \right) = -kt$
Half-life	$\tau = \frac{\ln(2)}{k} = \frac{0.693}{2}$