**Chemistry 1B**
2012W Midterm 2

- **Chapter 3.7-3.10**: Chemical Equations, Stoichiometry and Reaction Yields
- **Chapter 6**: Thermochemistry
- **Chapter 11**: Intermolecular Forces, Liquids and Solids
- **Chapter 12**: Physical Properties of Solutions

**Mid Term Exam Format**

- 45 min closed book exam.
- You may bring pens, pencils, calculator, and an info card.
- Graphics calculator may be used in midterm II and final.
- 4 multi-part problems on 4 pages.
- Can score 110 out of 100 points if you do it all correctly
- 10 minutes per page; pace yourself accordingly.
- No multiple choice questions.
- A problem involving calculations is considered solved when both the solution and the correct final answer with correct units and correct number of significant digits are provided.
- If explanations are requested (by explicitly asking you to “Explain.”) they must be provided.
- Solutions will not be posted but the grading rules will be.

**Exam Timing & Logistics**

- Arrive early (before 8 am) on Friday, March 2, 2012.
- Leave your belongings in front of the lecture hall.
- Seats will be assigned and posted in the back of the room.
- No cell phones or other telecom equipment allowed.
- Proctors will not answer any questions during examinations.
- Rigid timing:
  - 8.00 - 8.03 am: distributing the exams
  - 8.03 – 8.05 am: filling out the cover sheet
  - 8.05 am: exam starts ☑️ (may start earlier)
  - During: checking IDs
  - 8.50 am: exam ends ☑️
  - 8.50 – 8.55 am: Exams collected
  - 8.05 – 8.50 am: students may not enter or leave the room

**Exam Package**
The front page will look like this.

Fill out your name and bubble in your ID as soon as you get it, and then go to the next page.

**Exam Rules Page**
The second page will look like this:

Please fill in your name and ID number in it as well and wait for instructions to start the exam.

**Info Page**
The third page will contain information you will need during this exam. It will look something like this.

Study this page before the exam!!! A copy is available on the website.

If you think this is not enough info, make your own "3x5" index card with info. You will be allowed to bring one such card.
Summary
Things you should bring:
- Calculator
- Pen or pencil
- UCI student ID
- One 3 inch x 5 inch card with notes

Summary
Things you should leave at home:
- any other books or notes
- Cell phones

If any cell phone conversation is noticed, you will be asked to leave.

Significant Digits and Units
Actual grading instructions for midterm 2:
- **7 pts** if a solution is provided, the answer is correct, has significant digits, correct sign and units. Numerical value must be within 1% of the correct solution. The sign must be correct:
  - Remove 2 pts if the number of significant digits is incorrect
  - Remove 4 pts if the units are missing or incorrect (the answer must be in Joules, as requested)
- 0 pts if the solution is missing or the answer is incorrect (mistake in sign → 0 pts).

Significant Digits
How may significant digits do the following numbers have?

<table>
<thead>
<tr>
<th>Number</th>
<th>Significant Digits</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.02</td>
<td>Three digits</td>
</tr>
<tr>
<td>1.00</td>
<td>Three digits</td>
</tr>
<tr>
<td>1.2</td>
<td>Two digits</td>
</tr>
<tr>
<td>1.0</td>
<td>Two digits</td>
</tr>
<tr>
<td>1 (integer)</td>
<td>Infinite precision</td>
</tr>
<tr>
<td>0.0015</td>
<td>Two digits</td>
</tr>
<tr>
<td>0.0015000</td>
<td>Five digits</td>
</tr>
<tr>
<td>3.1×10^5</td>
<td>Two digits</td>
</tr>
<tr>
<td>2.780×10^4</td>
<td>Four digits</td>
</tr>
</tbody>
</table>

Concentration Units
Concentration can be expressed in terms of molarity, molality, and fractions:

- **M** = \( \frac{\text{moles of solute}}{\text{volume of solution}} \)
  - Units: \( \text{M} = \text{mol L}^{-1} \)
  - Also used: \( \text{mM}, \text{nM}, \ldots \)

- **m** = \( \frac{\text{moles of solute}}{\text{kg of solvent}} \)
  - Units: \( \text{m} = \text{mol kg}^{-1} \)

- **X** = \( \frac{\text{moles of solute}}{\text{moles of solute} + \text{solvent}} \)
  - Unitless

- **w** = \( \frac{\text{kg of solute}}{\text{kg of solution}} \)
  - Unitless
  - Sometimes expressed in:
    - \( \%w = w \times 100 \)
    - ppm = \( w \times 10^6 \)

What is the molality, weight fraction, and molar fraction of 10.0 mL of a 1.16 M solution of sucrose (molecular weight 342.3 g/mol) in water (molecular weight = 18.02 g/mol) with a density of 1.18 g/mL?

- Moles of solute:
  \[ n_{\text{solute}} = M \times V = 1.16 \times 10^{-2} \text{ mol} \]
- Mass of solute:
  \[ m_{\text{solute}} = n_{\text{solute}} \times MW_{\text{solute}} = 1.16 \times 10^{-2} \text{ mol} \times 342.3 \text{ g/mol} = 3.97 \text{ g} \]
- Mass of solution:
  \[ m_{\text{solution}} = m_{\text{solute}} + m_{\text{solvent}} = 11.80 \text{ g} - 3.97 \text{ g} = 7.83 \text{ g} \]
- Mass of solvent:
  \[ m_{\text{solvent}} = m_{\text{solution}} - m_{\text{solute}} = 7.83 \text{ g} / 18.02 \text{ g mol}^{-1} = 0.434 \text{ mol} \]
- Moles of solvent:
  \[ 1.16 \times 10^{-2} \text{ mol} / 0.434 \text{ mol} = \frac{1.48 \text{ mol}}{7.83 \times 10^{-2} \text{ kg}} \]
- \( w = \frac{3.97 \text{ g}}{11.80 \text{ g}} = 0.336 \)
- \( X = \frac{1.16 \times 10^{-2} \text{ mol}}{\frac{1.48 \text{ mol}}{7.83 \times 10^{-2} \text{ kg}}} = 0.0267 \)
Factors Affecting Solubility

What affects solubility?

Molecular structure of the solvent and solute (if molecules of solute and solvent can have strong intermolecular interactions with each other they are more likely to dissolve in each other; like dissolves like)

Temperature (solubility of gases decreases; solubility of solids and liquids usually increases)

Pressure (Henry's law for gases)

Henry's law: equilibrium concentration of a gas in a liquid is directly proportional to the partial pressure of this gas

\[
C \text{ [mol L}^{-1} \text{]} = k_H \text{ [mol L}^{-1} \text{ atm}^{-1} \text{]} \times P \text{ [atm]}
\]

\(k_H\) = Henry's constant

What is the weight fraction in ppm of \(O_2\) in an aqueous solution that is in equilibrium with air at 1.00 atm? The fraction of \(O_2\) in air is 21.0%, and \(k_H\) = 0.00130 mol L\(^{-1}\) atm\(^{-1}\). Molecular weight of oxygen is 32.0 g/mol.

Assume the density of the solution is 1.000 g/mL, same as water.

\[
P_{O_2} = 0.210 \times 1.00 \text{ atm} = 0.210 \text{ atm}
\]

\[
[O_2(aq)] = 0.210 \text{ atm} \times 0.00130 \text{ M atm}^{-1} = 2.73 \times 10^{-4} \text{ M}
\]

Let us assume that we have 1.000 L of solution = 1.000 kg 1000 g since (because the density is the same as for water)

\[
\text{mass}_{O_2} = 2.73 \times 10^{-4} \text{ M} \times 32.0 \text{ g mol}^{-1} = 0.00874 \text{ g}
\]

\[
\omega_{O_2} = 0.00874 \text{ g/1000 g} = 8.74 \times 10^{-6} = 8.74 \text{ ppm}
\]

Factors Affecting Solubility

Which compound is the most soluble in water?

\(\text{CH}_3-\text{CH}_2-\text{CH}_3\) \(\text{CH}_3-\text{CH}_2-\text{Cl}\) \(\text{CH}_3-\text{CH}_2-\text{OH}\)

What interactions with water are possible?

Dipole-induced Dipole-dipole Dipole-dipole Dispersion

Dispersion Hydrogen bond Dipole-induced dipole

Which interaction is likely to be the strongest?

Hydrogen bond > dipole-dipole > dispersion, dipole-induced dipole

Therefore \(\text{CH}_3-\text{CH}_2-\text{OH}\) will have the strongest interactions with water, and the largest solubility.

Colligative Properties

Properties of solutions that depend on the number of solute molecules but not on their nature.

Boiling point elevation:

\[
\Delta T_b = i K_b m
\]

Freezing point depression:

\[
\Delta T_f = i K_f m
\]

Osmotic pressure:

\[
P = i MRT
\]

Vapor pressure lowering:

\[
P_1 = X_1 P_1
\]

\(i\) = actual number of particles upon solvation

\(n\) = number of units dissolved in solution

The van 't Hoff factor indicates the actual number of particles that contribute to the colligative properties.

Colligative Properties

What mass of ethylene glycol (\(\text{C}_2\text{H}_6\text{O}_2\), molecular weight 62.1 g/mol) must be added to 1.00 L of water to produce a solution that freezes at -8.00 °C? Assume the density of water is 1.00 g/mL. \(K_f = 1.86 \text{ °C kg/mol}\)

\[
\Delta T = 8.00 \text{ °C} - 1.86 \text{ °C kg/mol} m = 4.30 \text{ mol/kg}
\]

\[
r_{\text{aqueous}} = m \times \text{mass}_{\text{aqueous}} = 4.30 \text{ mol/kg} \times (1.00 \text{ L} \times 1.00 \text{ kg/L}) = 4.30 \text{ mol}
\]

\[
\text{mass}_{\text{aqueous}} = 4.30 \text{ mol} \times 62.1 \text{ g/mol} = 267 \text{ g}
\]

Internal Energy, Heat and Work

The energy of the universe is constant.

\[
\Delta E = q + w
\]

\(w = -P \Delta V\)

Sign convention: if heat or work goes into a system, the sign is positive.

"What is the System?" – is the most important question you need to answer to get the signs right.
Oxidation of Na in dry air at 1.00 atm pressure generated 98.0 kJ of heat and consumed 25.0 L of oxygen. What is the internal energy change in the reaction system?

\[ 4 \text{Na(s)} + \text{O}_2 \text{(g)} \rightarrow 2 \text{Na}_2\text{O(s)} \]

Our “system” is Na(s) that is being converted into Na_2O by the reaction with atmospheric oxygen.

\[ q = -98.0 \text{kJ} \]

The sign is negative because the heat is released by the system and goes into the surroundings.

\[ w = -P \Delta V \]

The sign is positive because the surroundings do work on the system by supplying gas to it.

\[ (1 \text{ atm} \times 101.3 \text{ L}) \]

\[ \Delta E = q + w = -98.0 \text{kJ} + 2.53 \text{kJ} = -95.5 \text{kJ} \]

Negative sign means that the system lost internal energy during the reaction.

\[ \Delta H_{\text{evap}} = 40.80 \text{kJ/mol} \]

You have a beaker containing 20.00 g of ice cubes floating in 200.0 g of water (molecular weight = 18.02 g/mol at 0°C (normal melting point of ice)). You are adding 5.000 g of pure sulfuric acid to this mixture (molecular weight = 98.08 g/mol). The following information is provided to you:

- The heat of dissolution of sulfuric acid is \( \Delta H_{\text{soln}} = -96.30 \text{kJ/(mol H}_2\text{SO}_4) \).
- The specific heat of liquid water, \( C_{\text{m(liquid)}} = 4.184 \text{J/(g x K)} \)
- The specific heat of water vapor, \( C_{\text{m(vapor)}} = 1.990 \text{J/(g x K)} \)
- Heat of evaporation, \( \Delta H_{\text{evap}} = 40.80 \text{kJ/mol} \)
- Density of liquid water and solution: assume 1.000 g/mL

### Q1. How much heat is required to melt all the ice in the beaker?

**What is the system?** – it is our beaker with water and ice.

Moles of ice: \[ n_{\text{ice}} = \frac{m_{\text{ice}}}{MW_{\text{ice}}} = \frac{20.00 \text{g}}{18.02 \text{g/mol}} = 1.110 \text{mol} \]

Heat required to melt all of it:

\[ q_{\text{fus}} = n_{\text{ice}} \times \Delta H_{\text{fus}} = 1.110 \text{mol} \times 6.010 \text{kJ/mol} = 6.670 \text{kJ} \]

The sign is taken as positive since heat must go into the system = ice.

### Q2. How much heat will be released by dissolving sulfuric acid in the water/ice mixture?

Moles of \( \text{H}_2\text{SO}_4 \):

\[ n_{\text{SA}} = \frac{m_{\text{SA}}}{MW_{\text{SA}}} = \frac{5.000 \text{g}}{98.08 \text{g/mol}} = 0.05098 \text{mol} \]

Heat released by the dissolution:

\[ q_{\text{soln}} = n_{\text{SA}} \times \Delta H_{\text{soln}} = 0.05098 \text{mol} \times 96.30 \text{kJ/mol} = 4.909 \text{kJ} \]

The sign is taken as positive since heat goes into the system: our solution warms up as a result of the dissolution.

### Q3. Will there be any ice left in the beaker after adding sulfuric acid?

Heat released by the dissolution:

\[ q_{\text{soln}} = 4.909 \text{kJ} \]

Heat required to melt all ice:

\[ q_{\text{fus}} = 6.670 \text{kJ} \]

Conclusion: some ice will remain because there not enough heat supplied into the system by the dissolution of sulfuric acid to melt all of the ice.

### Q4. If your answer to the previous question is “yes” how much ice will be left? If your answer is “no” what will be the temperature of the solution?

The answer was "yes", some ice will remain. The moles of ice we can melt:

\[ n_{\text{melted}} = \frac{q_{\text{soln}}}{\Delta H_{\text{fus}}} = \frac{4.909 \text{kJ}}{6.010 \text{kJ/mol}} = 0.8168 \text{mol} \]

Mass of melted ice:

\[ m_{\text{melted}} = n_{\text{melted}} \times MW_{\text{ice}} = 0.8168 \text{mol} \times 18.02 \text{g/mol} = 14.70 \text{g} \]

Remaining ice:

\[ = 20.00 \text{g} - 14.70 \text{g} = 5.300 \text{g} \]
Energy, Enthalpy and Calorimetry

If a reaction happens under constant volume, then $\Delta V = 0$

$$\Delta E = q_v$$

The internal energy is a good state function for measuring heat during a chemical reaction. A 'bomb' calorimeter (V constant) measures $\Delta E = q_v$

If a reaction happens under constant pressure, then the internal energy is not a good state function, but enthalpy is:

$$\Delta H = q_p$$

A 'coffee cup' calorimeter (P constant) measures the enthalpy change, which corresponds to the heat during a chemical reaction.

Constant Pressure Calorimetry

In a coffee-cup calorimeter, 750 mL of 0.100 M silver nitrate reacts with 750 mL of 0.100 M sodium chloride to produce solid silver chloride. The two solutions were initially at 20.0°C, and after the reaction the mixture is at 20.8°C. Calculate the enthalpy of reaction in kJ/mol AgCl. Assume that the combined solution has specific heat capacity $C_p = 4.18 \text{J/}^\circ\text{C} \times g$ and a density of 1.0 g/mL.

1. Write a balanced chemical equation.

$$\text{Ag}^+ (aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl(s)}$$

2. Use the equation for $q_p$ and the info given to find the heat gained by the solution.

$$\Delta T = 20.8 - 20.0 = 0.8^\circ\text{C}$$

$$m_{water} = (750\text{mL} + 750\text{mL}) \times \frac{1\text{g}}{1\text{mL}} = 1500\text{g}$$

$$q_p = 4.18 \times \frac{J}{^\circ\text{C} \times g} \times 1500\text{g} \times 0.8^\circ\text{C} = 5021\text{J}$$

Constant Volume Calorimetry

If a reaction happens under constant pressure, then the internal energy is:

$$\Delta E = q_v$$

A. 0.1584 g sample of benzoic acid (C$_7$H$_6$O$_2$, molecular weight 112.1 g/mol) is burned in a bomb calorimeter with a heat capacity of 1.65 kJ/°C. The temperature of the calorimeter increases by 2.54 °C. Calculate the energy of combustion per mole benzoic acid.

1. Identify the heat transfer process

$$|q_v\text{ lost by benzoic acid}| = |q_v\text{ gained calorimeter}|$$

2. Calculate the heat gained by the calorimeter.

$$q_v(\text{ calorimeter}) = \frac{1.65\text{kJ}}{^\circ\text{C}} \times 2.54^\circ\text{C} = 4.19\text{kJ}$$

3. Calculate the $\Delta E$ in kJ/mol benzoic acid.

$$\Delta E = \frac{-4.19\text{kJ}}{\text{moles benzoic acid}} = \frac{-4.19\text{kJ}}{(0.1584\text{g}/112.1\text{g/mol})} = -3.23 \times 10^3 \text{kJ/mol}$$

Hess’s Law: State Functions are Additive

$$\text{CO(g)} + \frac{1}{2} \text{O}_2(g) \rightarrow \text{CO}_2(g) \quad \Delta H = -283\text{kJ}$$

$$\text{N}_2(g) + 2\text{O}_2(g) \rightarrow 2\text{NO}_2(g) \quad \Delta H = -180.6\text{kJ}$$

Use the information above to calculate $\Delta H$ for the following reaction:

$$\text{CO(g) + NO(g) } \rightarrow \frac{1}{2} \text{N}_2(g) + \text{CO}_2(g)$$

Reorder the sub-reactions with reactants to the left and products to the right, and correct for stoichiometries. Add $\Delta H$ values appropriately.

$$\text{CO(g) + NO(g) } \rightarrow \frac{1}{2} \text{N}_2(g) + \text{CO}_2(g) \quad \Delta H = -283\text{kJ}$$

$$\text{NO}(g) \rightarrow \frac{1}{2} \text{N}_2(g) + \frac{1}{2} \text{O}_2(g) \quad \Delta H = -\frac{1}{2}(180.6\text{kJ})$$

$$\text{CO(g) + NO(g) } \rightarrow \frac{1}{2} \text{N}_2(g) + \text{CO}_2(g) \quad \Delta H = -373\text{kJ}$$

Standard Enthalpies of Formation

$\Delta H^\circ$ = change in enthalpy occurring during the formation of one mole of a compound from its elements with all substances in their standard states. [kJ/mol]

Standard state for an element:

- The most stable state in which the element exists under conditions of $P = 1 \text{ atm}$ and $T = 25^\circ\text{C}$

For a compound:

- Gaseous substance: $P = 1 \text{ atm}$ and $T = 25^\circ\text{C}$
- Liquid or solid substance: pure liquid or solid
- Substance in solution: concentration of exactly 1 M.
Standard Enthalpies of Formation

Using the following enthalpies of formation:

<table>
<thead>
<tr>
<th>Substance</th>
<th>( \Delta H_f^° ) (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>NO(g)</td>
<td>90</td>
</tr>
<tr>
<td>N(_2)O(_5)(s)</td>
<td>-42</td>
</tr>
<tr>
<td>NO(_2)(g)</td>
<td>34</td>
</tr>
</tbody>
</table>

Note: there is no need to specify \( \Delta H_f^° \) for O\(_2\)(g) because it is 0, by definition of a standard state.

Calculate the standard reaction enthalpy of the following reaction:

\[
\text{NO(g) + NO}_2\text{(g) + O}_2\text{(g) \rightarrow N}_2\text{O}_5\text{(s)}
\]

\[
\Delta H_f^° = \sum n_p \Delta H_f^° \text{ (products)} - \sum n_r \Delta H_f^° \text{ (reactants)}
\]

\[
= (-1 \text{ mol} \times 42 \text{ kJ/mol}) - (1 \text{ mol} \times 90 \text{ kJ/mol} + 1 \text{ mol} \times 34 \text{ kJ/mol})
\]

\[
= -166 \text{ kJ/mol}
\]

Also Review all the MT 1 Material !!!!

Molecular Interactions

What are examples of van der Waals intermolecular forces?

Dipole-dipole (H-bonding is an example)

Ion - induced dipole

Dispersion forces

Molecular Interactions

What kind of forces hold the ions in NaCl together?

Coulombic (electrostatic) forces

Can there be attractive forces between two neutral and nonpolar molecules? Yes, these are called dispersion forces. Temporary fluctuations in the molecules create dipoles that can, in turn, induce temporary dipole moments in neighboring molecules.

Which species has a higher boiling point, MgCl\(_2\) or PCl\(_5\)? MgCl\(_2\) consists of a lattice held together by ionic bonding (Coulombic) forces. PCl\(_5\) molecules are held together by dispersion forces. Ionic forces are stronger than dispersion forces, so MgCl\(_2\) is expected to have a higher boiling point.

Solid Structures

<table>
<thead>
<tr>
<th>Atomic crystal</th>
<th>Ionic crystal</th>
<th>Molecular crystal</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metals</td>
<td>Solid salts</td>
<td>Hard materials</td>
</tr>
<tr>
<td>Atomic networks</td>
<td>Electrostatic interactions</td>
<td>High melting points</td>
</tr>
<tr>
<td>Covalent linkages</td>
<td>Van der Waals interactions</td>
<td>Soft materials, low melting points</td>
</tr>
</tbody>
</table>

Crystal structures

The density of palladium is 12.0 g/cm\(^3\) and the cell volume is 5.89 x 10\(^{-23}\) cm\(^3\). Determine the number of atoms in the unit cell. Note: the number of atoms in a unit cell is a whole number. Atomic weight of palladium is 106 g/mol.

Determine the mass of the unit cell:

\[ m = d \times V = 12.0 \text{ g/cm}^3 \times 5.89 \times 10^{-23} \text{ cm}^3 \]

\[ = 7.07 \times 10^{-22} \text{ g} \]

Determine the number of moles:

\[ n = \frac{7.07 \times 10^{-22} \text{ g}}{106 \text{ g/mol}} = 6.67 \times 10^{-24} \text{ mol} \]

Determine the number of atoms:

\[ n_{\text{atoms}} = \frac{6.67 \times 10^{-24} \text{ mol}}{6.022 \times 10^{23} \text{ mol}^{-1}} = 4.02 \]

\[ = 4 \]

Types of Solids

Give a qualitative explanation for the high density (mass per volume) of metals.

The valence electrons in metals occupy molecular orbitals that are delocalized. This allows rapid transport of charges across the lattice.

Give a qualitative explanation for the high electric conductivity in metals.

Why is diamond much harder than solid sulfur (S\(_8\))? In the diamond lattice all the atoms are covalently bonded, which is much stronger than the Vanderwaals interactions that hold the S\(_8\) lattice together.
**Phase Transitions**

**Molar heat of fusion** ($\Delta H_{\text{fus}}$) is the energy required to melt 1 mole of a solid substance at its freezing point.

**Molar heat of vaporization** ($\Delta H_{\text{vap}}$) is the energy required to evaporate 1 mole of a liquid substance at its boiling point.

**Molar heat of sublimation** ($\Delta H_{\text{sub}}$) is

$$\Delta H_{\text{sub}} = \Delta H_{\text{fus}} + \Delta H_{\text{vap}}$$

(Hess’s Law)

Clausius-Clapeyron equation:

$$\ln P = \frac{-\Delta H_{\text{vap}}}{RT} + C$$

---

**Phase Diagrams**

At what temperature do these substances change from the solid phase into another phase at 1 atm?

- **Water**: 0°C
- **CO₂**: -78°C

What are the differences between these phase transitions?

- Water: melting point decreases
- CO₂: sublimation point increases

How does the temperature of the phase transition change when the pressure is increased?

- Water: melting (solid to liquid)
- CO₂: sublimation (solid to vapor)

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See you on Friday!