LECTURE WORKSHEETS

Lecture Worksheets

The purpose of the lecture worksheets is to make you aware of some of the basics that will be covered in Chem 110 lecture, in the hopes that you will get more out of lecture through this kind of preparation. It can also serve as a guide in reading the text, increasing your understanding of what you read. Lecture Worksheets are to be turned in at the beginning of the Chem 110 lectures. There will be a box in the front of the lecture hall marked: Chem 108 Lecture Worksheets. Early lecture worksheets will be accepted. Anything past the indicated due date (the date of the Chem 110 lecture to which it corresponds) will not receive credit!

In general, Lecture Worksheets will be graded based on effort rather than results, i.e., if the assignment is done on time and with reasonable care, full credit will be given, even if there are some mistakes in the work.

PLEASE: do not disturb the lecture to turn in your Lecture Worksheet. Be considerate of your peers and Chem 110 lecturer. Turn in the worksheets prior to the beginning of lecture or immediately afterward.

THANK YOU!
Lecture Worksheet 1
(Chapter 1 and 2.1-4)
Name ___________________________
Chem 108 Section #________
Due: Weds. Jan. 16

1. Define the following:
   (a) mixture
   (b) element
   (c) compound

2. (a) Define physical and chemical property and give an example of each.

3. (a) Define physical and chemical change and give an example of each.

4. Using scientific notation, write the following numbers in meters:
   a) 1 km
   b) 26 cm
   c) 200 µm
   d) 760 nm

5. How many significant figures are in the following numbers:
   a) 10.06
   b) 0.00057
   c) 0.4500

6. Describe how you deal with significant figures in:
   (a) multiplication/division problems
   (b) addition/subtraction problems.

Chapter 2.1-4
1. What is an atom according to Dalton’s atomic theory?

2. What is the difference between Dalton’s view of an atom and the modern view? (What are the names of the subatomic particles in an atom and where are they?)

3. (a) What is the difference between the mass number and the atomic number?
   (b) In the symbol, $^{12}\text{C}$ what does the number 12 tell you? What does the number 6 tell you?
Chapter 5.1-3
1. For a chemical reaction in a laboratory, what is the system? What are the surroundings?

2. Define the following:
   (a) heat
   
   (b) State function
   
   (c) Enthalpy

Chapter 6.1-3
1. What is the value of the speed of light (include units)?

2. Give the relationship (an equation) between the speed of light, c, wavelength, λ, and frequency, ν.

3. What is a photon? How do we describe the energy of a photon?

4. What is the difference between a continuous spectrum and a line spectrum?

5. Write the equation that predicts the frequency of absorbed or emitted light when an electron in a hydrogen atom moves between energy levels?
1. If $\psi$ is a wave function of an electron in an atom, what is the physical interpretation of $\psi^2$?

2. What is the Heisenberg uncertainty principle? How does it apply to the structure of an atom?

3. (a) What are the names and letter designations of each of the 3 quantum numbers needed to designate an orbital? What do each of these quantum numbers tell you about the orbital?

   (b) What is meant by the term subshell?

   (c) Which quantum numbers are needed to define a subshell?

5. Draw the following:
   (a) a 1s orbital.

   (b) a 2px orbital next to a 2py orbital. (What is different about these two orbitals?)

   (c) a 2s orbital next to a 3s orbital. (What is different about these two orbitals?)
1. How does the presence of additional (more than one) electrons affect the orbital energy levels?

2. What is the fourth quantum number?

3. Define the Pauli Exclusion Principle.

4. (a) Define Hund's Rule.

   (b) What are two possible ground state electron configurations for a nitrogen atom? Which one is “correct” according to Hund’s rule?

5. What is meant by core and valence electrons?

6. (a) Write the full electron configuration of Br.

   (b) Write the electron configuration of Br using the noble gas shortcut.

7. Given these three atoms in their ground state, indicate the principal quantum number of the electron in the highest energy orbital (outermost occupied shell) for each.

   (a) Li  
   (b) Cl  
   (c) Sn

8. (a) Write the electron configuration of Cr and Cu.

   (b) Why are these anomalous? (See page 249 of the text.)
Chapter 2.5
On the periodic table, what is a period and a group?

Chapter 7.1-6
1. (a) What is the equation used to define the effective nuclear charge of a many electron atom? What do S and Z in the equation mean?

  (b) What would the effective nuclear charge be in a one electron atom or ion?

2. What is the periodic trend for increasing size of atoms in a family? In a period?

3. (a) What is the periodic trend for increasing size of ions of the same charge in a family?

  (b) What is the trend in size for ions with the same electron configuration (isoelectronic series)?

4. (a) Define ionization energy.

  (b) How can you tell from successive ionization energies that a core electron has been removed?

  (c) What is the periodic trend for increasing first ionization energy?

5. What is the electron configuration of Li$^{+}$ and Fe$^{2+}$?


7. Using the following list of elements: C, Na, F, Cs, Ba, Ni.

  (a) Which are metals?

  (b) Which one has the most metallic character? ______

  (c) Which one (from the complete list given above) has the least metallic character? ______

Name ___________________________
Chem 108 Section # ________
DUE: Weds. Jan. 30
Chapter 2.7-8
1. What would be the predicted ionic charge for Potassium? For sulfur? For Bromine?

2. How many electrons do Na\(^+\) and Cl\(^-\) have compared to the neutral atoms?

2. Name the following ions:
   (a) Al\(^{3+}\)  
   (b) S\(^{2-}\)  
   (c) NO\(_3^-\)  
   (d) SO\(_4^{2-}\)

Chapter 8.1-3
1. Define ionic bond and covalent bonds

2. What are valence electrons?

3. State the octet rule. Give an example of how it works in ionic bonding.

4. (a) Define lattice energy.

   (b) Define E, k, Q\(_1\), Q\(_2\), and d in the following equation:

   \[ E = \frac{k Q_1 Q_2}{d} \]

5. What is the electron configuration of Fe and Fe\(^{2+}\)? Is the octet rule satisfied for the Fe\(^{2+}\) ion?

6. What kind of bonds are present between the atoms in the polyatomic ions given in tables 2.4 and 2.5?

7. Compare the length and strength of double and triple bonds.
Chapter 8.4-6

1. Define bond polarity.

2. What is the definition of electronegativity?

3. Why is sodium fluoride written as Na\(^+\) F\(^-\), but hydrofluoric acid written as H\(\delta^+\) F\(\delta^-\)?

4. What is a dipole moment? (See figure 8.8)

5. Why is hydrogen at the end of the formula for MgH\(_2\) and at the beginning of H\(_2\)S? (What are the differences in these compounds that effect their names and molecular formulas?)

6. a) Illustrate the steps for drawing a Lewis structure for SO\(_4^{2-}\), writing down something for each step.

7. What are the formal charges of the S atom and the oxygen atoms in your Lewis structure of SO\(_4^{2-}\)? Be sure to show how you determine formal charge.

8. Draw all resonance structures for NO\(_3^-\). Compare the length and strength of the 3 NO bonds in NO\(_3^-\).

9. (a) Resonance is an important concept in describing the bonding of which class of hydrocarbons? What is the name and structure of the simplest molecule of this type?

(b) What is special about this bonding arrangement?
Lecture Worksheet 8  
Name ___________________________  
(Chapter 2.9, 8.7-8)  
Chem 108 Section # _______  
Due Fri. Feb. 8  

Chapter 2.9  
1. What is a hydrocarbon? Draw the structural formula of two simple hydrocarbons.

2. What is an alcohol? Draw the structure of two alcohols that are derived from the two hydrocarbons you just drew.

Chapter 8.7-8  
1. a) Draw the Lewis structure for BF$_3$.

   b) Does this molecule obey the octet rule?

   c) Why does BF$_3$ react with molecules that have a lone pair?

2. Explain why PCl$_5$, PCl$_3$, and NCl$_3$ are known compounds, whereas NCl$_5$ is not.

3. Consider any molecule made up of a central atom with an expanded octet surrounded by other atoms.

   (a) What are the characteristics of the central atom? (size, electronegativity, etc.)

   (b) What are the characteristics of the surrounding atoms?

4. (a) Define bond dissociation energy.

   (b) Why is this energy ALWAYS endothermic?

5. Considering only single bonds, how does a bond's length vary with its strength?
1. What is organic chemistry?

2. What is a hydrocarbon? Draw the structural formula of two simple hydrocarbons.

3. What is an alcohol? Draw the structure of two alcohols that are derived from the two hydrocarbons you just drew.

4. An alkane contains only C-C single bonds. What is the class of compounds called that contains C-C double bonds? Draw and name one example.

5. What is the class of compounds called that contains C-C triple bonds? Draw and name one example.
Chapter 2.6

1. What does the subscript in a molecular formula tell you?

2. How can you tell if a compound will be molecular?

3. What is the difference between a molecular formula and an empirical formula?

4. Give the molecular formula and structural formula of ethane.

Chapter 3.3-5

1. What is the formula weight of Mg(NO₃)₂? What is the % of Mg by mass in Mg(NO₃)₂?

2. What is Avogadro's number? What is a mole?

3. What is the procedure to convert:
   (a) grams to moles?

   (b) grams to molecules?

4. What is the difference between molecular formula and empirical formula? (Illustrate using benzene C₆H₆.)
1. What are the 5 fundamental geometries on which the molecular shapes are based? What are the bond angles associated with these geometries?

2. What is an electron domain? How many electron domains are there about the S in SO₂?

3. According to VSEPR theory, what is the best arrangement of electron domains?

4. (a) What is the difference between electron-domain geometry and molecular geometry?

   (b) Give the electron domain geometry and molecular geometry of NH₃. (Draw and name them.)

5. Why is the HOH angle in water 104.5° even though the electron-domain geometry is tetrahedral and the tetrahedral angle is 109.5°?

6. What role do triple and double bonds play in predicting geometry?
1. Define bond dipole:

2. What two features of a molecule with more than two atoms can be used to decide if the molecule has a dipole moment?

3. In the given molecules: (a) draw the Lewis structures; (b) indicate the electron domain (EDG) and molecular (MG) geometries; (c) show the bond dipoles (BDP) (using arrows); and (d) determine whether the molecule is polar or nonpolar.

<table>
<thead>
<tr>
<th>Lewis structure</th>
<th>EDG</th>
<th>MG</th>
<th>BDP</th>
<th>Polarity</th>
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</thead>
<tbody>
<tr>
<td>(a) CO₂</td>
<td></td>
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<td></td>
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<tr>
<td>(b) H₂O</td>
<td></td>
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<tr>
<td>(c) HCl</td>
<td></td>
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<tr>
<td>(d) BF₃</td>
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</tbody>
</table>
1. According to valence bond theory, what causes a covalent bond?

2. Mixing 3 atomic orbitals produces how many hybrid orbitals? __________

3. If one s atomic orbital and two p atomic orbitals are mixed to form hybrid orbitals, what is the name of the hybrid orbitals which form? __________

4. Does the geometry of the hybrid orbitals of an atom parallel its molecular geometry or its electron pair geometry? __________

4. What are the names of the hybrid orbitals use by atoms that have expanded octets?

5. (a) What are the three steps used to predict the hybrid orbitals used by an atom in bonding?

(b) Illustrate these steps by showing how you would determine the hybridization of N in NH$_3$?

6. (a) What is a sigma bond? (How is a sigma bond formed?)

(b) What is a pi bond? (How is a pi bond formed?)

7. What are delocalized bonds? (Which kind of bonds (sigma or pi) are used in forming delocalized bonds?)
1. Do intermolecular forces effect melting points (and boiling points)? If so how?

2. (a) Name three intermolecular attractive forces known to exist between neutral molecules.

   (b) What intermolecular force is important in solutions of ionic solids in polar liquids?

3. (a) What is the meant by an instantaneous dipole?

   (b) What is the name of the attractive interaction produced by instantaneous dipoles?

4. (a) Define polarizability.

   (b) How does a molecule's polarizability vary with its size?

5. Define hydrogen bonding.

6. Use the FLOW CHART on page 452 (Fig. 11.12) of your text to determine the intermolecular forces between the molecules:

   methanol (HO-CH₃), PF₅, NF₃, NH₄Cl, benzene (C₆H₆).
1. (a) Explain how a barometer measures atmospheric pressure.

(b) Explain how an open end manometer measures the pressure of an enclosed gas.

2. (a) What conditions for a gas must be defined for Boyle's law to be true?

(b) Under these conditions, how does the pressure of a gas relate to its volume? (Use both an equation and plot to show this.)

3. (a) Give an equation and plot that represents Charles's law.

(b) Give an equation and plot that represents Avogadro's law.

4. What conditions represent STP?

5. What is so special about 22.4 liters of gas at STP?

6. Under what conditions does $P_1 V_1 / T_1 = P_2 V_2 / T_2$?

7. The ideal-gas equation has many applications.
   (a) Give the equation for the density of a gas.

   (b) Give the equation for determining the molar mass of a gas.
1. Air is composed of a mixture of gases.
   (a) What did Dalton observe while studying the properties of air?

   (b) Give the equation that demonstrates his observation.

2. Define mole fraction:
   (a) in terms of moles.

   (b) in terms of partial pressure of gases.

3. (a) What property of a gas depends only on temperature?

   (b) How does this property depend on temperature?

4. How is the root-mean-square speed of a molecule related to its average kinetic energy?

5. (a) What is the difference between light and heavy molecules in terms of rms speed?

   (b) Give the equation that describes this phenomenon.

6. (a) State Graham's law of effusion (in words and by an equation).

   (b) How does rate of effusion relate to rms speed?

7. Under what conditions of pressure and temperature do real gases most closely approximate ideal-gas behavior?
1. (a) Name the regions of the atmosphere.

   (b) Why is it organized in this way?

2. (a) Give the chemical reaction for the **photodissociation** of \( \text{O}_2 \).

   (b) Give the chemical reaction for the **photoionization** of \( \text{N}_2 \).

   (c) Why are these thermosphere reactions important to life on earth?

2. Why is the concentration of \( \text{O}_3 \) fairly large in the stratosphere (the region between 30 and 90km)?

3. (a) Why is the presence of ozone in the stratosphere important to us?

   (b) Why does the presence of CFC's in the stratosphere interfere with this process?

4. (a) Give the chemical reaction that shows how troposphere \( \text{SO}_3 \) is responsible for acid rain.

   (b) Give at least two environmental problems caused by acid rain.

5. Why is carbon monoxide (CO) poisonous?

6. What atmospheric gases are responsible for photochemical smog?

7. (a) Briefly explain the greenhouse effect.

   (b) What two gases are responsible for it?
Chapter 5.5
1. Define:
   (a) Heat capacity
   (b) Specific heat
   (c) Calorimeter

Chapter 11.4
2. (a) List three phase changes that absorb energy. (These processes are endothermic.)

   (b) List three phase changes that give off energy. (These processes are exothermic.)

3. What is the name of the process where a solid is converted to a gas?
   Give the name of a common substance that undergoes this process at 1 atm pressure.

4. Define heat of fusion and heat of vaporization.

5. Why does a substance's heating curve have a plateau at the substance's melting and boiling point?

6. Define critical pressure and temperature.
1. (a) How does the viscosity of a liquid vary with the strength of its intermolecular forces?

   (b) What is surface tension? Does it change as intermolecular forces increase?

2. (a) Define dynamic equilibrium.

   (b) Give an example of a dynamic equilibrium.

3. (a) Define vapor pressure. How does vapor pressure vary with intermolecular forces?

   (b) What is the vapor pressure of H\textsubscript{2}O at 25\textdegree C? (Use the table in Appendix B and/or Figure 11.24 in your text.)

4. (a) Define normal boiling point and normal melting point.

   (b) What is the normal boiling point of diethyl ether?

5. Using the phase diagrams on page 464 in your text (figure 11.27) to answer the following:

   (a) What is the critical point of CO\textsubscript{2}? \\

   (b) What is the triple point of water?

   (c) What state is carbon dioxide in when the P= 1 atm and T = −100\textdegree C?

   (d) When P = 73 atm and T = 25\textdegree C?
1. Is the solution process exothermic or endothermic? What is the role of entropy (or disorder) in the solution process?

2. (a) Define saturated. How does dynamic equilibrium apply to saturated solutions? Give an example.

   (b) What is solubility?

3. What does "like dissolves like" mean?

4. How does the partial pressure of a gas affect the solubility of that gas?

4. (a) How does temperature affect the solubility of ionic salts in H₂O?

   (b) How does temperature affect the solubility of gases in water?

5. Give definitions of the following concentration units.
   (a) mass percent (%)  (b) parts per million (ppm)  (c) molality (m)

   (d) **Molarity** and **molality** are very similar: One is the moles per volume of solution, the other is moles per kg of solvent. What is the reason for these two different concentration units? (See page 545 of your text.)

**Chapter 4.1 & 5**

1. (a) What happens to an ionic compound when it dissolves in water?

   (b) What happens to a molecular compound when it dissolves in water? (READ this carefully! How is it DIFFERENT from what happens to an ionic compound?)

2. What is an electrolyte? What is the difference between a strong electrolyte, a weak electrolyte and a nonelectrolyte? Give an example of each.
1. (a) What is a colligative property (by definition)?

(b) What are the four colligative properties?

2. What happens to the vapor pressure of an aqueous solution when you dissolve an ionic salt in it?

3. (a) What are the expressions that are used to determine the temperature change of boiling point and freezing point for a solution?

(b) What are the concentration units used in these expressions?

(c) If a 1m solution of NaCl is made, what concentration is used to predict the boiling point of the solution?

4. What is a colloid? Give some examples.

5. What is the difference between hydrophobic and hydrophilic colloids?
1. What is the difference between a subscript and a coefficient in a chemical reaction? Which is changed when balancing a reaction?

2. Give an example of a combination reaction and a decomposition reaction. (Clearly label which is which.)

3. What are the products of the complete combustion of a hydrocarbon in oxygen gas (O_2)?

4. Circle the compounds that are soluble in aqueous solution.
   - KCl, BaSO_4, Mg(OH)_2, PbBr_2, (NH_4)_3PO_4, and Fe(NO_3)_2.

5. (a) Define metathesis reaction. Give an example.

6. (a) What is an acid? What is a base?
   (b) What is the difference between a strong acid (or base) and a weak acid (or base)?
   (c) List the 7 strong acids and the 8 strong bases. (Memorize them!)

7. Using the table on page 133 decide if the following substances are strong, weak or nonelectrolytes?
   (a) H_3PO_4 (phosphoric acid)   (c) CaCl_2
   (b) CH_3OH (methanol)           (d) HBr

8. For the reaction between aqueous hydrochloric acid and aqueous sodium hydroxide:
   (a) Write the molecular equation.
   (b) Write the ionic equation.
   (c) Write the net ionic equation and list the spectator ions.
Lecture Worksheet 23
(Chapter 4.4, 3.6-7)

Chapter 4.4
1. Define oxidation and reduction.

2. What is the oxidation state of silicon in SiCl\textsubscript{4}? of bromine in Br\textsubscript{2}? of nitrogen in NO\textsubscript{3}\textsuperscript{−}(aq)?

3. What is meant by an “active” metal. Given an example of one.

Chapter 3.6-7
1. List two ways of interpreting stoichiometric coefficients in a balanced chemical equation.

2. (a) Define limiting reagent.

   (b) A ham sandwich consist of 2 slices of bread and one piece of ham (2B + H → B\textsubscript{2}H). What is maximum number of identical sandwiches you can make with 10 slices of bread and 7 slices of ham? ________

   What is the limiting reagent? ________

   What is the excess reagent? ________

   (b) What is maximum amount (in moles) of water (H\textsubscript{2}O) that can be made when you start with 10 moles of hydrogen (H\textsubscript{2}) and 7 moles of oxygen (O\textsubscript{2})? ________

   HINT: write a balanced reaction first. It is NOT the same as the sandwich “reaction”.

   What is the limiting reagent? ________

   What is the excess reagent? ________

3. (a) Define theoretical yield.

   (b) Define actual yield.

   (c) Define percent yield.
Chapter 4.6

1. How is mass converted to moles? Show how 1g of NaOH can be converted to moles.

2. (a) Given molarity (M) and volume(V), how do you determine the moles of a solute in solution?

(b) If you have 100 mL of 12M H₂SO₄, how many moles of sulfuric acid would it contain?

(c) How many moles of acid (H⁺(aq)) do you have?

3. (a) What is the equivalence point of a titration?

(b) What is the endpoint of a titration?

Chapter 10.5

4. (a) How are the moles of gas related to experimentally measurable quantities? (What equation do you need?)

(b) How many moles of gas are in a 1.0L flask at STP?

5. Use the following balanced reaction to answer the next questions.

\[ 2\text{NaN}_3(s) \rightarrow 2\text{Na}(s) + 3\text{N}_2(g) \]

(a) If you know the mass of sodium azide (\text{NaN}_3) in a car, outline the procedure to find the volume of \text{N}_2 that will be produced when it decomposes, assuming \( T = 25^\circ \text{C} \) and \( P = 1 \text{ atm} \).

(b) If you know the volume of \text{N}_2 (at \( T = 25^\circ \text{C} \) and 1 atm) that you want in the airbag, outline the procedure to find the mass of sodium azide that must decompose to produce it.
1. What is the difference between enthalpy $\Delta H$ and internal energy $\Delta E$?

2. What is the equation that describes the enthalpy change for a chemical reaction?

3. (a) What is an extensive property?

   (b) Give an example of an extensive property other than enthalpy (see sec 1.3).

4. For the reaction $A + 2B \rightarrow C$ the $\Delta H_{\text{rxn}} = -310 \text{ kJ}$.
   (a) Is the reaction exothermic or endothermic?

   (b) If 2 moles of A react, how much heat will be given off?

   (c) $\Delta H_{\text{rxn}}$ for the reverse reaction is:

5. Define calorimetry.

6. What is the difference between molar heat capacity and specific heat capacity?

7. Using a coffee cup calorimeter, the heat absorbed by a solution is measured at constant pressure. What equation can be used to find the heat of a reaction from the measured heat given?

8. How is the heat of reaction determined when a bomb calorimeter (constant volume calorimetry) is used? (Use and equation to answer this, but be sure to define each variable.)
1. State Hess's Law.

2. What is meant by standard state? What temperature is used to report standard state enthalpies?

3. What is the standard enthalpy of formation of a compound($\Delta H^\circ_f$)?

4. What is the heat of formation ($\Delta H^\circ_f$) for any element in its most stable state at 25°C and standard atmospheric pressure?

5. How can ($\Delta H^\circ_{\text{rxn}}$) be determined from heats of formation?

Chapter 8.8

6. What is the equation needed to estimate $\Delta H_{\text{rxn}}$ from bond dissociation energies?
Lecture Worksheet 27  
(Sections 15.1 - 15.3)  
DUE: Weds. Apr. 23

1. (a) What is chemical equilibrium?

(b) Give an example of chemical equilibrium

2. (a) What is the law of mass action?

(b) What is the equilibrium constant expression ($K_c$) for the reaction:
   \[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \]?

(c) What is $K_p$ for this reaction? (Give an expression in terms of partial pressures.)

(d) How are $K_c$ and $K_p$ related? What is the meaning of $\Delta n$ in the equation?

3. If the numerical value of $K_c$ is greater than 1, what does that tell you about the chemical reaction?

4. If an equilibrium mixture of a reaction is predominately reactants, what does this tell you about $K_c$?

5. If you know $K_{eq}$ for the following reaction
   \[ 2\text{A} \rightleftharpoons \text{B} \]
   how can you use that number to find $K_{eq}$ for the reverse reaction:
   \[ \text{B} \rightleftharpoons 2\text{A} \]

How can you find $K_{eq}$ for the reaction:
   \[ \text{A} \rightleftharpoons \frac{1}{2} \text{B} \]
1. (a) Define homogeneous equilibria.

(b) Give an example of a homogeneous chemical reaction.

(c) What is the equilibrium constant expression for that reaction?

2. (a) Define heterogeneous equilibria.

(b) Give an example of a heterogeneous chemical reaction.

(c) What is the equilibrium constant expression for that reaction?

3. How can an equilibrium constant be determined if the initial concentrations of products and reactants are known along with the concentration of one component at equilibrium?
Lecture Worksheet 29
(Sections 15.6 - 15.7)

Name ___________________________

(Chem 108 Section # _________)

DUE: Tues. Apr. 29 in Chem 108 class

1. How can an equilibrium constant be determined if the initial concentrations of products and reactants are known along with the concentration of one component at equilibrium?

2. What is the reaction quotient?

3. What is the difference between K and Q?

4. (a) If the value of Q is greater than the value of K, what do you know about the direction of the chemical reaction?
   
   (b) If the value of Q is less than the value of K, what do you know about the direction of the chemical reaction?

5. What is Le Châtelier's Principle?

6. For the following reaction at equilibrium

   \[ \text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3 + \text{heat} \]

   how will the following disturbances affect this reaction?

   (a) increasing [N\textsubscript{2}]
   
   (b) increasing [NH\textsubscript{3}]
   
   (c) increasing the pressure by decreasing the volume
   
   (d) increasing the pressure adding Ar
   
   (e) heating up the mixture