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!
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 \( ^{12}_6 \text{C} \), 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) Enthalpy

Chapter 6.1-6
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. (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?

6. Draw a 2pₓ orbital next to a 2pᵧ orbital. (What is different about these two orbitals?)
Lecture 3 Worksheet
(Chapter 6. 7-9)
(Chapter 2.5, 7.1-6)

1. How does the presence of additional (more than one) electrons affect the orbital energy levels?

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

3. What is meant by core and valence electrons?

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

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

5. 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

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?

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) What is the periodic trend for increasing first ionization energy?
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, p. 315)

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 affect 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?
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 8.7-8

1. a) Draw the Lewis structure for BF₃.

   b) Does this molecule obey the octet rule?

   c) Why does BF₃ react with molecules that have a lone pair?

2. Explain why PCl₅, PCl₃, and NCl₃ are known compounds, whereas NCl₅ 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?
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. Give the names of the geometries of the following simple molecules? Draw them.
   (a) PF$_5$                                (b) CF$_4$                      (c) SF$_6$

2. (a) What is the difference between electron-domain geometry and molecular geometry?
   
   (b) Give the electron domain geometry and molecular geometry of NH$_3$. (Draw and name them.)

3. What are the steps needed to predict molecular geometries with the VSEPR (valence shell electron pair repulsion) model?

4. What role do double and triple bonds have in predicting geometry?

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

6. For a molecule with more than two atoms, what features effect the dipole moment?
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 domain geometry? ____________________________

5. What are the names of the hybrid orbitals used by atoms that have expanded octets?

6. (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 NH₃?

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

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

8. What are delocalized bonds? (Which kind of bonds (sigma or pi) are used in forming delocalized bonds?)

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

<table>
<thead>
<tr>
<th>Lewis structure</th>
<th>EPG</th>
<th>MG</th>
<th>BDP</th>
<th>Polarity</th>
</tr>
</thead>
<tbody>
<tr>
<td>(a) CO₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(b) H₂O</td>
<td></td>
<td></td>
<td></td>
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</tr>
</tbody>
</table>
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 FLOWCHART 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. Under what conditions of pressure and temperature do real gases most closely approximate ideal-gas behavior?

7. (a) Name the regions of the atmosphere.

   (b) Why is it organized in this way?

8. (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?
Lecture 14 Worksheet
(Chapter 5.5, 11.3-6)

From Chapter 5 section 5
1. Define: Heat capacity and Specific Heat

Chapter 11.3-6
2. Define heat of fusion and heat of vaporization.

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

4. Define dynamic equilibrium.

5. (a) Define vapor pressure.

(b) What is the vapor pressure of H₂O at 25°C? (Use the table in Appendix B of your text.)

6. Define normal boiling point and normal melting point.

7. Using the phase diagrams on page 464 in your text (figure 11.27) to answer the following:
   (a) What is the critical point of CO₂? _________________
   (b) What is the triple point of water? _________________
   (c) What state is CO₂ in when P = 1 atm and the T = -100°C? _________________
   (d) When P = 73 atm and T = 25°C? _________________
Lecture 15 Worksheet Name ___________________________
(Chapter 13.1-4, 4.1 & 5) Chem 6 Section # ______

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.

   (a) 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 (O₂)?

4. Circle the compounds that are soluble in aqueous solution.
   - KCl, BaSO₄, Mg(OH)₂, PbBr₂, (NH₄)₂PO₄, and Fe(NO₃)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 non-electrolytes.
   - (a) H₃PO₄ (phosphoric acid)
   - (b) CH₃OH (methanol)
   - (c) CaCl₂
   - (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.
Chapter 4.4
1. Define oxidation and reduction.

2. What is the oxidation state of silicon in SiCl₄? of bromine in Br₂? of nitrogen in NO₃⁻(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 → B₂H). 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 (H₂O) that can be made when you start with 10 moles of hydrogen (H₂) and 7 moles of oxygen (O₂)? _________

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 (NaN₃) in a car, outline the procedure to find the volume of N₂ that will be produced when it decomposes, assuming T = 25°C and P = 1atm.

   (b) If you know the volume of N₂ (at T = 25°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 ($\Delta H^\circ_{rxn}$) be determined from heats of formation?

Chapter 8.8
6. What is the equation needed to estimate $\Delta H_{rxn}$ from bond dissociation energies?
1. (a) What is chemical equilibrium? Give an example of chemical equilibrium.

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

   (b) What is the equilibrium constant expression (Keq) for the reaction:
   \[ N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \]?

3. If the numerical value of Keq is greater than 1, what does that say about the chemical reaction?

4. If you know Keq for the following reaction
   \[ 2A \rightleftharpoons B \]
   how can you use that number to find Keq for the reverse reaction:
   \[ B \rightleftharpoons 2A \]

   How can you find Keq for the reaction:
   \[ A \rightleftharpoons B \]

5. (a) Define homogeneous equilibria.

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

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

6. (a) Define heterogeneous equilibria.

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

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

2. What is the difference between K and Q?

3. (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?

4. What is Le Châtelier's Principle?

5. 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\text{\textsubscript{2}}]

(b) increasing [NH\text{\textsubscript{3}}]

(c) increasing the pressure by decreasing the volume

(d) increasing the pressure adding Ar

(e) heating up the mixture