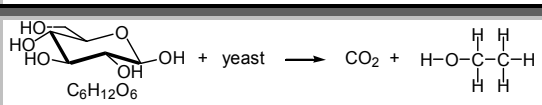
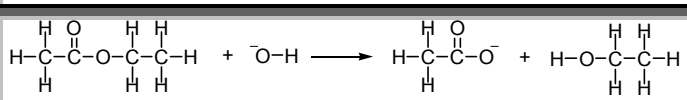
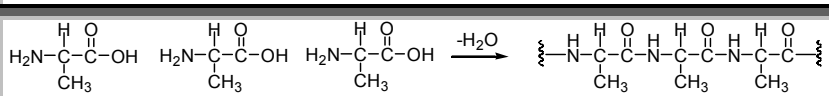
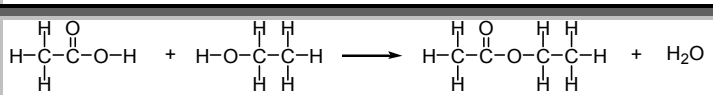
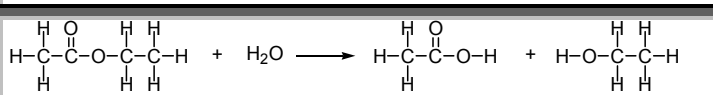
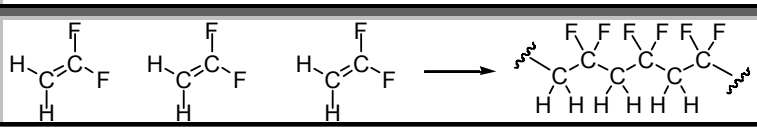
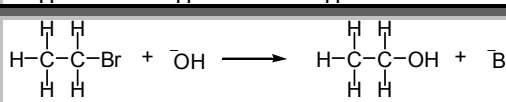
Part I: Organic Chemistry
A) Problems:

1. Label the types of **FIVE** of the following organic reactions based on the descriptions given to you in class and the information on the front:

2) There are FIVE isomers of C_6H_{14}. Draw them all:

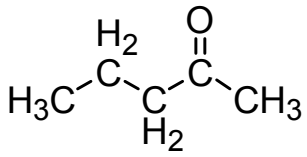
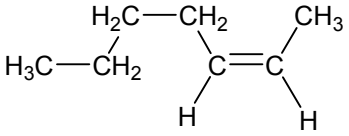
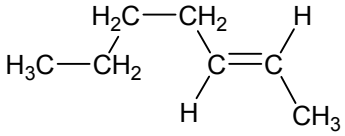
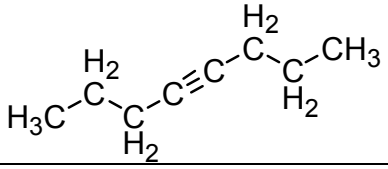
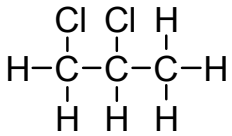
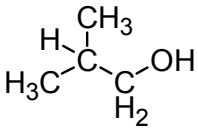
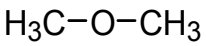
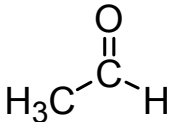
B) Tru/False/CE

ANSWER 5 out of 8 questions

Q	Statement I	Because	Statement II	T/T/CE
1.	Carbon is a nonmetal	Because	Carbon atoms can bond with each other	
2.	The hybrid orbital form of carbon in ethyne is believed to be the SP form	Because	It is a linear compound with a triple bond between carbons	
3.	Normal butanol and 2-butanol are isomers	Because	Isomers vary in the number of neutrons in the nucleus of the atom	
4.	The alkynes are considered a saturated series	Because	Saturated series have maximum number of hydrogen atoms on a carbon chain	
5.	Benzene is a poor electrolyte in water solution	Because	It does not ionize	
6.	Benzene does not have true single and double bonds between its carbon atoms in the ring	Because	It is composed of delocalized pi electrons in the ring giving rise to resonance structures	
7.	Long chain hydrocarbons are insoluble in water	Because	"like dissolves like" and water contains oxygen and no carbon and long chain hydrocarbons contain carbon, but no oxygen	
8.	Ethene (C_2H_4) has a higher carbon-carbon bond energy than ethyne	Because	Ethylene contains a double bond and acetylene has only a single bond between the carbons	F, F

C) Questions:

Name 5 of the following compounds

Structure	Name
(1) 	
(2) 	
(3) 	
(4) 	
(5) 	
(6) 	
(7) 	
(8) 	

Draw the structures for 5 of the following organic names:

(9)	butanal
(10)	*cis*-1,2-dichloroethene
(11)	diethyl ether
(13)	2-methyl-2-propanol
(14)	3-pentanone
(15)	2,2,5,5-tetremethyl-3-hexyne
(16)	butanoic acid
(17)	Benzene (common name for *1,3,5-cyclohexatriene*)

Part II: Electrochemistry:

A Galvanic / Voltaic cell driven by the oxidation of zinc and reduction of copper.

1. What **oxidation** (half reaction) is happening?

2. What **reduction** (half reaction) is happening?

3. What is the **overall reaction**?

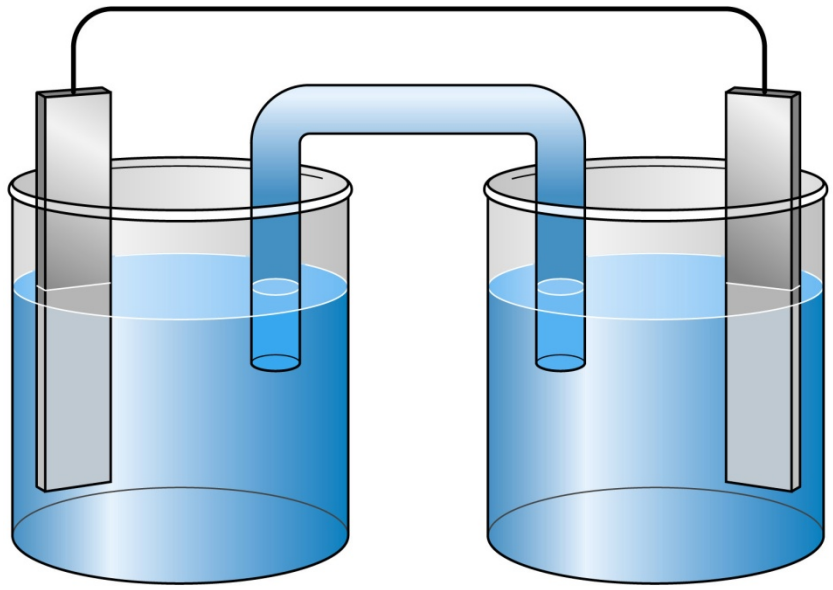
4. What is the **shorthand notation** for this cell?

5. Why are metals good conductors?

6. What is the purpose of the **salt bridge**?

7. **EXTRA CREDIT:** In the diagram below, label the following parts of an electrochemical cell:

Zn(s) electrode	salt bridge	direction of e^- flow	cathode
Zn^{2+}(aq) solution	Cu(s) electrode	direction of ion flow	(+) end
Cu^{2+}(aq) solution	wire & voltmeter	anode	(-) end



8. Balance **ONE** of the following reactions in ACID (ie. Add H_3O^+ to the left side) using the half reaction method:

   a. $BrO_3^- + Fe^{2+} \rightarrow Br^- + Fe^{3+}$

   b. $NO_2 \rightarrow HNO_3 + NO$

9. Write the oxidation number of **each element** for **each compound** below

   a. $MnSO_4$

   b. $CaCO_3$

   c. HCO_3^-

10. For each reaction, note the oxidation number of each element in each compound. Classify each *reactant* as an oxidizing agent (**ox**) or a reducing agent (**red**) based on how the oxidation numbers change.

   a. $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

   b. $3C + 2Fe_2O_3 \rightarrow 3CO_2 + 4Fe$

11. What's the potential of the reaction below given the following half-reaction potentials:

$$2Fe^{2+} + Cl_2 \rightarrow 2Fe^{3+} + 2Cl^-$$

$$Fe^{3+} + e^- \rightarrow Fe^{2+}; E = 0.77 \text{ V}$$
$$Cl_2 + 2e^- \rightarrow 2Cl^-; E = 1.36 \text{ V}$$

12. If the following reactions are used to make a galvanic cell, which species will be reduced and which species will be oxidized? What is the potential of the cell?

$$F_2 + 2e^- \rightarrow 2F^-(aq); E = +2.87 \text{ V}$$
$$Ca^+ + 2e^- \rightarrow Ca(s); E = -2.76 \text{ V}$$

13. Consider

$$Fe^{2+} + 2e^- \rightarrow Fe;$$
$$Au^{3+} + 3e^- \rightarrow Au$$

a. Which is a better reducing agent? Which is a better oxidizing agent? Which gets reduced better? Which gets oxidized better? EXPLAIN HOW YOU CAME TO THAT CONCLUSION (check the table!):

b. If iron was placed as the anode in an electrolytic cell, what minimum voltage would be needed to operate the cell?

$$3Fe + 2Au^{3+} \rightarrow 3Fe^{2+} + 2\ Au$$

14. Which half-reaction can occur at the anode in a voltaic cell? **Explain your reasoning:**

a. $Ni^{2+} + 2e^- \rightarrow Ni$
b. $Sn + 2e^{-+} \rightarrow Sn^{2+}$
c. $Zn \rightarrow Zn^{2+} + 2e^-$
d. $Fe^{3+} \rightarrow Fe^{2+} + e^-$

15. Regarding the anode & cathode, what is the <u>same</u> about an <u>electrolytic</u> and <u>galvanic</u> cell?

16. Regarding the anode & cathode, what is <u>different</u> about an <u>electrolytic</u> and <u>galvanic</u> cell?

Part III: Acids & Bases

1. What is the pH of a 0.2M solution of HCN, $K_a = 4.9 \times 10^{-10}$ at 25 ^oC?

2. What is the $[H_3O^+]$ of a 0.2M solution of HCN, $K_a = 4.9 \times 10^{-10}$ at 25 ^oC?

3. What is the $[OH^-]$ of a 0.2M solution of HCN, $K_a = 4.9 \times 10^{-10}$ at 25 ^oC?

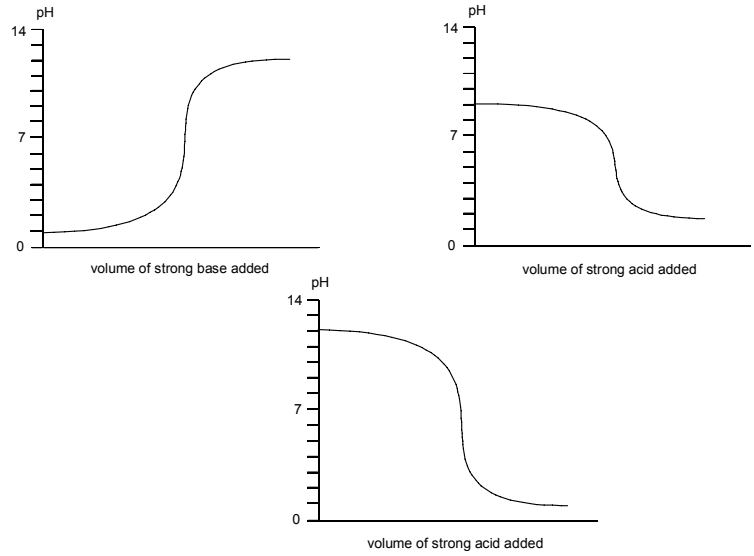
4. What is the pH of a 0.2M solution of HNO_3?

5. How many times stronger than HCN is HNO_3?

6. States the three definitions of Acids and Bases and give an example of each:

7. What volume of 1.420 M NaOH(aq) is needed to titrate 25.00 mL of a 2.430 M HCl(aq) solution?

8. Label the following titration curves as the titration of a: strong acid, strong base or weak base. Circle the equivalence point in each graph. Is the salt generated acidic, basic or neutral?



9. Answer **ONE** of the following two questions:
   A) A 0.00650M solution of a weak acid has a pH of 5.23. Calculate the K_a and state what concentration of un-dissociated hydrogen remains at equilibrium?

   B) What is K_a for HCO_2H (formic acid) if 0.10 M formic acid has an equilibrium concentration of $[H^+] = 0.0042$ M?

10. **EXTRA CREDIT:** What is the pH of a buffer solution containing 0.40 M $CH_3CO_2H(aq)$ ($K_a = 1.8 \times 10^{-5}$) and 0.20 M $CH_3CO_2Na(aq)$?

Part IV: Equilibrium, Solutions and Rates

1. How many *grams* of $C_{12}H_{22}O_{11}$ are in 2.00 L of a 0.0146 M solution?

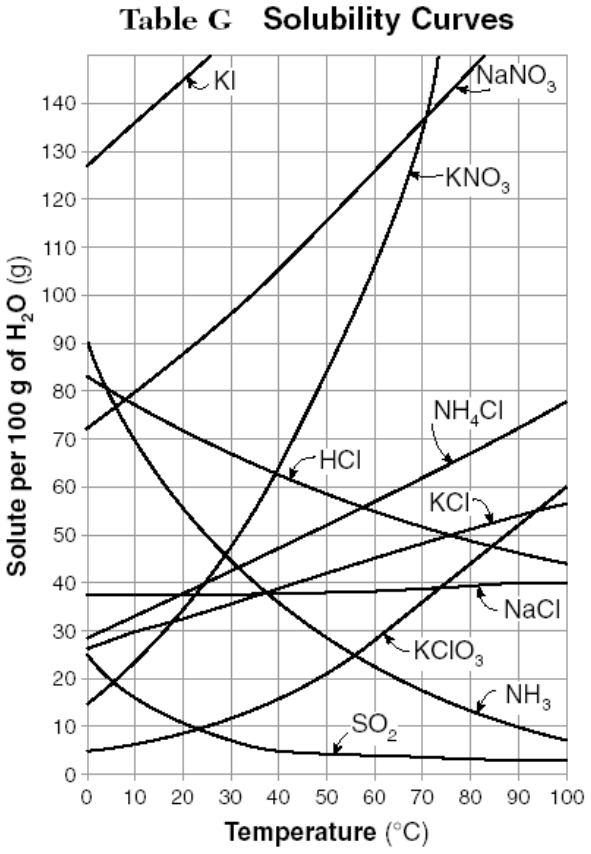
2. How many grams of ammonium sulfate are required to prepare 3.50 L of a 1.55 M solution?

3. **EXTRA CREDIT:** 50.0 mL of CH_3CH_2OH (d = 0.79 g/mL) is mixed with 550.0 mL of water. What *molarity* is the solution with correct number of significant figures?

4. 28 g KOH is mixed with 105 g H_2O. The total volume is 0.11 L. What's the *molality* of the solution?

5. How many *mililiters* of a 3.0 M NaOH solution should be used to make 1.0 L of a 1.0 M NaOH solution?

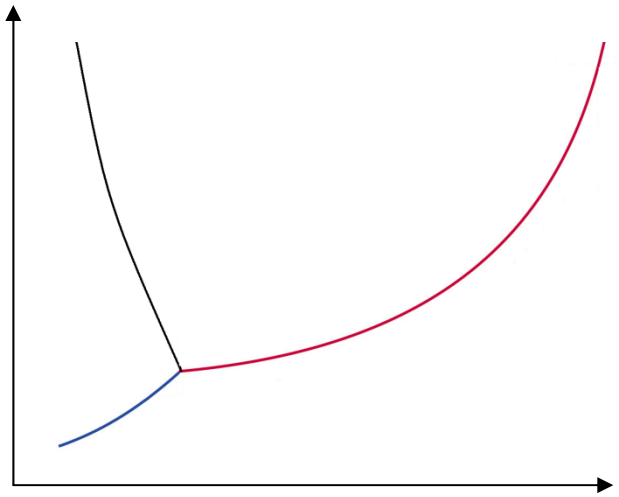
6. How many grams of potassium nitrate can dissolve in 50 g H_2O at 50 ^oC?



7. If 45 g KCl are added to 100 g H_2O at 80 ^oC and the mixture is then cooled to 42 ^oC, how many g KCl will precipitate?

8. Consider four different solutions listed below:
   A) Rank the following aqueous solutions from ***highest*** boiling to ***lowest*** boiling. (Hint: What happens to ionic species when they dissolve in water? What is the concentration of each species below in terms of *particles*?) Explain your reasoning:
      a. 1.0 *m* NaCl
      b. 2.0 *m* $FeCl_3$
      c. 3.0 *m* $KMnO_4$
      d. 1.0 *m* Na_3PO_4
   B) Rank the solutions in terms of highest to lowest vapor pressure:

9. Label the indicated components in each graph below.*(Phase diagram for water)*
   Temperature & pressure axes
   Normal melting & boiling point lines and sublimation line
   Critical point (what is it?)
   Gas, liquid and solid phase regions
   Triple point (what is it?)



10. For $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, $\Delta H < 0$, which way does the reaction shift in response to the listed change?

a. Add H_2

b. Remove NH_3

c. Remove H_2

d. Remove N_2

e. Add NH_3

f. Add N_2

g. Increase pressure

h. Increase temperature

11. For $2NO_2(g) \rightleftharpoons N_2O_4(g)$, $\Delta H < 0$, how is the concentration of N_2O_4 affected by:

    a. increasing temperature

    b. increasing pressure

12. The K_{sp} value for magnesium fluoride is 6.4×10^{-9}. What is the concentration of fluoride ion in a saturated solution of magnesium fluoride?

13. In a saturated solution of CuBr, $[Cu^+] = 7.9 \times 10^{-5}$ M. What is K_{sp} for CuBr?

14. Based on question 14, what would the solubility of CuBr be in a solution of 0.4M NaBr?

15. **EXTRA CREDIT:** Tooth enamel is made of *hydroxyapatite*, $Ca_5(PO_4)_3OH$, one of the strongest substances in the human body. Excessive exposure to acid (directly, or from food & bacteria) will dissolve enamel and cause cavities. One way to prevent this is by adding *fluoride* ion to enamel to make *fluorapatite*. The fluoride ion exchanges with the hydroxide ion.

a. Write the reaction of *hydroxyapatite* with *fluoride* to make *hydroxide* and *fluorapatite*: (a double-replacement rxn)

b. If $K_{sp} = 2.34 \times 10^{-59}$ for *hydroxyapatite* and $K_{sp} = 3.16 \times 10^{-60}$ for *fluorapatite*, why is using fluoride a good idea to prevent tooth decay?

16. Given: $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$; $K_{eq} = 55.0$.

a. If an *equilibrium position* occurs at: $[H_2] = 0.0048$ M and $[Cl_2] = 0.0021$ M, what is [HCl] at this position?

b. What is the *value* of K_{eq} for $2HCl(g) \rightleftharpoons H_2(g) + Cl_2(g)$?

c. What is the *value* of K_{eq} for $4HCl(g) \rightleftharpoons 2Cl_2(g) + 2H_2(g)$?

17. For: $A + 2B + C \rightarrow 2D + E$

a. What is the rate law?

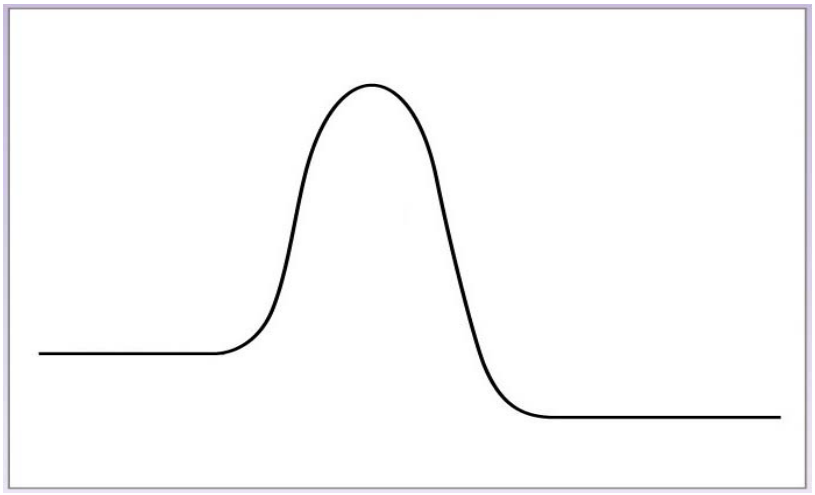
b. What is the value of R_5?

Exp	$[A]_o$	$[B]_o$	$[C]_o$	Initial rate (M/sec)
1	1.40	1.40	1.00	$R_1 = 1.000$
2	0.70	1.40	1.00	$R_2 = 0.500$
3	0.70	0.70	1.00	$R_3 = 0.125$
4	1.40	1.40	0.50	$R_4 = 2.000$
5	0.70	0.70	0.50	$R_5 = ?$

Part V: Thermodynamics

1. Explain factors that affect rate of reactions and state which step in a reaction determines the rate (slow step or fast step) and why:

2. On the diagram below, label:
   - energy of reactants
   - energy of products
   - activation energy
   - transition state
   - Then sketch what happens to the diagram when a catalyst is added to this reaction



3. Use Hess' law to cacculate the enthalpy change:

$C(s) + 2H_2(g) \rightarrow CH_4(g)$ $\quad \Delta H^\circ = -17.8$ kcal

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(g)$ $\quad \Delta H^\circ = -57.8$ kcal

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta H^\circ = -94.0$ kcal

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ $\quad \Delta H^\circ = ?$

4. Answer **ALL** of the following True / False / Correct Explanation questions:

Exothermic reactions have a positive ΔH	Because	Heat must be added to a system for any reaction to occur
The entropy of a solid decreases when dissolved	Because	Solutions are more disordered than solids
Ice melting is an endothermic process	Because	Heat must be absorbed by ice if it is to melt
An exothermic reaction is always spontaneous	Because	The universe favors a negative enthalpy change
An endothermic reaction can be spontaneous	Because	Both enthalpy and entropy changes affect Gibbs free energy
Covalent bonds are broken when a liquid boils	Because	Heat must be released for a liquid to change into a gas
As ice absorbs heat and begins to melt, its temperature remains constant	Because	The absorbed heat is consumed by the breaking of intermolecular interactions
When a sample of water freezes, the process is exothermic	Because	Ice is at a lower potential energy state than water
When a salt sample dissolves in water, $\Delta S > 0$	Because	Aqueous ions have greater entropy than ions in a solid state
Exothermic reactions absorb heat	Because	Breaking covalent bonds always requires energy

5. What temperature range will make a process spontaneous if: $\Delta H = -114.1$ kJ and $\Delta S = -147$ J/K?

6. 87.0 g of iron at 31 ^{o}C is heated to 543 ^{o}C by adding 4900 calories. What is the specific heat of iron in J/g^{o}C?

7. Consider: $2CH_4(g) + 4O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$;
   Use ΔG of formation values to calculate the overall energy of the reaction. Then State whether or not the reaction is spontaneous:

Part VI: Stoichiometry and Limiting Reactants and Gases

1. How many moles of H_2 fill a flask that's 1.00 L, 100 ^oC, 1.50 atm?

2. Suppose 3.5 g Cu reacts with 6.0 g $AgNO_3$ to make $Cu(NO_3)_2$ and Ag.

   a) Write a balanced chemical equation for this reaction.

   b) Which reactant is limiting?

   c) How many g $Cu(NO_3)_2$ can be made?

   d) Suppose 4.5 g of $Cu(NO_3)$ are produced; what is the percent yield?

3. 24.3 g Mg is reacted with excess oxygen to make MgO. Assuming a complete reation, what mass of MgO is produced?

4. How many moles of Mg are needed to react completely with 45 L O_2 at STP?

5. Aluminum chloride reacts with silver nitrate to make silver chloride and aluminum nitrate.
   a. Write the balanced chemical equation for this reaction.
   b. How many g AgCl are produced from 200. g $AlCl_3$ and excess $AgNO_3$?

6. The composition of t-butylmercaptan is 53.27% C, 11.18% H and 35.55% S.
   a) What is the empirical formula of t-butylmercaptan?
   b) How many moles of C are there in 100g of the above mentioned compound?

7. What is the mass percent of oxygen in caffeine, $C_8H_{10}N_4O_2$?

8. How many moles of Cl^- are there in 209g of $AlCl_3$?

Part VII: Nuclear Chemistry and The Periodic Table

1. What are the possible values of m for an electron with l = 1, meaning in a *p* orbital?

2. Is $1s^2 2s^2 2p^5 3s^2$ the configuration of an atom in the *ground* state or *excited* state? What element is it?

3. What is the number of filled energy levels in lithium in the ground state?

4. Write the electron configuration and orbital system of the element in group6 and fifth period:

5. Write the gamma emission of ^{238}U.

a. What is the initial neutron/proton ratio?

b. What is the neutron/proton ratio after decay?

6. Answer the following:
   a) Write the beta decay of ^{131}I.

   b) Write the alpha decay of ^{226}Ra.

   c) What are beta + and electron capture decays?

7. Compare bond energy and bond length in a single bond, a double bond and a triple bond and resonance bon (as in benzene):

8. As we move from top to bottom and right to left in the periodic table, what happens to the following:
   a) Electronegativity
   b) Electron affinity
   c) Atomic radius
   d) Ionic radius for metals and nonmetals
   e) Ionization energy

9. List all types of bonds you know and all the intermolecular forces and order them in terms of strongest to weakest

10. Draw the Lewis structure and STAE THE GEOMETRY of:
    a) $AsSe_3$

    b) H_2Se

    c) CO_2

    d) $(BF_3)(NH_3)$ "coordinate covalent bond"

11. If 80 mg of a radioactive isotope decays to 10 mg in 30 min, what is the element's half-life in minutes?

12. After 62.0 hours, 1.0 g remains from a sample of K-42. If the half life is 12.4 hours, how much K-42 was in the original sample?

Have a Nice Summer!

-Ali

Reduction Potentials		(1M, 25 °C, 1 atm)	E^{o}(V)
$F_2(g) + 2e^-$	→	$2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^-$	→	$2H_2O(l)$	+1.78
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	→	$Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$Au^{3+}(aq) + 3e^-$	→	Au(s)	+1.50
$Cl_2(g) + 2e^-$	→	$2Cl^-(aq)$	+1.36
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^-$	→	$2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$O_2(g) + 4H^+(aq) + 4e^-$	→	$2H_2O(l)$	+1.23
$4H^+(aq) + MnO_2(s) + 2e^-$	→	$Mn^{2+}(aq) + 2H_2O(l)$	+1.22
$Br_2(l) + 2e^-$	→	$2Br^-(aq)$	+1.09
$Ag^+(aq) + e^-$	→	Ag(s)	+0.80
$Hg_2^{2+}(aq) + 2e^-$	→	2Hg(s)	+0.80
$Fe^{3+}(aq) + e^-$	→	$Fe^{2+}(aq)$	+0.77
$I_2(s) + 2e^-$	→	$2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^-$	→	$4OH^-(aq)$	+0.40
$Cu^+(aq) + e^-$	→	Cu(s)	+0.52
$Cu^{2+}(aq) + 2e^-$	→	Cu(s)	+0.34
$4H^+(aq) + SO_4^{2-}(aq) + 2e^-$	→	$SO_2(g) + 2H_2O(l)$	+0.17
$Cu^{2+}(aq) + e^-$	→	$Cu^+(aq)$	+0.15
$Sn^{4+}(aq) + 2e^-$	→	$Sn^{2+}(aq)$	+0.15
$2H^+(aq) + 2e^-$	→	$H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^-$	→	Pb(s)	-0.13
$Sn^{2+}(aq) + 2e^-$	→	Sn(s)	-0.14
$Ni^{2+}(aq) + 2e^-$	→	Ni(s)	-0.26
$Co^{2+}(aq) + 2e^-$	→	Co(s)	-0.28
$Cd^{2+}(aq) + 2e^-$	→	Cd(s)	-0.40
$Fe^{2+}(aq) + 2e^-$	→	Fe(s)	-0.45
$Cr^{3+}(aq) + 3e^-$	→	Cr(s)	-0.74
$Zn^{2+}(aq) + 2e^-$	→	Zn(s)	-0.76
$2H_2O(l) + 2e^-$	→	$H_2(g) + 2OH^-(aq)$	-0.83
$Mn^{2+}(aq) + 2e^-$	→	Mn(s)	-1.19
$Al^{3+}(aq) + 3e^-$	→	Al(s)	-1.66
$Mg^{2+}(aq) + 2e^-$	→	Mg(s)	-2.37
$Na^+(aq) + e^-$	→	Na(s)	-2.71
$Ca^{2+}(aq) + 2e^-$	→	Ca(s)	-2.87
$Sr^{2+}(aq) + 2e^-$	→	Sr(s)	-2.89
$Ba^{2+}(aq) + 2e^-$	→	Ba(s)	-2.91
$Cs^+(aq) + e^-$	→	Cs(s)	-2.92
$K^+(aq) + e^-$	→	K(s)	-2.93
$Rb^+(aq) + e^-$	→	Rb(s)	-2.98
$Li^+(aq) + e^-$	→	Li(s)	-3.04