Chemistry 103, Dr. Hamers
Practice Problems that may help you prepare for Exam II
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Here are some problems to try. If you get bogged down on a particular questions: call a friend, ask a buddy, but don't work on one problem for hours. Good Luck!

Chapter 5

1. Which of the following processes are exothermic?
   I. Boiling water.
   II. Freezing water.
   III. Evaporating water.
   IV. Melting ice.
   V. The combustion of methanol, CH_3OH.
   VI. Breaking the H-H bond in a hydrogen gas molecule.

2. In an endothermic reaction with work done *by* the system, it is expected that
a. ΔU_{sys} increases
b. ΔU_{sys} decreases
c. ΔU_{sys} stays the same
d. ΔU_{sys} depends on the magnitude of q and w

3. If a 15.0 g piece of Au needed 100 J to increase from 20.0˚C to 71.7˚C, what is its specific heat capacity?

4. If the specific heat capacity of copper metal is 0.385J/gK, what is the molar heat capacity of copper metal?

5. When work is done *on* the system, it is expected that
a. ΔU_{sys} increases and ΔV increases
b. ΔU_{sys} decreases and ΔV decreases
c. ΔU_{sys} increases and ΔV decreases
d. ΔU_{sys} decreases and ΔV increases

6. 289.8g of Au at 150.˚C is dropped into 100.g of water at 20.0˚C (specific heat capacity = 4.184J/g˚C). A final temperature of 30.6˚C is reached. What is the specific heat capacity of the gold?

7. Hydrogen sulfide, H_2S, is a foul-smelling gas. It burns to form sulfur dioxide.

$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(g) \qquad \Delta H_{rxn} = -1037\ kJ$

Calculate the enthalpy change when 25 grams of hydrogen sulfide (H_2S) is burned.

8. What is the heat of reaction for the following reaction: $CH_4(g) + NH_3(g) \rightarrow 3H_2(g) + HCN(g)$
Use the following thermodynamic data.

Rxn I: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \qquad \Delta H_{rxn} = -91.8\ kJ$
Rxn II: $C(graphite) + 2H_2 \rightarrow CH_4(g) \qquad \Delta H_{rxn} = -74.9\ kJ$
Rxn III: $H_2(g) + 2C(graphite) + N_2 \rightarrow 2HCN(g) \qquad \Delta H_{rxn} = 270.3\ kJ$

9. Given the following reactions with their respective ΔH_{rxn},

Rxn I: $H_2O(l) \rightarrow H_2(g) + {}^1/_2O_2(g) \qquad \Delta H_{rxn} = 285.8kJ$
Rxn II: $C_2H_6(g) \rightarrow 2C(s) + 3H_2(g) \qquad \Delta H_{rxn} = 84.7kJ$
Rxn III: $2CO_2(g) + 3H_2O(l) \rightarrow C_2H_6(g) + {}^7/_2O_2(g) \qquad \Delta H_{rxn} = 1560.5kJ$

calculate the ΔH_{rxn} for the following reaction: $CO_2(g) \rightarrow C(s) + O_2(g)$

(take a breather)

10. How much heat is released when 10.0g of water vapor at 200.˚C is cooled to water at 30.0˚C? The specific heat of ice is 2.03J/g ˚C, the specific heat of water is 4.184 J/g˚C, the specific heat of $H_2O(g)$ is 2.00J/g˚C, the heat of fusion for ice is 333J/g, and the heat of vaporization of water is 2256J/g.

11. Write the reaction that would correspond to the ΔH_f of each chemical listed below:
a. $Mn_2O_3(s)$ b. $CaSO_4(s)$ c. $CH_3CH_2OH(l)$ (ethanol)

12. Calculate ΔH°_{rxn} for the reaction below using the standard enthalpies of formation given.

$$2H_2S(g) + 3O_2 \rightarrow 2H_2O(l) + 2SO_2(g)$$

ΔH°_f for $H_2S(g)$ = -20.2 kJ/mol, ΔH°_f for $H_2O(l)$ = -285.8 kJ/mol, ΔH°_f for $SO_2(g)$ = -296.8 kJ/mol

13. When methanol (CH_3OH) reacts as shown below, the change in enthalpy for the reaction is –1277.4 kJ. Given the ΔH_f° of $H_2O(g)$ is –241.8kJ/mol and ΔH_f° of $CO_2(g)$ is –393.5kJ/mol, what is the ΔH_f° for methanol?

$$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)$$

14. a. Write an appropriate equation illustrating the standard enthalpy of formation for $MnO_2(s)$. b. Given the two reactions below, calculate ΔH_f° for $MnO_2(s)$.

$MnO_2(s) \rightarrow MnO(s) + {}^1/_2\, O_2(g)$ $\quad \Delta H = 136.0$ kJ

$MnO_2(s) + Mn(s) \rightarrow 2MnO(s)$ $\quad \Delta H = -248.9$ kJ

15. a. Write an appropriate equation illustrating the standard enthalpy of formation for $SO_2(g)$. b. Given the two reactions below, calculate ΔH_f° for $SO_2(g)$. c. Draw an energy level diagram for this series of reactions.

$S(s) + {}^3/_2\, O_2(g) \rightarrow SO_3(g)$ $\quad \Delta H = -395.2$ kJ

$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$ $\quad \Delta H = -198.2$ kJ

16. What is the final temperature of a solution if 12.5g $NH_4NO_3(s)$ is dissolved in an insulated coffee cup containing 150.g of water at 50.0˚C? The heat of solution (dissolving) $NH_4NO_3(s)$ is 26.4kJ/mol, and the heat capacity of the solution is 4.184J/g˚C.

$$NH_4NO_3(s) \rightarrow NH_4NO_3(aq)$$

17. 25.0 ml of 1.0M HCl is combined with 35.0 ml of 0.5M NaOH. The initial temperatures of the solutions are 25˚C, the density of the solution is 1.0g/ml, the specific heat capacity of the solution is 4.184 J/g˚C, the reaction is completed in an insulated coffee cup, and the standard enthalpy of reaction for $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$ is –56kJ/mol. What is the final temperature of the solution?

18. What is the heat of solution, ΔH_{soln}, for CsOH(s) if when 5.75g of CsOH were dissolved in an insulated coffee cup containing 75.0g of water at 22.5˚C the final temperature was 31.2˚C? The heat capacity of the solution is 4.184J/g˚C.

$$CsOH(s) \rightarrow CsOH(aq)$$

19. When 2.25g of methane, CH_4, is burned in a constant volume "bomb" calorimeter the temperature increases from 21.56 to 71.63. The calorimeter contains 425g water (C = 4.184J/g˚C) and the vessel has a calorimeter constant of 475J/˚C. What is the ΔH_{rxn} (in kJ)?

(take a nap)

20. If 50.0g metal (C = 4.00J/g˚C) initially at –80.˚C is placed in 25.0g water at 55.0˚C, what is the final temperature and how many grams of water, if any, freezes? The heat capacity of water = 4.184J/g˚C, heat capacity of ice = 2.0J/g˚C, and heat of fusion for water is 333J/g.

21. A beaker contains 50g of water (C = 4.184J/g˚C) at 20˚C and a 10gram piece of metal (C = 5.75J/g˚C) at 450˚C is added to it. How many grams of water boiled and what is the final temperature in the beaker? The heat of fusion for water is 333J/g and the heat of vaporization is 2256J/g.

22. A container has 10g of ice at 0˚C and 50g of water at 20˚C. What is the final temperature and will all the ice melt? The specific heat of ice is 2.03J/g˚C, the specific heat of water is 4.184 J/g˚C, the specific heat of $H_2O(g)$ is 2.0J/g˚C, the heat of fusion of ice is 333J/g, and the heat of vaporization of water is 2256J/g.

23. A 100g piece of Fe (C = 0.449J/g˚C) at -100.0˚C is combined with 125g water at 75.0˚C and 75.0g ice at 0˚C. What is the final temperature of the system? The specific heat of ice is 2.03J/g˚C, the specific heat of water is 4.184 J/g˚C, the specific heat of $H_2O(g)$ is 2.0J/g˚C, the heat of fusion of ice is 333J/g, and the heat of vaporization of water is 2256J/g.

24. A 125g piece of Al (C = 0.897J/g˚C) at 200.0˚C is combined with 115g water at 50.0˚C and 50.0g ice at 0˚C. What is the final temperature of the system? The specific heat of ice is 2.03J/g˚C, the specific heat of water is 4.184 J/g˚C, the specific heat of $H_2O(g)$ is 2.0J/g˚C, the heat of fusion of ice is 333J/g, and the heat of vaporization of water is 2256J/g.

Chapter 6

25. A photon of light has frequency of $5.5\text{x } 10^9 s^{-1}$. a. What is the energy (J) and wavelength (m) of this light? b. What is the energy of a mole of this light?

26. Which of the following showed that light had a particle nature to it?
a. Photoelectric Effect b. UV Catastrophe c. Schrodinger Equation
d. de Broglie equation e. Blackbody Radiation

27. If a C=O double bond requires 745kJ/mol to be broken, what wavelength could break a C=O double bond?

28. a. Which of the following regions in the electromagnetic spectrum has the greatest wavelength?
I. UV II. IR III. x-ray IV.red light
b. Which region in the electromagnetic spectrum has a wavelength smaller than red light?
I. microwave II. radio wave III. infrared IV. green light
c. Which color of visible light has the greatest frequency?
I. orange II. red III. green IV. blue
d. Of the following orbitals, which have the same *l* value?
I. 3d II. 4s III. 3p IV. 7s

29. An electron in the 4th energy level can undergo 6 different transitions of which 5 are shown below. Which transition will have the shortest wavelength?
a. $4 \rightarrow 3$ b. $4 \rightarrow 2$ c. $4 \rightarrow 1$ d. $3 \rightarrow 1$ e. $2 \rightarrow 1$

30. What is the change in energy if an electron in the 6th energy level falls to the 2nd energy level in a H atom?
a. $-{}^1/_3$Rhc b. $-{}^2/_9$Rhc c. $-{}^1/_4$Rhc d. $-{}^1/_{36}$Rhc e. $-{}^1/_{32}$Rhc

(watch some TV!)

31. What frequency of light is required for an electron in the 3rd energy level of hydrogen to transition to the 5th energy level?

32. If an electron has a wavelength of 1 x 10^2 nm and a mass of 9.11 x 10^{-28} g, how fast is it traveling?
a. 7.27 x 10^{-9} m/s b. 7.27 x 10^{-6} m/s c. 7.27 x 10^{-3} m/s d. 7.27 e. 7.27 x 10^3 m/s

33. There is one valid set of quantum numbers (n, l, m_l, m_s) below. Identify which is the **valid** set.
a. (2, 3, 2, $+^1/_2$) b. (4, 4, 2, $+^1/_2$) c. (4, 3, 4, $+^1/_2$) d. (4, 3, 2, $+1^1/_2$) e. (7, 3, 2, $+^1/_2$)

34. a. Draw the $3p_x$ orbital. Show the nodal surfaces. b. Draw the $d_{x^2-y^2}$ orbital. Show the nodal surfaces.

35. a. The quantum number l describes the ___________ of the orbital.
b. The quantum number m_l describes the __________ of the orbital.
c. Describe the Heisenberg Uncertainty Principle.
d. Light has both wave and particle properties. Give an experimental result that supported the wave nature of light and give an experimental result that supported the particle nature of light.
e. Electrons have both wave and particle properties. Give an experimental result that supported the wave nature of light and give an experimental result that supported the particle nature of light.

36. I. What is the maximum number of ***orbitals*** that these quantum numbers represent?
a. $n = 4$ b. $n = 3, l = 2$ c. $n = 2, l = 1$ d. $n = 5, l = 5$ e. $n = 2, l = 1, m_l = 1$
II. What is the maximum number of ***electrons*** that can have these quantum numbers:
a. $n = 4$ b. $n = 5, m_s = +^1/_2$ c. $n = 3, l = 2$

Chapter 7
37. Answer the following questions about quantum numbers and orbitals.
a. Which subshell is not allowed?
I. 2s II. 2p III. 2d IV. 3d V. 7s
b. Which subshell has the lowest energy?
I. 3s II. 3p III. 2p IV. 3d V. 2s
c. Which orbital penetrates the closest to the nucleus?
I. 2s II. 2p III. 3s IV. 3p V. 3d
d. How many nodes does the 4p orbitals have?
I. 0 II. 1 III. 2 IV. 3 V. 4
e. How many nodal surfaces does the 3d orbitals have?
I. 0 II. 1 III. 2 IV. 3 V. 4
f. Which electron in an atom of chlorine would experience the greatest effective nuclear charge?
I. 2s II. 2p III. 3s IV. 3p V. 3d

38. I. a. Write the electron configuration for Mn. Do not use abbreviations.
b. Draw an orbital box diagram (or use simply lines) for this *using the noble gas abbreviation.*
c. How many unpaired electrons does Mn have?
d. How many unpaired electrons does Mn^{+3} have?
II. a. Write the electron configuration for Co. Do not use abbreviations.
b. Draw an orbital box diagram (or use simply lines) for this *using the noble gas abbreviation.*
c. How many unpaired electrons does Co have?
d. How many unpaired electrons does Co^{+3} have?

(eat some pizza)

39. a. What is the electron configuration of Fe^{+2}? Do not use abbreviations.
b. What is the electron configuration of P^{-2}? Do not use abbreviations.
c. What is the electron configuration for Cu. Do not use abbreviations.

40.

a.	Which atom has the **largest** atomic radii?	Ca, C, Rb, Mg, B
b.	Which atom or ion is **smallest** in size?	Sb^{-3}, Te^{-2}, I^{-}, Xe, Xe^{+}
c.	Which atom has the **highest first** ionization energy?	S, Cl, Ar, Ne, Na
d.	Which atom has the **lowest first** ionization energy?	F, Cl, Br, I
e.	Which atom has the **highest first** ionization energy?	Na, Mg, Al, Si
f.	Which atom has the **highest first** ionization energy?	B, C, N, O
g.	Which atom has the **highest second** ionization energy?	Na, Mg, Al
h.	Which of the following ions/molecules are isoelectronic with CNO^{-}?	O_3, N_3^{-}, CH_4, CO
i.	Which atom has the most negative electron affinity (more exo).	N, O, F, Cl
j.	Which atom or ion is **largest**?	O^{-2}, S^{-2}, Se^{-2}, Br^{-}, Kr

41. Place the following atoms in order of **decreasing** atomic radii: Cl, Se, Ge, Sn
a. Se > Ge > Sn > Cl
b. Sn > Ge > Se > Cl
c. Cl > Se > Ge > Sn
d. Cl > Ge > Se >Sn
e. None of the above.

42. Which atom or ion has the most unpaired electrons in the ground state?
a. Si
b. Mn^{+2}
c. N^{-}
d. Pb^{+2}
e. P

43. Rank the atoms from smallest to highest **second** ionization energy? Ca, Rb, K
a. Ca < Rb < K
b. Rb < Ca < K
c. Rb < K < Ca
d. K < Rb < Ca
e. Ca < K < Rb

44. Which chemical below will require the most energy if an electron is removed from it?
a. Na
b. P
c. S^{+}
d. O^{-}
e. Si

45. Which material would most likely be the least densest?
a. sodium(s)
b. aluminum(s)
c. magnesium(s)
d. need additional info

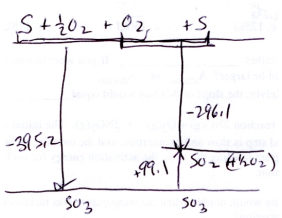
46. Which atom would most make sense to have the following ionization energies:
IE_1 = 500kJ/mol IE_2 = 900kJ/mol IE_3 = 1500kJ/mol IE_4 = 7200kJ/mol
a. Li
b. Mg
c. C
d. Al
e. none of the above

47. Short Answer questions.
a. The chemicals that react within a beaker are called the __________ (system/surroundings) in thermodynamics.
b. A "bomb" calorimeter has constant __________ .
c. The 3 quantum numbers, n, l, and m_l are a solution to the __________ equation.
d. As the speed of a subatomic particle like a proton increases, the wavelength associated with it (increases/decreases) __________.
e. According to the first law of thermodynamics, what is conserved? __________
f. If the enthalpy of fusion for a metal is 250J/g, then the enthalpy for this metal going from a solid to a liquid would be equal to __________.
g. The name of the family of elements with a valance shell = ns^1 __________ .

(yea done!)

ANSWERS

1. II, V
2. d {when $q > 0 \rightarrow \Delta U_{sys}$ increases; work *by* the system $\rightarrow w < 0$; $\Delta U_{sys} = q + w$ so whether ΔU_{sys} increases or decreases depends on the sizes of q and w}
3. 0.129 J/g ˚C {Heat = mCΔT; C = 100J/(15g x [71.7 - 20.0]C) = 0.129J/gC}
4. 24.46J/molK {0.385J/gK x (63.54g/1mol) = 24.46J/molK}
5. c {when work *on* the system $\rightarrow w > 0 \rightarrow \Delta U_{sys}$ increases; $w > 0 \rightarrow \Delta V$ decreases}
6. 0.128J/g˚C {heat lost = -heat gained: $C_{Au}(289.8)(30.6 - 150) = -(4.184)(100)(30.6 - 20) \rightarrow$ $C_{Au} = -(4.184)(100)(30.6 - 20)/[(289.8)(30.6 - 150)] = 0.128$J/g˚C}
7. -381 kJ {proportionality of heats: -1037kJ/2mol H_2S = xkJ/25g $H_2S \rightarrow$ -1037kJ/68g H_2S = xkJ/25g H_2S}
8. 256.0kJ {Rxn I: reverse, divide by 2; Rxn II: reverse; Rxn III: divide by 2; $\Delta H_{rxn} = [-(-91.8)/2] + [-(-74.9)] + [(270.3/2)]$}
9. 394kJ {Rxn I: reverse and multiply by 3/2; Rxn II: multiply by 1/2; Rxn III: multiply by 1/2; $\Delta H_{rxn} = [-(285.8)(1.5)] + [(84.7)(0.5)] + [(1560.5) \times 0.5] = 393.9$kJ}
10. –27,500J {Heat = heat released when vapor goes from 200 to 100˚C + heat released when vapor condenses to water + heat released when water cools from 100 to 30˚C = $(C_{H2O(g)} \times mass_{H2O(s)} \times \Delta T_{H2O(g)}) + (\Delta H_{vap} \times g_{H2O(g)}) + (C_{H2O(l)} \times mass_{H2O(l)} \times \Delta T_{H2O(l)})$ = (2.0 x 10 x (100-200)) + (2256J/g x 10) + (4.184 x 10 x (30-100)) = -27,489J}
11. a. $2Mn(s) + {}^3/_2O_2(g) \rightarrow Mn_2O_3(s)$ {produce 1 mol of substance from elements in natural states}
    b. $Ca(s) + S(s) + 2O_2(g) \rightarrow CaSO_4(s)$ c. $2C(s) + 3H_2(g) + {}^1/_2O_2(g) \rightarrow CH_3CH_2OH(l)$
12. -1125 kJ {$\Delta H_{rxn} = [\Sigma\Delta H_{f\ prod} \times mole_{prod}] - [\Sigma\Delta H_{f\ react} \times mole_{react}]$ = [2(-285.8) + 2(-296.8)] - [2(-20.2) + 3(0)] = -1125}
13. -238.4kJ/mol {use $\Delta H_{rxn} = [2 \times \Delta H_f^\circ\ CO_2 + 4 \times \Delta H_f^\circ\ H_2O] – [2 \times \Delta H_f^\circ\ CH_3OH + 4 \times \Delta H_f^\circ\ O_2]$; plug in values and solve for $\Delta H_f^\circ\ CH_3OH$; -1277.4 = [2(-393.5) + 4(-241.8)] – [2x + 3(0)]; -1277.4 = [-1754.2] – [2x]; solve for x; x = $\Delta H_f^\circ\ CH_3OH$ = -238.4kJ/mol}
14. a. $Mn(s) + O_2(g) \rightarrow MnO_2(s)$ b. -520.9 kJ {Rxn I: rev; x2; Rxn II: unchanged; ΔH_{rxn} = (-136.0)(2) + (-248.9)}
15. a. $S(s) + O_2(g) \rightarrow SO_2(g)$
    b. -296.1kJ {manipulate the 2 reactions given to end up with the ΔH_f reaction: reaction I: leave as it is; reaction II: reverse and multiply by $^1/_2$; add the modified ΔH_{rxns}: (-395.2) + (-1)(0.5)(-198.2) = -296.1kJ}

    c.
16. 43.4˚C {1mol NH_4NO_3/26.4kJ = 12.5g NH_4NO_3/x; 1mol NH_4NO_3/26.4kJ = 0.1563mol NH_4NO_3/x; x = 4.125kJ was absorbed by the NH_4NO_3 as it dissolved; $q_{sys} = -q_{surr}$; q_{surr} = -4.125kJ = -4125J; q_{soln} = CmΔT; -4125 = (4.184)(150)(T_f – 50); T_f = 43.4˚C}
17. 28.9˚C {reaction: HCl + NaOH $\rightarrow$ H_2O + NaCl; find mol that react; mol HCl = 0.025L x 1.0M = 0.025mol HCl; mol NaOH = 0.035L x 0.5M = 0.0175mol NaOH; 0.0175mol reacts since NaOH is limiting reagent; set up proportionality: 1mol NaOH/-56kJ = 0.0175mol NaOH/x; solve for x; x = -0.980kJ x (1000J/1kJ) = -980J; since the system is the HCl and NaOH and the water is the surroundings, -980 becomes +980; another way to think of this is that the acid/base gave off the heat (negative sign) and the water absorbed the heat (positive sign); find grams of being heated up; this will be entire solution, not just the HCl and NaOH, since all of the water is also heating up; 25ml + 35ml = 60ml; convert to grams using D = mass/vol; solve for mass; mass = D x vol = 1 x 60 = 60grams; use heat = CmΔT; 980 = 4.184 x 60 x (T_f – 25); solve for T_f; T_f = 28.9˚C}

18. -71.2kJ {$q_{soln} = Cm\Delta T = (4.184)(75)(31.2 – 22.5) = 2730.1J$; $q_{sys} = -q_{surr}$; $q_{sys} = -2730J$;
1mol CsOH/xJ = 5.75g CsOH/-2730J; 1mol CsOH/xJ = 0.03836mol CsOH/-2730; x = $\Delta H_{rxn} = -71170J = -71.2kJ$}

19. –802kJ {write reaction: $CH_4 + 2O_2$ $CO_2 + 2H_2O$; $q_{heat\ water} + q_{heat\ calorimeter} + q_{reaction} = 0$;
(4.184)(425)(71.63-21.56) + (475)(71.63-21.56) = $-q_{rxn}$ = 89034 + 23783 = 112817J; q_{rxn} = -112817J; set up a proportionality:
2.25g CH_4/-112817J = 1molCH_4/ΔH_{rxn}; convert 1 mol CH_4 to grams: 1mol CH_4x (16g CH_4/1mol CH_4) = 16g CH_4;
2.25g CH_4/-112817J = 16gCH_4/ΔH_{rxn}; 2.25ΔH_{rxn} = -1805076J; ΔH_{rxn} = -802256J = -802kJ}

20. -7.7˚C and all the water freezes {$q_{heat\ metal} + q_{cool\ water} + q_{freeze\ water} + q_{cool\ ice} = 0$; started by assuming T_f = 0˚C: $q_{cool\ water}$ = (4.184)(25)(0-55) = -5753J; $q_{heat\ metal}$ to 0˚C = (4)(50)(0 – (-80)) = 16000J; since the metal will cool the water down to 0˚C and the metal at this time will still be lower than 0˚C, some water freezes; assume all water freezes: $q_{freeze\ water}$ = (-333)(25) = -8325J; since cooling and freezing all the water yields -5753 + (-8325) = -14078J, and it requires 16000J to heat the metal to 0˚C, the ice then cools below 0˚C → $q_{heat\ metal} + q_{cool\ water} + q_{freeze\ water} + q_{cool\ ice}$ =
(4)(50)(T_f – (-80)) + (-5753) + (-8325) + (2.0)(25)(T_f – 0) = 0; 200T_f + 16000 - 5753 – 8325 + 50T_f = 0; 250T_f = -1922;
T_f = -7.69˚C}

21. 1.50g boils; T_f = 100˚C {$q_{cool\ metal} + q_{heat\ water} + q_{boil\ water} + q_{heat\ water(g)} = 0$; $q_{heat\ water}$ to 100˚C =
(4.184)(50)(100-20) = 16736J needed; assume metal cools to 100˚C: $q_{metal\ cooling}$ to 100˚C = (5.75)(10)(100-450) = -20125J; sufficient to raise water to 100˚C and some/all water boils; assume all the water boils: $q_{boil\ water}$ = 2256(50) = 112,800J; not sufficient amount of heat from the metal to boil all the water so only some water boils and T_f = 100˚C →
$q_{cool\ metal} + q_{heat\ water} + q_{boil\ water} = 0$; -20125 + 16736 + (2256)(x); x = 1.50g boils; T_f = 100˚C}

22. T_f = 3.4˚C, and all the ice melts {$q_{melt\ ice} + q_{heat\ water\ from\ ice} + q_{cool\ water} + q_{freeze\ water} = 0$; assume T_f = 0˚C → all ice melts: $q_{melt\ ice}$ = 333 x 10 = 3330J; $q_{cool\ water}$ to 0˚C = (4.184)(50)(0 – 20) = -4184J; this means that there is sufficient energy to melt all the ice and T_f > 0˚C → $q_{melt\ ice} + q_{heat\ water\ from\ ice} + q_{cool\ water} = 0$;
3330 + (4.184)(10)(T_f – 0) + (4.184)(50)(T_f – 20) = 0; 3330 + 41.84T_f + 209.2T_f – 4184 = 0; 251.04T_f = 854; T_f = 3.40˚C}

23. 11.1˚C {$q_{heat\ metal} + q_{cool\ water} + q_{freeze\ water} + q_{ice\ from\ water\ cool} + (q_{melt\ ice} + q_{heat\ water\ from\ ice}$ or $q_{cool\ ice}) = 0$; T_{range}: -100 to 75; assume T_f = 0˚C → $q_{heat\ metal}$ = (0.449)(100)(0 – (-100)) = 4490J; $q_{cool\ water}$ to 0˚C =
(4.184)(125)(0 – 75) = -39925J; $q_{melt\ ice}$ = (333)(75) = 24975J; this means that there is sufficient energy in the warm water to melt all the ice and heat the metal to 0˚C and raise T_f > 0˚C → $q_{heat\ metal} + q_{melt\ ice} + q_{heat\ water\ from\ ice} + q_{cool\ water} = 0$;
(0.449)(100)(T_f – (-100)) + (333)(75) + (4.184)(75)(T_f – 0) + (4.184)(125)(T_f – 75) = 0;
44.9T_f + 4490 + 24975 + 313.8T_f + 523T_f – 39225 = 0; 881.7T_f = 9760; T_f = 11.1˚C}

24. 37.2˚C {$q_{cool\ metal} + (q_{melt\ ice} + q_{heat\ water\ from\ ice} + q_{boil\ water\ from\ ice} + q_{heat\ water\ gas\ from\ ice}) + (q_{heat\ water} + q_{boil\ water} + q_{heat\ water\ gas}) = 0$; T_{range}: 0 to 200; assume T_f = 100˚C →
$q_{cool\ metal}$ = (0.897)(125)(100 – 200) = 11212.5J; $q_{melt\ ice}$ = (333)(50) = 16650J; these numbers say that T_f < 100˚C since there is not enough energy released from the metal as it cools to 100˚C to melt all of the ice so the water must cool off as well to melt some ice; $q_{cool\ water}$ = (4.184)(115)(0 – 50) = -24058J which is sufficient to melt all of the ice → 0 < T_f < 100;
$q_{cool\ metal} + q_{melt\ ice} + q_{heat\ water\ from\ ice} + q_{cool\ water} = 0$;
(0.897)(125)(T_f – 200) + (333)(50) + (4.184)(50)(T_f – 0) + (4.184)(115)(T_f – 50) = 0;
112.125T_f - 22425 + 16650 + 209.2T_f + 481.16T_f – 24058 = 0; 802.5T_f = 29833; T_f = 37.2˚C}

25. a. 3.644×10^{-24} J; 5.45×10^{-2}m {$E_{photon} = h\nu = (6.626 \times 10^{-34}Js)(5.5 \times 10^{9}s^{-1}) = 3.644 \times 10^{-24}$ J; $\nu = c/\lambda$;
$\lambda = (3 \times 10^{8}$ m/s)/[$5.5 \times 10^{9}s^{-1}$)] = 5.45×10^{-2}m}

b. 2.19J {3.644×10^{-24} J/photon x (6.022×10^{23}photons/1mol photons = 2.19J/mol photons}

26. a

27. 1.61×10^{-7}m {745kJ/mol x (1000J/kJ) x (1mol/6.022×10^{23}) = 1.24×10^{-18}J; E = hc/λ; λ = hc/E =
[6.626×10^{-34})(3×10^{8})]/[1.24×10^{-18}] = 1.61×10^{-7}m}

28. a. II {The lower the energy, the longer the wavelength; IR has the lowest energy.} b. IV c. IV d. II and IV

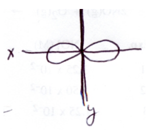
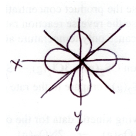
29. c {as the energy levels get to higher n values, they get closer together in energy; hence, the 2 furthest energy levels would be 4 and 1. The further they are, the larger the energy difference → the larger the ΔE → the shorter the wavelength since $E \propto 1/\lambda$}

30. b {$\Delta E = -Rhc[^1/(n_f)^2 - {}^1/(n_i)^2] = -Rhc[^1/(2)^2 - {}^1/(6)^2] = -Rhc[^1/4 - {}^1/36] = -Rhc[^9/36 - {}^1/36] = -Rhc[^8/36] = -Rhc[^2/9]$}

31. $2.34 \times 10^{14}\ s^{-1}$ {$\Delta E = -2.18 \times 10^{-18}\ J\,[1/(n_f)^2 - 1/(n_i)^2] = -2.18 \times 10^{-18}\ J\,[1/(5)^2 - 1/(3)^2] = 1.55 \times 10^{-19}\ J$; $E = h\nu \rightarrow$ $\nu = E/h = 1.55 \times 10^{-19}\ J/6.626 \times 10^{-34} Js = 2.34 \times 10^{14}\ s^{-1}$}

32. e {$\lambda = h/mv \rightarrow v = h/m\lambda = (6.626 \times 10^{-34} Js)/[(9.11 \times 10^{-28} g \times (1kg/1000g)][1 \times 10^2\ nm \times (1m/1 \times 10^9 nm)] = 7.27 \times 10^3\ m/s$}

33. e {(2, 3, 2, $+^1/_2$) and (4, 4, 2, $+^1/_2$): l value invalid: must be $n - 1$ or less; (4, 3, 4, $+^1/_2$): m_l invalid because m_l can not be greater than l quantum number; (4, 3, 2, $+1^1/_2$): m_s invalid it must be $+^1/_2$ or $-^1/_2$}

34. a.  b. 

35. a. shape b. orientation

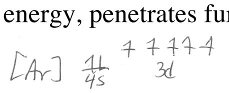
c. The Heisenberg Uncertainty Principle states that *both* the position and energy of electrons can not be known. If one of these characteristics is known exactly, the other characteristic can not be known exactly.

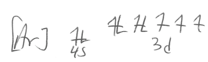
d. Light can be diffracted with a slit and diffraction is a wave property. Light with high enough frequency can eject electrons from metals and only light with this threshold frequency can eject electrons independent of intensity. This is a particle characteristic and this experiment is called the Photoelectric effect.

e. Electrons can be diffracted; diffraction is a wave property. Electrons have mass; mass is a particle property.

36. I. a. 16 b. 5 c. 3 d. 0 {when n = 5, $l \neq 5$} e. 1 II. a. 32 b. 25 c. 10

37. a. III b. V {order of filling from lowest to highest E: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, etc.}
   c. I {the subshell fills first has lower energy and penetrates closer to the nucleus}
   d. IV {total nodes = n–1 = 4-1 = 3} e. III {nodal surfaces = l = 2}
   f. I {the subshell fills first has lower energy, penetrates further, and its electrons experience a greater effective nuclear charge}

38. I. a. $1s^22s^22p^63s^23p^64s^23d^5$ b.  c. 5 d. 4

II. a. $1s^22s^22p^63s^23p^64s^23d^7$ b.  c. 3 d. 4

39. a. $1s^22s^22p^63s^23p^63d^6$ b. $1s^22s^22p^63s^23p^5$
   c. $1s^22s^22p^63s^23p^64s^13d^{10}$ {This is an exception; you'd expect from the Periodic Table $[Ar]4s^23d^9$ but 1 of the 4s electrons moves into the 3d subshell to give a half filled 4s and completely filled 3d subshell which is more stable.}

40. a. Rb {Radii increase as you go to the left and down on the Periodic Table}
   b. Xe^+ {Cations < Anions → Xe^+; Xe^+ in the 4th period versus the other choices which are in the 5th – hence, smaller}
   c. Ne {IE increase as you go to the right and up on the Periodic Table.}
   d. I {IE increases as you go up and to the right on the Periodic Table.}
   e. Si {Mg has a higher IE than Al because it has a filled 3s subshell which stabilizes it}
   f. N {this is an exception because N has a half-filled subshell; more stable}
   g. Na {Look at positions of the +1 ions, Na^+, Mg^+, Al^+ since this this the IE2 which involve removing the 2nd electron from the +1 ion. IE2 trend is higher as you go up and to the right on the Periodic Table. Na^+ is in the Ne position.}
   h. N_3^- {The number of electrons in CNO^- is 6 + 7 +8 +1 = 22; only N_3^- has that number: 3(7) +1 = $22e^-$}
   i. Cl {This is an exception; Cl has the greatest EA}
   j. Se^{-2} {greater the negative charge, usually the larger it is; given three -2 anions, the largest is the one furthest down the PT}

41. b {Radii increase as you go to the left and down on the Periodic Table}

42. b {Mn: $[Ar]4s^23d^5$; Mn^{+2}: $[Ar]3d^5$ since s electrons are removed 1st}

43. a {Look at positions of the +1 ions, Ca^+, Rb^+, K^+ since this is the IE2 which involve removing the 2nd electron from the +1 ion. IE2 trend is higher as you go up and to the right on the Periodic Table. K^+ is in the Ar position.}

44. c {removing an electron from a cation is more difficult than removing an electron from an anion or a neutral atom}

45. a {D = m/V; Na would be the largest in size and has the smallest mass so would be the least dense}

46. d {Since the IE_4 is so large as compared to the other IE, the removal of the 4th electron probably breaks a noble gas configuration. Only Al breaks a noble gas configuration during IE_4: $Al^{+3} \rightarrow Al^{+4} + e^-$}

47. a. system b. volume c. Schrodinger d. decreases e. energy f. –250J/g g. alkali metals