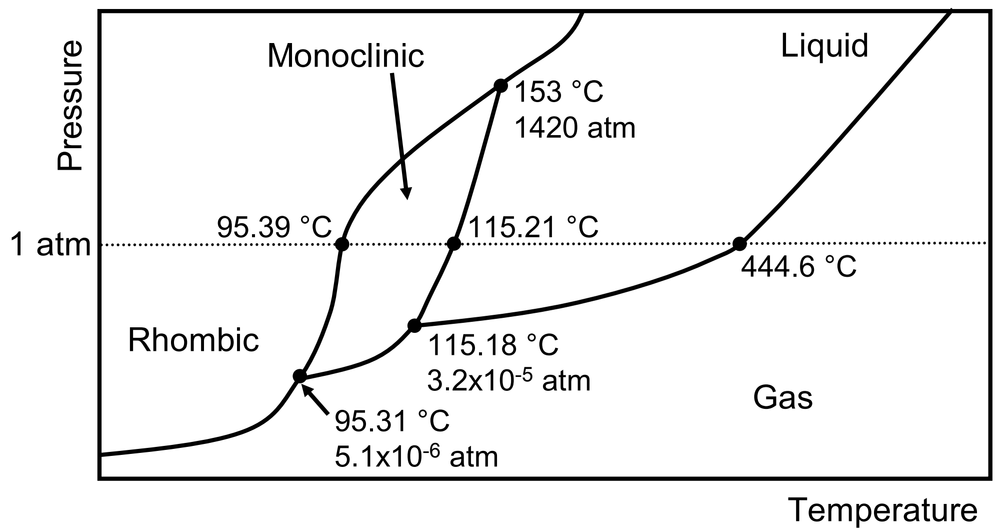
**CHEM 121**
**Spring 2006**
**Quiz 8**

Name: ______________________________________________

1. Consider the phase diagram for sulfur shown below. Note that the scales are distorted.




a. (1 Point) What phase is stable at room temperature and 1.0 atm pressure?

**The rhombic phase is the most stable under those conditions.**

b. (3 Points) Can monoclinic sulfur exist in equilibrium with sulfur vapor? If so, under what conditions?

**Yes it can. The conditions are denoted by the line that runs from the triple point at 95.31 °C and $5.1 \times 10^{-6}$ atm of pressure to the triple point 115.18 °C and $3.2 \times 10^{-6}$ atm of pressure.**

c. (2 Points) What are the normal melting point and normal boiling point of sulfur?

**The normal melting point is 115.21 °C and the normal boiling point is 444.6 °C.**

d. (1 Point) Which is the denser solid phase, monoclinic or rhombic sulfur?

**Rhombic sulfur is denser than monoclinic sulfur.**

*At constant temperature a pressure increase will force the particles in the substance closer together, and thus the density will increase (same mass in a smaller volume). If we start in a region where monoclinic is the stable solid phase and increase the pressure at constant temperature, we will eventually come to a phase transition to the rhombic form, which means the rhombic form is more dense. In general, if the line that separates two* *phases on a phase diagram has a positive slope, then the phase that is stable on the high pressure side has a higher density. If the slope is negative, then this relationship is reversed. Note that for most substances the line that describes the solid-liquid phase transition has a positive slope, and so the solid is denser than the liquid. Water is the one common substance which has a negative slope for the line that describes the solid-liquid phase transition (ice is less dense than liquid water).*

2. (8 Points) Determine the molarity of a 1.00 m solution of acetone dissolved in ethanol. Assume that the final volume equals the sum of the volumes of acetone and ethanol. The molar mass of acetone is 58.08 g/mole and its density is 0.788 g/cm$^3$. The molar mass of ethanol is 46.07 g/mole and its density is 0.789 g/cm$^3$.

*Definition of molarity is moles of solute per liter of solution. We have molality, which is defined as moles of solute per kilogram of solvent. We need to find the volume of the 1.00 m solution, which we will assume is the sum of the volume of acetone and the volume of ethanol.*

*Since ethanol is the solvent, we can write the following.*

$$1.00\text{x}10^3 \text{ g ethanol}\left(\frac{1.000\text{x}10^{-3} \text{ L ethanol}}{0.789 \text{ g ethanol}}\right) = 1.26_7 \text{ L ethanol}$$

*The volume of acetone is determined as follows.*

$$1.00 \text{ mole acetone}\left(\frac{58.08 \text{ g acetone}}{1 \text{ mole acetone}}\right)\left(\frac{1.000\text{x}10^{-3} \text{ L acetone}}{0.788 \text{ g acetone}}\right) = 0.0737_0 \text{ L acetone}$$

*The total volume is*

$$0.0737_0 \text{ L acetone} + 1.26_7 \text{ L ethanol} = 1.34_0 \text{ L solution}$$

*The molarity is*

$$\frac{1.00 \text{ mole acetone}}{1.34_0 \text{ L solution}} = 0.746 \text{ M}$$

**The molarity of a 1.00 m solution of acetone in ethanol is 0.746 M.**

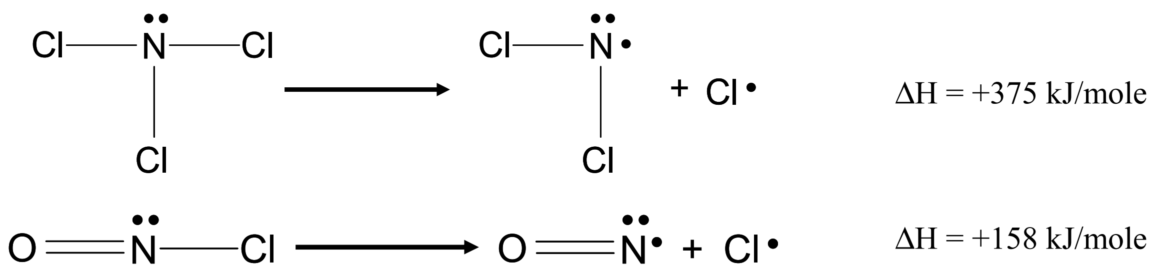
3. (4 Points) Explain why the solubility of all gases decreases as temperature increases.

**Forcing a gas to dissolve in a liquid decreases the number of energy states available to the gas, or in other words, increases the order within the system. Thus, the equilibrium gas + liquid $\rightleftharpoons$ solution has $\Delta S < 0$. For any equilibrium $\Delta G = \Delta H - T\Delta S$, and for the equilibrium under consideration $\Delta S$ is negative. So, the $T\Delta S$ term will tend to make $\Delta G$ positive, which means this equilibrium lies toward the left (gas and liquid separate, not in solution). As the temperature increases the $T\Delta S$ makes $\Delta G$ more positive and drives the equilibrium further to the left and less gas will dissolve in the solvent (solubility decreases).**

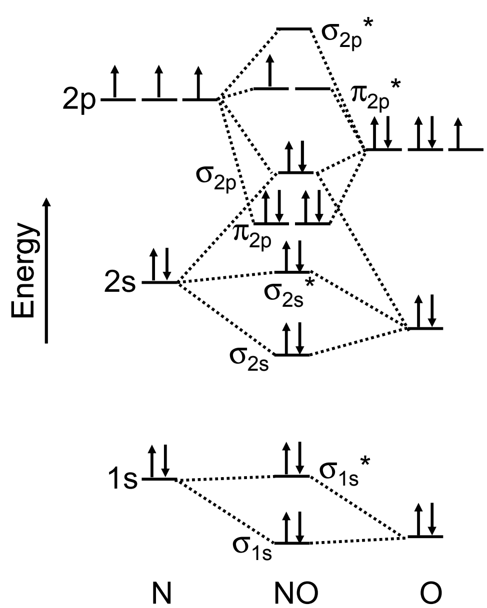
*Note that we expect that most gases will exhibit weak intermolecular interactions either between their own particles or between its particles and solvent particles. It is likely, therefore, that ΔH for the equilibrium between a free gas and a gas in solution is likely to be very small and make only a minimal contribution to ΔG for the equilibrium.*

4. (5 Points) Attach problem 14-54 to this sheet.

*The Lewis dot structures for each species involved in the reactions are given below (note that lone pairs are only shown on the N atoms.*




**In the gas phase ΔH for a reaction should be the sum of the bond dissociation energies for the bonds broken minus the sum of the bond dissociation energies of the bonds formed. Both of these reactions apparently only break the N–Cl bond, and should have the same ΔH. However, ΔH for the reaction of ONCl is much less than that of $NCl_3$, which suggests that either there is a bond being formed in ONCl or that an additional bond is broken in $NCl_3$. We could draw resonance structures for $NCl_3$ that have N=Cl, but these are not expected to be very important. As we could also write similar resonance structures for ONCl, we discard resonance structures of this type as the explanation for the difference in ΔH.**



**We must conclude that a bond is formed when ONCl dissociates Cl. The Lewis dot structure of NO does not give us any insight into where this additional bond is. However, the MO diagram of NO (shown above) shows that there is one electron in the $\pi_{2p}$*. This means that the bond order is 2.5 in NO, which is an increase in the bond order from the double bond present in ONCl indicated by the Lewis dot structure. This increase in bond order stabilizes the products and releases energy as heat. This makes the overall reaction less endothermic and accounts for the observed differences in ΔH.**

5. (5 Points) Attach problem 12-66 to this sheet.

a. **There are 32 electrons with $n = 4$.**

*The possible values of $\ell$ with n = 4 are 0, 1, 2, 3, corresponding to 4s, 4p, 4d and 4f orbitals. There is only one 4s orbital, but there are three 4p, five 4d and seven 4f, each corresponding to the allowed $m_\ell$ values for each $\ell$. So, with n = 4 there are 16 possible orbitals, each of which can hold two electrons, for a total of 32 electrons.*

b. **There are 8 electrons with n = 5 and $m_\ell = +1$.**

With $n = 5$ the possible $\ell$ values are 0, 1, 2, 3, and 4. Only $\ell = 1, 2, 3$ and 4 can have $m_\ell = +1$ and each of these will have only one orbital with $m\ell = +1$ (for a total of four orbitals). Each of these orbitals can hold two electrons each, for a total of eight.

c. **There are 25 electrons with $n = 5$ and $m_s = +1/2$.**

*With n = 5 the possible $\ell$ values are 0, 1, 2, 3, and 4. There is one 5s orbital, three 5p, five 5d, seven 5f and nine 4g orbitals. These twenty-five orbitals can hold 50 electrons, but only half of them can have $m_s = +1/2$.*

d. **There are 10 electrons with $n = 3$ and $\ell = 2$.**

*These quantum numbers describe the five 3d orbitals, which can hold 10 electrons in total.*

e. **There are 6 electrons with $n = 2$ and $\ell = 1$.**

*These quantum numbers describe the three 2p orbitals, each of which can hold 2 electrons for a total of 6.*

f. **There are no electrons with $n = 0$ because this is not an allowed quantum number.**

g. **These four quantum numbers uniquely define 1 electron.**

h. **There are 18 electrons with $n = 3$.**

*The possible orbitals are 3s, 3p and 3d. There is one 3s, three 3p and five 3d orbitals. Each orbital holds 2 electrons, for a total of 18.*

i. **There are no electrons with these quantum numbers because $\ell = 2$ is not possible when $n = 2$.**

j. **There are two electrons with the three quantum numbers $n = 1$, $\ell = 0$ and $m_\ell = 0$.**

*These three quantum numbers describe the 1s orbital, which can hold 2 electrons.*