Name:_____________

(4 points)

Chemistry 114
Second Hour Exam

Remember- Show all work for partial credit. Each problem is worth 12 points.

1. Define the following terms:

Octahedral hole -used to describe an ionic lattice where he cation occupies the space between 3 atoms in one layer and 3 atoms in the next layer. A diagram really helps this definition.

Heat of Fusion - The heat energy required to change one mole of solid to a liquid.

Substitutional alloy - An alloy refers to a mixture of a metal with another element. In a substitutional alloy the new element substitutes for an atom of the metal in the metals normal lattice.

Superheating - Heating a liquid above its normal boiling point, but the liquid has not yet turned into a vapor.

Triple point - The point on a phase diagram where 3 different phases can coexist at one temperature and pressure.

ccp - Cubic closest packed. The arrangement observed when spheres are paked into an a-b-c arrangement.

2A. PF_3 is a gas, PCl_3 is a liquid with a b.p. of 74°C, PBr_3 is a liquid with a b.p. of 175°C and PI_3 is a solid. Explain.

If these compounds were polar, then PF_3, with the most polar bonds would be the solid. However in this case it is a gas, arguing that all 4 compounds are non polar. In nonpolar compounds the London force is proportional to the size of the compound, so the largest compound PI_3 will have the largest intermolecular forces and be the solid, while PF_3 will have the weakest forces and be a gas. And the other compounds wil have intermediate forces and be liquids.

2B. H_2O has a b.p. of 100°C, H_2S has a b.p. of -60 °C, H_2Se has a b.p. of -40°C, and H_2Te has a b.p. of -10°C. Explain.

You can use the same argument that was used in 2A to explain H_2S,H_2Se, and H_2Te. However H_2O with the very high bp does not fit. In this case the polar hydrogen bonding nature of water add the hydrogen bonding forces to make this molecule the one with the largest intermolecular forces and the highest boiling point.

3. The normal boiling point of acetone is 56.5 C. The boiling point in Spearfish when the pressure is .88 atm is 52.9. What is the ΔH_{vap} for acetone?

$$\ln\left(\frac{P_1}{P_2}\right) = \frac{\Delta H}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right);$$

$$\ln\left(\frac{.88}{1}\right) = \frac{\Delta H}{8.3145}\left(\frac{1}{56.5 + 273.2} - \frac{1}{52.9 + 273.2}\right)$$

$$\ln(.88) \times 8.3145 = \Delta H(.003033 - .003067)$$

$$-1.063 = \Delta H(-.000034)$$

$$\Delta H = 313 kJ / mol$$

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4. Br_2 (Bromine) has a normal melting point of -7.2°C and a normal boiling point of 59°C. The triple point for Br_2 is -7.3°C @ 40 torr, and the critical point is 320°C and 100 atm.

A.) Sketch a phase diagram for Br_2.

Sorry its hard to sketch a phase diagram on a computer. It should look like figure 10.52 from your text, with a triple point at -7.3°C,40 torr, a critical point at 320°C and 100 torr, and at 1 atm it will be a solid below -7.2°C, a liquid between 72. And 59°C, and a gas above 59°C.

B.) Based on your phase diagram, order the three phases from most dense to least dense.

S→l→g

C.) What is the stable phase of Br_2 at room temperature and 1 atm pressure?

Liquid

D.) Under what temperature conditions can liquid bromine never exist

I was looking for >320°C, the critical T, but below -7.3°C also works.

E.) What phase changes occur as the temperature of a sample of bromine at .1 atm is decrease from 200°C to -50°C

g→l→s

5. The concentration of a solution, expressed as weight % is 5% $MgCl_2$.

A.) What is the mole fraction of $MgCl_2(aq)$ in this solution?

100 g of a 5% $MgCl_2$ solution would have 5 g of $MgCl_2$ and 95 g of water.

This would be 5/(24.31+2(35.45) or .05252 moles of $MgCl_2$

and

95/18 or 5.28 moles of water

The mole fraction $MgCl_2$ would then be .05252/(.05252+5.28)

Or χ = .00985

B.) What is the mole fraction of $Cl^-(aq)$ in this solution?

This is just a bit trickier

If we have .05252 moles of $MgCl_2$, then we have .05252 moles of Mg^{2+} and 2(.05252) or .105 moles of Cl^-

The mole fraction of Cl^- is then .105/(.105+.05252+5.28)

and χ = .00193

6. When the pressure of N_2 gas is 1atm, its concentration in water, expressed as a mole fraction χ is 1.386 x10^{-5}.

A. What is this concentration in units of molarity? (Assume the density of this solution is 1.00g/ml)

1.386×10^{-5} = mole N_2/(moles N_2 + moles H_2O)

Since we want to turn this into molarity we want 1 liter of solution. 1 liter of water would weigh 1000g and would contain 1000g/18 g/mol or 55.56 moles of water

$1.386\text{X}10^{-5} = X/(X+55.56)$
$1.386\times10^{-5}X + 7.7\times10^{-4} = X$
$7.7\times10^{-4} = X-.00001386X$
$7.7\times10^{-4} = .99999X$
$X=7.7\times10^{-4}$ moles / liter

B. What is the Henry's law constant for this gas? (The k in the equation C=kP)

7.7×10^{-4}mol/liter = k 1atm

$k=7.7\times10^{-4}$ mol/l·atm

7. A solution is prepared by mixing 5.81 g of acetone (C_3H_6O, molar mass 58.1) and 11.9 g chloroform ($HCCl_3$, molar mass 119.4 g/mol) at 35° this solution has a vapor pressure of 260 torr. The vapor pressure of pure acetone and pure chloroform is 345 and 293, respectively. Is this an ideal solution? Do you think the solution got warmer or colder when it was mixed together? Why?

Using Raoult's law the vapor pressure of this solution is

$\chi_{acetone}$ V.P.$_{acetone}$ + $\chi_{chloroform}$ + V.P.$_{chloroform}$

This solution contains 5.81g /58.1 g/mol or .1 moles acetone
The solution contains 11.9g/119.4 g/mol or .1 moles chloroform
so the mole fraction for both species is .1/(.1+.1) or .5

V.P. solution = .5(345) + .5(293) or 319 torr.

Since the actual V.P of 260 torr is lower than out theoretical value, we have a negative deviation from Raoult's law indicating that there are strong interactions between these two liquids in the solution. This would correlate with a negative $\Delta H_{solution}$ and this would meant that the solution got warmer as it was mixed.

8. I am going to spread $CaCl_2$ to melt the ice on my sidewalk. Assuming there is 30 liters of water on my sidewalk weighing 30 kg, what is the minimum amount of $CaCl_2$ I need to throw on the sidewalk to make all the ice melt at -5 °C. (K_f for water is 1.86 °C·kg/mol)

The freezing point depression equation is : $\Delta T = K_f\ m$

But in this case $CaCl_2$ is ionics, and makes 3 ions for each mole of solute
The proper equation to use is thus: $\Delta T = K_f\ 3\ m$

ΔT = 0-(-5) = 5
5=1.86(3)m
m=.896 mole/kg solvent
since I want to melt 30 kg of water
.896 = X moles/30 kg
and I want 26.88 mole or 26.88(40.08+2(35.45)) = 2,980 g or 2.98 kg of $CaCl_2$