Solutions to Problem Set 8

1. water saturated with H_2 and O_2 gases at a total pressure of 5 atm with (2:1 = H_2:O_2)

Equation	Basis for the equation	Eq. #
$5\text{ atm} = p_{H2}+p_{O2}$; $p_{H2}/p_{O2} = 2/1$ $p_{H2} = (2/3)\ 5\text{ atm} = 3.333\text{ atm}$ $p_{O2} = (1/3)\ 5\text{ atm} = 1.666\text{ atm}$ $p_{H2} = K_H(H_2)\ x_{H2}$ $3.333\text{ atm} = 7.76\times10^4\text{ atm } x_{H2}$ $x_{H2} = 3.333/7.76\times10^4 = 4.295\times10^{-5}$ $p_{O2} = K_H(O_2)\ x_{O2}$ $1.666\text{ atm} = 4.58\times10^4\text{ atm } x_{O2}$ $x_{O2} = 1.666/4.58\times10^4 = 3.639\times10^{-5}$ mole ratio of H_2:O_2 in solution was $4.295\times10^{-5}/3.639\times10^{-5} = 1.18$ or 54.1 mole% H_2 45.9 mole% O_2 in the gases driven off **Answer**	Given Henry's law applied to H_2 gas solubility Concentration in solution is directly proportional to the partial pressure of the gas above the solution p_{H2} = partial pressure of H_2 in the overhead gas x_{H2} = mole fraction of H_2 in the aqueous solution K_H (in atm) is given in provided table Solubility of these gases do not strictly follow Henry's law, since the apparent Henry's law constant from experiment varies with overhead partial pressure! Use the K_H at the closest partial pressure (as we have done here). Better yet, make a least squares fitting of apparent K_H vs p and use the equation to find the appropriate K_H to use at the partial pressures of this problem.	1 2 3 4

2. vapor pressure lowering by a solute

Equation	Basis for the equation	Eq. #
$p_{Bz}/p^*_{Bz} = x_{Bz}$ $74.01/74.66 = x_{Bz} = 0.9913$	Assuming ideal solution for the solvent, Raoult's law holds	1
$100\text{ g } C_6H_6/78\text{ g mol}^{-1} = 1.282\text{ mol Bz}$		2
$1.282/[1.282\text{ mol} + n] = 0.9913$ $n = 0.01125$ mol solute 2 g/MW g mol-1 = 0.01125 MW = 178 178 = c 12 + h 1 c 12/178 = 0.944 c = 14 178 = 14(12) + h(1) h = 10 The molecular formula is $C_{14}H_{10}$ **Answer**	Find n number of moles of solute in 100 g Bz Find molecular weight of solute For molecular formula C_cH_h weight fraction of carbon in this formula is 94.4% Solve for c Substitute c into Eq 6	3 4 5 6 7 8

3. Freezing point depression by a solute

$$H_2O(liq) \Leftrightarrow H_2O(s)$$

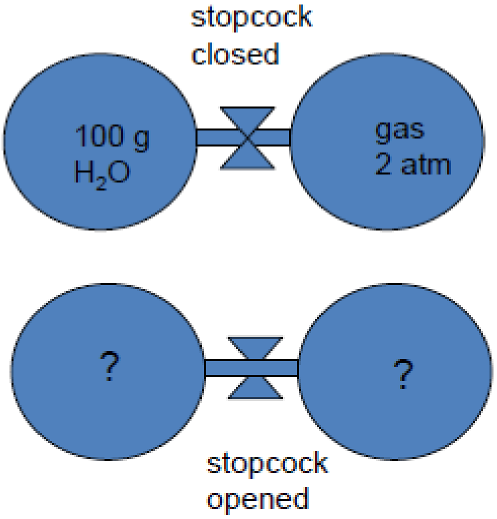
Equation	Basis for the equation	Eq. #
$\mu^*_{\mathrm{H2O,liquid}}(T,p^*) = \mu^*_{\mathrm{H2O,vap}}(T,p^*)$	Pure liquid H_2O in equilibrium with its vapor, * means pure	1
$\mu^*_{\mathrm{H2O,vap}}(T,p^*) = \mu^{\ominus}_{\mathrm{H2O,vap}}(T) + RT\ln(p^*/1\mathrm{atm})$	For the vapor at a given temperature, behaving like an ideal gas Recall that we used d**G** = Vdp when dT = 0 and replaced V by RT/p to get this. If non-ideal vapor use RT ln (f/1atm) instead	2
$\mu_{\mathrm{H2O,liquid\ solution}}(T,p) = \mu_{\mathrm{H2O,vap}}(T,p)$ $\mu_{\mathrm{H2O,vap}}(T,p) = \mu^{\ominus}_{\mathrm{H2O,vap}}(T) + RT\ln(p_{\mathrm{H2O}}/1\mathrm{atm})$	H_2O in aqueous solution in equilibrium with its vapor. As above, except that partial pressure p_{H2O} of H_2O in the vapor in equilibrium with liquid solution is less than p*	3 4
$\mu_{\mathrm{H2O,liquid\ solution}}(T,p) - \mu^*_{\mathrm{H2O,liquid}}(T,p^*) = RT\ln(p_{\mathrm{H2O}}/p^*)$	Eq. (3+4) minus Eq (1+2)	5
$p_{\mathrm{H2O}}/p^* = x_{\mathrm{H2O}}$	Assume Raoult's law holds for the liquid solution	6
$\mu_{\mathrm{H2O,liquid\ solution}}(T,p) - \mu^*_{\mathrm{H2O,liquid}}(T,p^*) = RT\ln x_{\mathrm{H2O}}$	Substituting Eq 6 into Eq 5	7
For $H_2O(liq) \Leftrightarrow H_2O(s)$ at equilibrium at T^*_f and 1 atm, $\mu^*_{\mathrm{H2O,solid}}(T^*_f,1\ \mathrm{atm}) = \mu^*_{\mathrm{H2O,liquid}}(T^*_f,1\ \mathrm{atm})$	Pure liquid water in equilibrium with ice at the normal freezing point T^*_f	8
For H_2O(liq solution) $\Leftrightarrow H_2O(s)$ at equilibrium at T and 1 atm, $\mu^*_{\mathrm{H2O,solid}}(T,1\ \mathrm{atm}) = \mu_{\mathrm{H2O,liquid\ solution}}(T,1\ \mathrm{atm})$	Aqueous solution in equilibrium with pure ice at the new freezing point T	9
$\mu^*_{\mathrm{H2O,solid}}(T,p) = \mu^*_{\mathrm{H2O,liquid}}(T,p) + RT\ln x_{\mathrm{H2O}}$	Applying Eq 10 and Eq 7	10
$\ln x_{\mathrm{H2O}} = [\mu^*_{\mathrm{H2O,solid}}(T,p) - \mu^*_{\mathrm{H2O,liquid}}(T,p)]/RT$ $\ln x_{\mathrm{H2O}} = -[\Delta_{\mathrm{fus}}\mathbf{G}]/RT$	Rearrange Eq 11	11 12
$(\partial\ln x_{\mathrm{H2O}}/\partial x_{\mathrm{H2O}}) = -(1/R)\bullet[\partial(\Delta_{\mathrm{fus}}\mathbf{G}/T)/\partial T]_p\bullet(\partial T/\partial x_{\mathrm{H2O}})_p$	To discover how new freezing point T depends on the water concentration in the liquid solution x_{H2O}, p being constant, we need to evaluate $(\partial T/\partial x_{\mathrm{H2O}})_p$ Apply $(\partial/\partial x_{\mathrm{H2O}})$ to Eq 13	13
$[\partial(\Delta_{\mathrm{fus}}\mathbf{G}/T)/\partial T]_p = -\Delta_{\mathrm{fus}}\mathbf{H}/T^2$	Gibbs-Helmholtz equation which we derived previously	14
$(\partial\ln x_{\mathrm{H2O}}/\partial x_{\mathrm{H2O}}) = [\Delta_{\mathrm{fus}}\mathbf{H}/RT^2]\bullet(\partial T/\partial x_{\mathrm{H2O}})_p$	Substitute Eq 15 into Eq 14	15

$1/ x_{H2O} = [\Delta_{fus}\mathbf{H}/RT^2]\bullet (\partial T/\partial x_{H2O})_p$ $dx_{H2O}/ x_{H2O} = (\Delta_{fus}\mathbf{H}/R)\, dT/T^2$		16 17
$\int dx_{H2O}/ x_{H2O} = \int (\Delta_{fus}\mathbf{H}/R)\, dT/T^2$	Integrating	18
$\int_1^{xH2O} dx_{H2O}/ x_{H2O}$ $= \int_{T^*f}^{T} (\Delta_{fus}\mathbf{H}/R)\, dT/T^2$	Between the limits of 1 mole fraction (pure liquid water) and x_{H2O}, and corresponding temperatures T^*_f and new freezing temperature T.	19
$\ln x_{H2O} = (\Delta_{fus}\mathbf{H}/R)\{ - T^{-1} + T^*_f{}^{-1}\}$	Assume $\Delta_{fus}\mathbf{H}$ = 80 cal g^{-1} and is only mildly dependent on T over this tiny temperature range	20
$\ln x_{H2O} = (80/1.9872)\{ - 263^{-1} + 273^{-1}\}$ $= - 0.00561$ $x_{H2O} = 0.9944$	For T = 263 K	21
$n_{H2O} = 4000/18 = 222.22$ $x_{H2O} = n_{H2O}/(n_{H2O}+n_{solute})$ $0.9944 = 222.22/(222.22 +n_{solute})$ $n_{solute} = 1.25$ mol	Using density of water = 1 g mL^{-1}, given 4 L H_2O weighs 4000 g	22
g CH_3OH = 1.25(32) = 40 g $C_2H_4(OH)_2$ = 1.25(62) = 77.5 **Answers**	Weights of methanol and ethylene glycol to be dissolved in 4 L water to prevent ice formation at 263 K	23

4. Freezing point depression by a solute and osmotic pressure

Equation	Basis for the equation	Eq. #
$\ln x_{H2O} = (\Delta_{fus}\mathbf{H}/R)\{ - T^{-1} + T^*_f{}^{-1}\}$	Following same derivation as in problem 3 above	1
$\ln x_{H2O} =(80/1.9872)\{ -272.74^{-1} +273^{-1}\}$ $= - 2.016\times10^{-4}$	Substituting $\Delta_{fus}\mathbf{H}$ = 80 cal g^{-1} and T^*_f = 273.15 for water and given t = -0.406°C = 272.74 K	2
$V\pi = - RT \ln x_{H2O}$	following derivation in Problem Set 7 prob. 1(b)	3
$V\pi = - RT(- 2.016\times10^{-4})$	Substituting Eq 2 into Eq 3	4
$\pi = - (8.31451 \text{Pa m}^3\text{K}^{-1}\text{mol}^{-1})\ 298$ $\bullet(- 2.016\times10^{-4})/ 18\times10^{-6}\ \text{m}^3$ π = 27750 Pa **Answer**	Using the molar volume of water = 18×10^{-6} m^3 based on molar mass of 18g, density of 1 g cm^{-3} at 298 K. This answer is inaccurate because we assumed that NaCl is undissociated in solution and the solution is an ideal solution. To do the problem correctly, we will need to use activities of the + and - ions in the solution, all of which contribute to the osmotic pressure. We will find out more about ionic activities in Lecture notes part 9.	5
Alternatively, we also could have done this. $\ln x_{H2O} = \ln(1-x_s) \approx -x_s -(1/2)x_s^2 - (1/3)x_s^3 - ...$	Use the sum of mole fractions = 1 and expand the *ln* in a series	4a
$V\pi = RT\, x_s$	The relation between molar volume of liquid water the osmotic pressure and the mole fraction of salt	5a

5.



10 L bulbs, T = 298 K throughout
Vapor pressure of water = 24 mm Hg = 0.0316 atm
solubility of gas in water is 2 moles per 1000 g water

Equation	Basis for the equation	Eq. #
2 atm•10 L = n_{gas} (0.0820578 L atm K^{-1} mol^{-1}) 298 K n_{gas} = 0.818 mol	*Initially* with stopcock closed, the left bulb has water and water vapor. Volume of liquid water is very close to100 mL. Available volume for water vapor is (10. – 0.100) = 19.90 L. If behaving as an ideal gas Right bulb only has 0.818 mol gas in the 10 L volume.	1
	At equilibrium after stopcock is opened, some gas goes into the left bulb into solution, liquid water remains in the left bulb and the gas phase now occupies a total of 19.90 L containing water vapor plus gas	1
mol of gas that went into 100 g liquid water is 0.20 mol Gas phase therefore has 0.818-0.20 = 0.618 mol gas in 19.90 L plus that amount of water vapor that corresponds to a partial pressure of 0.0316 atm and 19.90 L	First assume amount of liquid water that has turned into water vapor is much much less than 100 g. then later test whether this is true or not. Using the given gas solubility	1

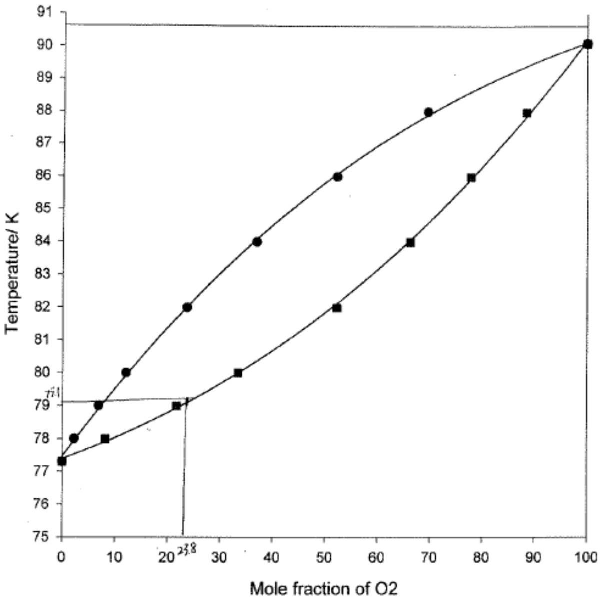
0.0316 atm•19.90 L = n (0.0820578 L atm K^{-1} mol^{-1}) 298 K $n_{\text{water vapor}}$ = 0.0257 mol g water vapor = 0.0257mol(18 g mol^{-1}) = 0.46 g	Assuming water vapor behaves as an ideal gas The mass of liquid water that this corresponds to is 0.46 g, occupying a volume of 0.46 mL So our assumption that the amount of liquid water that has turned into water vapor is much, much less than 100 g is quite good since 0.46 << 100.0. Also, the volume available to the gas phase is 19.90046, that is, 19.90 is therefore close enough.	1
In summary, equilibrium condition has 99.54 g of liquid water in the left bulb, and in the gas phase spread into both bulbs are 0.0257 mol water vapor and 0.618 mol of the gas. The partial pressure of the gas is 0.618 mol(0.0820578 L atm K^{-1} mol^{-1}) 298 K/ 19.9 L = 0.759 atm The total pressure in the gas phase is therefore 0.759 + 0.0316 = 0.791 atm **Answer**		2

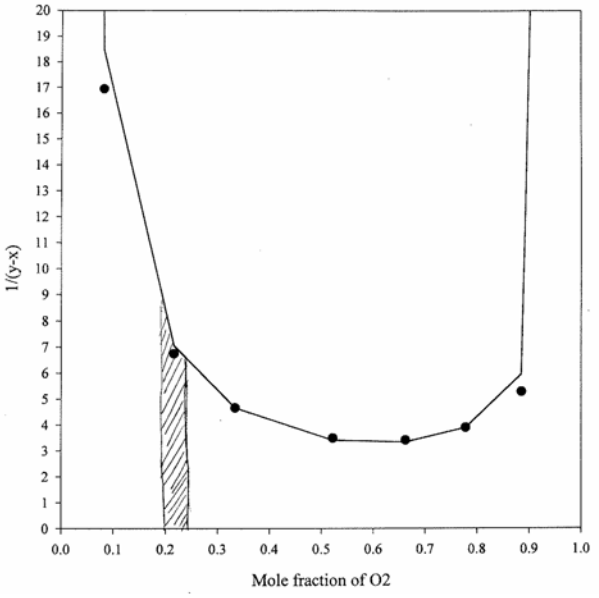
6. At 40°C 1 g of CO_2 will dissolve under 1 atm pressure in 1000 g H_2O.
At 20°C 1.7 g of CO_2 will dissolve under 1 atm pressure in 1000 g H_2O.

Equation	Basis for the equation	Eq. #
At 40°C 1 g of CO_2 will dissolve under 1 atm pressure in 1000 g H_2O. At 40°C 2 g of CO_2 will dissolve under 2 atm pressure in 1000 g H_2O.	Given We are allowed to assume Henry's law holds $p_{CO2} = K_H(CO_2)\, x_{CO2}$ at 40°C	1
Thus, we must have no more than 2 g of CO_2 in 1000 g H2O at 20°C At 20°C 1.7 g of CO_2 will dissolve under 1 atm pressure in 1000 g H_2O. this corresponds to a pressure of (2/1.7) = 1.176 atm of CO_2 at 20°C **Answer**	Using the given 2 atm pressure for the bursting limit (and we are neglecting the vapor pressure of water so that total pressure is entirely due to CO_2) Given Using $p_{CO2} = K_H(CO_2)\, x_{CO2}$ at 20°C	2
1.7 g is 1.7/44 = 0.0386 mol CO_2 1000 g is 1000/18 = 55.555 mol H_2O 1 atm = $K_H(CO_2)$ 0.0386/[55.555 + 0.0386] = $K_H(CO_2)$ 0.000695 $K_H(CO_2)$ = 1439.3 atm	Using molar masses of CO_2 and H_2O Find the $K_H(CO_2)$ at 20°C Applying $p_{CO2} = K_H(CO_2)\, x_{CO2}$ at 20°C	3

1.176 atm of CO_2 at 20°C therefore corresponds to $x_{CO2} = 1.176/1439 = 0.00082$ **Answer**	Substituting the value of K_H into $p_{CO2} / K_H(CO_2) = x_{CO2}$ at 20°C	

7. T vs X diagram for O_2 - N_2 mixture at 1 atm.





Let L be the amount of liquid remaining in the still at a given instant during the distillation. After distilling a differential amount dL, the amount of liquid will be (L - dL) and the amount of vapor formed will be dL. At the instant, the composition of the liquid is X_{O2} and the composition of the vapor is Y_{O2}. Consider material balance of O_2 component: O_2 overall = O_2 in the remaining liquid + O_2 in the vapor

$(L)(X_{O2}) = (L-dL)(X_{O2}-dX_{O2}) + (dL)(Y_{O2}+dY_{O2})$

Neglecting products of infinitesimals,

$(L)(X_{O2}) = (L)(X_{O2}) - (X_{O2})(dL) -(L)(dX_{O2}) + (dL)(Y_{O2})$ or $(Y_{O2} - X_{O2})(dL) = (L)(dX_{O2})$

$$\frac{(dL)}{(L)} = \frac{(dX_{O2})}{(Y_{O2} - X_{O2})}$$

LIQUID: the overall composition is 0.20 O_2 and 90 % is distilled, that is 10 % remaining

Integrating, from 10% up to 100%, $\ln (L_2/L_1) = \ln (1.00/0.10) = 2.303 = \int_{X}^{0.20} \frac{(dX_{O2})}{(Y_{O2} - X_{O2})}$

from molefraction 0.20 O_2 to X_{O2}. Plot the curve $(Y_{O2} - X_{O2})^{-1}$ vs X_{O2} (see figure above right). Now find the value of X_{O2} at which the integral underneath the curve between X_{O2} and 0.20 is equal to 2.303.

We read $X_{O2} = 0.23$ is the composition of the remaining liquid. **Answer**

And (reading T for which $X_{O2} = 0.23$ in the figure above left) this occurs at T = 79.1 K.

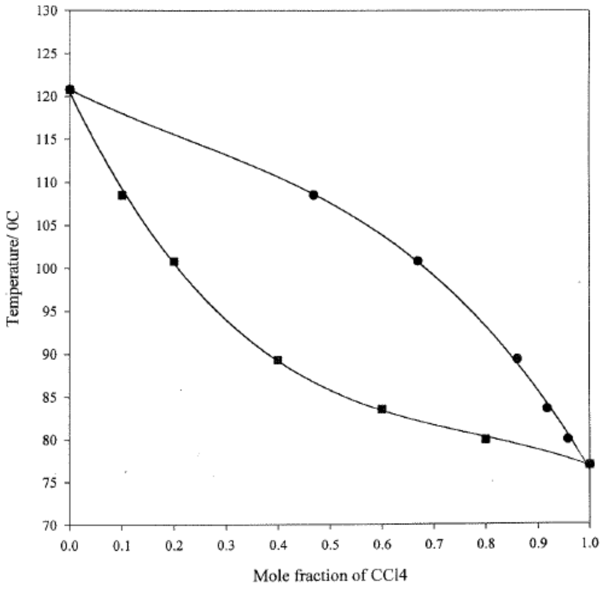
Answer

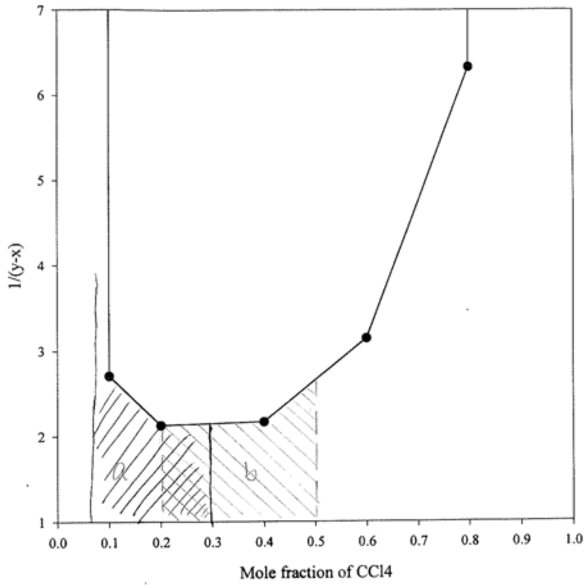
Calculate activities: $\mathbf{a}_{O2} = p_{O2}/p_{O2}^*$ $Y_{O2} = p_{O2}/p_{tot}$ $p_{tot} = 1.0$ atm $\mathbf{a}_{O2} = Y_{O2} \bullet (1 atm/p_{O2}^*{}_T)$

We can get p^*_T using the Clausius Clapeyron Eq

X_{O2}	0.0	0.081	0.216	0.334	0.522	0.662	0.778	0.885	1.0
$\mathbf{a}_{O2}$	0	0.022 $/p_{O2}^*{}_{78K}$	0.068 $/p_{O2}^*{}_{79K}$	0.12 $/p_{O2}^*{}_{80K}$	0.236 $/p_{O2}^*{}_{82K}$	0.369 $/p_{O2}^*{}_{84K}$	0.522 $/p_{O2}^*{}_{86K}$	0.696 $/p_{O2}^*{}_{88K}$	1.0

8. T vs X diagram for CCl_4 - C_2Cl_4 mixture at 1 atm.





(a) LIQUID: the overall composition is 0.30 CCl_4 and 50 % is distilled, that is 50 % remaining

Integrating, from 50% up to 100%, $\ln (L_2/L_1) = \ln (1.00/0.50) = 0.693 = \int_{X}^{0.30} \frac{(dX)}{(Y - X)}$

from molefraction 0.30 CCl_4 to X. Plot the curve $(Y - X)^{-1}$ vs X (see figure above right). Now find the value of X at which the integral underneath the curve between X and 0.30 (hatched area a) is equal to 0.693.

We read X = 0.044 is the CCl_4 composition of the remaining liquid. To find the Y (in the collected distillate), apply mass balance for CCl_4:

CCl_4 overall = CCl_4 in the remaining liquid + CCl_4 *in the collected distillate*

0.30 L = 0.044(L/2) + Y(L/2); Solve for Y, Y = 0.556 **Answer**

(b) LIQUID: the overall composition is 0.50 CCl_4 and is distilled until remaining liquid is 0.20 CCl_4. If the distillate is removed continuously, then at the instant that the liquid is 0.20 CCl_4, *the vapor coming off* is 0.67 CCl_4 (as seen in the table). **Answer**

On the other hand, if looking for Y composition in the total collected distillate, first find the area underneath the curve (figure above right) between X = 0.20 and X = 0.50, and

$\ln (1.00/f) = \int_{0.20}^{0.50} \frac{(dX)}{(Y - X)}$

The hatched area b provides the value of the fraction f of the liquid remaining.

We find the area = 0.93 = ln (1/f); 1/f = 2.53; thus f = 0.395

Mass balance of CCl_4 is applied, knowing f = 0.395, to find Y the *CCl_4 composition of the total collected distillate*.

0.50 L = 0.20L(0.395) + Y L(0.605) Solve for Y, Y = 0.70 **Answer**

This is larger than 0.67. Since CCl_4 is the lower boiling component, at the beginning where the liquid is 0.50 CCl_4, the first vapor to come off is about 0.90 CCl_4, and this Y value diminishes to the point where the remaining liquid is 0.20 CCl_4 the vapor coming off is 0.67 CCl_4, as we saw in the table. Therefore the total collected vapor has to have Y CCl_4 greater than 0.67

9. 1 g of antibiotic substance MW = 10000 g mol^{-1} is dissolved in 100 g H_2O.

Equation	Basis for the equation	Eq. #
1 g is 1 /10000 = 0.0001 mol substance 100 g is 100/18 = 5.5555 mol H_2O $x_s = 0.0001/(5.5555+.0001) = 1.8\times10^{-5}$ $1-x_s = x_{H2O}$ $\ln x_{H2O} = \ln(1-x_s) \approx -x_s -(1/2)x_s^2 -(1/3)x_s^3 - ...$ $\ln x_{H2O} = -x_s$ is close enough	Expansion of $\ln(1-x_s)$ for $x_s <<< 1$, as in this case	1 2 3
$p_A /p^*_A = x_A$	Raoult's law holds for the solution, here solvent A is H_2O	4
Since x_s is very small here, x_{H2O} is very close to 1 so p_A will be quite close to p^*_A thus, this is not a practical test.		
$\ln x_A = \Delta_{fus}H/R\left[-\frac{1}{T}+\frac{1}{T^*}\right]$ $-x_s = (80/1.9872)\ (T-T^*)/T^{*2}$ $-1.8\times10^{-5} = 40.26[-(T^*-T)/273^2]$ $(T^*-T) = 0.033$ deg **Answer** freezing point depression is small, not a very practical test	freezing point depression following derivation in problem 3 above	5 6
$\ln x_A = \Delta_{vap}H/R\left[\frac{1}{T}-\frac{1}{T^*}\right]$	boiling point elevation, by an analogous derivation	7
$-x_s = (539.43/1.9872)\ (T^*-T)/373^2$ $-1.8\times10^{-5} = 271.45[-(T^*-T)/373^2]$ $(T-T^*) = 0.0092$ deg **Answer** boiling point elevation is very small, not a very practical test		8
$V\pi = RT\ x_s$	osmotic pressure following derivation in Problem Set 7 prob. 1(b) and problem 4 above	9
$\pi = RT\ x_s/V$ $\pi = (8.31451 Pa\ m^3K^{-1}mol^{-1})\ 298 \bullet(1.8\times10^{-5})/\ 18\times10^{-6}\ m^3$ = 2477 Pa or 18.57 mm Hg **Answer** This is easily measurable and this is the preferred test, especially since only a few mg of substance is available for testing. **Answer**	Using the molar volume of water= $18\times10^{-6}\ m^3$	10

10. vapor pressure of zinc at 773 K, $T_{boil} = 1180$ K

Equation	Basis for the equation	Eq. #
$\ln(p/1\text{atm})=(\mathbf{H}_{vap}-\mathbf{H}_{liq})/RT_{boil}]$ $\bullet[1-T_{boil}/T]$ If we know $\Delta_{vap}\mathbf{H}$ at T_{boil} ,we can find vapor pressure p at temperature T $\Delta_{vap}\mathbf{H}$ for the phase change at 1180 K is 21 (1180) = 24780 cal mol^{-1}	As we derived in Problem Set 7 prob 6 when we could assume that the vapor behaved like an ideal gas. (the Clausius Clapeyron equation) Since this value is not given, we estimate the value of $\Delta_{vap}\mathbf{H}$ at T_{boil} Trouton's rule: $\Delta_{vap}\mathbf{H}/T_{boil} \approx 21$ cal mol^{-1} K^{-1}	1 2
For any temperature T, we need to correct this $\Delta_{vap}\mathbf{H}$ value by using $(\partial\mathbf{H}/\partial T)_p\, dT$ integrated from 1180 K to T for both the liquid and the vapor $\Delta_{vap}\mathbf{H}_T = \Delta_{vap}\mathbf{H}_{1180\text{ K}} + \int_{1180}^{T} [C_p(\text{vap}) - C_p(\text{liq})]\, dT$ $\Delta_{vap}\mathbf{H}_T = \Delta_{vap}\mathbf{H}_{1180\text{ K}} + \int_{1180}^{T} [4.968 - 7.09 + 1.15\text{x}10^{-3}\, T]\, dT$ $\Delta_{vap}\mathbf{H}_T = \Delta_{vap}\mathbf{H}_{1180\text{ K}} + 2.122(T-1180) + 1.15\text{x}10^{-3}\,(1/2)[T^2 - 1180^2]$ $\Delta_{vap}\mathbf{H}_{773} = \Delta_{vap}\mathbf{H}_{1180\text{ K}} - 2.122(773-1180) + 0.000575\,[773^2-1180^2]$ $= \Delta_{vap}\mathbf{H}_{1180\text{ K}} + 863.6 - 457.0$ $= \Delta_{vap}\mathbf{H}_{1180\text{ K}} + 406.5$ $\Delta_{vap}\mathbf{H}_{773} = 24780 + 406.5 = 25186$ As expected, $\Delta_{vap}\mathbf{H}$ is higher at the lower temperature. Our answer would be in error if we had assumed $\Delta_{vap}\mathbf{H}$ is independent.of T	Here we are allowed to assume that C_p (vap) = (5/2)R = 4.968 cal mol^{-1} and we are given $C_p(\text{liq}) = 7.09 - 1.15\text{x}10^{-3}\, T$ in cal mol^{-1}	3
$\ln(p/1\text{atm}) = (\Delta_{vap}\mathbf{H}_{773}/R)\bullet [1-(1180/773)]/1180$ $\ln(p/1\text{atm}) = 12672\bullet -4.46\times10^{-4} = -5.65$ p/1 atm = exp(-5.65) = 0.00352 **Answer** On the other hand, if we used $\Delta_{vap}\mathbf{H}_{1180\text{ K}} = 24780$ $\ln(p/1\text{atm}) = 12468\bullet -4.46\times10^{-4} = -5.56$ p/1 atm = exp(-5.56) = 0.00385 The vapor pressure would be 9% too high if we assume that $\Delta_{vap}\mathbf{H}$ is independent.of T **Answer**	Using Eq 1 (the Clausius Clapeyron equation) Actually our largest error is in assuming that Trouton's rule holds.	4