Practice Exercise page 24

Calculate the molality and mole fraction of NH_3 if it is in a solution composed of 30.6 g NH_3 in 81.3 g of H_2O. The density of the solution is 0.982 g/mL, density of water is 1.0 g/mL. (N = 14, H = 1, O = 16)

Answer:

Molality (m) = $n_{(solute)}$/ mass $_{(solvent\ in\ Kg)}$,

$n_{(solute)}$ = its mass / molar mass

so = $n_{(NH3)}$ = 30.6 / 17 = 1.8 moles

mass of solvent in Kg = 81.3 / 1000 = 0.0813 kg

so m = 1.44 / 0.0813 = 22.1 mole / kg solvent

mole fraction of NH_3 (X_{NH3}) = n_{NH3} / $n_{(total)}$

$n_{total} = n_{(NH3)} + n_{(H2O)}$

$n_{(H2O)}$ = its mass / molar mass = 81.3 /18 = 4.52 moles

so $n_{(total)} = n_{(NH3)} + n_{(H2O)}$ = 1.8 + 4.52 = 6.32 moles

then $X_{(NH3)}$ = 1.8/6.32 = 0.285

Given:

Mass of solute = 30.6 g

Mass of solvent = 81.3 g

molar mass$_{(NH3)}$ = 14 + 3x1 = 17 g/mol

molar mass$_{(H2O)}$ = 2x1+16 = 18

Practice Exercise page 34

Oxygen is much less soluble in water than carbon dioxide, 0.00412 g/100 mL at 20 °C and 760 mm Hg. Calculate the solubility of oxygen gas in water at 20 °C and a pressure of 1150 mm Hg.

Answer:

$m = k P$

since m is the solubility of gas in a liquid, and k: is Hennery's constant, and P: is the pressure of gas above liquid.

$m_1 / m_2 = P_1/P_2$ (1)

Apply in (1) by the given values: m_2 = 0.0062 g/100 mL

Given:

m_1 = 0.00412 g/100 ,

P_1 = 760 mg Hg

m_2 = ?? ,

P2 = 1150 mm Hg

Practice Exercise page 43

Benzene and toluene form an ideal solution at all proportion. The vapor pressure of pure benzene and pure toluene at 25 ^oC are 95.1 and 28.4 mmHg, respectively. What is the vapor pressure of a solution that contains 50.0 g benzene (mw 78.11) and 50.0 g toluene (mw 92.13)? What is the composition of the vapor that is in equilibrium with this solution at 25 ^oC?

Answer:

Since solution obeys Raoult's law, then, $\mathbf{P_t = X_A \,.\, P^o_A + X_B \,.\, P^o_B}$ **(1)**

P^o_A and P^o_B are the vapor pressure of the pure solvents, and XA, XB are their respective mole fractions.

Let benzene = A, and toluene = B

$n_A = \frac{m_A}{m.w_A} = \frac{50}{78.11} = 0.640\ mole$

$n_A = \frac{m_B}{m.w_B} = \frac{50}{92.13} = 0.543\ mole$

$X_A = \frac{n_A}{n_A + n_B} = \frac{0.640}{0.640+0.543} = 0.541$

$X_B = 1.0 - 0.541 = 0.459$

apply in (1),

$P_{(soln)} = 0.541 \times 95.1 + 0.459 \times 28.4 = 64.48$ mmHg

Given:

Masses of components:

$m_{(A)} = 50.0$ g, $m_{(B)} = 50.0$ g

molar masses of components (g/mol):

$m.w_{(A)} = 78.11$, $m.w_{(B)} = 92.13$

$P^o_A = 95.1$ mmHg,

$P^o_B = 28.4$ mmHg

to solve the second part, apply in Dalton's law of gases mixture which states that the partial pressure of gas in a gases mixture equals its mole fraction multiplied by the total pressure of that mixture:

$P_A = Y_A \,.\, P_{(t)}$, since Y_A is the mole fraction of A in vapor state.

$P_A = 0.541 \times 95.1 = 51.45$ mmHg , $P_B = 0.459 \times 28.4 = 13.04$ mmHg , P (t) = 64.48

So $Y_A = 64.48 / 51.45 = 0.8$, $Y_B = 0.2$

Practice Exercise page 60

How many grams of nonvolatile compound B (molar mass= 97.80 g/mol) would need to be added to 180.0 g of water to produce a solution with a vapor pressure of 40.362 torr? The vapor pressure of water at this temperature is 42.362 torr.

Answer:

$$\mathbf{P_{solution} = (\chi_{solvent})\,(P^\circ_{solvent})}$$

$40.362 = X_A \times 42.362$, then $X_A = 0.953$

So $X_B = 1.0 - 0.953 = 0.047$

$X_B = \frac{n_B}{n_A + n_B}$, $\quad 0.047 = \frac{n_B}{180/18 + n_B}$ so

$0.47 + 0.047\, n_B = n_B$ so $n_B = 0.493$ moles

n_B = mass/molar mass

so mass of B = $0.493 \times 97.80 = 48.2$ g

Given:

molar mass of B= 97.80 g/mol

mass of A (water) = 180 g

molar mass of water = 18 g/mol

V.P of solution = 40.362 torr

V.P of water (P^O_A) = 42.362 torr

Practice Exercise page 71

Find the boiling point of a solution of 92.1 g of iodine, I_2, in 800.0 g of chloroform, $CHCl_3$, assuming that the solution is ideal and the iodine is non-volatile non-electrolyte. (atomic mass of I = 126.9 amu).

Answer:

For non-volatile non-electrolyte solute, $\mathbf{\Delta T_b = k_b . m}$

n(solute) = mass /molar mass = 92.1/ 253.8 = 0.363 moles

m = n(solute) / kg of solvent

= 0.363 / 0.8 = 0.454 mol/Kg

$\Delta T_b = 3.63 \times 0.454 = 1.65$ °C

So B.P (solution) = ΔT_b + B.P (solvent)

= 1.65 + 61.26 = 62.91 °C

Mass of solute = 92.1 g

Mass of solvent = 800.0 g

Molar mass of I_2 = 126.9 × 2 = 253.8

B.P of solvent = 61.26 °C

K_b of solvent = 3.63

Practice Exercise page75

What is the osmotic pressure (atm) of a solution with a volume of 0.750 L and containing 5.0 g of methanol, CH_3OH, in water at 37 °C?

Answer:

$$\pi = M . R . T$$

since π is the osmotic pressure of a solution of molarity (M), at temperature of (T = t + 273), and R = universal gas constant = 0.0821 L . atm / mol. K

T = 37 + 273 = 310

M = n(solute) / V (soln) , n(solute) = mass / molar mass

molar mass of CH_3OH = 12 + 4 x 1.0 +16 = 32.0 g/mol

so n(solute) = 5.0 / 32.0 = 0.156 mol

M = 0.156/ 0.750 = 0.208 mol/ liter

So π = 0.208 x 0.0821 x 310 = 5. 3 atm