Sheet 1 (Let 3)

1) A mixture of water (Mw= 18 gram/mole) and acetone (Mw= 58 gram/mole) at 756 mmHg boils at 70°C. Calculate the molar fraction of each using the following table:

Temperature °C	Vapor pressure (atm) Acetone	Vapor pressure (atm) Water
60	1.14	0.198
70	1.58	0.312
80	2.12	0.456
90	2.81	0.694

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Sheet 1 / Solution

1) According to Raoult's law:

$P = x_{acetone} P_{acetone} + x_{water} P_{water}$

$P = \frac{756 \text{ mmHg}}{760 \text{ mmHg/atm}} = 0.995 \text{ atm}$

by substituting the values at 70°C we have

$0.995 = x_{acetone}\ 1.58 + x_{water}\ 0.312$

The sum of molar fractions is 1

$x_{aceton} + x_{water} = 1 \rightarrow x_{water} = 1 - x_{acetone}$

This

$x_{acetone} * 1.58 + 0.312\ (1 - x_{acetone}) = 0.995$

$x_{acetone} * 1.58 + 0.312 - 0.312\ x_{acetone} = 0.995$

$1.26\ x_{acetone} + 0.312 = 0.995$

$x_{acetone} = \frac{0.995 - 0.312}{1.26} = 0.54$

$x_{water} = 1 - 0.54 = 0.46$

2) A mixture of 40.0 g of oxygen (Mw 31.9988 g/mol) and 40.0 g of helium (Mw 4.0026 g/mol) has a total pressure of 0.900 atm. What is the partial pressure of each gas?

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2) Calculate moles of each gas

$$n_{He} = \frac{m_{He}}{Mw_{He}} = \frac{40\ g}{4.0026\ g/mol} = 9.9935\ mol$$

$$n_{O_2} = \frac{m_{O_2}}{Mw_{O_2}} = \frac{40\ g}{31.9988\ g/mol} = 1.25005\ mol$$

$$n_{total} = n_{He} + n_{O_2} = 9.9935 + 1.25005 = 11.24355\ mol$$

Calculate mole fraction

$$y_{He} = \frac{n_{He}}{n_{total}} = \frac{9.9935\ mol}{11.24355\ mol} = 0.88882$$

Calculate partial pressure

$$y_{He} = \frac{P_{He}}{P_{total}} \rightarrow P_{He} = y_{He} * P_{total}$$

$$P_{He} = 0.9\ atm * 0.88882 = 0.79938\ atm$$
$$\simeq 0.8\ atm$$

either

$$P_{total} = P_{He} + P_{O_2} \Rightarrow P_{O_2} = P_{total} - P_{He}$$
$$P_{O_2} = 0.9 - 0.8 = 0.1\ atm$$

or

$$y_{O_2} = \frac{n_{O_2}}{n_{total}} = \frac{1.25005\ mol}{11.24355\ mol} \simeq 0.11118$$

$$y_{O_2} = \frac{P_{O_2}}{P_{total}} \rightarrow P_{O_2} = y_{O_2} * P_{total} = 0.11118 * 0.9\ atm$$
$$P_{O_2} \simeq 1\ atm$$

3) A sample of 1.43 g of helium (Mw 4.0026 g/mol) and an unweighed quantity of O_2 (Mw 31.9988 g/mol) are mixed in a flask at room temperature. The partial pressure of helium in the flask is 42.5 torr, and the partial pressure of oxygen is 158 torr. What mass of O_2 is in the sample?

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3) $n_{He} = \frac{m_{He}}{Mw_{He}} = \frac{1.43\ g}{4.0026\ g/mol} = 0.35727\ mol$

$P_{total} = P_{He} + P_{O_2} = 42.5 + 158 = 200.5\ torr$

$y_{He} = \frac{P_{He}}{P_{total}} = \frac{42.5}{200.5} = 0.21197$

$y_{He} + y_{O_2} = 1 \Rightarrow y_{O_2} = 1 - y_{He} = 1 - 0.21197$

$y_{O_2} = 0.78803$

$y_{O_2} = \frac{n_{O_2}}{n_{total}} = \frac{n_{O_2}}{n_{He} + n_{O_2}}$

$0.78803 = \frac{n_{O_2}}{0.35727 + n_{O_2}}$

$n_{O_2} = 0.282 + 0.78803\ n_{O_2}$

$n_{O_2} - 0.78803\ n_{O_2} = 0.282$

$0.21197\ n_{O_2} = 0.282 \Rightarrow n_{O_2} = \frac{0.282}{0.21197}$

$n_{O_2} = 1.33\ mol$

$n_{O_2} = \frac{m_{O_2}}{Mw_{O_2}} \Rightarrow m_{O_2} = n_{O_2} * Mw_{O_2}$

$m_{O_2} = 1.33 * 31.9988 = 42.558\ gram$

4) At 60 C the vapor pressures of pure benzene and toluene are 0.513 and 0.185 bar, respectively. For a solution with 0.60 mole fraction toluene, what are the partial pressures of toluene and benzene, and what is the mole fraction of toluene in the vapor?

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4) $x_{tol.} + x_{ben.} = 1$

$x_{ben.} = 1 - x_{tol}$

$x_{ben} = 1 - 0.6 = 0.4$

$P_{tol.} = x_{tol.} P^{\circ}_{tol.} = 0.6 * 0.185 = 0.111$ bar

$P_{ben.} = x_{ben.} P^{\circ}_{ben.} = 0.4 * 0.513 = 0.205$ bar

$P_{total} = P_{tol.} + P_{ben.} = 0.111 + 0.205 = 0.316$ bar

either

$$y_{tol.} = \frac{P_{tol.}}{P_{total}} = \frac{0.111}{0.316} = 0.351$$

or

$$y_{tol} = \frac{x_{tol} P^{\circ}_{tol}}{P^{\circ}_{ben} + (P^{\circ}_{tol} - P^{\circ}_{ben}) x_{tol.}}$$

$$= \frac{0.6 * 0.185}{0.513 + (0.185 - 0.513)(0.6)}$$

$= 0.351$

5) The vapor pressure of 1-propanol is 10.0 torr at 14.7 °C. Calculate the vapor pressure at 52.8 °C. Heat of vaporization of 1-propanol = 47.2 kJ/mol.

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5) $\ln\left(\frac{P_1}{P_2}\right) = \frac{-\Delta H_{vap}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$

$\Delta H_{vap} = 47.2 \frac{KJ}{mol} * 1000 \frac{J}{KJ} = 47200 \frac{J}{mol}$

$T_K = T_{°C} + 273.5$

$T_1 = 14.7 + 273.5 = 288.2\ K$

$T_2 = 52.8 + 273.5 = 326.3\ K$

$\ln\left(\frac{10}{P_2}\right) = \frac{-47200}{8.314}\left(\frac{1}{288.2} - \frac{1}{326.3}\right)$

$\ln\left(\frac{10}{P_2}\right) = -5677.17104(0.00346981 - 0.00306466)$

$\ln\left(\frac{10}{P_2}\right) = -2.3$

Take exp of both side

$\exp\left(\ln\left(\frac{10}{P_2}\right)\right) = \exp(-2.3)$

$\frac{10}{P_2} = 0.10026$

$0.10026\ P_2 = 10 \rightarrow P_2 = \frac{10}{0.10026} = 99.74$ torr