

SET A:

1) Find the molarity of all ions in a solution that contains 0.165 moles of aluminum chloride in 820. ml solution.



Answer: $[\text{Al}^{3+}] = 0.201 \text{ M}$, $[\text{Cl}^-] = 0.603 \text{ M}$.

$$[\text{Al}^{3+}] = \frac{0.165 \text{ mole AlCl}_3}{0.820 \text{ L}} \left(\frac{1 \text{ mole Al}^{3+}}{1 \text{ mole AlCl}_3} \right) = 0.201 \frac{\text{mole Al}^{3+}}{\text{L}}$$

$$[\text{Cl}^-] = \frac{0.165 \text{ mole AlCl}_3}{0.820 \text{ L}} \left(\frac{3 \text{ moles Cl}^-}{1 \text{ mole AlCl}_3} \right) = 0.603 \frac{\text{moles Cl}^-}{\text{L}}$$

2) Find the molarity of each ion present after mixing 27 ml of 0.25 M HNO_3 with 36 ml of 0.42 M $\text{Ca}(\text{NO}_3)_2$. (Note: There is no reaction taking place.)

Answer: $[\text{H}^+] = 0.11 \text{ M}$, $[\text{NO}_3^-] = 0.58 \text{ M}$, $[\text{Ca}^{2+}] = 0.24 \text{ M}$.

1) Find # moles HNO_3 : $M_{\text{HNO}_3} V_{\text{HNO}_3} = 0.25 \frac{\text{mole}}{\text{L}} \times 0.027 \text{ L} = 6.8 \times 10^{-3} \text{ mole}$

2) Find # moles $\text{Ca}(\text{NO}_3)_2$: $M_{\text{Ca}(\text{NO}_3)_2} V_{\text{Ca}(\text{NO}_3)_2} = 0.42 \frac{\text{mole}}{\text{L}} \times 0.036 \text{ L} = 1.5 \times 10^{-2} \text{ mole}$

$$[\text{H}^+] = \frac{6.8 \times 10^{-3} \text{ mole H}^+}{0.063 \text{ L}} = 0.11 \text{ M}$$

$$[\text{Ca}^{2+}] = \frac{1.5 \times 10^{-2} \text{ mole Ca}(\text{NO}_3)_2}{0.063 \text{ L}} \left(\frac{1 \text{ mole Ca}^{2+}}{1 \text{ mole Ca}(\text{NO}_3)_2} \right) = 0.24 \text{ M}$$

moles NO_3^- from HNO_3 :

$$6.8 \times 10^{-3} \text{ mole HNO}_3 \left(\frac{1 \text{ mole NO}_3^-}{1 \text{ mole HNO}_3} \right) = 6.8 \times 10^{-3} \text{ mole NO}_3^-$$

moles NO_3^- from $\text{Ca}(\text{NO}_3)_2$:

$$1.5 \times 10^{-2} \text{ mole Ca}(\text{NO}_3)_2 \left(\frac{2 \text{ moles NO}_3^-}{1 \text{ mole Ca}(\text{NO}_3)_2} \right) = 3.0 \times 10^{-2} \text{ mole NO}_3^-$$

$$[\text{NO}_3^-] = \frac{(6.8 \times 10^{-3} + 3.0 \times 10^{-2}) \text{ mole NO}_3^-}{0.063 \text{ L}} = 0.58 \text{ M}$$

3) Find the molarity of each ion present after mixing 35 ml of 0.42 M K_2SO_4 with 27 ml of 0.17 M K_3PO_4 .

Answer: $[K^+] = 0.71M$, $[SO_4^{2-}] = 0.24 M$, $[PO_4^{3-}] = 0.074 M$.

Find moles K_2SO_4 :-

$$M_{K_2SO_4} V_{K_2SO_4} = 0.42 \frac{\text{mole}}{l} \times 0.035 l = .015 \text{ mole } K_2SO_4$$

Find moles of K_3PO_4 :-

$$M_{K_3PO_4} V_{K_3PO_4} = 0.17 \frac{\text{mole}}{l} \times 0.027 l = 4.6 \times 10^{-3} \text{ mole } K_3PO_4$$

Find conc. of SO_4^{2-} :-

$$.015 \text{ mole } K_2SO_4 \left(\frac{1 \text{ mole } SO_4^{2-}}{1 \text{ mole } K_2SO_4} \right) = .015 \text{ mole } SO_4^{2-}$$

$$[SO_4^{2-}] = \frac{.015 \text{ mole}}{.062 l} = \boxed{.24 M}$$

Find conc. of K^+ :-

$$.015 \text{ mole } K_2SO_4 \left(\frac{2 \text{ mole } K^+}{1 \text{ mole } K_2SO_4} \right) = .030 \text{ mole } K^+$$

$$4.6 \times 10^{-3} \text{ mole } K_3PO_4 \left(\frac{3 \text{ moles } K^+}{1 \text{ mole } K_3PO_4} \right) = .014 \text{ mole } K^+$$

$$[K^+] = \frac{(.030 + .014) \text{ mole } K^+}{.062 l} = \boxed{0.71 M}$$

Find conc. of PO_4^{3-} :-

$$[PO_4^{3-}] = \frac{.0046 \text{ mole } PO_4^{3-}}{.062 L} = \boxed{.074 M}$$

4) Calculate the concentration of each ion and the mass of any precipitate when a 0.300 mole of aluminum hydroxide is added to 50.0 ml of 2.5 M nitric acid solution (Assume that there is no volume change upon the addition of the aluminum hydroxide to the solution).

Hint: Write a balanced equation for the reaction taking place.

Answer: 20g Al(OH)_3 left over, $[\text{Al}^{3+}] = 0.83\text{M}$, $[\text{NO}_3^-] = 2.5\text{M}$



Find limiting reagent:

$$\frac{\# \text{ moles HNO}_3}{M_{\text{HNO}_3} V_{\text{HNO}_3}} = \frac{2.5 \text{ mole} \times 0.0500 \text{ L}}{\# \text{ moles HNO}_3} = 0.125 \text{ mole HNO}_3$$

Theoretical ratio of moles:

$$\frac{\# \text{ moles HNO}_3}{\# \text{ moles Al(OH)}_3} = \frac{3}{1}$$

Two sig fig

Available ratio of moles:

$$\frac{0.125 \text{ mole HNO}_3}{0.300 \text{ mole Al(OH)}_3} = \frac{.42}{1}$$

HNO_3 is the limiting reagent.

Find mass of Al(OH)_3 reacting

$$0.125 \text{ mole HNO}_3 \left(\frac{1 \text{ mole Al(OH)}_3}{3 \text{ moles HNO}_3} \right) = 0.0417 \text{ mole Al(OH)}_3 \text{ reacting}$$

Find mass of Al(OH)_3 left over:-

$$0.300 \text{ mole Al(OH)}_3 \text{ available} - 0.0417 \text{ mole Al(OH)}_3 \text{ reacting} = 0.258 \text{ mole Al(OH)}_3 \text{ left over}$$

$$0.258 \text{ mole Al(OH)}_3 \left(\frac{78.0 \text{ g Al(OH)}_3}{1 \text{ mole Al(OH)}_3} \right) = 20.2 \text{ g Al(OH)}_3 \text{ left over}$$

Find moles $\text{Al(NO}_3)_3$ produced:

$$0.125 \text{ mole HNO}_3 \left(\frac{1 \text{ mole Al(NO}_3)_3}{3 \text{ moles HNO}_3} \right) = 0.0417 \text{ mole Al(NO}_3)_3$$

$$0.0417 \text{ mole Al(NO}_3)_3 \left(\frac{1 \text{ mole Al}^{3+}}{1 \text{ mole Al(NO}_3)_3} \right) = 0.0417 \text{ mole Al}^{3+}$$

$$[\text{Al}^{3+}] = \frac{0.0417 \text{ mole Al}^{3+}}{0.0500 \text{ L}} = 0.83 \text{ M}$$

$$0.0417 \text{ mole Al(NO}_3)_3 \left(\frac{3 \text{ moles NO}_3^-}{1 \text{ mole Al(NO}_3)_3} \right) = 0.125 \text{ mole NO}_3^-$$

$$[\text{NO}_3^-] = \frac{0.125 \text{ mole}}{0.0500 \text{ L}} = 2.5$$

5) A solution consists of 3.88 g benzene, C_6H_6 , and 2.45 g toluene, $C_6H_5CH_3$. The vapor pressure of pure benzene at $20.^\circ C$ is 75 mm Hg and that of toluene at $20.^\circ C$ is 22 mm Hg. Assume that Raoult's law holds for each component of the solution, calculate the mole fraction of benzene in the vapor. (molar mass of benzene = 78.0 g/mole and toluene = 92.0 g/mole.)

Answer = 0.87

$$\begin{aligned}
 \textcircled{1} \quad P_{\text{total}} &= P_{\text{benzene}} + P_{\text{toluene}} \\
 &= X_{\text{benzene in solution}} P_{\text{benzene}}^{\circ} + X_{\text{toluene in solution}} P_{\text{toluene}}^{\circ} \\
 &= \frac{3.88 \text{ g}}{78.0 \text{ g/mole}} (75 \text{ torr}) + \frac{2.45 \text{ g}}{92.0 \text{ g/mole}} (22 \text{ torr}) \\
 &= \frac{3.88 \text{ g} + 2.45 \text{ g}}{78.0 \text{ g/mole} + 92.0 \text{ g/mole}} (75 \text{ torr}) + \frac{2.45 \text{ g} + 3.88 \text{ g}}{92.0 \text{ g/mole} + 78.0 \text{ g/mole}} (22 \text{ torr}) \\
 &= \frac{.0497 \text{ mole}}{(.0497 + .0266) \text{ mole}} (75 \text{ torr}) + \frac{.0266 \text{ mole}}{.0266 \text{ mole} + .0497 \text{ mole}} (22 \text{ torr}) \\
 &= .651 (75 \text{ torr}) + .349 (22 \text{ torr}) \\
 &= 49 \text{ torr} + 7.8 \text{ torr} \\
 &= 56.8 \text{ torr} = 57 \text{ torr}
 \end{aligned}$$

$$\begin{aligned}
 \textcircled{2} \quad X_{\text{benzene in vapor}} &= \frac{P_{\text{benzene}}}{P_{\text{total}}} \\
 &= \frac{49 \text{ torr}}{57 \text{ torr}} \\
 &\approx 0.86
 \end{aligned}$$

6) The freezing point of a glucose solution ($C_6H_{12}O_6$; molar mass = 180.0 g/mole) is $-10.3^\circ C$. The density of the solution is 1.50 g/ml. What is the molarity of the glucose solution? (K_f for water is $1.86^\circ C \cdot \text{kg/mole}$)

Answer: 4.16 mole/L

$$\textcircled{1} \quad \Delta T_{FP} = K_{FP} \text{ molality}$$

$$10.3^\circ C = 1.86 \frac{^\circ C \cdot \text{kg}}{\text{mole}} (\text{molality})$$

$$\text{molality} = \frac{10.3^\circ C}{1.86^\circ C \cdot \text{kg/mole}} = 5.54 \frac{\text{mole}}{\text{kg}}$$

$$\textcircled{2} \quad \text{mass of solute} = 5.54 \text{ mole glucose} \left(\frac{180.0 \text{ g glucose}}{1 \text{ mole glucose}} \right) = 997 \text{ g glucose}$$

in 1000 g H_2O

$$\textcircled{3} \quad \text{mass of solution} = 1000 \text{ g } H_2O + 997 \text{ g glucose} = 1997 \text{ g solution}$$

containing 5.54 mole glucose

$$\textcircled{4} \quad \text{volume of solution: } 1997 \text{ g solution} \left(\frac{1 \text{ ml solution}}{1.50 \text{ g solution}} \right) = 1331 \text{ ml solution} = 1.331 \text{ L solution}$$

containing 5.54 mole glucose

$$\textcircled{5} \quad \text{molarity} = \frac{5.54 \text{ mole glucose}}{1.331 \text{ L solution}} = 4.16 \text{ mole/L}$$

7) What is the normal boiling point of a 2.70 M solution of KBr that has a density of 1.80 g/ml? (K_B for H_2O is $0.512^\circ C \cdot kg/mole$)
 Answer: $101.9^\circ C$

$$2.70 \text{ mole KBr} \left(\frac{119 \text{ g KBr}}{1 \text{ mole KBr}} \right) = 321 \text{ g KBr}$$

in 1000ml solution

$$1000 \text{ ml solution} \left(\frac{1.80 \text{ g solution}}{1 \text{ ml solution}} \right) = 1800 \text{ g solution}$$

$$= 1.80 \times 10^3 \text{ g solution}$$

mass of H_2O :

$$1800 \text{ g solution} - 321 \text{ g KBr} = 1479 \text{ g } H_2O = 1.48 \times 10^3 \text{ g}$$

$$= 1.47 \text{ kg } H_2O$$

containing 2.70 mole K

$$\text{molality} = \frac{2.70 \text{ mole KBr}}{1.479 \text{ kg } H_2O} = 1.83 \text{ mole/kg}$$

$$\Delta T_{BP} = i \cdot K_{B.P} \cdot \text{molality}$$

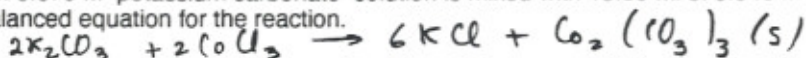
$$= 2 \frac{\text{particles}}{\text{formula units}} \cdot 0.512^\circ C \frac{\text{kg}}{\text{mole}} \times 1.83 \frac{\text{mole}}{\text{kg}}$$

$$= 1.87^\circ C$$

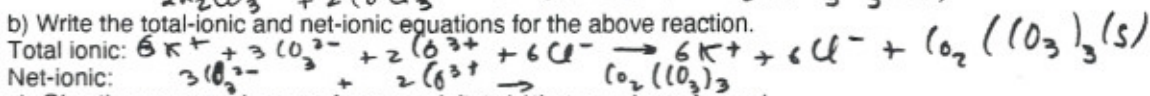
$$\text{normal B.P} = 100 + 1.87 = 101.9^\circ C$$

g) 28.00 ml of 0.670 M potassium carbonate solution is mixed with 15.00 ml of 0.940 M cobalt(III) chloride

a) Write a balanced equation for the reaction.



b) Write the total-ionic and net-ionic equations for the above reaction.



Net-ionic: $3CO_3^{2-} + 2Co^{3+} \rightarrow Co_2(CO_3)_3$

c) Give the name and mass of any precipitate(s) that may have formed.

Answer: 1.87 g of $Co_2(CO_3)_3$ precipitate.

Calculate the molar concentration of each ion remaining in solution after the reaction is complete.

Answer: concentration of potassium ions = 0.874 M,
concentration of cobalt (III) ions = 0.0372 M
concentration of carbonate ions = 0 M
concentration of chloride ions = 0.986 M

c) Find limiting reagent: -

$$\text{no. of moles } K_2CO_3 = M_{K_2CO_3} \times V_{K_2CO_3} = 0.0280L \times 0.670 \frac{\text{mole}}{L} = 0.0188 \text{ mole}$$

$$\text{no. of moles } CoCl_3 = M_{CoCl_3} \times V_{CoCl_3} = 0.01500L \times 0.940 \frac{\text{mole}}{L} = 0.0141 \text{ mole available}$$

theoretical ratio of moles:

$$\frac{\text{moles } K_2CO_3}{\text{moles } CoCl_3} = \frac{3}{1}$$

Available ratio of moles:

$$\frac{\text{moles } K_2CO_3}{\text{moles } CoCl_3} = \frac{0.0188 \text{ mole}}{0.0141 \text{ mole}} = 1.33$$

Therefore, K_2CO_3 is the limiting reagent

$$0.0188 \text{ mole } K_2CO_3 \left(\frac{1 \text{ mole } Co_2(CO_3)_3}{3 \text{ moles } K_2CO_3} \right) \left(\frac{297.89 \text{ g } Co_2(CO_3)_3}{1 \text{ mole } Co_2(CO_3)_3} \right) = 1.87 \text{ g of } Co_2(CO_3)_3 \text{ precipitate}$$

$$\text{d) } \underline{\text{moles of } CoCl_3 \text{ reacting:}} - 0.0188 \text{ mole } K_2CO_3 \left(\frac{2 \text{ moles } CoCl_3}{3 \text{ moles } K_2CO_3} \right) = 0.0125 \text{ moles } CoCl_3 \text{ reacting}$$

moles $CoCl_3$ left over:

$$0.0141 \text{ mole } CoCl_3 \text{ available} - 0.0125 \text{ mole } CoCl_3 \text{ reacting} = 0.0016 \text{ mole } CoCl_3 \text{ left over}$$

$$0.0188 \text{ mole } K_2CO_3 \left(\frac{6 \text{ moles } KCl}{3 \text{ moles } K_2CO_3} \right) = 0.0376 \text{ mole } KCl \text{ forming}$$

$$[Co^{3+}] = \frac{0.0016 \text{ mole}}{0.0430L} = 0.0372M$$

$$[K^+] = \frac{0.0376 \text{ mole}}{0.0430L} = 0.874M$$

$$[Cl^-] = \frac{(0.0016 \times 3) + 0.0376}{0.0430L} \text{ mole} = \frac{(0.0048 + 0.0376) \text{ mole}}{0.0430L}$$

$$= 0.986M$$

$$[CO_3^{2-}] = 0M$$

SET B:

1) A solution that contains 12.6 g of a nonvolatile nondissociating solute in 400. g of benzene freezes at 3.6 °C. The normal freezing point of benzene is 5.5 °C. What is the molar mass of the solute? (K_f for benzene = 4.96 °C · kg/mole)

Answer: 82 g/mole

$$\textcircled{1} \quad \Delta T_f = K_f \text{ molality}$$

$$\xrightarrow{\text{Two sig fig.}} 1.9^\circ\text{C} = \left(4.96 \frac{^\circ\text{C} \cdot \text{kg}}{\text{mole}}\right) \text{ molality}$$

$$\textcircled{2} \quad \# \text{ moles of solute} = 0.383 \frac{\text{mole}}{\text{kg}} \times 0.400 \text{ kg} = 0.153 \text{ mole}$$

$$\textcircled{3} \quad \text{molar mass} = \frac{12.6 \text{ g}}{0.153 \text{ mole}} = 82 \frac{\text{g}}{\text{mole}}$$

↑
Two sig fig

2) Chloroform and methanol form an ideal solution. The solution boils at 22 °C and 0.255 atm. At 22 °C, the vapor pressure of pure methanol is 0.192 atm and the vapor pressure of pure chloroform is 0.311 atm.

Answer: 0.529

What is the mole fraction of chloroform in solution?

$$P_{\text{total}} = X_{\text{methanol}} P_{\text{methanol}}^0 + X_{\text{chloroform}} P_{\text{chloroform}}^0$$

$$0.255 \text{ atm} = X_{\text{methanol}} (0.192 \text{ atm}) + (1 - X_{\text{methanol}}) 0.311 \text{ atm}$$

$$0.255 = 0.192 X_{\text{methanol}} + 0.311 - 0.311 X_{\text{methanol}}$$

$$0.311 X_{\text{methanol}} - 0.192 X_{\text{methanol}} = 0.311 - 0.255$$

$$0.119 X_{\text{methanol}} = 0.056$$

$$X_{\text{methanol}} = \frac{0.056}{0.119} = 0.470$$

$$X_{\text{chloroform}} = 1 - 0.470 = 0.529$$

3) What is the normal boiling point of 1.21 M solution of CaI_2 that has a density of 1.92 g/ml? (K_B for H_2O = $0.512^\circ\text{C} \cdot \text{kg/mole}$)

Answer: 101.2°C

① Find molality:

$$1.21 \text{ mole } \text{CaI}_2 \xrightarrow{\text{in}} 1 \text{ liter}$$

a) mass of solution: $1000 \text{ ml solution} \left(\frac{1.92 \text{ g solution}}{1 \text{ ml solution}} \right) = 1920 \text{ g solution}$

b) mass of solute:

$$1.21 \text{ mole } \text{CaI}_2 \left(\frac{294 \text{ g } \text{CaI}_2}{1 \text{ mole } \text{CaI}_2} \right) = 356 \text{ g } \text{CaI}_2$$

c) mass of H_2O :

$$1920 \text{ g solution} - 356 \text{ g } \text{CaI}_2 = 1564 \text{ g } \text{H}_2\text{O}$$

d) molality:

$$\frac{1.21 \text{ mole } \text{CaI}_2}{1.564 \text{ kg } \text{H}_2\text{O}} = 0.774 \frac{\text{mole}}{\text{kg}}$$

$$\Delta T_{\text{B.P}} = i K_B \text{ molality}$$

$$= 3 \left(0.512 \frac{^\circ\text{C} \cdot \text{kg}}{\text{mole}} \right) \cdot 0.774 \frac{\text{mole}}{\text{kg}} = 1.2^\circ\text{C}$$

$$\text{B.P} = 100 + 1.2 = 101.2^\circ\text{C}$$

4) Calculate the freezing point of a 36.0% by mass Na_3PO_4 solution. (K_f for H_2O = $1.86^\circ\text{C} \cdot \text{kg/mole}$)

Answer: -25.5°C

$$36.0 \text{ g } \text{Na}_3\text{PO}_4 \xrightarrow{\text{in}} 64.0 \text{ g } \text{H}_2\text{O}$$

a) $36.0 \text{ g } \text{Na}_3\text{PO}_4 \left(\frac{1 \text{ mole } \text{Na}_3\text{PO}_4}{164 \text{ g } \text{Na}_3\text{PO}_4} \right) = 0.220 \text{ mole } \text{Na}_3\text{PO}_4$

b) molality = $\frac{0.220 \text{ mole } \text{Na}_3\text{PO}_4}{0.0640 \text{ kg solvent}} = 3.43 \frac{\text{mole}}{\text{kg}}$

$$\Delta T_{\text{f.p}} = i K_f \text{ molality}$$

$$= 4 \left(1.86 \frac{^\circ\text{C} \cdot \text{kg}}{\text{mole}} \right) 3.43 \frac{\text{mole}}{\text{kg}}$$

$$= 25.5^\circ\text{C}$$

$$T_{\text{f.p}} = 0 - 25.5^\circ\text{C} = -25.5^\circ\text{C}$$

5) 32.00 ml of 0.311 M aluminum nitrate is mixed with 64.00 ml of 0.177 M sodium carbonate and allowed to react.

a) Write a balanced equation for the reaction. $2 \text{Al}(\text{NO}_3)_3(\text{aq}) + 3 \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{Al}_2(\text{CO}_3)_3(\text{s}) + 6 \text{NaNO}_3(\text{aq})$

b) Write total-ionic and net-ionic equations for the above reaction.

Total-ionic:

Net-ionic:

c) Give the name and mass of any precipitate that may have formed.

Answer: 0.884 g of $\text{Al}_2(\text{CO}_3)_3$ precipitate

d) Calculate the molar concentration of each ion remaining in solution after reaction is complete.

Answer: Concentration of carbonate ions = 0 M

Concentration of aluminum ions = 0.0252 M

Concentration of nitrate ions = 0.312 M

Concentration of sodium ions = 0.236 M

i) moles of $\text{Al}(\text{NO}_3)_3$ available: $M_{\text{Al}(\text{NO}_3)_3} V_{\text{Al}(\text{NO}_3)_3} = 0.311 \frac{\text{mole}}{\text{L}} \times 0.03200 \text{ L}$
 $= 9.95 \times 10^{-3} \text{ mole Al}(\text{NO}_3)_3$

ii) moles of Na_2CO_3 available: $M_{\text{Na}_2\text{CO}_3} V_{\text{Na}_2\text{CO}_3} = 0.177 \frac{\text{mole}}{\text{L}} \times 0.06400 \text{ L}$
 $= 0.01133 \text{ mole Na}_2\text{CO}_3$

iii) find limiting reagent: $\frac{\text{mole Na}_2\text{CO}_3}{\text{mole Al}(\text{NO}_3)_3} = \frac{3}{2} = \frac{1.5}{1}$
theoretical ratio of moles

A available ratio of moles: $\frac{\text{mole Na}_2\text{CO}_3}{\text{mole Al}(\text{NO}_3)_3} = \frac{0.01133 \text{ mole}}{0.00995 \text{ mole}} = \frac{1.14}{1}$

Na_2CO_3 is the limiting reagent.

$0.01133 \text{ mole Na}_2\text{CO}_3 \left(\frac{1 \text{ mole Al}_2(\text{CO}_3)_3}{3 \text{ moles Na}_2\text{CO}_3} \right) \left(\frac{234 \text{ g Al}_2(\text{CO}_3)_3}{1 \text{ mole Al}_2(\text{CO}_3)_3} \right) = 0.884 \text{ g Al}_2(\text{CO}_3)_3$
 precipitate

Find moles of excess reagent, $\text{Al}(\text{NO}_3)_3$ left over:
 $0.01133 \text{ mole Na}_2\text{CO}_3 \left(\frac{2 \text{ moles Al}(\text{NO}_3)_3}{3 \text{ moles Na}_2\text{CO}_3} \right) = 0.00753 \text{ mole Al}(\text{NO}_3)_3$

moles of $\text{Al}(\text{NO}_3)_3$ left over = $9.95 \times 10^{-3} \text{ mole} - 7.53 \times 10^{-3} \text{ mole} = 2.42 \times 10^{-3} \text{ mole Al}(\text{NO}_3)_3$
 left over

Find moles NaNO_3 formed:
 $0.01133 \text{ mole Na}_2\text{CO}_3 \left(\frac{6 \text{ moles NaNO}_3}{3 \text{ moles Na}_2\text{CO}_3} \right) = 0.0227 \text{ mole NaNO}_3$ formed

Find concentration of ions in solution:

a) moles of Al^{3+} from $\text{Al}(\text{NO}_3)_3$ left over:
 $0.00242 \text{ mole Al}(\text{NO}_3)_3 \left(\frac{1 \text{ mole Al}^{3+}}{1 \text{ mole Al}(\text{NO}_3)_3} \right) = 0.00242 \text{ mole Al}^{3+}$

$[\text{Al}^{3+}] = \frac{0.00242 \text{ mole}}{0.0960 \text{ L}} = 0.0252 \text{ M}$; $[\text{Na}^+] = \frac{0.02260 \text{ mole}}{0.0960 \text{ L}} = 0.236 \text{ M}$

$[\text{NO}_3^-] = \frac{(0.00242 \times 3 + 0.0227) \text{ mole}}{0.09600 \text{ L}} = 0.312 \text{ M}$

$[\text{CO}_3^{2-}] = 0 \text{ M}$

SET C:

1) What is the molarity of an aqueous solution of $C_6H_{12}O_6$ that has a normal boiling point of $101.40^\circ C$ and density of 1.68 g/ml ? K_B for water is $0.512^\circ C \cdot \text{kg/mole}$. ($C_6H_{12}O_6$ is a nonvolatile nondissociating solute.)

Answer: 3.07 mole/L

① $\Delta T_B = K_B \text{ molality}$

$1.40^\circ C = 0.512^\circ C \cdot \text{kg} \cdot \text{molality}$

$$\text{molality} = \frac{1.40^\circ C}{0.512 \frac{^\circ C \cdot \text{kg}}{\text{mole}}} = 2.73 \frac{\text{mole solute}}{\text{kg } H_2O}$$

② Change 2.73 mole $C_6H_{12}O_6$ into mass:

$$2.73 \text{ mole } C_6H_{12}O_6 \left(\frac{180.09 \text{ g } C_6H_{12}O_6}{1 \text{ mole } C_6H_{12}O_6} \right) = 491 \text{ g } C_6H_{12}O_6$$

③ Find mass of solution:

$$491 \text{ g } C_6H_{12}O_6 + 1000 \text{ g } H_2O = 1491 \text{ g solution}$$

④ Find volume of solution:

$$1491 \text{ g solution} \left(\frac{1 \text{ ml solution}}{1.68 \text{ g solution}} \right) = 888 \text{ ml solution}$$

⑤ Molarity = $\frac{\text{no. of moles of solute}}{\text{Liters of solution}} = \frac{2.73 \text{ mole } C_6H_{12}O_6}{0.888 \text{ L}} = 3.07 \frac{\text{mole}}{\text{L}}$

2) Calculate the normal freezing point of a 0.6837 M aqueous solution of $C_{12}H_{22}O_{11}$ that has a density of 1.35 g/ml . ($C_{12}H_{22}O_{11}$ is a nonvolatile nondissociating solute.) The molal freezing point depression constant of water is $1.86^\circ C \cdot \text{kg/mole}$.

Answer: Freezing point = $-1.14^\circ C$

1) Change $0.6837 \text{ mole } C_{12}H_{22}O_{11}$ into mass $C_{12}H_{22}O_{11}$:

$$0.6837 \text{ mole } C_{12}H_{22}O_{11} \left(\frac{342.2 \text{ g } C_{12}H_{22}O_{11}}{1 \text{ mole } C_{12}H_{22}O_{11}} \right) = 234.0 \text{ g } C_{12}H_{22}O_{11}$$

2) mass of solution:

$$1000 \text{ ml solution} \left(\frac{1.35 \text{ g solution}}{1 \text{ ml solution}} \right) = 1350 \text{ g solution}$$

3) mass of H_2O : $1350 \text{ g solution} - 234.0 \text{ g solute} = 1116 \text{ g } H_2O$

4) molality = $\frac{\text{no. moles of solute}}{\text{kg of } H_2O} = \frac{0.6837 \text{ mole solute}}{1.116 \text{ kg } H_2O} = 0.6126 \frac{\text{mole}}{\text{kg}}$

5) $\Delta T_F = K_f \text{ molality}$

$$= 1.86 \frac{^\circ C \cdot \text{kg}}{\text{mole}} \left(0.6126 \frac{\text{mole}}{\text{kg}} \right)$$

$$= 1.14^\circ C$$

6) Freezing point = $-1.14^\circ C$

3) Heptane, C_7H_{16} , and octane, C_8H_{18} , form ideal solutions. At $40.^\circ C$, the vapor pressure of pure heptane is 0.522 atm , and the vapor pressure of pure octane is 0.238 atm . A solution is made of 5.32 g heptane and 8.80 g octane. Calculate the mole fraction of octane in the vapor at the above temperature.

Answer: 0.398

$$\begin{aligned} \text{Total vapor pressure} &= X_{\text{heptane}} P_{\text{heptane}}^{\circ} + X_{\text{octane}} P_{\text{octane}}^{\circ} \\ &= \frac{5.32 \text{ g}}{100.1 \text{ g/mole}} (.522 \text{ atm}) + \frac{8.80 \text{ g}}{114.2 \text{ g/mole}} (.238 \text{ atm}) \\ &= \frac{5.32 \text{ g}}{100.1 \text{ g/mole}} + \frac{8.80 \text{ g}}{114.2 \text{ g/mole}} \\ &= \frac{0.0531 \text{ mole}}{(0.0531 + 0.0771) \text{ mole}} (.522 \text{ atm}) + \frac{0.0771 \text{ mole}}{(0.0531 + 0.0771) \text{ mole}} (.238 \text{ atm}) \\ &= \frac{0.0531}{0.130} (.522 \text{ atm}) + \frac{0.0771}{0.130} (.238 \text{ atm}) \\ &= (.408) (.522 \text{ atm}) + .593 (.238) \\ &= .213 \text{ atm} + 0.141 \text{ atm} = 0.354 \text{ atm} \\ \text{P}_{\text{octane in vapor}} &= X_{\text{octane in vapor}} P_{\text{total}} ; 0.141 \text{ atm} = X_{\text{octane in vapor}} (.354 \text{ atm}) \\ X_{\text{octane in vapor}} &= \frac{0.141}{.354} = 0.398 \end{aligned}$$

4) What is the molar mass and molecular formula of a nondissociating compound whose empirical formula is C_4H_2N , if 3.84 g of the compound in $500. \text{ g}$ benzene give a freezing point depression of $0.307^\circ C$? (The molal freezing point depression constant for benzene is $5.12^\circ C \cdot \text{kg/mole}$.)

Answer: 128 g/mole ; $C_8H_4N_2$

$$\textcircled{1} \Delta T = K_F \text{ molality}$$

$$0.307^\circ C = 5.12 \frac{^\circ C \cdot \text{kg}}{\text{mole}} \text{ molality}$$

$$\text{molality} = 0.0600 \frac{\text{mole}}{\text{kg}}$$

$$\textcircled{2} \text{ molality} = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

$$\text{* moles of solute: } 0.0600 \frac{\text{mole}}{\text{kg}} \times .500 \text{ kg solvent} = .0300 \text{ mole solute}$$

$$\textcircled{3} \text{ molar mass: } \frac{3.84 \text{ g solute}}{.0300 \text{ mole solute}} = 128 \text{ g/mole}$$

$$\textcircled{4} \text{ molar mass of empirical formula, } C_4H_2N = 64.0 \text{ g/mole}$$

$$\frac{\text{mass of molecular formula}}{\text{mass of empirical formula}} = \frac{128 \text{ g/mole}}{64. \text{ g/mole}} = \frac{2}{1}$$

5) Liquids A and B form an ideal solution. The vapor pressure of pure A is 0.700 atm at the normal boiling point of a solution prepared from 0.250 mole of B and 0.650 mole of A. What is the vapor pressure of pure B at this temperature?

Answer: 1.77 atm

$$P_{\text{total}} = P_A + P_B$$

$$= X_A P_A^{\circ} + X_B P_B^{\circ}$$

$$1 \text{ atm} = \frac{.650}{.250 + .650} (.700 \text{ atm}) + \frac{.250}{(.250 + .650)} P_B^{\circ}$$

$$1 \text{ atm} = .722 (.700 \text{ atm}) + .278 P_B^{\circ}$$

$$1 \text{ atm} = .506 + .278 P_B^{\circ}$$

$$1.00 - .506 = .278 P_B^{\circ}$$

$$.494 = .278 P_B^{\circ}$$

$$P_B^{\circ} = \frac{.494}{.278} = 1.77 \text{ atm}$$

6) A 0.900 L aqueous solution contains 30.0 g of a protein. The osmotic pressure of the solution is 12.7 torr at 25 °C. What is the molar mass of the protein?

Answer: $4.88 \times 10^4 \text{ g/mole}$

$$\textcircled{1} \quad \Pi = M R T$$

$$12.7 \text{ torr} \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = M \left(.0821 \frac{\text{L atm}}{\text{mole K}} \right) 298 \text{ K}$$

$$M = 6.83 \times 10^{-4} \frac{\text{mole}}{\text{L}}$$

$$\textcircled{2} \quad n_{\text{of moles of protein}} = M \times V$$

$$= 6.83 \times 10^{-4} \frac{\text{mole}}{\text{L}} \times .900 \text{ L}$$

$$= 6.15 \times 10^{-4} \text{ mole}$$

$$\textcircled{3} \quad \text{molar mass} = \frac{30.0 \text{ g protein}}{6.15 \times 10^{-4} \text{ mole protein}} = 4.88 \times 10^4 \text{ g/mole}$$

7) Acetone and methanol form ideal solution. At 25 °C, the vapor pressures of pure acetone and pure methanol are 0.342 atm and 0.188 atm respectively. Calculate the mole fraction of methanol in a solution that boils at 25 °C and 0.248 atm.

Answer: $X = 0.610$

$$X_{\text{methanol}} + X_{\text{acetone}} = 1$$

$$P_{\text{total}} = X_{\text{acetone}} P^{\circ}_{\text{acetone}} + X_{\text{methanol}} P^{\circ}_{\text{methanol}}$$

$$= (1 - X_{\text{methanol}}) (0.342 \text{ atm}) + X_{\text{methanol}} (0.188)$$

$$0.248 \text{ atm} = (1 - X_{\text{methanol}}) 0.342 + X_{\text{methanol}} (0.188)$$

$$X_{\text{methanol}} = 0.610$$