

UNIT 6 – CHAPTER 11 STUDENT NOTES: SOLUTIONS AND IMFs

****Remember: solute – gets dissolved; solvent – does the dissolving; solution – solute + solvent****

Expressions of concentration

1.00 grams of ethanol (C_2H_5OH) is dissolved in 100 grams of water (assume the density of the solution is 1 g/mL). Express the concentration of this solution in terms of:

Molarity (M) = $\frac{\text{moles solute}}{\text{liters of solution}}$

$$\frac{1.00g C_2H_5OH}{146.07g C_2H_5OH} \times \frac{1 mol C_2H_5OH}{46.07g C_2H_5OH} = \frac{0.0217 mol}{0.101 L} = 0.215 M$$

100 mL + 1 g (mL)

Mass % = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$

$$\frac{1.00g C_2H_5OH}{101g Sol'n} = 0.990\% C_2H_5OH$$

(100 mL (g) + 1 g)

Mole fraction = $\frac{\text{mole A}}{\text{mole A} + \text{mole B}}$

$$\frac{1.00g C_2H_5OH}{146.07g C_2H_5OH} \times \frac{1 mol C_2H_5OH}{46.07g C_2H_5OH} = \frac{0.0217 mol C_2H_5OH}{(0.0217 + 5.56)} = 0.00389 C_2H_5OH$$

$$\frac{100g H_2O}{18g H_2O} = 5.56 mol H_2O$$

Molality (m) = $\frac{\text{moles solute}}{\text{kg of solvent}}$

$$\frac{0.0217 mol C_2H_5OH}{0.100 kg H_2O} = 0.217 m C_2H_5OH$$

EX 1: A 1.5-liter solution of a 4.25 M $LiCl_2$ solution is produced. If the density of the solution is 1.126 g/mL, calculate the: a) mass %, b) molality, and c) mole fraction of the solution.

EX 2: The electrolyte in automobile lead storage batteries is a 3.75 M sulfuric acid solution that has a density of 1.230 g/mL. Calculate the a) mass %, b) molality, c) mole fraction of the sulfuric acid.

EX 3: Common commercial acids are sold with certain properties. For sulfuric acid they are:

$$\text{H}_2\text{SO}_4 \text{ density} = 1.84 \text{ g/cm}^3 \quad \text{mass \%} = 95 \rightarrow \frac{95 \text{ g H}_2\text{SO}_4}{(95 \text{ g H}_2\text{SO}_4 + 5 \text{ g H}_2\text{O})}$$

$1 \text{ cm}^3 = 1 \text{ mL} = .001 \text{ L}$

Calculate the a) molarity, b) molality, c) mole fraction of the solution.

$$\text{a) } 95\% = \frac{95 \text{ g H}_2\text{SO}_4}{100 \text{ g SOLN}} \left| \frac{1 \text{ mol H}_2\text{SO}_4}{98.09 \text{ g H}_2\text{SO}_4} \right| = \frac{.9694 \text{ mol H}_2\text{SO}_4}{100 \text{ g SOLN}} \left| \frac{1.84 \text{ g}}{1.001 \text{ L}} \right| = \boxed{17.8 \text{ M}}$$

$$\text{b) } \frac{.9694 \text{ mol H}_2\text{SO}_4}{5 \text{ g H}_2\text{O} = .005 \text{ kg H}_2\text{O}} = \boxed{194 \text{ m H}_2\text{SO}_4}$$

$$\text{c) } \frac{5 \text{ g H}_2\text{O}}{18 \text{ g H}_2\text{O}} = .3 \text{ mol H}_2\text{O} \quad \frac{.9694 \text{ mol H}_2\text{SO}_4}{(.9694 \text{ mol H}_2\text{SO}_4) + (.3 \text{ mol H}_2\text{O})} = \boxed{0.76 \text{ H}_2\text{SO}_4}$$

Dilutions – are problems that require you to produce a certain concentration. They are of two types:

Solid dilutions -

* How many grams of salt to make 750 mL of 0.2M solution NaCl? *

$$0.2 \text{ M} = \frac{.2 \text{ mol NaCl}}{1 \text{ L SOLN}} \left| \frac{.750 \text{ L}}{1 \text{ mol NaCl}} \right| = 0.05 \text{ mol NaCl} \left| \frac{58.5 \text{ g NaCl}}{1 \text{ mol NaCl}} \right| = \boxed{8.8 \text{ g NaCl}}$$

Liquid dilutions - $M_1 \cdot V_1 = M_2 \cdot V_2$

* How much stock solution is needed to make 1.50 L of 0.03 M solution from above? *

$$\underbrace{(0.03 \text{ M})}_{\text{DILUTE}} (1.50 \text{ L}) = \underbrace{(0.20 \text{ M})}_{\text{STOCK}} (X \text{ L})$$

$$\boxed{X = 0.225 \text{ L STOCK SOLUTION NEEDED}}$$

Raoult's Law: describes that the presence of a nonvolatile solute lowers the vapor pressure of a solute.

$$P_{\text{solution}} = X_{\text{solvent}} \cdot P^{\circ}_{\text{solvent}}$$

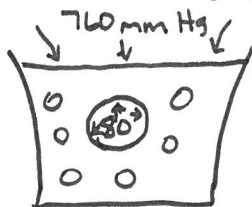
X_{solvent} = mole fraction of solvent

$P^{\circ}_{\text{solvent}}$ = vapor pressure solvent

Equilibrium vapor pressure – pressure of a liquid in equilibrium with its vapor

Liquid + energy $\longleftrightarrow$ vapor

Boiling liquids are dependent on their equilibrium vapor pressure



80.5°C → 355.1 mmHg
90°C → 525.8 mmHg
100°C → 760 mmHg
105°C → 780 mmHg
110°C → 800 mmHg

Pascal's law – pressure on the surface of a liquid acts the same throughout the liquid

EX 4: Calculate the vapor pressure at 25°C for a solution prepared by dissolving 158.0 grams of table sugar ($C_{12}H_{22}O_{11}$ = 342 g/mole) in 643.5 cm³ of water. At 25°C the density of water is 0.9971 g/cm³ and the vapor pressure is 23.76 torr.

$$P_{\text{solution}} = X_{H_2O} \cdot P^{\circ}_{H_2O}$$

$$158g C_{12}H_{22}O_{11} \left| \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342g C_{12}H_{22}O_{11}} \right| = 0.462 \text{ mol } C_{12}H_{22}O_{11}$$

$$643.5 \text{ cm}^3 H_2O \left| \frac{0.9971 g}{1 \text{ cm}^3} \right| = 641.6 g H_2O \left| \frac{1 \text{ mol } H_2O}{18 g H_2O} \right| = 35.6 \text{ mol } H_2O$$

$$\text{MOLE FRACTION } (X_{H_2O}) = \frac{35.6 \text{ mol } H_2O}{(35.6 \text{ mol } H_2O + 0.462 \text{ mol } C_{12}H_{22}O_{11})} = 0.987 H_2O$$

$$P_{\text{solution}} = (0.987) \cdot (23.76 \text{ Torr}) = \boxed{23.45 \text{ Torr}}$$

EX 5: A mixture of 10 mL of octane (C_8H_{18} , density = 0.75 g/mL) at a pressure of 150 torr with 15 mL butane (C_4H_{10} , density = 0.98 g/mL) at a pressure of 100 torr is mixed in a closed container. Calculate the a) vapor pressure of the solution, and b) mole fraction of the two gases in the vapor.

$$a) P_{\text{solution}} = P_{\text{octane}} + P_{\text{butane}}$$

$$\text{DETERMINE } X: \text{OCTANE} \rightarrow 10 \text{ mL} \left| \frac{0.75 g}{1 \text{ mL}} \right| = 7.5 g \text{ OCTANE} \left| \frac{1 \text{ mol } C_8H_{18}}{114 g C_8H_{18}} \right| = 0.066 \text{ mol } C_8H_{18}$$

$$\text{BUTANE} \rightarrow 15 \text{ mL} \left| \frac{0.98 g}{1 \text{ mL}} \right| = 14.7 g \text{ BUTANE} \left| \frac{1 \text{ mol } C_4H_{10}}{58.0 g C_4H_{10}} \right| = 0.25 \text{ mol } C_4H_{10}$$

$$X_{\text{octane}} = \frac{0.066 \text{ mol } C_8H_{18}}{(0.066 + 0.25)} = 0.209 \text{ OCTANE}$$

$$X_{\text{butane}} = \frac{0.25 \text{ mol } C_4H_{10}}{(0.25 + 0.066)} = 0.791 \text{ BUTANE}$$

$$P_{\text{solution}} = (0.209 \times 150 \text{ Torr}) + (0.791 \times 100) = \boxed{110.45 \text{ Torr}}$$

$$b) P_{\text{gas}} = X_{\text{solvent}} \cdot P^{\circ}_{\text{total}}$$

$$\text{OCTANE} \rightarrow 31.35 \text{ Torr} = X_{\text{octane}} \cdot 110.45 \text{ Torr} = \boxed{0.284}$$

$$\text{BUTANE} \rightarrow 79.10 \text{ Torr} = X_{\text{butane}} \cdot 110.45 \text{ Torr} = \boxed{0.716}$$



EX 6: Predict the vapor pressure of a solution prepared by mixing 35.0 grams of solid Na_2SO_4 (molar mass = 142 g/mole) with 175 grams of water at 25°C . (Vapor pressure of water @ 25°C = 23.76 torr)

$$P_{\text{solution}} = X_{\text{H}_2\text{O}} \cdot P_{\text{H}_2\text{O}}^0$$

DETERMINE X: $\text{H}_2\text{O} \rightarrow 175\text{g H}_2\text{O} \frac{1 \text{ mol H}_2\text{O}}{18\text{g H}_2\text{O}} = 9.72 \text{ mol H}_2\text{O}$

$$\text{Na}_2\text{SO}_4 \rightarrow 35.0\text{g} \frac{1 \text{ mol Na}_2\text{SO}_4}{142\text{g Na}_2\text{SO}_4} = 0.246 \text{ mol} \frac{3 \text{ PARTICLES}}{1 \text{ mol Na}_2\text{SO}_4} = 0.738$$

$$X_{\text{H}_2\text{O}} = \frac{9.72 \text{ mol H}_2\text{O}}{(9.72 \text{ mol H}_2\text{O} + 0.738 \text{ mol PARTICLES})} = 0.929$$

$$P_{\text{solution}} = (0.929) \cdot (23.76 \text{ TORR}) = \boxed{22.1 \text{ TORR}}$$

END 6.3

BEGIN 6.2

Colligative properties: are properties determined by the number of particles, not the type or chemistry of the solute.

Boiling point elevation - $\Delta T_b = K_b \cdot m$

K_b = boiling point constant ($0.51^\circ\text{C}/m$ for water)

m = molality of solution

SOLUTE RAISES BP OF SOLVENT
BY INCREASING PRESSURE ON
BUBBLE

Freezing point depression - $\Delta T_f = K_f \cdot m$

K_f = freezing pt constant ($1.86^\circ\text{C}/m$ for water)

m = molality of solution

SOLUTE LOWERS FP OF SOLVENT
BY HINDERING CRYSTAL FORMATION

TABLE 11.5 Molal Boiling-Point Elevation Constants (K_b) and Freezing-Point Depression Constants (K_f) for Several Solvents

Solvent	Boiling Point ($^\circ\text{C}$)	K_b ($^\circ\text{C} \cdot \text{kg/mol}$)	Freezing Point ($^\circ\text{C}$)	K_f ($^\circ\text{C} \cdot \text{kg/mol}$)
Water (H_2O)	100.0	0.51	0	1.86
Carbon tetrachloride (CCl_4)	76.5	5.03	-22.99	30.
Chloroform (CHCl_3)	61.2	3.63	-63.5	4.70
Benzene (C_6H_6)	80.1	2.53	-5.5	5.12
Carbon disulfide (CS_2)	46.2	2.34	-111.5	3.83
Ethyl ether ($\text{C}_4\text{H}_{10}\text{O}$)	34.5	2.02	-116.2	1.79
Camphor ($\text{C}_{10}\text{H}_{16}\text{O}$)	208.0	5.95	179.8	40.

Osmotic pressure - $\pi = i \cdot M \cdot R \cdot T$

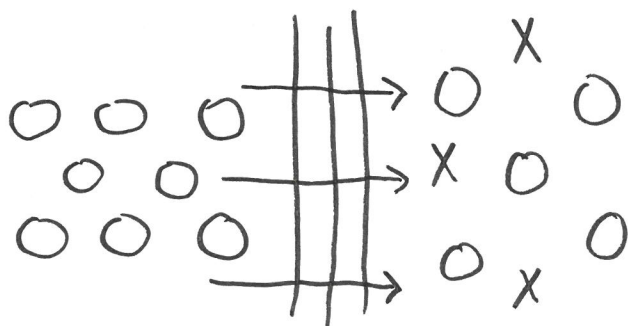
π = pressure in atmospheres

i = van't Hoff factor

M = Molarity of solution

$R = 0.08206 \text{ L-atm/K} \cdot \text{mol}$

T = Kelvin



EX 7: Calculate the boiling point elevation and freezing point depression of a solution that contains 2.0 grams of sucrose ($C_{12}H_{22}O_{11}$; molar mass = 343 grams) in 100 grams of solution.

The van't Hoff factor (i) – is the effect an ionic solute has on the colligative properties of a solution.

EX 8: Calculate the boiling point elevation and freezing point depression of a solution that contains 2.0 grams of NaCl (molar mass = 58.5 grams) in 100 grams of solution.

EX 9: A solution is prepared by dissolving 18.0 grams of a nonvolatile solute in 150.0 grams of water. The resulting solution was found to have a boiling point of $100.34^{\circ}C$. Calculate the molar mass of glucose.

EX 10: What mass of ethylene glycol ($C_2H_6O_2$; molar mass = 62.1 g/mol), the main component in antifreeze, must be added to 10.0 liters of water to produce a solution for use in a car radiator that freezes at $-23.3^{\circ}C$? Assume the water density is 1 g/mL.

$C_2H_6O_2$ 62.1 g/mol
 $-23.3^{\circ}C$

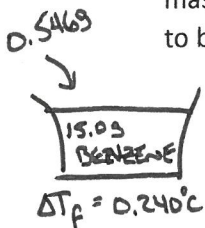
$\Delta T_f = K_f \cdot m \cdot \#ions$
 $23.3^{\circ}C = 1.86^{\circ}C \cdot m \cdot 1$
 $m = 12.5 \text{ moles } C_2H_6O_2 / 10 \text{ Kg } H_2O = 125 \text{ moles } C_2H_6O_2$

$125 \text{ moles } C_2H_6O_2 \cdot 62.1 \text{ g } C_2H_6O_2 / 1 \text{ mol } C_2H_6O_2 = 7,762.5 \text{ g } C_2H_6O_2$

10.0 L H_2O * 1 L = 1 Kg H_2O
 1 mL = 1 g H_2O
 $D = 1.0 \text{ g/mL}$

* $C_2H_6O_2$ IS MOLECULAR,
 "NON-VOLATILE".
 (1 PARTICLE)

EX 11: A chemist is trying to identify a human hormone by determining its molar mass. A sample with a mass of 0.546 grams is dissolved in 15.0 g of benzene and the freezing point depression was determined to be 0.240°C. Calculate the molar mass of the hormone.



$$\Delta T_f = K_f \cdot m \cdot \#ions$$

$$0.240^\circ C = 5.12 \cdot m \cdot 1 \rightarrow m = \frac{0.0469 \text{ g Compd}}{1 \text{ kg Benzene}} = 0.0007035 \text{ mol Compd}$$

$$\frac{15 \text{ g Compd}}{0.0007035} = 7809 \text{ g/mol}$$

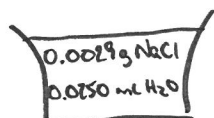
Osmotic pressure – is caused by passing water through a semipermeable membrane.

$$\pi = i \cdot M \cdot R \cdot T$$

EX 12: A normal human cell will burst if the pressure inside the cell exceeds 90 atm of pressure. Calculate the osmotic pressure of a cell if the contents have 0.0029 grams of NaCl dissolved in 0.0250 mL of water at 25°C and determine if the cell can survive.

$$\pi = (2)(1.98M)(0.08206)(298) \rightarrow \pi = 96.9 \text{ atm} > 90 \text{ atm} \therefore \text{IT WILL BURST}$$

$$M = \frac{0.0029 \text{ g NaCl}}{146.3 \text{ g NaCl}} = \frac{0.000049572 \text{ mol NaCl}}{1 \text{ L Sol'n}} = 1.98 \text{ M Sol'n}$$



* CELL BURSTS @ 90 ATM *

$$K = ^\circ C + 273 \rightarrow K = 25^\circ C + 273 = 298$$

NaCl $\rightarrow$ Na⁺ + Cl⁻ 2 PARTICLES

EX 13: To determine the molar mass of a certain protein, 0.001 grams of it is dissolved in enough water to make 1.00 mL of solution. The osmotic pressure of this solution was found to be 1.12 torr at 25°C. Calculate the molar mass of the protein.



0.001g PROTEIN
 1.00 mL SOLUTION
 1.12 Torr @ 25°C
 $K = ^\circ C + 273$
 $K = 298$

$$\frac{1.12 \text{ Torr}}{760 \text{ Torr}} = 0.00147 \text{ ATM}$$

$$\pi = i \cdot M \cdot R \cdot T \quad (\text{PROTEIN} = \text{NONVOLATILE, 1 PARTICLE})$$

$$0.00147 \text{ ATM} = (1)(M)(0.08206)(298)$$

$$M = \frac{0.000060 \text{ mol PROTEIN}}{1 \text{ L Sol'n}} = 6.01 \times 10^{-8} \text{ MOLES}$$

$$\frac{0.001 \text{ g}}{6.01 \times 10^{-8} \text{ MOLE}} = 16,639 \text{ g/mol}$$

Intermolecular and intramolecular forces

Intramolecular forces – forces that hold the atoms of compounds together, 3 types

1. Metallic bonds – metallic atoms in a “sea” of electrons
2. Ionic bonds – transfer of electrons
3. Covalent bonds – sharing of electrons

COULOMB'S CONSTANT
 CHARGE

$$F = K \cdot \left(\frac{q_1 \cdot q_2}{r^2} \right)$$

 RADIUS / DISTANCE

END 6.2

BEGIN 6.4

Intermolecular forces (IMFs) – those forces that hold molecules together.

Two important characteristics of IMFs

1) Boiling point

2) Vapor pressure

1. London dispersion forces

Molecule	State @ room temp	Boiling Point
F – F	Gas	-188°C
Cl – Cl	Gas	-35°C
Br – Br	Liquid	58°C
I – I	Solid	184°C

He

Ar

CH₄

C₈H₁₈

2. Dipole-Dipole forces

CO

Acetone (C₂H₆O)

Hydrogen bonding – special dipole-dipole forces

H – F, O, N

H – F

H₂O

NH₃

List the intermolecular forces present in CH₃OH, Ar, Ne, CH₃OCH₃ and then rank from lowest to highest boiling point

What determines the strength of the following forces:

Dipole-Dipole forces

Hydrogen bonding forces

London dispersion forces

Draw a picture of an ion-induced dipole:

Identify the main type of intermolecular force in the following:

Ne _____

NH₃ _____

KCl _____

PH₃ _____

BCl₂ _____

NaCl in water _____

C₂H₆ _____

Which of the following molecule pairs are NOT involved in hydrogen bonding?

1) HCOOH, H₂O

3) CH₃OH, CH₃COOH

2) H₂O, NH₃

4) H₂ and I₂

Which of the following molecules interact primarily through London-dispersion forces?

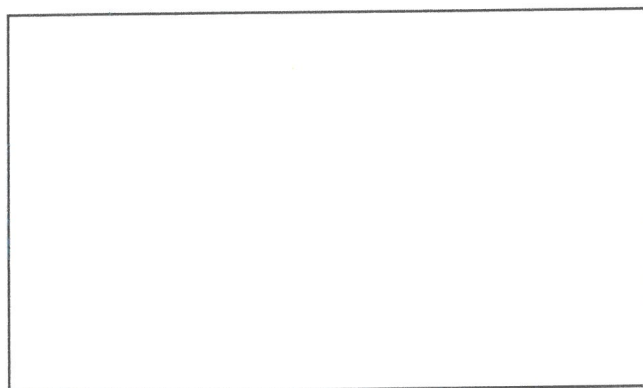
1) SO₂

3) CH₂Cl₂

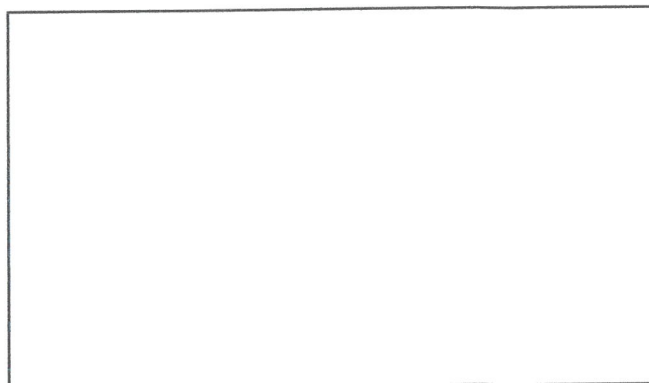
2) CCl₄

4) H₂S

Draw diagrams showing 3 molecules of carbon dioxide in the state of matter indicated below in the box.
On the diagram, show where the intermolecular and intramolecular forces are.



Liquid



Gas

In the box provided, draw a picture of a water molecule dissolving a picture of a crystal of NaCl

