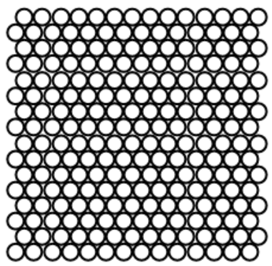
Name ________________________________ **Chem 162, Section**: _____ **Group Number:** ______

ALE 24. Colligative Properties (Part 2)

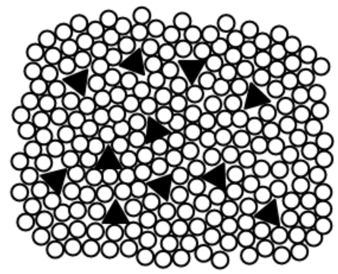
(Reference: 13.6 *Silberberg 5th edition*)

Why is calcium chloride spread on highways in the North during the Winter?

The Model: Intermolecular Forces and Melting Point



Compound "A": a pure crystalline solid
One Molecule of Compound A = ○



A mixture of solids A and B:
One molecule of compound B = ▲
One molecule of compound A = ○

Key Questions

1a. In which crystal (that of the pure compound or that of the mixture) would you expect there to be more (and, therefore, stronger) intermolecular forces of attraction between neighboring molecules of compound A? (Circle your response.)

b. As intermolecular forces increase in strength, what happens to the amount of thermal energy that is needed to separate the molecules? Does the amount of thermal energy decrease, increase, or remain the same? (Circle your response.)

c. Which should have the higher melting point: the pure crystal of compound A or the solid mixture of A and B? (Circle your response.)

2. To ensure that it will be safe to drive on the roadways where it snows, calcium chloride is often spread on the pavement in the wintertime to prevent water from freezing. Does this make sense in light of your answer to Question 1c? (What is the difference between the freezing point and the melting point?) *Explain your answer.*

The Model: Boiling Point Elevation and Freezing Point Depression

In the previous ALE, we learned that the boiling point of a liquid is elevated when it becomes the solvent of a nonvolatile solute. The extent to which the boiling point is raised is given by the equation:

$$\Delta T_b = K_b\, m \qquad \textbf{(1)}$$

How much a freezing point is affected ("ΔT_f") is directly proportional to the concentration (*i.e.*, the molality) of the solute particles. The equation that allows one to calculate just how much a liquid's freezing point is changed is nearly identical to eq **1** for boiling point elevation. The only difference is that subscript *f*'s are used instead of *b*'s.

Key Questions

3. Finish the equation below that allows for the calculation of how much the freezing point of a liquid is changed when it serves as the solvent of a solution of nonvolatile solute particles with a molality equal to *m*.

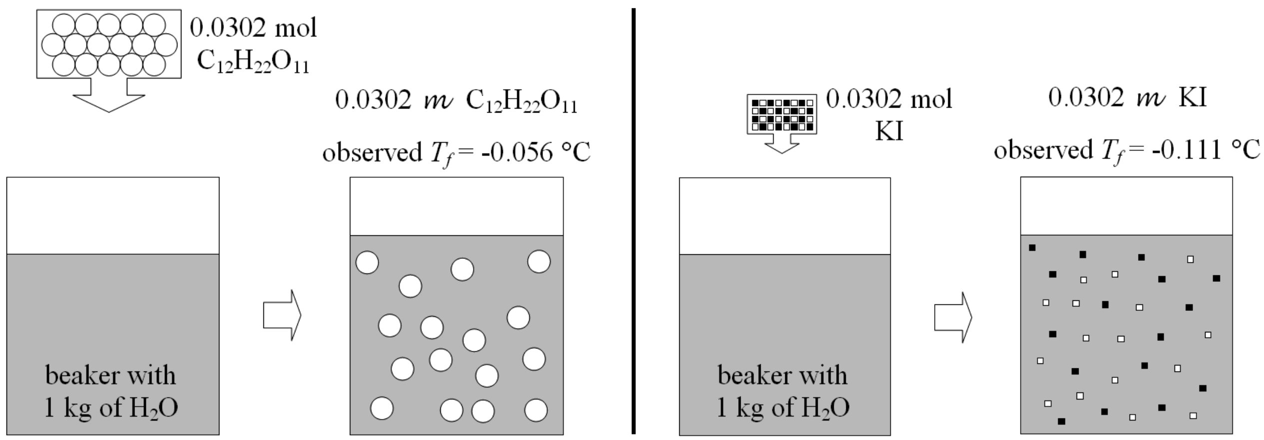
   $$\Delta T_f =$$

4. A certain aqueous solution of ethylene glycol has both a freezing point of -54.0 °C and a boiling point of 110.0 °C. Use this information to show that the value of K_f is not equivalent to the value of K_b. (*Hint*: You do NOT need to know the molality of the solution to do this.)

Exercise

A. On a hot summer day, a frozen sugar water treat hits the spot. 182 g of sucrose (sucrose: $C_{12}H_{22}O_{11}$ = 342.3 g/mol) is dissolved in enough water so that the volume of the solution is 473 mL. This aqueous solution of sugar has a density of 1.14 g/mL. If this sugar-water solution is to be frozen, what is the maximum temperature that one's freezer can be? *Hints*: ① How many moles of sucrose are in the solution? ② What is the mass of the solution? ③ What is the mass of water in kilograms? ④ What is the molality of the solution? ⑤ What is the freezing point depression constant of water? (Look up the value in your textbook.) ⑥ What is the freezing point depression of the solution? ⑦ What is the freezing point of the solution?

The Model: Ionic Solutes and Colligative Properties



A **colligative property** is a property that is affected by the concentration of solute *particles* dissolved in solution.

Key Question

5. Use the Model to explain why the freezing point of 0.0302 *m* KI(*aq*) is -0.111 °C instead of -0.056 °C as it is for 0.0302 *m* $C_{12}H_{22}O_{11}$(*aq*).

Exercise

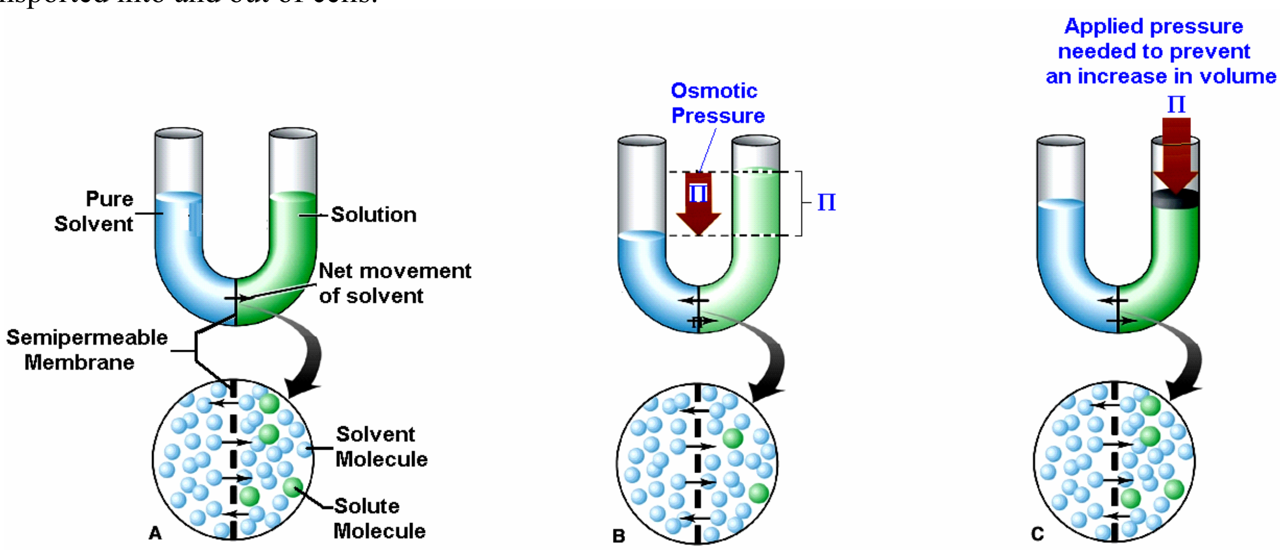
B. If 9.0 g of calcium chloride salt is spread onto a wet road for every 91.0 g of water on the road, at what temperature will the resulting solution on the road freeze? Show your calculations. *Hints:* ① What is the formula of calcium chloride and what is its molar mass? ② What is the molality of calcium chloride in the solution? ③ What is the molality of ions in the solution? ④ What is the freezing point depression of the solution? ⑤ What is the freezing point of the solution?

The Model: Osmotic Pressure

Osmosis is the diffusion of water through a semi-permeable membrane *from a solution of low solute concentration to a solution with high solute*. A **semi-permeable membrane** allows water molecules to pass through, but not the solute molecules. Osmosis is passive process in which a solvent moves without the input of energy from the surroundings. Osmosis releases energy, and can be made to do work, as when a growing tree-root splits a stone or concrete driveway.

Osmosis involves a net movement of solvent is *from the less-concentrated (**hypotonic**) to the more-concentrated (**hypertonic**) solution*, which tends to reduce the difference in concentrations. This effect can be countered by increasing the pressure on the hypertonic solution, with respect to the hypotonic as seen in figure C, below. The **osmotic pressure** is defined to be the pressure required to maintain equilibrium, with no net movement of solvent. Osmotic pressure is a **colligative property**, meaning that the property depends on the concentration of the solute but not on its identity

Osmosis is important in biological systems as many biological membranes are semipermeable. In general, these membranes are impermeable to organic solutes with large molecules, such as polysaccharides, while permeable to water and small, uncharged solutes. Permeability may depend on solubility properties, charge, or polarity as well as solute size. Osmosis provides the primary means by which water is transported into and out of cells.



The osmotic pressure, **Π,** of a solution can be calculated using equation 2:.

Where

$\Pi = MRT$ **(2)**

Π is the osmotic pressure of the solution in atm
M is the molarity, mol solute *particles* per liter of solution
R is the gas constant, where $R = 0.08206\ \mathrm{L \cdot atm \cdot mol^{-1} \cdot K^{-1}}$
T is the temperature in Kelvin

Key Questions

6. Why is it sensible that the osmotic pressure in eqn **2** should be directly proportional to the concentration of solute particles in the solution? (Consider what the Model says is the reason why there is a net flow of water molecules from the body of pure water to the solution.)

7a. Why is it sensible that the osmotic pressure in eqn **2** should be directly proportional to the temperature of the system? (Consider what happens to the number of collisions between water molecules and the semipermeable membrane when temperature is increased.)

b. Why must the temperature in eqn **2** be put on the *absolute* ("Kelvin") scale? Hint: Consider at what temperature molecular motion would stop and if the temperature is below 0 ^{o}C.

Exercises

C. If blood is subjected to a solution with different osmotic pressure than the solution that is within blood cells, the cells could swell (i.e. in a hypotonic solution) or shrivel (i.e. in a hypertonic solution)—in either case, the cells could die. A 0.15 *M* NaCl aqueous solution has the same osmotic pressure as blood—i.e. 0.15 *M* NaCl is **isotonic** with blood. What is the osmotic pressure of 0.15 *M* NaCl at 37 °C (*i.e.* body temperature) in mmHg? *Hints*: ① What is the molarity of solute *particles*? ② What is the absolute temperature? ③ What is the osmotic pressure of the solution in atm? In mmHg?

D. How many moles of solvent particles are present in 1.0 mL of each of the following aqueous solutions? Show your work using dimensional analysis—circle your answers.

a.) 0.010 *M* $CuSO_4$

b.) 0.0050 *M* $Ca(OH)_2$

c.) 0.060 *M* glucose, $C_6H_{12}O_6$

d.) 0.050 *M* $(NH_4)_2CO_3$

E. Given the following three aqueous solutions:

(I) 0.02 *m* Urea, $(NH_2)_2C{=}O$; (II) 0.02 *m* NaCl; (III) 0.02 *m* $CuSO_4$

Rank the aqueous solutions, above, in order of *decreasing:*

a.) Osmotic pressure ____________________________________

b.) Boiling point ____________________________________

c.) Freezing point ____________________________________

d.) Vapor pressure at 298 K ____________________________________

F. Calculate the minimum mass of ethylene glycol ("antifreeze," $C_2H_6O_2$ or $HOCH_2CH_2OH$) that must be dissolved in 14.5 kg of water to prevent the solution from freezing at –10.0 ^{o}F. Assume ideal behavior. $^{\circ}C = \frac{5}{9}(^{\circ}F - 32)$