

CH 222 Chapter Eleven Concept Guide

1. Molality

Question

A 4.5 M nitric acid solution contains 65.0 g of HNO_3 in 288 g of solution. What is the molality of this solution?

Approach

Molality is calculated by dividing the moles of solute by the kilograms of solvent.

Solution

Step 1. Calculate the number of moles of HNO_3 .

$$\left(65.0 \text{ g HNO}_3\right)\left(\frac{1 \text{ mol HNO}_3}{63.01 \text{ g HNO}_3}\right) = 1.03 \text{ mol HNO}_3$$

Step 2. Calculate the molality of the solution.

There are 223 g of solvent: $288 \text{ g solution} - 65.0 \text{ g solute} = 223 \text{ g solvent}$.

Molality = moles of solute/kilograms of solvent

$$\text{Molality} = \frac{1.03 \text{ mol HNO}_3}{0.223 \text{ kg}} = 4.62 \text{ mol/kg}$$

2. Weight percent

Question

What is the weight percent of methanol in a solution of 1.0 L of methanol in 2.5 L of diethyl ether? The density of methanol is 0.79 g/mL and the density of diethyl ether is 0.71 g/mL.

Approach

To find the weight percent of methanol, we must know the masses of each component in the solution. Then, we can calculate the weight percent by dividing the mass of methanol by the mass of solution, and then multiplying by 100.

Solution

Step 1. Calculate the mass of methanol and the mass of diethyl ether.

$$\left(1.0 \text{ L methanol}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right)(0.79 \text{ g/mL}) = 790 \text{ g methanol}$$

$$\left(2.5 \text{ L diethyl ether}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right)(0.71 \text{ g/mL}) = 1800 \text{ g diethyl ether}$$

Step 2. Calculate the mass of the solution.

$$790 \text{ g methanol} + 1800 \text{ g diethyl ether} = 2590 \text{ g solution}$$

Step 3. The weight percent of methanol in the solution is calculated by dividing the mass of methanol by the mass of the solution, then multiplying by 100.

$$\left(\frac{790 \text{ g methanol}}{2600 \text{ g solution}} \right) (100) = 30. \% \text{ methanol}$$

3. Solution Concentration

Problem

Commercial vinegar is an aqueous solution that must contain at least 4 weight percent acetic acid. The density of this solution is 1.0058 g/mL and the volume of the sample is 100.00 g. Calculate the concentration of acetic acid, if this vinegar is exactly 4.000% acetic acid, in terms of (a) mole fraction, (b) molality, and (c) molarity. Acetic acid is $\text{CH}_3\text{CO}_2\text{H}$ and has a molar mass of 60.05 g.

Approach

To find the mole fraction, we need to calculate the numbers of moles of solute and solvent in 100.00 g of solution. Then, divide the number of moles of acetic acid by the total number of moles of solution. To find the molality, we will need to use a ratio of moles of acetic acid to kilograms of solvent. Last, to find the molarity of the solution, simply divide the moles of acetic acid by the number of liters of solution.

Solution

(a) Mole Fraction

Step 1. Calculate the mass and moles of both acetic acid and water.

$$4 \text{ mass percent } \text{CH}_3\text{CO}_2\text{H} = \frac{\text{mass } \text{CH}_3\text{CO}_2\text{H}}{\text{mass solution}}$$

$$\text{Mass } \text{CH}_3\text{CO}_2\text{H} = 0.04 \times 100.00 = 4.00 \text{ g } \text{CH}_3\text{CO}_2\text{H}$$

$$\text{Moles of } \text{CH}_3\text{CO}_2\text{H} = \left(4.00 \text{ g } \text{CH}_3\text{CO}_2\text{H} \right) \left(\frac{1 \text{ mol } \text{CH}_3\text{CO}_2\text{H}}{60.05 \text{ g } \text{CH}_3\text{CO}_2\text{H}} \right) = 0.0666 \text{ mol } \text{CH}_3\text{CO}_2\text{H}$$

$$\begin{aligned} \text{Mass of } \text{H}_2\text{O} &= \text{mass of solution} - \text{mass of } \text{CH}_3\text{CO}_2\text{H} \\ &= 100.00 \text{ g solution} - 4.00 \text{ g } \text{CH}_3\text{CO}_2\text{H} = 96.00 \text{ g } \text{H}_2\text{O} \end{aligned}$$

$$= \left(96.00 \text{ g } \text{H}_2\text{O} \right) \left(\frac{1 \text{ mol } \text{H}_2\text{O}}{18.015 \text{ g } \text{H}_2\text{O}} \right) = 5.329 \text{ mol } \text{H}_2\text{O}$$

Step 2. Calculate the mole fraction of acetic acid.

$$\begin{aligned} \text{Mole fraction } \text{CH}_3\text{CO}_2\text{H} &= \frac{\text{moles } \text{CH}_3\text{CO}_2\text{H}}{\text{moles solution}} \\ &= \frac{0.0666 \text{ mol } \text{CH}_3\text{CO}_2\text{H}}{0.0666 \text{ mol } \text{CH}_3\text{CO}_2\text{H} + 5.329 \text{ mol } \text{H}_2\text{O}} = 0.0123 \end{aligned}$$

(b) Molality

$$\begin{aligned}\text{Molality} &= \frac{\text{moles CH}_3\text{CO}_2\text{H}}{\text{kg solution}} \\ &= \frac{0.0666 \text{ mol CH}_3\text{CO}_2\text{H}}{0.09600 \text{ kg H}_2\text{O}} \\ &= 0.694 \text{ mol/kg H}_2\text{O}\end{aligned}$$

(c) Molarity

Step 1. Calculate the volume of the solution.

$$\left(100 \text{ g solution}\right)\left(\frac{1 \text{ mL}}{1.0058 \text{ g}}\right) = 99.42 \text{ mL solution}$$

Step 2. Calculate the molarity by dividing the moles of solute by the liters of solution.

$$= \frac{0.0666 \text{ mol CH}_3\text{CO}_2\text{H}}{0.09942 \text{ L solution}} = 0.670 \text{ mol/L}$$

4. Parts Per Million

Problem

You have a solution consisting of 2.665 g of solute and 151.78 g of solution. Convert this to units of ppm.

Approach

To calculate ppm, convert grams of solute to milligrams, and grams of solution to kilograms.

Solution

$$\frac{2665 \text{ mg solute}}{0.15178 \text{ kg solution}} = 17,560 \text{ ppm}$$

5. Solubility and Henry's Law

Question

A soft drink has an aqueous CO₂ concentration of 0.0511 M at 25 °C. What is the pressure of CO₂ gas in the drink? Henry's law constant for CO₂ is 4.48 x 10⁻⁵ M/mm Hg at 25 °C.

Approach

We need to use Henry's law:

$$S_g = k_H P_g$$

where S_g is gas solubility, k_H is Henry's law constant, and P_g is the partial pressure of CO₂. Substituting the gas solubility and Henry's law constant for CO₂ into the equation will yield the pressure of CO₂.

Solution

$$P_g = \frac{S_g}{k_H}$$

$$P_g = \frac{0.0511 \text{ M}}{4.48 \times 10^{-5} \text{ M/mm Hg}} = 1.14 \times 10^3 \text{ mm Hg}$$

6. Solubility and Henry's Law

Question

The partial pressure of O₂ in a person's lungs varies from 22 mm Hg to 40 mm Hg. How much O₂ can dissolve in water at 25 °C if the partial pressure of O₂ is 35 mm Hg? Henry's law constant for O₂ is 1.66 x 10⁻⁶ M/mm Hg at 25 °C.

Solution

We need to use Henry's law

$$S_g = k_H P_g$$

where S_g is gas solubility, k_H is Henry's law constant, and P_g is the partial pressure of O₂. Substituting Henry's law constant for O₂ and the partial pressure into the equation will yield the solubility of O₂ in water.

$$S_g = \left(\frac{1.66 \times 10^{-6} \text{ M}}{\text{mm Hg}} \right) (35 \text{ mm Hg}) = 5.8 \times 10^{-5} \text{ M}$$

7. Solubility and Henry's Law

Question

What is the concentration of O₂ (in grams of O₂ per liter of water) in a freshwater stream in equilibrium with air at 25 °C? The atmospheric pressure is 1.0 atm and Henry's law constant for O₂ is 1.66 x 10⁻⁶ M/mm Hg at 25 °C. Assume air contains 21% oxygen.

Approach

Henry's Law can be used to calculate the solubility of oxygen. First, calculate the partial pressure of oxygen in air (21% of air is oxygen and the mole fraction of O₂ is 0.21). Then, calculate the solubility of oxygen using Henry's law.

Solution

Step 1. Calculate the partial pressure of oxygen.

$$P(\text{O}_2) = \left(1.0 \text{ atm} \right) \left(\frac{760 \text{ mm Hg}}{1 \text{ atm}} \right) (0.21) = 160 \text{ mm Hg}$$

Step 2. Calculate the solubility of oxygen in units of grams of oxygen per liter of water.

$$\left(\frac{1.66 \times 10^{-6} \text{ mol/L}}{1 \text{ mm Hg}} \right) (160 \text{ mm Hg}) = 2.7 \times 10^{-4} \text{ mol/L}$$

$$(2.7 \times 10^{-4} \text{ mol/L})(32. \text{ g/mol}) = 0.0085 \text{ g/L}$$

8. Vapor Pressure and Raoult's Law

Question

What is the vapor pressure at 25 °C of a benzene-toluene solution of composition $X_{\text{benz}} = 0.40$ and $X_{\text{tol}} = 0.60$? The vapor pressures of the pure substances are 73 Torr for benzene, C_6H_6 , and 27 Torr for toluene, C_7H_8 , at 25 °C. Assume that the benzene and toluene form an ideal solution.

Approach

According to Raoult's law, the partial pressure of each component in the vapor phase is directly proportional to its mole fraction in the solution:

$$P_A = X_A P_A^\circ$$

The total vapor pressure of the solution is the sum of the partial pressures.

Solution

Step 1. Calculate the partial pressures of benzene and toluene.

$$P_{\text{benz}} = X_{\text{benz}} P_{\text{benz}}^\circ = (0.40)(73 \text{ Torr}) = 29 \text{ Torr}$$

$$P_{\text{tol}} = X_{\text{tol}} P_{\text{tol}}^\circ = (0.60)(27 \text{ Torr}) = 16 \text{ Torr}$$

Step 2. Calculate the total vapor pressure of the solution.

$$P_{\text{soln}} = P_{\text{benz}} + P_{\text{tol}} = 29 \text{ Torr} + 16 \text{ Torr} = 45 \text{ Torr}$$

9. Colligative Properties: Boiling Point Elevation and Freezing Point Depression

Question

A solution contains a mixture of 5 sugars: 0.50 mol fructose, 0.60 mol glucose, 0.50 mol lactose, 0.60 mol maltose, and 0.60 mol sucrose dissolved in 1.00 kg of water. What are the boiling point and the freezing point of this solution? The boiling point elevation constant for water is $0.512^\circ\text{C} / \text{m}$, and the freezing point depression constant for water is $-1.86^\circ\text{C} / \text{m}$.

Approach

We need to first calculate the total number of moles of solute in the solution. Then, we will need to use the following relationship to calculate the boiling point elevation:

$$\Delta T_{\text{bp}} = K_{\text{bp}} \cdot m_{\text{solute}}$$

Similarly, for freezing point depression, we will need to use the relationship:

$$\Delta T_{\text{fp}} = K_{\text{fp}} \cdot m_{\text{solute}}$$

The solution's boiling point can be calculated by adding the change in temperature found to the boiling point of water. The solution's freezing point can be calculated by subtracting the change in temperature found from the freezing point of water.

Solution

Step 1. Calculate the total number of moles of solute in the solution and the total concentration of solute.

$0.50 \text{ mol fructose} + 0.60 \text{ mol glucose} + 0.50 \text{ mol lactose} + 0.60 \text{ mol maltose} + 0.60 \text{ mol sucrose} = 2.80 \text{ mol solute}$

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kg of solvent}} = \frac{2.80 \text{ mol}}{1.00 \text{ kg H}_2\text{O}} = 2.80 \text{ mol/kg}$$

Step 2. Calculate the boiling point elevation of the solution and the boiling point of the solution.

$$\Delta T_{\text{bp}} = K_{\text{bp}} * m_{\text{solute}} = (0.512 \text{ }^{\circ}\text{C} * \text{kg/mol})(2.80 \text{ mol/kg}) = 1.43 \text{ }^{\circ}\text{C}$$

$$T_{\text{bp, solution}} = T_{\text{bp, solvent}} + \Delta T_{\text{bp}} = 100.00 \text{ }^{\circ}\text{C} + 1.43 \text{ }^{\circ}\text{C} = 101.43 \text{ }^{\circ}\text{C}$$

Step 3. Calculate the freezing point depression of the solution and the freezing point of the solution.

$$\Delta T_{\text{fp}} = K_{\text{fp}} * m_{\text{solute}} = (1.86 \text{ }^{\circ}\text{C} * \text{kg/mol})(2.80 \text{ mol/kg}) = -5.21 \text{ }^{\circ}\text{C}$$

$$T_{\text{fp, solution}} = T_{\text{fp, solvent}} - \Delta T_{\text{fp}} = 0.00 \text{ }^{\circ}\text{C} - 5.21 \text{ }^{\circ}\text{C} = -5.21 \text{ }^{\circ}\text{C}$$

10. Osmotic Pressure

Problem

An aqueous solution contains 77.1 g of insulin ($\text{C}_6\text{H}_{10}\text{O}_5$)_x, a high molecular mass sugar, per liter of solution. The osmotic pressure at 20 °C of this solution is 0.58 atm. Calculate the molar mass of insulin.

Approach

The molarity of the solution can be calculated from the osmotic pressure, using the following relationship: $\pi = cRT$, where π is the osmotic pressure, c is concentration (in moles per liter), R is the gas constant, and T is the absolute temperature. Then, the molar mass can be found from the mass and molarity.

Solution

Step 1. Calculate the solution concentration.

$$\pi = cRT$$

$$c = \frac{0.58 \text{ atm}}{(0.0821 \text{ L atm/Kmol})(293 \text{ K})} = 0.024 \text{ mol/L}$$

Step 2. Calculate the molar mass of insulin.

$$\left(\frac{77.1 \text{ g}}{1 \text{ L}} \right) \left(\frac{1 \text{ L}}{0.024 \text{ mol}} \right) = 3200 \text{ g/mol}$$