Some sample colligative properties and concentration problems ..

What is the freezing point of a 41% solution of urea in water?

$(NH_2)_2CO$: urea , FW = 60.062 g/mol

$\Delta T_f = K_f \times C_m$; $C_m = \dfrac{\text{mol urea}}{\text{kg water}}$; p509: $K_f = 1.858\ ^{\circ}C/m$; $T_{f,H_2O} = 0.00^{\circ}C$

We need to find molal concentration of urea, Cm. We have the mass percent urea - 41%

41% urea : $\dfrac{41\text{ g urea}}{100\text{ g solution}}$

We need to convert 41% urea to molality, so let's assume a basis of 100 g solution! This makes the amount of urea 41 grams. Convert to moles:

$41\text{ g urea} \times \dfrac{\text{mol urea}}{60.062\text{ g urea}} = 0.6826279511\text{ mol urea}$

Subtract to find mass water:

$100\text{ g solution} - 41\text{ g urea} = 59\text{ g } H_2O = 0.059\text{ kg } H_2O$

Find Cm:

$$C_m = \frac{\text{mol urea}}{\text{kg water}} = \frac{0.6826279511 \text{ mol urea}}{0.059 \text{ kg } H_2O} = 11.5699652 7 \text{ m urea}$$

We can now find delta T:

$$\Delta T_f = K_f \times C_m$$

K_f = 1.858 °C/m (p 504)

$$\Delta T_f = (1.858\ °C/m)(11.56996527\ m) = 21\ °C$$

To get the freezing temperature of the solution, subtract the freezing point depression from the original freezing point of water:

$$T_{f,\ H_2O} = 0.00\ °C$$

$$T_{f,\ solution} = 0.00\ °C - 21\ °C = \boxed{-21\ °C}$$

0.2436 g of an unknown substance is dissolved in 20.0 mL of cyclohexane, C_6H_{12}
If the freezing point depression of this solution is 2.5 C, what is the molecular weight of the unknown? The density of cyclohexane at the temperature the cyclohexane volume was measured is 0.779 g/mL.

$\Delta T_f = K_f \cdot C_m$

ΔT_f = 2.5°C, K_f = 20.0°C/m (p509)

$C_m = \dfrac{\text{mol unknown}}{\text{kg } C_6H_{12}}$ ← We need to find Cm because that lets us find out the moles unknown so that we can calculate molecular weight!

Calculate Cm:

$2.5°C = (20.0°C/m) \cdot C_m$

$C_m = 0.125\,m$ unknown

Caution: Cm is in MOLALITY units, not MOLES. We will need to calculate moles from molality, but they're not the same number!

To get moles, we'll need to multiply Cm by the kilograms of cyclohexane actually used in the experiment... So, convert the volume of cyclohexane to mass using the density given:

$0.779\text{ g } C_6H_{12} = \text{mL } C_6H_{12}$

$$20.0\text{ mL } C_6H_{12} \times \frac{0.779\text{ g } C_6H_{12}}{\text{mL } C_6H_{12}} = 15.58\text{ g } C_6H_{12} = 0.01558\text{ kg } C_6H_{12}$$

$$C_m = \frac{\text{mol unknown}}{\text{kg } C_6H_{12}}$$

definition of molality

$$C_m = 0.125 \text{ m unknown}$$

calculated molality from freezing point depression

$$0.01558 \text{ kg } C_6H_{12}$$

mass of 20.0 mL cyclohexane (solvent)

Calculate moles unknown:

$$0.125 \text{ m} = \frac{\text{mol unknown}}{0.01558 \text{ kg } C_6H_{12}}$$

$$\text{mol unknown} = 0.0019475 \text{ mol unknown}$$

Once we have found the moles unknown, we can find the molecular weight:

$$MW = \frac{\text{mass unknown}}{\text{mol unknown}} = \frac{0.2436 \text{ g}}{0.0019475 \text{ mol}} = \boxed{130 \text{ g/mol}}$$

Two significant figures justified ... we only knew the freezing point depression to two significant figures!

Commercial sulfuric acid is 18.0 M. If the density of the acid is 1.802 g/mL, what is the molality?

H_2SO_4, FW = 98.096 g/mol

$$\frac{18.0 \text{ mol } H_2SO_4}{\text{L solution}} \longrightarrow \frac{? \text{ mol } H_2SO_4}{? \text{ kg solvent}}$$

molarity (definition) — molality (definition)

ASSUME A BASIS of 1L of solution. This means that the moles of sulfuric acid has to be 18.0 mol. All we have to do is figure out the corresponding amount of SOLVENT. Start by caclulating the mass of solution, since we know the density.

1.802 g solution = mL solution

$$1000 \text{ mL} \times \frac{1.802 \text{ g}}{\text{mL}} = 1802 \text{ g solution}$$

(1 L)

We know the mass of SOLUTION, but we need the mass of SOLVENT. Find the mass of solvent by subtracting the mass of sulfuric acid. So we need to know the mass of sulfuric acid!

$$18.0 \text{ mol } H_2SO_4 \times \frac{98.096 \text{ g } H_2SO_4}{\text{mol } H_2SO_4} = 1765.728 \text{ g } H_2SO_4$$

Subtract to find mass of solvent:

$$1802\ g\ \text{solution} - 1765.728\ g\ H_2SO_4 = 36.272\ g\ \text{solvent}$$

$$= 0.036272\ kg\ \text{solvent}$$

And the molality is:

$$\frac{\text{mol } H_2SO_4}{kg\ \text{solvent}} = \frac{18.0\ \text{mol } H_2SO_4}{0.036272\ kg\ \text{solvent}} = \boxed{496\ m\ H_2SO_4}$$

Note: The molality in this example is really high due to the fact that there is very little solvent (water) in concentrated sulfuric acid. Since there is so little solvent, the denominator in the molality calculation is very small - giving us the high molality.