Some sample colligative propoerties 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

$\Delta T_f = K_f \times C_m$ $\quad C_m = \frac{\text{mol urea}}{\text{kg water}}$

p509

$K_{f,w} = 1.858\ ^{\circ}C/m$

$T_{f,w} = 0.000\ ^{\circ}C$

We need to find Cm ... and for that we need moles urea and kilograms water.

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

We need mass WATER (not mass solution), so subtract out urea!

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

Since we've assumed a basis of 100g solution, we can calculate the moles urea and then find Cm ...

$41\text{g urea} \times \frac{\text{mol urea}}{60.062\text{g urea}} = 0.682627951\text{ mol urea}$

Now, find Cm:

$$C_m = \frac{0.682627951 \text{ mol urea}}{0.059 \text{ kg } H_2O} = 11.56996527 \text{ m urea}$$

Now, we can find delta T

$$\Delta T_f = \underline{k_f} \times C_m$$

$\llcorner$ 1.858 °C/m (p509)

$$\Delta T_f = (1.858\ °C/m)(11.56996527 \text{ m urea})$$
$$= 21\ °C$$

To get the Tf of the solution, subtract Tf for the pure solvent and denta T:

$T_{f,H_2O} = 0.000\ °C$ (p509)

$$T_{f,41\%\text{ urea}} = 0.000\ °C - 21\ °C = \boxed{-21\ °C}$$

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.

$$\underbrace{\Delta T_f}_{2.5^\circ C} = \underbrace{K_f}_{20.0^\circ C/m} \times C_m \quad \text{(p509)}$$

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

First, calculate Cm:

$$2.5^\circ C = (20.0^\circ C/m) \times C_m$$

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

We want to find moles unknown (we need it for formula weight). To do that, we'll have to first find out the amount of solvent used ... in kilograms.

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

$$20.0\ mL\ C_6H_{12} \times \frac{0.779 g\ C_6H_{12}}{mL\ C_6H_{12}} = 15.58 g\ C_6H_6$$

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

Fnd moles unknown:

$$0.01558\ \text{kg}\ C_6H_{12} \times \frac{0.125\ \text{mol unknown}}{\text{kg}\ C_6H_{12}} = 0.0019475\ \text{mol unknown}$$

Now we can get molecular weight:

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

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　　　　　　　　　　　　molality

ASSUME A BASIS of 1 L solution....

$$1 \text{ L solution} \times \frac{18.0 \text{ mol } H_2SO_4}{\text{L}} = 18.0 \text{ mol } H_2SO_4$$

We know the moles sulfuric acid ... all that's left to do is to figure out the mass of the solvent. Start by using the volume and density of the SOLUTION.

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

To find the mass of SOLVENT ... we need to subtract out the mass of SULFURIC ACID:

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

So the mass of solvent is ...

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

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

... and the molal concentration is ...

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