Some sample colligative properties and concentration problems ...

(1) What is the freezing point of an aqueous urea solution which contains 41 grams urea per 100. grams of solution?

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

$\Delta T_f = K_f \times C_m$ — molal concentration

K_f = 1.86 °C/m

OpenStax p619:
Kf = 1.86 C/m
Tf = 0.0 C

Use the info in the problem to find out what the MOLAL CONCENTRATION of the urea solution is...

$$\frac{\text{mol urea} \ (1)}{\text{kg } H_2O \ (2)}$$

1) Convert mass urea to moles using formula weight.

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

2) Subtract to find mass of water.

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

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

$\Delta T_f = K_f \times C_m$

K_f = 1.86 °C/m

C_m = 11.5699652 m urea

$\Delta T_f = (1.86\ °C/m)(11.5699652\ m) = 21.52013541\ °C$

$= 22\ °C$ <- This is CHANGE in freezing point, so subtract it from the original freezing point to get ...

new $T_f = 0.0\ °C - 22\ °C$

$= -22\ °C$

(2) 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. Kf for cyclohexane is 20.0 C/m.

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

ΔT_f = 2.5°C; K_f = 20.0°C/m

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

First, let's calculate Cm...

$$(2.5°C) = (20.0°C/m)\,C_m$$

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

To continue, we need to find out how many moles of unknown there are. We know the moles PER KILOGRAM OF CYCLOHEXANE, so if we find out how much cyclohexane we have, we can find the actual moles of unknown!

Find mass cyclohexane

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

Find moles unknown..

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

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

(2 sig figs!)

(3) 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

$$\frac{18.0 \text{ mol } H_2SO_4}{\text{L solution}} \longrightarrow \frac{?\text{ mol } H_2SO_4}{\text{Kg Solvent}}$$

Definition of molarity

Definition of molality

Assume a basis of 1 liter of solution. This means that the number of moles of sulfuric acid is 18.0 moles. Since we know how many moles of sulfuric there is, all we have to do is find the mass of solvent.

Start off by finding the mass of SOLUTION... 1 L = 1000 mL

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

If we subtract the mass of sulfuric acid, we'd have the mass of solvent. Convert 18.0 moles sulfuric acid to mass.

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

$$1802 \text{ g solution} - 1765.728 \text{ g } H_2SO_4 = 36.272 \text{ g solvent} = 0.036272 \text{ kg}$$

Now we can find molality...

$$m = \frac{18.00 \text{ mol } H_2SO_4}{0.036272 \text{ kg Solvent}} = \boxed{496 \; \underline{m} \; H_2SO_4}$$