An Application of the van't Hoff Factor in Acid/Base Equilibrium

The freezing point of 0.10 *m* acetic acid is -0.19°C. What is the van't Hoff factor for acetic acid at this concentration and what fraction (in percentage) of the acetic acid molecules are ionized?

Hints:

$$i = \frac{\text{number of particles in solution}}{\text{number of particles that dissolved}} = \frac{\Delta T_{measured}}{\Delta T_{theoretical}}$$

$$\text{Percentage ionization} = \frac{[\mathrm{H}^+]}{C_{\text{total acetic acid}}} \times 100 = \frac{[\mathrm{CH_3CO_2^-}]}{C_{\text{total acetic acid}}} \times 100$$

$$\Delta T = -iK_f c \quad \text{so...} \quad i = -\frac{\Delta T}{K_f c} = -\frac{-0.19°\mathrm{C}}{\left(1.86\,\frac{°\mathrm{C}}{\mathrm{m}}\right)\left(0.10\ \mathrm{m}\right)} = 1.02$$

$$1.02 = \frac{n_{\mathrm{CH_3CO_2H}} + n_{\mathrm{CH_3CO_2^-}} + n_{\mathrm{H^+}}}{n_{\mathrm{CH_3CO_2H\ original}}}$$

$$n_{\mathrm{CH_3CO_2H}} + n_{\mathrm{CH_3CO_2^-}} = n_{\mathrm{CH_3CO_2H\ original}}$$

The volume of the solution remains unchanged between initial condition and equilibrium. So it's possible to pick any volume of solution to determine $n_{\mathrm{CH_3CO_2H\ original}}$.

$$1.02 = \frac{n_{\mathrm{CH_3CO_2H\ original}} + n_{\mathrm{H^+}}}{n_{\mathrm{CH_3CO_2H\ original}}} \qquad \text{For } n_{\mathrm{CH_3CO_2H\ original}} = 0.1\ \text{mol}, \quad n_{\mathrm{H^+}} = 0.002\ \text{mol}$$

$$\text{Percentage ionization} = \frac{[\mathrm{H}^+]}{C_{\text{total acetic acid}}} \times 100 = \frac{0.002/V}{0.1/V} \times 100 = 2\%\ \text{ionized}$$