Acids and Bases

#7

Calculate the pH of a solution of 1.116 g KOAc in 250 mL solution.

Solution

KOAc:

$$1.116\ \text{g} = \frac{1.116\ \text{g}}{98.1\ \frac{\text{g}}{\text{mol}}} = 0.0114\ \text{mol}$$

$$\frac{0.0114\ \text{mol}}{0.250\ \text{L}} = 0.456\ \frac{\text{mol}}{\text{L}}$$

KOAc:

= salt, completely dissociated

$KOAc(aq) \rightarrow K^+(aq) + OAc^-(aq)$

K^+ = weak acid, weaker than water

OAc^- = weak base

$OAc^-(aq) + H_2O \rightarrow HOAc(aq) + OH^-(aq)$

$$K_{b_{OAc^-}} = \frac{[HOAc] \times [OH^-]}{[OAc^-]} = \frac{1.0 \times 10^{-14}}{K_{a_{HOAc}}} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$\frac{\text{mol}}{\text{L}}$	OAc^-	HOAc	OH^-
Start	0.456	0	0*
Δ	– x	+ x	+ x
Equilibrium	0.456 – x	x	x

(*) The hydroxide ions delivered by the ionization of water are neglected. The weak base OAc^- ($K_b = 5.6 \times 10^{-10}$) is much stronger than the weak base H_2O ($K_b = 1.0 \times 10^{-14}$).

$$K_{b_{OAc^-}} = \frac{x^2}{0.456 - x} = 5.6 \times 10^{-10}$$

$$\Rightarrow x^2 + 5.6 \times 10^{-10} x - 2.5 \times 10^{-10} = 0$$

$$\Rightarrow x = 1.6 \times 10^{-5}$$

$\frac{\text{mol}}{\text{L}}$	OAc^-	HOAc	OH^-
Equilibrium	0.131	1.6×10^{-5}	1.6×10^{-5}

$\Rightarrow$ pOH = 4.80 $\Rightarrow$ **pH = 9.20**