Marks
8

- Solution A consists of a 0.50 M aqueous solution of HF at 25 °C. Calculate the pH of Solution A. The pK_a of HF is 3.17.

As HF is a weak acid, $[H_3O^+]$ must be calculated using a reaction table:

	HF	H₂O	$\rightleftharpoons$	H₃O⁺	HF
initial	0.50	large		0	0
change	-x	negligible		+x	+x
final	0.50 –x	large		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{x^2}{0.50 - x}$$

As pK_a = -log$_{10}$$K_a$, $K_a = 10^{-3.17}$ and is very small, 0.50 – x ~ 0.50 and hence:

$$x^2 = 0.50 \times 10^{-3.17} \quad \text{or} \quad x = 1.84 \times 10^{-2} \text{ M} = [H_3O^+]$$

Hence, the pH is given by:

$$pH = -log_{10}[H_3O^+] = -log_{10}[(1.84 \times 10^{-2})] = 1.74$$

pH = **1.74**

ANSWER CONTINUES ON THE NEXT PAGE

At 25 °C, 1.00 L of Solution B consists of 12.97 g of lithium fluoride, LiF, dissolved in water. Calculate the pH of Solution B.

The molar mass of LiF is 6.941 (Li) + 19.00 (F) = 25.941. Hence, the number of moles in 12.97 g is:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{12.97}{25.941} = 0.5000 \, \text{mol}$$

As this is dissolved in 1.00 L, $[F^-]$ = 0.500 M.

F^- is a weak base so $[H_3O^+]$ must be calculated via the calculation of $[OH^-]$ from the reaction table:

	F^-	H_2O	$\rightleftharpoons$	HF	OH^-
initial	0.500	large		0	0
change	-x	negligible		+x	+x
final	0.500 - x	large		x	x

The equilibrium constant K_b is given by:

$$K_b = \frac{[HF][OH^-]}{[F^-]} = \frac{x^2}{0.500 - x}$$

As pK_a + pK_b = 14.00, pK_b = 14.00 − 3.17 = 10..83. As pK_b = -$\log_{10}K_b$, K_b = $10^{-10.83}$ and is very small. Hence, 0.500 – x ~ 0.500 and hence:

$$x^2 = 0.500 \times 10^{-10.83} \quad \text{or} \quad x = 2.72 \times 10^{-6} \, M = [OH^-]$$

Hence, the pOH is given by:

$$pOH = -\log_{10}[OH^-] = -\log_{10}[(2.72 \times 10^{-6})] = 5.57$$

Finally, as pH + pOH = 14.00, pH = 14.00 – 5.57 = 8.43

pH = **8.43**

ANSWER CONTINUES ON THE NEXT PAGE

Solution B (1.00 L) is poured into Solution A (1.00 L) and allowed to equilibrate at 25 °C. Calculate the pH of the final solution.

The solution consists of a weak acid (HF) and its conjugate base (F^-). The Henderson -Hasselbalch equation can be used:

$$pH = pK_a + \log_{10}\left(\frac{[base]}{[acid]}\right)$$

[base] = [F^-] = 0.500 M and [acid] = [HF] = 0.50 M. Hence,

$$pH = pK_a + \log_{10}\left(\frac{[base]}{[acid]}\right) = 3.17 + \log_{10}\left(\frac{0.500}{0.50}\right) = 3.17$$

pH = **3.17**

If you wanted to adjust the pH of the mixture of Solution A and Solution B to be exactly equal to 4.00, which component in the mixture would you need to increase in concentration?

pH needs to increase: increase [base] (i.e. [LiF])