Marks 6

- Write the balanced chemical equation for the dissolution of solid $Ca(CH_3CO_2)_2$ in water.

$Ca(CH_3CO_2)_2(s) \rightarrow Ca^{2+}(aq) + 2CH_3CO_2^-(aq)$

What is the pH of a solution that has 158.2 g of $Ca(CH_3CO_2)_2$ dissolved in 1.000 L of water? The pK_a of acetic acid, CH_3COOH, is 4.76.

The molar mass of $Ca(CH_3CO_2)_2$ is:

molar mass = (40.08 (Ca) + 4×12.01 (C) + 6×1.008 (H) + 4×16.00 (O)) g mol^{-1}
= 158.168 g mol-1

Thus, 158.2 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{158.2\ \text{g}}{158.168\ \text{g mol}^{-1}} = 1.000\ \text{mol}$$

When $Ca(CH_3CO_2)_2$ is dissolved, it produces $Ca^{2+}(aq) + 2CH_3CO_2^-$. Hence, if 1.000 mol of $Ca(CH_3CO_2)_2$ is dissolved in 1.0 L, $[CH_3CO_2^-]_{initial} = 2.000$ M.

As $CH_3CO_2^-$ is a weak base, $[CH_3CO_2^-]$ must be calculated by considering the equilibrium:

	$CH_3CO_2^-$	H_2O	⇌	CH_3CO_2H	OH^-
initial	**2.000**	**large**		**0**	**0**
change	**-*x***	**negligible**		**+*x***	**+*x***
final	**2.000 – *x***	**large**		***x***	***x***

The equilibrium constant K_b is given by:

$$K_b = \frac{[CH_3CO_2H][OH^-]}{[CH_3CO_2^-]} = \frac{x^2}{(2.000 - x)}$$

For an acid and its conjugate base:

pK_a + pK_b = 14.00 so pK_b = 14.00 – 4.76 = 9.24

As pK_b = 9.24, $K_b = 10^{-9.24}$. K_b is very small so 2.000 – *x* ~ 2.000 and hence:

$x^2 = 2.000 \times 10^{-9.24}$ or $y = 0.0000393$ M = $[OH^-]$

Hence, the pOH is given by:

pOH = $-\log_{10}[OH^-] = \log_{10}[0.0000393] = 4.47$

Finally, pH + pOH = 14.00 so

pH = 14.000 – 4.47 = 9.53

pH = **9.53**

ANSWER CONTINUES ON THE NEXT PAGE

Calculate the pH of this solution after the addition of 0.250 mol of HCl gas?

The solution contain 2.000 mol of $CH_3CO_2^-$. This will react with 0.250 mol of HCl(g) to produce 0.250 mol of CH_3CO_2H, leaving (2.000 – 0.250) mol = 1.750 mol of $CH_3CO_2^-$ in unreacted.

As the solution has a volume of 1.000 L, $[CH_3CO_2H] = 0.250$ M and $[CH_3CO_2^-] = 1.750$ M.

The solution contains a weak acid and its conjugate base. The pH of this buffer solution can be calculated using the Henderson-Hasselbalch equation;

$$\mathbf{pH = pK_a + \log\frac{[base]}{[acid]} = 4.76 + \log\frac{1.750}{0.250} = 5.61}$$

pH = **5.61**