Marks 8

- Solution A consists of a 0.020 M aqueous solution of propionic acid, $C_3H_6O_2$, at 25 °C. Calculate the pH of Solution A. The pK_a of propionic acid is 4.87.

As $C_3H_6O_2$ is a weak acid, $[H^+]$ must be calculated by considering the equilibrium:

	$C_3H_6O_2$	$\rightleftharpoons$	$C_3H_5O_2^-$	H^+
initial	**0.020**		**0**	**0**
change	$-x$		$+x$	$+x$
final	$0.020 - x$		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[C_3H_5O_2^-][H^+]}{[C_3H_6O_2]} = \frac{x^2}{(0.020 - x)}$$

As pK_a = 4.87, $K_a = 10^{-4.87}$. K_a is very small so $0.020 - x \sim 0.020$ and hence:

$$x^2 = 0.020 \times 10^{-4.87} \quad \text{or} \quad x = 0.000519 \text{ M} = [H^+]$$

Hence, the pH is given by:

$$pH = -\log_{10}[H^+] = -\log_{10}[0.000519] = 3.28$$

Answer: **pH = 3.28**

At 25 °C, 1.00 L of Solution B consists of 2.24 g of potassium propionate ($KC_3H_5O_2$) dissolved in water. Calculate the pH of Solution B.

The molar mass of $KC_3H_5O_2$ is:

molar mass = (39.10 (K) + 3×12.01 (C) + 5×1.008 (H) + 2×16.00 (O)) g mol^{-1}
= 112.17 g mol-1

Thus, 2.24 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{2.24 \text{ g}}{112.17 \text{ g mol}^{-1}} = 0.0200 \text{ mol}$$

If this is dissolved in 1.0 L, $[C_3H_5O_2^-]_{initial}$ = 0.0200 M.

As $C_3H_5O_2^-$ is a weak base, $[C_3H_5O_2^-]$ must be calculated by considering the equilibrium:

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	$C_3H_5O_2^-$	H_2O	$\rightleftharpoons$	$C_3H_6O_2$	OH^-
initial	**0.0200**	**large**		**0**	**0**
change	**-*y***	**negligible**		**+*y***	**+*y***
final	**0.0200 – *y***	**large**		***y***	***y***

The equilibrium constant K_b is given by:

$$K_b = \frac{[C_3H_6O_2][OH^-]}{[C_3H_5O_2^-]} = \frac{y^2}{(0.0200 - y)}$$

For an acid and its conjugate base:

$pK_a + pK_b = 14.00$

$pK_b = 14.00 - 4.87 = 9.13$

As $pK_b = 9.13$, $K_b = 10^{--9.13}$. K_b is very small so $0.0200 - y \sim 0.0200$ and hence:

$y^2 = 0.0200 \times 10^{-9.13}$ **or** $y = 0.000000385$ M $= [OH^-]$

Hence, the pOH is given by:

$pOH = -\log_{10}[OH^-] = \log_{10}[0.000000385] = 5.41$

Finally, pH + pOH = 14.00 so

$pH = 14.00 - 5.41 = 8.59$

Answer: **pH = 8.59**

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

Combining the two solutions will double the overall volume, to 2.00 L. As a result, the concentration of both the acid and base will halve: [acid] = 0.010 M and [base] = 0.0100 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:

$$pH = pK_a + \log\frac{[base]}{[acid]} = 4.87 + \log\frac{0.0100}{0.010} = 4.87$$

Answer: **pH = 4.87**

If you wanted to adjust the pH of Solution C to be exactly equal to 5.00, which component in the mixture would you need to increase in concentration?

More base is needed: add $KC_3H_5O_2$