Marks 7

- Solution A consists of a 0.020 M aqueous solution of aspirin (acetylsalicylic acid, $C_9H_8O_4$) at 25 °C. Calculate the pH of Solution A. The pK_a of aspirin is 3.52.

As $C_9H_8O_4$ is a weak acid, $[H^+]$ must be calculated by considering the equilibrium:

	$C_9H_8O_4$	$\rightleftharpoons$	$C_9H_7O_4^-$	H^+
initial	**0.020**		**0**	**0**
change	$-x$		$+x$	$+x$
final	$0.020 - x$		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[C_9H_7O_4^-][H^+]}{[C_9H_7O_{24}]} = \frac{x^2}{(0.020 - x)}$$

As pK_a = 3.52, $K_a = 10^{-3.52}$. K_a is very small so $0.020 - x \sim 0.020$ and hence:

$x^2 = 0.020 \times 10^{-3.52}$ **or** $x = 0.00246$ **M** $= [H^+]$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[H^+] = -\log_{10}[0.00246] = 2.61$$

Answer: **2.61**

At 25 °C, 1.00 L of Solution B consists of 4.04 g of sodium acetylsalicylate ($NaC_9H_7O_4$) dissolved in water. Calculate the pH of Solution B.

The molar mass of $NaC_9H_7O_4$ is:

molar mass = (22.99 (Na) + 9×12.01 (C) + 7×1.008 (H) + 4×16.00 (O)) g mol^{-1}
= 202.136 g mol-1

Thus, 4.04 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{4.04 \text{ g}}{202.136 \text{ g mol}^{-1}} = 0.0200 \text{ mol}$$

If this is dissolved in 1.0 L, $[C_9H_7O_4^-]_{initial}$ = 0.0200 M.

As $C_9H_7O_4^-$ is a weak base, $[C_9H_7O_4^-]$ must be calculated by considering the equilibrium:

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	$C_9H_7O_4^-$	H_2O	$\rightleftharpoons$	$C_9H_8O_4$	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:

$$K_b = \frac{[C_9H_8O_4][OH^-]}{[C_9H_7O_4^-]} = \frac{y^2}{(0.0200-y)}$$

For an acid and its conjugate base:

$pK_a + pK_b = 14.00$

$pK_b = 14.00 - 3.52 = 10.48$

As $pK_b = 10.48$, $K_b = 10^{-10.48}$. K_b is very small so $0.0200 - y \sim 0.0200$ and hence:

$y^2 = 0.0200 \times 10^{-10.48}$ **or** $y = 0.000000814$ **M** $= [OH^-]$

Hence, the pOH is given by:

$pOH = -\log_{10}[OH^-] = \log_{10}[0.000000814] = 6.09$

Finally, pH + pOH = 14.00 so
pH = 14.00 – 6.09 = 7.91

Answer: **7.91**

Solution B (200.0 mL) is mixed with Solution A (400.0 mL) and water (200.0 mL) to give Solution C. Calculate the pH of Solution C after equilibration at 25 °C.

400.0 mL of solution A (the acid) contains:

number of moles = concentration × volume $= (0.0200 \text{ mol L}^{-1}) \times (0.4000 \text{ L}) = 0.00800$ **mol**

200.0 mL of solution B (the base) contains:

number of moles = concentration × volume $= (0.0200 \text{ mol L}^{-1}) \times (0.2000 \text{ L}) = 0.00400$ **mol**

The final solution has a total volume of (200.0 + 400.0 + 200.0) mL = 800.0 mL.

The concentrations of acid and base in the final solution are:

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$$\textbf{concentration of acid} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.00800 \text{ mol}}{0.8000 \text{ L}} = 0.0100 \text{ M}$$

$$\textbf{concentration of base} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.00400 \text{ mol}}{0.8000 \text{ L}} = 0.00500 \text{ M}$$

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

$$\text{pH} = \text{p}K_a + \log\frac{[\text{base}]}{[\text{acid}]} = 3.52 + \log\frac{0.00500}{0.0100} = 3.22$$

Answer: **3.22**

If you wanted to adjust the pH of Solution C to be exactly equal to 3.00, which component in the mixture would you need to increase in concentration?

To lower the pH, more acid is required: solution A