Marks 8

- Solution A consists of a 0.15 M aqueous solution of nitrous acid (HNO_2) at 25 °C. Calculate the pH of Solution A. The pK_a of HNO_2 is 3.15.

Nitrous acid is a weak acid so $[H_3O^+]$ must be calculated:

	HNO_2	H_2O	$\rightleftharpoons$	H_3O^+	NO_2^-
initial	**0.15**	**large**		**0**	**0**
change	**-x**	**negligible**		**+x**	**+x**
final	**0.15 – x**	**large**		x	x

The equilibrium constant K_a is given by:

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{x^2}{0.15 - x}$$

As $K_a = 10^{-3.15}$ is very small, $0.15 - x \sim 0.15$ and hence:

$$x^2 = 0.15 \times 10^{-3.15} \quad \text{or} \quad x = 1.03 \times 10^{-2}\ \text{M} = [H_3O^+]$$

Hence, the pH is given by:

$$\text{pH} = -\log_{10}[H_3O^+(aq)] = -\log_{10}[(1.03 \times 10^{-2})] = 1.99$$

pH = **1.99**

ANSWER CONTINUES ON THE NEXT PAGE

At 25 °C, 1.00 L of Solution B consists of 13.8 g of sodium nitrite ($NaNO_2$) dissolved in water. Calculate the pH of Solution B.

The formula mass of $NaNO_2$ is (22.99 (Na) + 14.01 (N) + 2 × 16.00 (O)) = 69. Therefore, 13.8 g corresponds to:

$$\text{number of moles of } NaNO_2 = \frac{\text{mass}}{\text{formula mass}} = \frac{13.8}{69.0} = 0.200\,\text{mol}$$

As this is dissolved in 1.00 L, the concentration is 0.200 M.

NO_2^- is a weak base so $[OH^-(aq)]$ must be calculated from the equilibrium:

	NO_2^-	H_2O	$\rightleftharpoons$	OH^-	HNO_2
initial	0.200	large		0	0
change	$-y$	negligible		$+y$	$+y$
final	$0.200 - y$	large		y	y

The equilibrium constant K_b is given by:

$$K_a = \frac{[OH^-][HNO_2]}{[NO_2^-]} = \frac{y^2}{0.2 - y}$$

For an acid and its conjugate base, $pK_a + pK_b = 14.00$ so:

$$pK_b = 14.00 - 3.15 = 10.85$$

As $pK_b = 10.85$, $K_b = 10^{-10.85}$. K_b is very small so $0.200 - y \sim 0.200$ and hence:

$$y^2 = 0.200 \times 10^{-10.85} \quad \text{or} \quad y = 1.68 \times 10^{-6}\ \text{M} = [OH^-(aq)]$$

Hence, the pOH is given by pOH = $-\log_{10}[OH^-(aq)] = -\log_{10}[1.68 \times 10^{-6}] = 5.77$

Finally, pH + pOH = 14 so pH = (14.00 – 5.77) = 8.23

pH = **8.23**

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.

Solution A, HNO_2(aq), is 0.15 M. When 1.00 L of A is added to 1.00 L of B, a 2.00 L solution is formed. This dilution halves the concentration to 0.075 M. Similarly, solution B, KNO_2(aq), which is initially 0.200 M is diluted to 0.100 M by the addition of solution A.

The combined solution contains a mixture of a weak acid (HNO_2) and its conjugate base (NO_2^-) so acts as a buffer and the Henderson-Hasselbalch equation can be used:

$$\mathrm{pH} = \mathrm{p}K_a + \log_{10}\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$ **with [base] = [NO_2^-(aq)] and [acid] = [HNO_2(aq)].**

As pK_a for HNO_2 = 3.15, [NO_2^-(aq)] = 1.00 M and [HNO_2(aq)] = 0.075 M:

$$\mathrm{pH} = 3.15 + \log_{10}\left(\frac{[0.100]}{[0.075]}\right) = 3.27$$

pH = **3.27**

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

acid (HNO_2)