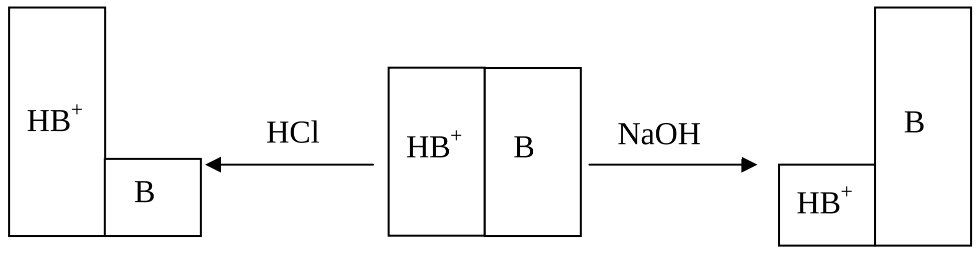
Marks 7

- A buffer system with a weak base B and its conjugate acid HB^+ is shown in the diagram below with equal concentrations. Complete the diagram by showing the relative concentrations after the addition of some HCl or NaOH.



Write down the balanced net ionic equations for both these reactions.

$HB^+(aq) + OH^-(aq) \rightarrow B + H_2O(l)$
$B(aq) + H^+(aq) \rightarrow HB^+(aq)$

Calculate the pH of a buffer if it contains 0.200 mol of $NaNO_2$ and 0.300 mol of HNO_2 in 1.00 L of water. The pK_a of HNO_2 is 3.15.

The concentrations of acid (HNO_2) and base (NO_2^-) are:

concentration of acid = number of moles / volume
= (0.300 mol / 1.00 L) = 0.300 mol L^{-1}

concentration of base = (0.200 mol / 1.00 L) = 0.200 mol L^{-1}

The pH of this buffer can then be calculated using the Henderson-Hasselbalch equation:

$$pH = pK_a + \log\frac{[\text{base}]}{[\text{acid}]} = 3.15 + \log\frac{0.200}{0.300} = 2.97$$

pH = **2.97**

What is the pH if (a) 0.05 mol of HCl(g) and (b) 0.25 mol of HCl(g) is added?

HCl will react with the NO_2^- to form HNO_2.

(a) 0.05 mol of HCl(g) will react with 0.05 mol of the NO_2^- present to form an additional 0.05 mol of HNO_2. The amount of NO_2^- will decrease by 0.05 mol:

concentration of acid = (0.35 mol / 1.00 L) = 0.35 mol L^{-1}

concentration of base = (0.15 mol / 1.00 L) = 0.15 mol L^{-1}

ANSWER CONTINUES ON THE NEXT PAGE

Hence:

$$\mathbf{pH = p}\boldsymbol{K}_{\mathbf{a}}\mathbf{ + log}\frac{[\mathbf{base}]}{[\mathbf{acid}]}\mathbf{ = 3.15 + log}\frac{\mathbf{0.15}}{\mathbf{0.35}}\mathbf{ = 2.78}$$

(b) 0.25 mol of HCl(g) will react *all* of the NO_2^-. As there is only 0.200 mol of NO_2^- present, 0.05 mol of HCl will remain unreacted. As HCl is a strong acid, this will completely ionize to give $[H^+(aq)]$ = 0.05 mol L^{-1}. Hence:

$$\mathbf{pH = -log_{10}[H^+(aq)] = -log_{10}(0.05) = 1.30}$$

(a) pH = **2.78**	(b) pH = **1.30**