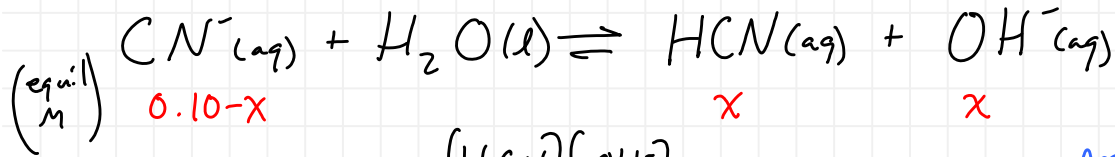
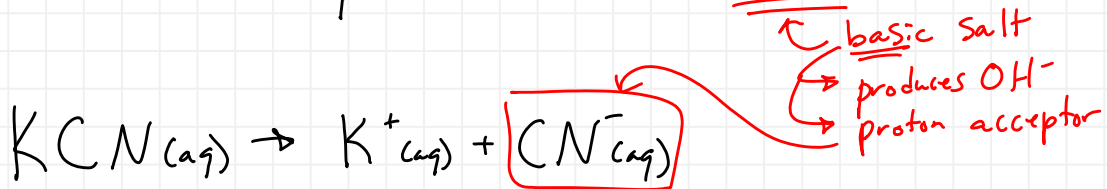


Find the pH of 0.10 M KCN.



$$\frac{K_w}{K_a} = K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]}$$

Assume
x << 0.10 *Assumption*
valid

$$\frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.04 \times 10^{-5} = \frac{x^2}{0.10-x}$$

$$\sqrt{2.04 \times 10^{-6}} = \sqrt{x^2} = x = 0.00143 \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.00143) = 2.85$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 2.85 = \boxed{11.15}$$

Find the pH of 0.10 M NaCl.

$$\boxed{\text{pH} = 7.00}$$

Find the pH of 0.10 M HNO₃.

$$[\text{H}_3\text{O}^+] = 0.10 \text{ M} \quad -\log(0.10) = \boxed{1.00}$$

strong acid

Common Ion Effect

Find the pH of a solution that is

both ~~0.20 M HC₂H₃O₂~~ & 0.10 M HCl.

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.7 \times 10^{-5} = \frac{(x)(0.10 - x)}{0.20 + x}$$

Assume
 $x \ll 0.10$

$$\frac{1.7 \times 10^{-5}}{0.50} = \frac{0.50x}{0.50} = x = [H_3O^+] = 3.4 \times 10^{-5} M$$

$$pH = -\log[H_3O^+] = -\log(3.4 \times 10^{-5}) = \boxed{4.47}$$

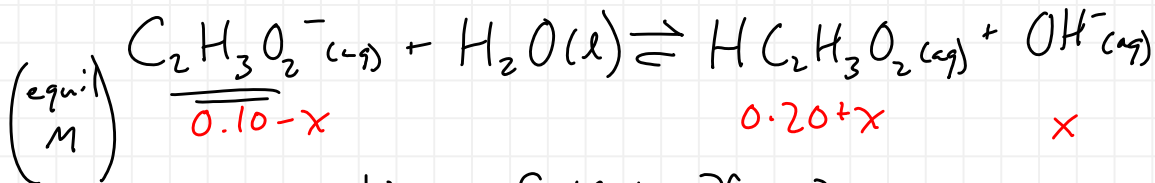
Find the pH of a solution that is both

0.10 M NaC₂H₃O₂ & 0.20 M HC₂H₃O₂.

basic salt

conjugate

weak acid



$$\frac{K_w}{K_a} = K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

Assume
 $x \ll 0.10$

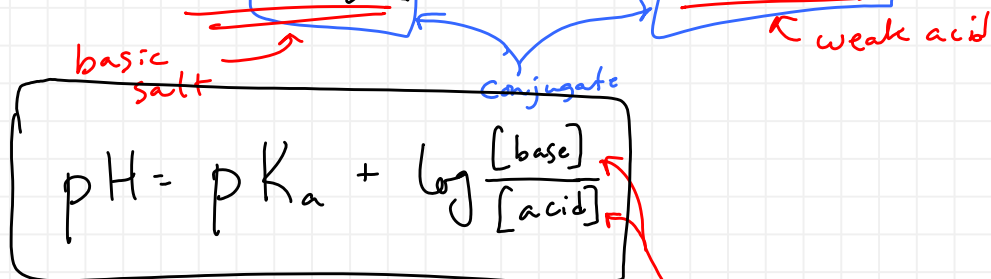
$$\frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} = 5.88 \times 10^{-10} = \frac{(0.20 + x)(x)}{(0.10 - x)}$$

$$\frac{5.88 \times 10^{-10}}{2.0} = \frac{2.0x}{2.0} = 2.94 \times 10^{-10} = [OH^-]$$

$$pOH = -\log[OH^-] = -\log(2.94 \times 10^{-10}) = 9.53$$

$$pH = 14.00 - pOH = 14.00 - 9.53 = \boxed{4.47}$$

Find the pH of a solution that is both
 0.10M $\text{NaC}_2\text{H}_3\text{O}_2$ & 0.20M $\text{HC}_2\text{H}_3\text{O}_2$.



$$\text{pH} = -\log(1.7 \times 10^{-5}) + \log\left(\frac{0.10}{0.20}\right) = 4.47$$

Find the ratio of $[\text{NH}_3]$ to $[\text{NH}_4^+]$ in a
 buffer solution that has a pH = 10.00.
 \rightarrow pOH = 4.00

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{A}]}{[\text{B}]}$$

$$4.00 = -\log(1.8 \times 10^{-5}) + \log \frac{[\text{A}]}{[\text{B}]}$$

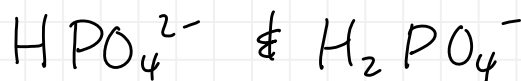
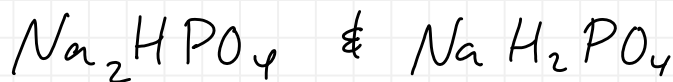
$$4.00 = 4.745 + \log \frac{[\text{A}]}{[\text{B}]}$$

$$-0.745 = \log \frac{[\text{A}]}{[\text{B}]}$$

10

10

$$0.18 = \frac{[\text{A}]}{[\text{B}]} \Rightarrow \frac{[\text{B}]}{[\text{A}]} = 5.6 = \frac{[\text{NH}_3]}{[\text{NH}_4^+]}$$



\leftarrow conjugate pairs \rightarrow

What is the pH of a solution prepared by mixing 65.0 mL of 0.10 M NH_3 with 25.0 mL of 0.20 M NH_4Cl ?

$$\textcircled{\text{NH}_3} \quad \frac{C_1 V_1}{V_2} = \frac{C_2 \cancel{V_2}}{\cancel{V_2}} = \frac{(0.10 \text{ M})(65.0 \text{ mL})}{(90.0 \text{ mL})} = 0.0722 \text{ M NH}_3$$

$$\textcircled{\text{NH}_4^+} \quad \frac{C_1 V_1}{V_2} = C_2 = \frac{(0.20 \text{ M})(25.0 \text{ mL})}{(90.0 \text{ mL})} = 0.0556 \text{ M NH}_4^+$$

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{A}]}{[\text{B}]}$$

$$= -\log(1.8 \times 10^{-5}) + \log \left(\frac{0.0556 \text{ M}}{0.0722 \text{ M}} \right) = 4.63$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 4.63 = \boxed{9.37}$$

What is the pH of this buffer solution after the addition of 5.0 mL of 0.15 M HCl ?
 90.0 mL of 0.0722 M NH_3 & 0.0556 M NH_4^+
 0.15 M H_3O^+ (strong acid, produce H_3O^+)
 react with each other

$$\text{H}_3\text{O}^+(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l})$$

(mol)	$\text{H}_3\text{O}^+(\text{aq})$	$\text{NH}_3(\text{aq})$	$\text{NH}_4^+(\text{aq})$	$\text{H}_2\text{O}(\text{l})$
Init	0.00075	0.00650	0.00500	
Change	-0.00075	-0.00075	+0.00075	
End	0	0.00575	0.00575	

$$\text{H}_3\text{O}^+ \quad 5.0 \text{ mL} \times \frac{0.15 \text{ mol H}_3\text{O}^+}{1000 \text{ mL}} = 0.00075 \text{ mol H}_3\text{O}^+ \text{ initially}$$

$$\text{NH}_3 \quad 90.0 \text{ mL} \times \frac{0.0722 \text{ mol NH}_3}{1000 \text{ mL}} = 0.00650 \text{ mol NH}_3 \text{ initially}$$

$$\text{NH}_4^+ \quad 90.0 \text{ mL} \times \frac{0.0556 \text{ mol NH}_4^+}{1000 \text{ mL}} = 0.00500 \text{ mol NH}_4^+ \text{ initially}$$

$$\text{pOH} = \text{p}K_b + \log \left(\frac{[\text{A}]}{[\text{B}]} \right) \quad \leftarrow \begin{array}{l} \text{mole} \\ \text{ratio} = \text{molarity} \\ \text{ratio} \end{array}$$

$$= -\log(1.8 \times 10^{-5}) + \log \left(\frac{0.00575}{0.00575} \right)$$

$$\text{pOH} = 4.74 + 0$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 4.74 = \boxed{9.26}$$