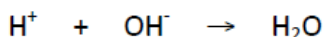


IB Chemistry II Study Worksheet 18.2 KEY

- (i) $\text{NH}_3 + \text{NH}_4\text{Cl}$ (weak base + salt of weak base and strong acid) ; (1)
(or $\text{HCl} + \text{NH}_3$ strong acid + excess weak base)

- (ii) pH changes very little / most of the acid is neutralized by the base ; (1)
equation ; (1)

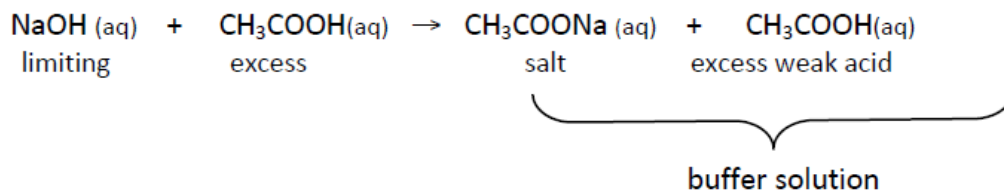


If a small amount of H^+ from a strong acid is added they combine with OH^- ions to form water, decreasing the concentration of OH^- ions causing the pH to increase. Alternatively, according to Le Chateliers Principle, the position of equilibrium will shift to the right to replace the OH^- ions that have been used up causing the pH to remain unchanged.

- 1.
2. B
3.
 - a) A **buffer solution** is a solution which will resist changes in pH ; (1)
when a small amount of a strong acid or base is added ; (1)
(do not accept pH does not change)
 - b)
 - (i) not a buffer solution ; (1)
after reaction the mixture contains - 0.10 mol NH_4Cl and 0.10 mol HCl / NH_3 needs to be in excess to make a basic buffer; (1)
 - (ii) buffer solution ; (1)
after reaction the mixture contains - 0.10 mol NH_3 and 0.10 mol NH_4Cl ; (1)
4. B
If you add more salt (conjugate base) the buffer solution becomes more basic & pH will become closer to 14. If you add more acid the buffer solution becomes more acidic, the pH will become closer to 1.
5.

Fewer moles of NaOH than CH_3COOH ; (1)

CH_3COOH should be in excess so the final solution contains both CH_3COO^- and CH_3COOH ; (1)



6.

Calculate the pH of a mixture of 50cm³ of ammonia solution of concentration 0.10 mol dm⁻³ and 50cm³ of hydrochloric acid solution of concentration 0.050 mol dm⁻³.
pK_b (NH₃) = 4.75

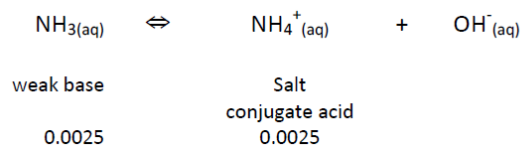
$$\begin{aligned}n(\text{HCl}) &= c \times v \\&= 0.050 \times (50 / 1000) \\&= 0.0025 \text{ mol} \\&= 0.0025 \text{ mol dm}^{-3}\end{aligned}$$

$$\begin{aligned}n(\text{NH}_3) &= c \times v \\&= 0.10 \times (50 / 1000) \\&= 0.0050 \text{ mol} \\&= 0.0050 \text{ mol dm}^{-3} ;\end{aligned}$$

NH₃ is in excess

$$n(\text{NH}_3) \text{ remaining} = 0.0050 - 0.0025 = 0.0025 \text{ mol} = 0.0025 \text{ mol dm}^{-3}$$

From mole ratio in the equation assume [NH₄⁺] = [NH₃] = 0.0025 mol dm⁻³ because they are in a 1:1 mole ratio ;



$$K_b = 10^{-\text{pK}_b} = 10^{-4.75} = 1.78 \times 10^{-5}$$

$$K_b = \frac{[\text{NH}_4^+] \times [\text{OH}^-]}{[\text{NH}_3]}$$

$$[\text{OH}^-] = \frac{K_b \times [\text{NH}_3]}{[\text{NH}_4^+]}$$

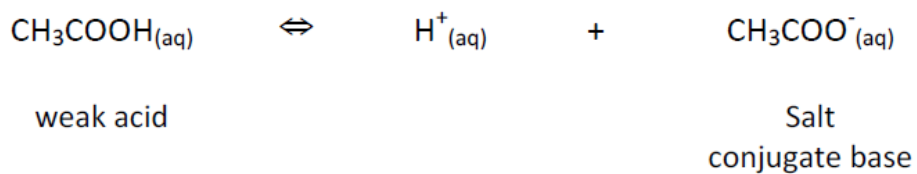
$$[\text{OH}^-] = \frac{1.78 \times 10^{-5} \times 0.0025}{0.0025}$$

$$[\text{OH}^-] = 1.78 \times 10^{-5} \text{ mol dm}^{-3} ;$$

$$\text{pOH} = 4.75$$

$$\text{pH} = 9.25 \text{ (allow 9.2 to 9.3) ;}$$

7.



$$K_a = \frac{[\text{H}^+] \times [\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

$$[\text{CH}_3\text{COOH}] = 0.0500 \text{ mol dm}^{-3}$$

$$[\text{CH}_3\text{COO}^-] = 0.100 \text{ mol dm}^{-3}$$

$$K_a = 1.74 \times 10^{-5}$$

$$[\text{H}^+] = \frac{K_a \times \text{CH}_3\text{COOH}}{\text{CH}_3\text{COO}^-}$$

$$[\text{H}^+] = \frac{1.74 \times 10^{-5} \times 0.0500}{0.100}$$

$$[\text{H}^+] = 8.70 \times 10^{-6} \text{ mol dm}^{-3}$$

$$\begin{aligned} \text{pH} &= -\log_{10} [\text{H}^+] \\ &= -\log_{10} 8.70 \times 10^{-6} \\ &= 5.06 ; \text{ (3SF)} \end{aligned}$$

(accept answer in the range 5.0 to 5.1)

- 8. A. Buffer solution: $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$

➔ If an acid is added, the H^+ will bond with CH_3COO^- , thus preventing additional “free” H^+ ions from decreasing the pH of the solution.

B. (Similar to example #2 from PowerPoint) pH = 4.46