IB Chemistry II Study Worksheet 18.2 KEY

(i) NH_3 + NH_4Cl (weak base + salt of weak base and strong acid) ; (1)
(or HCl + NH_3 strong acid + excess weak base)

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

$NH_3(aq)$	+	H_2O	$\leftrightarrow$	$NH_4^+(aq)$	+	$OH^-(aq)$
Weak base		acid	$\longleftarrow$	Strong conj acid		Conj base

$$H^+ + OH^- \rightarrow H_2O$$

If a <u>small amount</u> 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 Chataliers 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 <u>small amount</u> 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 / NH3 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)

$$\underset{\text{limiting}}{NaOH\,(aq)} + \underset{\text{excess}}{CH_3COOH(aq)} \rightarrow \underbrace{\underset{\text{salt}}{CH_3COONa\,(aq)} + \underset{\text{excess weak acid}}{CH_3COOH(aq)}}_{\text{buffer solution}}$$

.

6.

Calculate the pH of a mixture of 50cm^3 of ammonia solution of concentration 0.10 moldm^{-3} and 50cm^3 of hydrochloric acid solution of concentration 0.050 mol dm^{-3}.
pK_b (NH_3) = 4.75

n (HCl) = c x v
= 0.050 x (50 / 1000)
= 0.0025 mol
= 0.0025 mol dm^{-3}

n(NH_3)= c x v
= 0.10 x (50 /1000)
= 0.0050 mol
= 0.0050 mol dm^{-3} ;

NH_3 is in excess

n(NH_3) remaining = 0.0050 – 0.0025 = 0.0025 mol = 0.0025 mol dm^{-3}

From mole ratio in the equation assume [NH_4^+] = [NH_3] = 0.0025 mol dm^{-3} because they are in a 1:1 mole ratio ;

$$NH_{3(aq)} \Leftrightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

$NH_{3(aq)}$		$NH_4^+{}_{(aq)}$
weak base		Salt conjugate acid
0.0025		0.0025

K_b = 10^{-pKb} = $10^{-4.75}$ = 1.78×10^{-5}

$$K_b = \frac{[NH_4^+] \times [OH^-]}{[NH_3]}$$

$$[OH^-] = \frac{K_b \times [NH_3]}{[NH_4^+]}$$

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

$[OH^-]$ = 1.78×10^{-5} moldm^{-3} ;

pOH = 4.75

pH = 9.25 (allow 9.2 to 9.3) ;

7.

$$CH_3COOH_{(aq)} \Leftrightarrow H^+_{(aq)} + CH_3COO^-_{(aq)}$$

weak acid

Salt
conjugate base

$$K_a = \frac{[H^+] \times [CH_3COO^-]}{[CH_3COOH]}$$

$[CH_3COOH] = 0.0500 \text{ mol dm}^{-3}$

$[CH_3COO^-] = 0.100 \text{ mol dm}^{-3}$

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

$$[H+] = \frac{K_a \times CH_3COOH}{CH_3COO^-} ;$$

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

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

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

$$= -\log_{10} 8.70 \times 10^{-6}$$

$$= 5.06 \; ; \; (3SF)$$

(accept answer in the range 5.0 to 5.1)

- 8. A. Buffer solution: $CH_3COOH \leftrightarrow CH_3COO- + H^+$

➔ 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