

**CHEM 1412**  
**Chapter 19 Homework Answers**

1. Which combination of chemical has the potential of making a good buffer solution? Explain your decision in each case.

(a) NaOH and KOH

This is not buffer material because it includes two strong bases.

(b) CH<sub>3</sub>NH<sub>2</sub> and CH<sub>3</sub>NH<sub>3</sub>Cl

Yes, this could be a buffer. It includes a weak base and the salt of that weak base.

(c) NaH<sub>2</sub>PO<sub>4</sub> and Na<sub>2</sub>HPO<sub>4</sub>

Yes, this could be a buffer. There are two ways we could interpret the contents:  
weak acid (NaH<sub>2</sub>PO<sub>4</sub>) and salt/conjugate base of that weak acid (Na<sub>2</sub>HPO<sub>4</sub>)  
or weak base (Na<sub>2</sub>HPO<sub>4</sub>) and the salt/conjugate acid of that weak base (NaH<sub>2</sub>PO<sub>4</sub>)

(d) HBr and NaBr

Not buffer material - buffers cannot be made from strong acids.

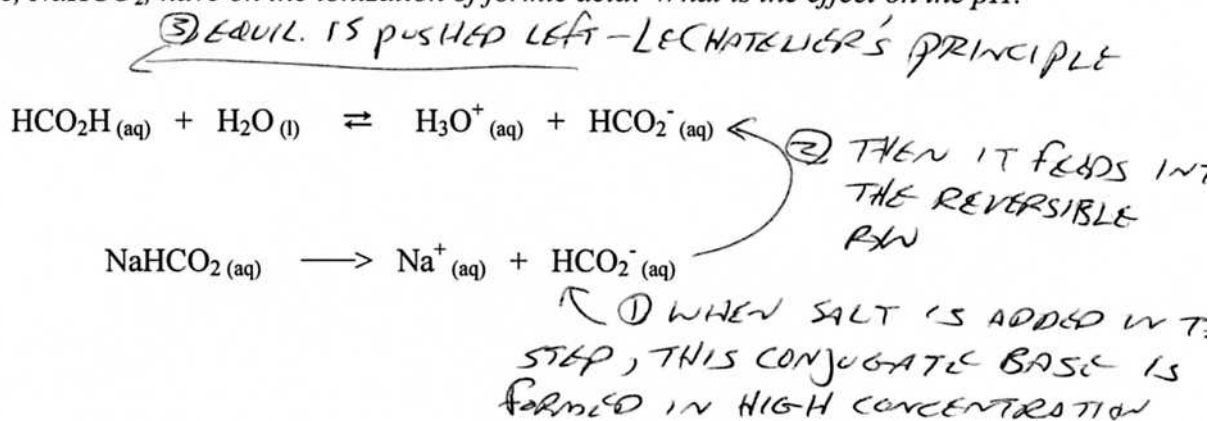
(e) HCN and KCl

Not buffer material - there is no common ion.

(f) NaCH<sub>3</sub>CO<sub>2</sub> and CH<sub>3</sub>CO<sub>2</sub>H

Yes, this could be a buffer. There is a weak acid and the salt of that weak acid.

2. Write the equation for ionization of formic acid, HCO<sub>2</sub>H. What effect does the addition of sodium formate, NaHCO<sub>2</sub>, have on the ionization of formic acid? What is the effect on the pH?



Because the equilibrium in the first reaction is pushed to the left, the [H<sub>3</sub>O<sup>+</sup>] decreases. That is, the solution becomes *less* acidic, which means that pH of the solution increases.

3. Calculate the pH for a buffer solution containing 0.025 M  $\text{CH}_3\text{CO}_2\text{H}$  and 0.025 M  $\text{Ba}(\text{CH}_3\text{CO}_2)_2$ .

There's a bit of a trick here. Notice, the salt in this buffer comes from a diprotic base.

	$\text{Ba}(\text{CH}_3\text{CO}_2)_2$ (aq)	$\longrightarrow$	$\text{Ba}^{2+}$ (aq)	$+$	$2 \text{CH}_3\text{CO}_2^-$ (aq)
init.	0.025		0		0
$\Delta$	<u>-0.025</u>		<u>+0.025</u>		<u>+2(0.025)</u>
final	0		0.025		0.050

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]}\right)$$

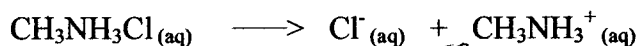
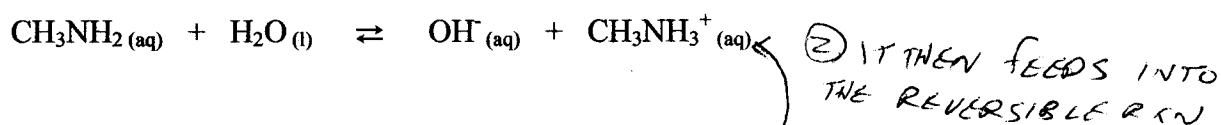
$$\text{pH} = -\log(1.8 \times 10^{-5}) + \log\left(\frac{0.050}{0.025}\right)$$

$$\text{pH} = 4.74 + 0.30$$

$$\boxed{\text{pH} = 5.04}$$

4. Write the equation for ionization of methylamine,  $\text{CH}_3\text{NH}_2$ . What effect does the addition of methyl ammonium chloride,  $\text{CH}_3\text{NH}_3\text{Cl}$ , have on the ionization of methylamine? What is the effect on the pH?

③ EQUIL. IS PUSHED LEFT - LE CHATELIER'S PRINCIPLE



② IT THEN FEEDS INTO THE REVERSIBLE RXN  
 WHEN SALT IS ADDED IN THIS STEP, THIS CONJUGATE ACID IS FORMED IN HIGH CONCENTRATION

Because the equilibrium in the first reaction is pushed to the left, the  $[\text{OH}^-]$  decreases. That is, the solution becomes *less basic* or more acidic. The pH of the solution decreases.

5. Calculate the pH for a buffer solution containing 0.15 M  $\text{NH}_3$  and 0.20 M  $(\text{NH}_4)_2\text{SO}_4$ .

The same trick again. Notice, the salt in this buffer comes from a diprotic acid.

	$(\text{NH}_4)_2\text{SO}_4$ (aq)	$\longrightarrow$	$2 \text{NH}_4^+$ (aq)	$+$	$\text{SO}_4^{2-}$ (aq)
init.	0.20		0		0
$\Delta$	<u>-0.20</u>		<u>+2(0.20)</u>		<u>+0.20</u>
final	0		0.40		0.20

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_3]}$$

$$\text{pOH} = -\log(1.8 \times 10^{-5}) + \log \frac{0.40}{0.15}$$

$$\text{pOH} = 4.74 + 0.42 = 5.16$$

$$\boxed{\text{pH} = 8.84}$$

6. A buffer solution of pH 5.30 is to be prepared from propionic acid,  $\text{C}_2\text{H}_5\text{CO}_2\text{H}$ , and sodium propionate,  $\text{NaC}_2\text{H}_5\text{CO}_2$ . The concentration of sodium propionate must be 0.50 M. What should be the concentration of the acid?  $K_a = 1.3 \times 10^{-5}$  for  $\text{C}_2\text{H}_5\text{CO}_2\text{H}$ .

$$\text{pH} = \text{pK}_a + \log \frac{[\text{NaC}_2\text{H}_5\text{CO}_2]}{[\text{C}_2\text{H}_5\text{CO}_2\text{H}]}$$

$$5.30 = -\log(1.3 \times 10^{-5}) + \log \left( \frac{0.50}{[\text{C}_2\text{H}_5\text{CO}_2\text{H}]} \right)$$

$$5.30 = 4.89 + \log \left( \frac{0.50}{[\text{C}_2\text{H}_5\text{CO}_2\text{H}]} \right)$$

$$0.41 = \log \left( \frac{0.50}{[\text{C}_2\text{H}_5\text{CO}_2\text{H}]} \right)$$

$$10^{0.41} = 2.57 = \left( \frac{0.50}{[\text{C}_2\text{H}_5\text{CO}_2\text{H}]} \right)$$

$$[\text{C}_2\text{H}_5\text{CO}_2\text{H}] = \frac{0.50}{2.57} = \boxed{0.19 \text{ M}}$$

7. How many grams of sodium benzoate,  $\text{NaC}_6\text{H}_5\text{CO}_2$ , must be added to make 2.0 L of a solution that is 0.045 M in benzoic acid,  $\text{C}_6\text{H}_5\text{CO}_2\text{H}$ , so that the pH is 5.00?

$$\text{pH} = \text{pK}_a + \log \left( \frac{[\text{NaC}_6\text{H}_5\text{CO}_2]}{0.045} \right)$$

$$5.00 = -\log(6.3 \times 10^{-5}) + \log \left( \frac{[\text{NaC}_6\text{H}_5\text{CO}_2]}{0.045} \right)$$

$$5.00 = 4.20 + \log \left( \frac{[\text{NaC}_6\text{H}_5\text{CO}_2]}{0.045} \right)$$

$$0.80 = \log \left( \frac{[\text{NaC}_6\text{H}_5\text{CO}_2]}{0.045} \right)$$

$$10^{0.80} = 6.31 = \left( \frac{[\text{NaC}_6\text{H}_5\text{CO}_2]}{0.045} \right)$$

$$[\text{NaC}_6\text{H}_5\text{CO}_2] = (6.31)(0.045 \text{ M}) = 0.28 \text{ M}$$

$$\frac{2.0 \text{ L}}{1 \text{ L}} \left| \frac{0.28 \text{ mol}}{1 \text{ mol}} \right| \frac{144 \text{ g}}{1 \text{ mol}} = \boxed{81 \text{ g NaC}_6\text{H}_5\text{CO}_2}$$

8. Which of the following solutions are effective buffers? Explain your decision in each case.

(a) 0.10 M HCN and 0.10 M NaCN

This would make a good buffer. Weak acid + conjugate base, both with reasonable concentrations, and concentrations that are similar/same.

(b) 0.10 M NaOH and 0.10 M NaCl

Not a good buffer combination. Buffers must be weak acid + conjugate base or weak base + conjugate acid. This is a strong base and a salt.

(c) 0.10 M NH<sub>3</sub> and 0.25 M NH<sub>4</sub>Br

This would make a good buffer. Weak base + conjugate acid, both with reasonable concentrations, and concentrations that are similar/same.

(d) 0.25 M NaF and 0.15 M HF

This would make a good buffer. Weak acid + conjugate base, both with reasonable concentrations, and concentrations that are similar/same.

(e)  $1.0 \times 10^{-5}$  M CH<sub>3</sub>CO<sub>2</sub>H and  $1.0 \times 10^{-5}$  M NaCH<sub>3</sub>CO<sub>2</sub>

This would not make a good buffer. The identity of the components is good, but the concentrations need to be higher; at least 0.01 M, or so.

(f) 0.10 M HCl and 0.22 M NaCl

This is not a buffer. Buffers cannot be made from strong acids.

(g) 0.15 M C<sub>5</sub>H<sub>5</sub>N and 0.0015 M C<sub>5</sub>H<sub>5</sub>NHCl

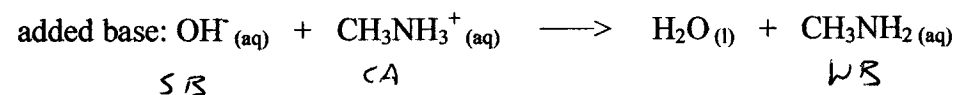
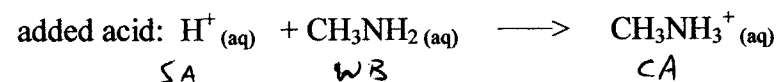
This is not an effective buffer. The ratio of concentrations of the components must be between 0.1 and 10. In this case,  $0.15/0.0015 = 100$ .

(h) 0.25 M HF and 0.20 M KCl

This is not a buffer. There is no common ion.

9. Briefly explain why the pH of a buffer solution of CH<sub>3</sub>NH<sub>2</sub>/CH<sub>3</sub>NH<sub>3</sub>Cl will remain nearly constant when either small amounts of HCl or NaOH are added. Use equations to illustrate your answer.

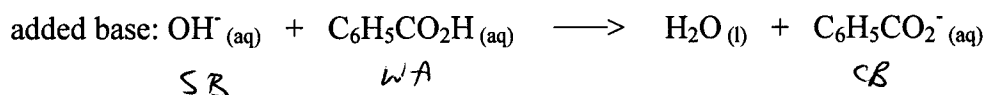
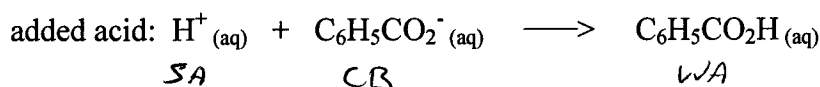
In either case, the buffer will react with and consume the added HCl or NaOH.



pH change is caused by the presence of H<sub>3</sub>O<sup>+</sup> or OH<sup>-</sup>. If the buffer consumes these species, then there can be no significant change in the pH

10. Briefly explain why the pH of a buffer solution of  $C_6H_5CO_2H/NaC_6H_5CO_2$  will remain nearly constant when either small amounts of HCl or NaOH are added. Use equations to illustrate your answer.

The explanation is the same as in Problem 9; however, the equations are a bit different.



11. Calculate the pH change that occurs when 0.050 mol HCl is added to 1.0 L of a buffer containing 0.45 M  $HCO_2H$  and 0.55 M  $NaHCO_2$ .

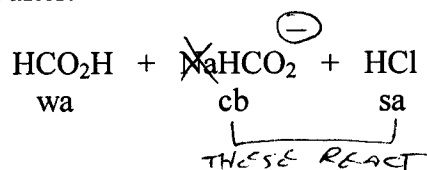
(a) pH before:

$$pH = -\log(1.8 \times 10^{-4}) + \log\left(\frac{0.55}{0.45}\right)$$

$$pH = 3.74 + 0.087$$

$$pH = 3.83$$

(b) pH after:



	$HCO_2^-_{(aq)}$	$+ H^+_{(aq)}$	$\longrightarrow$	$HCO_2H_{(aq)}$
init.	0.55	0.050		0.45
$\Delta$	<u>-0.050</u>	<u>-0.050</u>		<u>+0.050</u>
final	0.50	0		0.50

$$pH = -\log(1.8 \times 10^{-4}) + \log\left(\frac{0.50}{0.50}\right)$$

$$pH = 3.74 + 0 = 3.74$$

$$pH \text{ change } (\Delta pH) = 3.83 - 3.74 = 0.09$$

12. Calculate the pH change that occurs when 0.020 mol of NaOH is added to 1.0 L of a buffer containing 0.50 M HCO<sub>2</sub>H and 0.40 M NaHCO<sub>2</sub>.

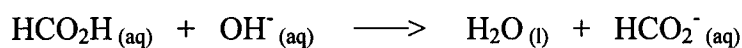
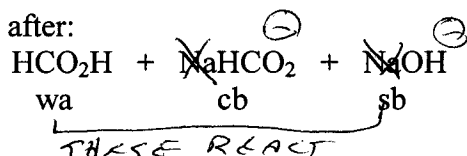
(a) pH before:

$$\text{pH} = -\log(1.8 \times 10^{-4}) + \log\left(\frac{0.40}{0.50}\right)$$

$$\text{pH} = 3.74 - 0.10$$

$$\boxed{\text{pH} = 3.64}$$

(b) pH after:



init.	0.50	0.020	---	0.40
Δ	-0.020	-0.020	-----	+0.020
final	0.48	0		0.42

$$\text{pH} = -\log(1.8 \times 10^{-4}) + \log\left(\frac{0.42}{0.48}\right)$$

$$\text{pH} = 3.74 - 0.058 = \boxed{3.68}$$

$$\Delta\text{pH} = 3.68 - 3.64 = \boxed{0.04}$$

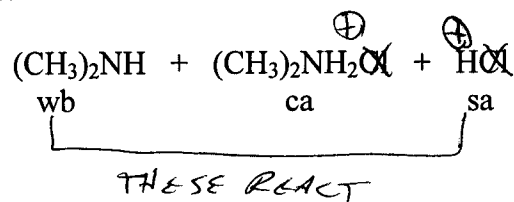
13. (a) Calculate the pH of a buffer containing 0.25 M (CH<sub>3</sub>)<sub>2</sub>NH and 0.15 M (CH<sub>3</sub>)<sub>2</sub>NH<sub>2</sub>Cl.

$$\text{pOH} = -\log(7.4 \times 10^{-4}) + \log\left(\frac{0.15}{0.25}\right)$$

$$\text{pOH} = 3.13 - 0.22 = 2.91$$

$$\boxed{\text{pH} = 11.09}$$

(b) Calculate the pH after 0.010 mol of HCl is added to 1.0 L of the buffer in (a), without change in volume.



$$(\text{CH}_3)_2\text{NH}_{(\text{aq})} + \text{H}^+_{(\text{aq})} \longrightarrow (\text{CH}_3)_2\text{NH}_2^+_{(\text{aq})}$$

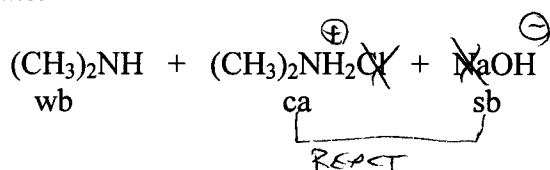
init.	0.25	0.010	0.15
$\Delta$	-0.010	-0.010	+0.010
final	0.24	0	0.16

$$\text{pOH} = -\log(7.4 \times 10^{-4}) + \log\left(\frac{0.16}{0.24}\right)$$

$$\text{pOH} = 3.13 - 0.18 = 2.95$$

$$\text{pH} = 11.05$$

(c) Calculate the pH after 0.040 mol of NaOH is added to 1.0 L of the buffer in (a), without change in volume.



$$(\text{CH}_3)_2\text{NH}_2^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \longrightarrow \text{H}_2\text{O}_{(\text{l})} + (\text{CH}_3)_2\text{NH}_{(\text{aq})}$$

init.	0.15	0.040	---	0.25
$\Delta$	-0.040	-0.040		+0.040
final	0.11	0		0.29

$$\text{pOH} = -\log(7.4 \times 10^{-4}) + \log\left(\frac{0.11}{0.29}\right)$$

$$\text{pOH} = 3.13 - 0.42 = 2.71$$

$$\text{pH} = 11.29$$

14. A 25.00 mL sample of 0.125 M HCl is titrated with 0.125 M NaOH. Calculate the pH of the solution after the addition of (a) 20.00 mL NaOH, (b) 22.50 mL NaOH, (c) 25.00 mL NaOH, (d) 26.00 mL NaOH, (e) 29.00 mL NaOH.

(a) 20.00 mL NaOH

(i) new molarities

$$M(\text{HCl}) = \frac{(0.125 \text{ M})(25.00 \text{ mL})}{45.00 \text{ mL}} = 0.0694 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.125 \text{ M})(20.00 \text{ mL})}{45.00 \text{ mL}} = 0.0556 \text{ M}$$

(ii) eqn. and table

	$\text{HCl}_{(aq)}$	$+$	$\text{NaOH}_{(aq)}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	$+$	$\text{NaCl}_{(aq)}$
init.	0.0694		0.0556		---		0
$\Delta$	<u>-0.0556</u>		<u>-0.0556</u>				<u>+0.0556</u>
final	0.0138		0				0.0556

(iii) calc. pH

$$[\text{HCl}] = [\text{H}_3\text{O}^+] = 0.0138 \text{ M}$$

$$\text{pH} = -\log(0.0138)$$

$$\boxed{\text{pH} = 1.860}$$

STRONG  
ACID

IGNORE - SALT OF  
[SA+SB]

(b) 22.50 mL NaOH

(i) new molarities

$$M(\text{HCl}) = \frac{(0.125 \text{ M})(25.00 \text{ mL})}{47.50 \text{ mL}} = 0.0658 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.125 \text{ M})(22.50 \text{ mL})}{47.50 \text{ mL}} = 0.0592 \text{ M}$$

(ii) eqn. and table

	$\text{HCl}_{(aq)}$	$+$	$\text{NaOH}_{(aq)}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	$+$	$\text{NaCl}_{(aq)}$
init.	0.0658		0.0592		---		0
$\Delta$	<u>-0.0592</u>		<u>-0.0592</u>				<u>+0.0592</u>
final	0.0066		0				0.0592

(iii) calc. pH

$$[\text{HCl}] = [\text{H}_3\text{O}^+] = 0.0066 \text{ M}$$

$$\text{pH} = -\log(0.0066)$$

$$\boxed{\text{pH} = 2.18}$$

(c) 25.00 mL NaOH

(i) new molarities

$$M(\text{HCl}) = \frac{(0.125 \text{ M})(25.00 \text{ mL})}{50.00 \text{ mL}} = 0.0625 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.125 \text{ M})(25.00 \text{ mL})}{50.00 \text{ mL}} = 0.0625 \text{ M}$$



(ii) eqn. and table

	$\text{HCl}_{(aq)}$	+	$\text{NaOH}_{(aq)}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	+	$\text{NaCl}_{(aq)}$
init.	0.0625		0.0625		---		0
$\Delta$	<u>-0.0625</u>		<u>-0.0625</u>				<u>+0.0625</u>
final	0		0				0.0625

(iii) calc. pH

The only species remaining is NaCl, the salt of a [strong acid + strong base]. Because this salt does not undergo hydrolysis, it does not affect pH.

$$\text{pH} = 7.00$$

(d) 26.00 mL NaOH

(i) new molarities

$$M(\text{HCl}) = \frac{(0.125 \text{ M})(25.00 \text{ mL})}{51.00 \text{ mL}} = 0.0613 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.125 \text{ M})(26.00 \text{ mL})}{51.00 \text{ mL}} = 0.0637 \text{ M}$$

(ii) eqn. and table

	$\text{HCl}_{(aq)}$	+	$\text{NaOH}_{(aq)}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	+	$\text{NaCl}_{(aq)}$
init.	0.0613		0.0637		---		0
$\Delta$	<u>-0.0613</u>		<u>-0.0613</u>				<u>+0.0613</u>
final	0		0.0024				0.0613

(iii) calc. pH

$$[\text{NaOH}] = [\text{OH}^-] = 0.0024 \text{ M}$$

$$\text{pOH} = -\log(0.0024) = 2.62$$

$$\text{pH} = 11.38$$

(e) 29.00 mL NaOH

(i) new molarities

$$M(\text{HCl}) = \frac{(0.125 \text{ M})(25.00 \text{ mL})}{54.00 \text{ mL}} = 0.0579 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.125 \text{ M})(29.00 \text{ mL})}{54.00 \text{ mL}} = 0.0671 \text{ M}$$

(ii) eqn. and table

	$\text{HCl}_{(aq)}$	+	$\text{NaOH}_{(aq)}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	+	$\text{NaCl}_{(aq)}$
init.	0.0579		0.0671		---		0
$\Delta$	<u>-0.0579</u>		<u>-0.0579</u>				<u>+0.0579</u>
final	0		0.0092				0.0579

(iii) calc. pH

$$[\text{NaOH}] = [\text{OH}^-] = 0.0092 \text{ M}$$

$$\text{pOH} = -\log(0.0092) = 2.04$$

$$\text{pH} = 11.96$$

15. A 50.00 mL sample of 0.200 M  $\text{CH}_3\text{CO}_2\text{H}$  is titrated with 0.200 M NaOH. Calculate the pH of the solution after the addition of (a) 20.00 mL NaOH, (b) 30.00 mL NaOH, (c) 50.00 mL NaOH, (d) 55.00 mL NaOH.

(a) 20.00 mL NaOH

(i) new molarities

$$M(\text{CH}_3\text{CO}_2\text{H}) = \frac{(0.200 \text{ M})(50.00 \text{ mL})}{70.00 \text{ mL}} = 0.143 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.200 \text{ M})(20.00 \text{ mL})}{70.00 \text{ mL}} = 0.0571 \text{ M}$$

(ii) eqn. and table

	$\overset{WA}{\text{CH}_3\text{CO}_2\text{H}_{(aq)}}$	+	$\overset{SB}{\text{NaOH}_{(aq)}}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	+	$\overset{CB}{\text{NaCH}_3\text{CO}_2_{(aq)}}$
init.	0.143		0.0571		---		0
$\Delta$	<u>-0.0571</u>		<u>-0.0571</u>				<u>+0.0571</u>
final	0.086		0				-0.0571

(iii) calc. pH

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{NaCH}_3\text{CO}_2]}{[\text{CH}_3\text{CO}_2\text{H}]}\right)$$

$$\text{pH} = -\log(1.8 \times 10^{-5}) + \log\left(\frac{0.0571}{0.086}\right)$$

$$\text{pH} = 4.74 + \log(0.66)$$

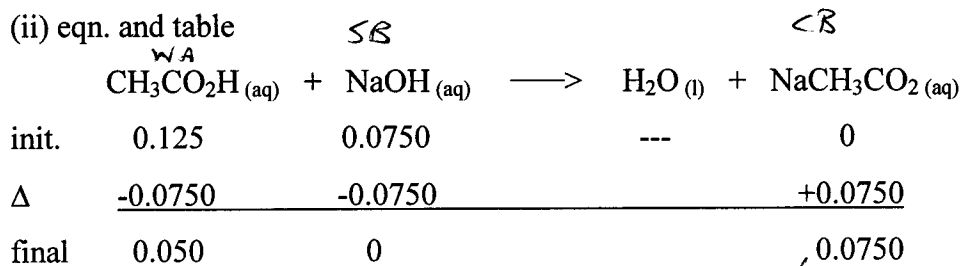
$$\text{pH} = 4.74 - 0.18$$

$$\text{pH} = 4.56$$

(b) 30.00 mL NaOH

(i) new molarities

$$M(\text{CH}_3\text{CO}_2\text{H}) = \frac{(0.200 \text{ M})(50.00 \text{ mL})}{80.00 \text{ mL}} = 0.125 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.200 \text{ M})(30.00 \text{ mL})}{80.00 \text{ mL}} = 0.0750 \text{ M}$$



(iii) calc. pH

$$pH = pK_a + \log \left( \frac{[NaCH_3CO_2]}{[CH_3CO_2H]} \right)$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left( \frac{0.0750}{0.050} \right)$$

$$pH = 4.74 + \log(1.5)$$

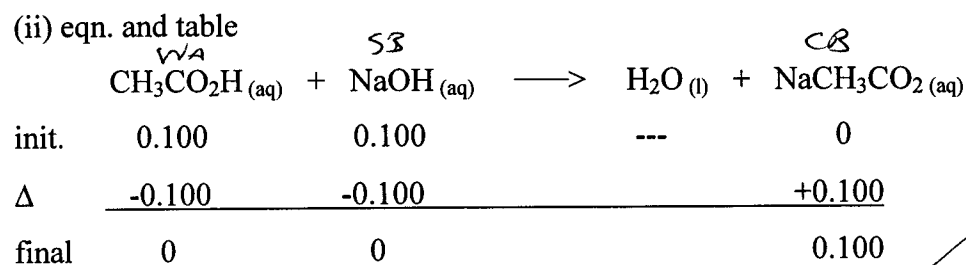
$$pH = 4.74 + 0.18$$

$$pH = 4.92$$

(c) 50.00 mL NaOH

(i) new molarities

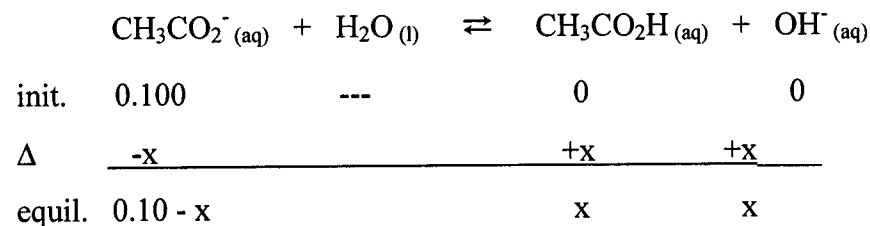
$$M(CH_3CO_2H) = \frac{(0.200 \text{ M})(50.00 \text{ mL})}{100.00 \text{ mL}} = 0.100 \text{ M}; \quad M(NaOH) = \frac{(0.200 \text{ M})(50.00 \text{ mL})}{100.00 \text{ mL}} = 0.100 \text{ M}$$



SALT of [WA + SB] -  
UNDERGOES  
HYDROLYSIS

(iii) calc. pH

The only remaining species is a salt of [weak acid + strong base]. The salt undergoes hydrolysis, and thus affects the pH.



$$K_b = \frac{[\text{CH}_3\text{CO}_2\text{H}][\text{OH}^-]}{[\text{CH}_3\text{CO}_2^-]} \quad K_b = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.100 - x}$$

$$\text{assume: } 5.6 \times 10^{-10} = \frac{x^2}{0.100}$$

$$x^2 = (0.100)(5.6 \times 10^{-10}) = 5.6 \times 10^{-11}$$

$$x = 7.5 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = -\log(7.5 \times 10^{-6}) = 5.12$$

$$\boxed{\text{pH} = 8.88}$$

check assumption:

$$\frac{7.5 \times 10^{-6}}{0.100} \times 100 = 7.5 \times 10^{-3} \%$$

assumption is o.k.

(d) 55.00 mL NaOH

(i) new molarities

$$M(\text{CH}_3\text{CO}_2\text{H}) = \frac{(0.200 \text{ M})(50.00 \text{ mL})}{105.00 \text{ mL}} = 0.0952 \text{ M}; \quad M(\text{NaOH}) = \frac{(0.200 \text{ M})(55.00 \text{ mL})}{105.00 \text{ mL}} = 0.105 \text{ M}$$

(ii) eqn. and table

	$\overset{WA}{\text{CH}_3\text{CO}_2\text{H}}_{(aq)}$	$\overset{SB}{+ \text{NaOH}}_{(aq)}$	$\longrightarrow$	$\text{H}_2\text{O}_{(l)}$	$+ \overset{CB}{\text{NaCH}_3\text{CO}_2}_{(aq)}$
init.	0.0952	0.105		---	0
$\Delta$	<u>-0.0952</u>	<u>-0.0952</u>			<u>+0.0952</u>
final	0	0.010			0.0952

(iii) calc. pH

Because of the presence of the strong base, the pH change caused by the salt is negligible. We need only use the [strong base] to determine the pH of the solution.

$$[\text{NaOH}] = [\text{OH}^-] = 0.010 \text{ M}$$

$$\text{pOH} = -\log(0.010) = 2.00$$

$$\boxed{\text{pH} = 12.00}$$

16. A 60.00 mL sample of 0.120 M  $\text{NH}_3$  is titrated with 0.120 M  $\text{HCl}$ . Calculate the pH of the solution after the addition of (a) 35.00 mL  $\text{HCl}$ , (b) 45.00 mL  $\text{HCl}$ , (c) 60.00 mL  $\text{HCl}$ , (d) 70 mL  $\text{HCl}$ .

(a) 35.00 mL  $\text{HCl}$

(i) new molarities

$$M(\text{NH}_3) = \frac{(0.120 \text{ M})(60.00 \text{ mL})}{95.00 \text{ mL}} = 0.0758 \text{ M}; \quad M(\text{HCl}) = \frac{(0.120 \text{ M})(35.00 \text{ mL})}{95.00 \text{ mL}} = 0.0442 \text{ M}$$

(ii) eqn. and table

	$\text{NB}$	$\text{SA}$	$\longrightarrow$	$\text{CA}$	
	$\text{NH}_3(\text{aq})$	$+$	$\text{HCl}(\text{aq})$	$\longrightarrow$	$\text{NH}_4\text{Cl}(\text{aq})$
init.	0.0758		0.0442		0
$\Delta$	<u>-0.0442</u>		<u>-0.0442</u>		<u>+0.0442</u>
final	0.0316		0		0.0442

$\swarrow$  *buffer*  $\searrow$

(iii) calc. pH

$$\text{pOH} = \text{p}K_b + \log\left(\frac{[\text{NH}_4\text{Cl}]}{[\text{NH}_3]}\right)$$

$$\text{pOH} = -\log(1.8 \times 10^{-5}) + \log\left(\frac{0.0442}{0.0316}\right)$$

$$\text{pOH} = 4.74 + 0.14 = 4.88$$

$$\text{pH} = 9.12$$

(b) 45.00 mL  $\text{HCl}$

(i) new molarities

$$M(\text{NH}_3) = \frac{(0.120 \text{ M})(60.00 \text{ mL})}{105.00 \text{ mL}} = 0.0686 \text{ M}; \quad M(\text{HCl}) = \frac{(0.120 \text{ M})(45.00 \text{ mL})}{105.00 \text{ mL}} = 0.0514 \text{ M}$$

(ii) eqn. and table

	$\text{NB}$	$\text{SA}$	$\longrightarrow$	$\text{CA}$	
	$\text{NH}_3(\text{aq})$	$+$	$\text{HCl}(\text{aq})$	$\longrightarrow$	$\text{NH}_4\text{Cl}(\text{aq})$
init.	0.0686		0.0514		0
$\Delta$	<u>-0.0514</u>		<u>-0.0514</u>		<u>+0.0514</u>
final	0.0172		0		0.0514

$\swarrow$  *buffer*  $\searrow$

(iii) calc. pH

$$\text{pOH} = \text{p}K_b + \log\left(\frac{[\text{NH}_4\text{Cl}]}{[\text{NH}_3]}\right)$$

$$\text{pOH} = -\log(1.8 \times 10^{-5}) + \log\left(\frac{0.0514}{0.0172}\right)$$

$$\text{pOH} = 4.74 + 0.475 = 5.22$$

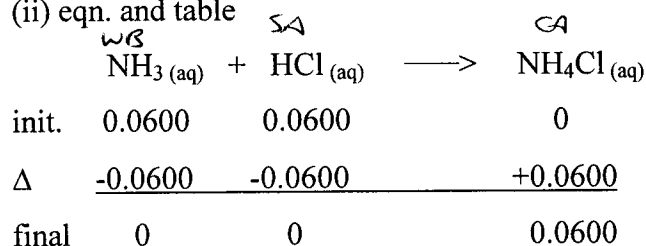
$$\text{pH} = 8.78$$

(c) 60.00 mL HCl

(i) new molarities

$$M(\text{NH}_3) = \frac{(0.120 \text{ M})(60.00 \text{ mL})}{120.00 \text{ mL}} = 0.0600 \text{ M}; \quad M(\text{HCl}) = \frac{(0.120 \text{ M})(60.00 \text{ mL})}{120.00 \text{ mL}} = 0.0600 \text{ M}$$

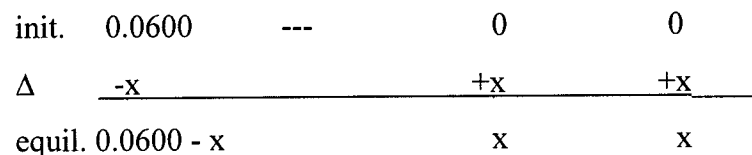
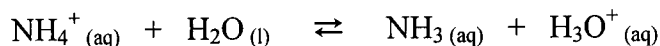
(ii) eqn. and table



SALT of [SA + WB]  
UNDERGOES HYDROLYSIS

(iii) calc. pH

The only remaining species is a salt of [strong acid + weak base]. The salt undergoes hydrolysis, and thus affects the pH.



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} \quad K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.0600 - x}$$

$$\text{assume: } 5.6 \times 10^{-10} = \frac{x^2}{0.0600}$$

$$x^2 = (0.0600)(5.6 \times 10^{-10}) = 3.4 \times 10^{-11}$$

$$x = 5.8 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(5.8 \times 10^{-6})$$

$$\text{pH} = 5.24$$

check assumption:

$$\frac{5.8 \times 10^{-6}}{0.0600} \times 100 = 9.7 \times 10^{-3} \%$$

assumption is o.k.

(d) 70.00 mL HCl

(i) new molarities

$$M(\text{NH}_3) = \frac{(0.120 \text{ M})(60.00 \text{ mL})}{130.00 \text{ mL}} = 0.0554 \text{ M}; \quad M(\text{HCl}) = \frac{(0.120 \text{ M})(70.00 \text{ mL})}{130.00 \text{ mL}} = 0.0646 \text{ M}$$

(ii) eqn. and table

	$\text{NH}_3$	+	$\text{HCl}$	$\longrightarrow$	$\text{NH}_4\text{Cl}$
	(aq)		(aq)		(aq)
init.	0.0554		0.0646		0
$\Delta$	<u>-0.0554</u>		<u>-0.0554</u>		<u>+0.0554</u>
final	0		0.0092		0.0554

XS SA - CONTROLS  
pH

(iii) calc. pH

Because of the presence of the strong acid, the pH change caused by the salt is negligible. We need only use the [strong acid] to determine the pH of the solution.

$$[\text{HCl}] = [\text{H}_3\text{O}^+] = 0.0092 \text{ M}$$

$$\text{pH} = -\log(0.0092)$$

$$\text{pH} = 2.04$$