ACIDS & BASES

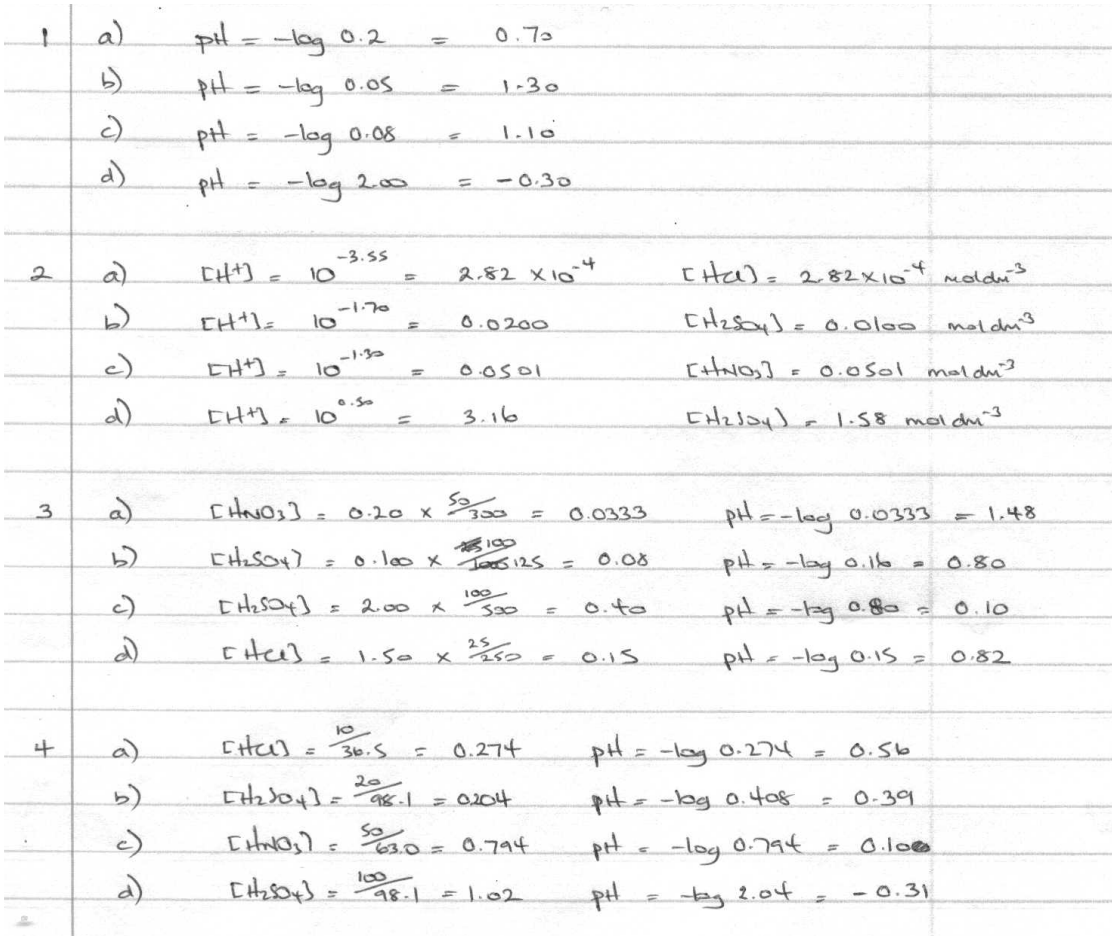
CHEMSHEETS.co.uk

TASK 1 – Bronsted-Lowry acids & bases

1 Acid = H_2O, base = NH_3
2 Acid = HCl, base = H_2O
3 Acid = HCOOH, base = KOH
4 Acid = HCl, base = CH_3COOH
5 Acid = HCl, base = NH_3
6 Acid = HCO_3^-, base = OH^-
7 Acid = H^+, base = HCO_3^-
8 Acid = H_2SO_4, base = HNO_3

TASK 2 – pH of strong acids

1	a	0.70	b	1.30	c	1.10	d	-0.30
2	a	2.82×10^{-4}	b	0.0100	c	0.0501	d	1.58 mol dm^{-3}
3	a	1.48	b	0.80	c	0.10	d	0.82
4	a	0.56	b	0.39	c	0.10	d	-0.31



TASK 3 – pH of strong bases

1	a	13.18	b	12.70	c	13.60
2	a	2.00	b	0.0158	c	0.501 mol dm^{-3}
3	a	12.30	b	13.20	c	13.00
4	a	13.70	b	14.25	c	12.22

1 a) $[OH^-] = 0.15$ $[H^+] = \frac{10^{-14}}{0.15} = 6.67 \times 10^{-14}$ pH = 13.18

b) $[OH^-] = 0.05$ $[H^+] = \frac{10^{-14}}{0.05} = 2 \times 10^{-13}$ pH = 12.70

c) $[OH^-] = 0.40$ $[H^+] = \frac{10^{-14}}{0.40} = 2.5 \times 10^{-14}$ pH = 13.60

2 a) $[H^+] = 10^{-14.30} = 5.01 \times 10^{-15}$ $[OH^-] = \frac{10^{-14}}{5.01 \times 10^{-15}} = 2.00$ $[NaOH] = 2.00$

b) $[H^+] = 10^{-12.50} = 3.16 \times 10^{-13}$ $[OH^-] = \frac{10^{-14}}{3.16 \times 10^{-13}} = 0.0316$ $[Ba(OH)_2] = 0.0158$

c) $[H^+] = 10^{-13.70} = 2.00 \times 10^{-14}$ $[OH^-] = \frac{10^{-14}}{2.00 \times 10^{-14}} = 0.501$ $[KOH] = 0.501$

3 a) $[NaOH] = 0.100 \times \frac{25}{125} = 0.0200$ $[OH^-] = 0.0200$

$[H^+] = \frac{10^{-14}}{0.0200} = 5 \times 10^{-13}$ pH = 12.30

b) $[Ba(OH)_2] = 0.100 \times \frac{100}{125} = 0.080$ $[OH^-] = 0.160$

$[H^+] = \frac{10^{-14}}{0.160} = 6.25 \times 10^{-14}$ pH = 13.20

c) $[KOH] = 1.00 \times \frac{100}{100} = 0.100$ $[OH^-] = 0.100$

$[H^+] = \frac{10^{-14}}{0.100} = 1 \times 10^{-13}$ pH = 13.00

4 a) mol NaOH = $\frac{20}{40.1} = 0.499$ $[OH^-] = 0.499$

$[H^+] = \frac{10^{-14}}{0.499} = 2.00 \times 10^{-14}$

pH = 13.70

b) mol KOH = $\frac{100}{56.1} = 1.783$ $[OH^-] = 1.78$

$[H^+] = \frac{10^{-14}}{1.783} = 5.62 \times 10^{-15}$

pH = 14.25

c) mol $Sr(OH)_2$ = $\frac{1.00}{121.6} = 0.00822$ $[OH^-] = 0.0164$

$[H^+] = 10^{-14} / 0.0164 = 6.08 \times 10^{-13}$

pH = 12.22

TASK 4 – pH of mixtures of strong acids and strong bases

1	2.00	2	1.48	3	13.12	4	13.15	5	1.60
6	new pH = 1.00, increase by 0.30			7	new pH = 1.60, decrease by 11.58				

1

$\text{mol } H^+ = 0.100 \times 20/1000 = 0.002$

$\text{mol } OH^- = 0.050 \times 30/1000 = 0.0015$

$\text{XS } H^+ = 0.0005$

$[H^+] = 0.0005 / 50/1000 = 0.01$

$pH = 2.00$

2

$\text{mol } H^+ = 0.150 \times 25/1000 \times 2 = 0.0075$

$\text{mol } OH^- = 0.100 \times 50/1000 = 0.005$

$\text{XS } H^+ = 0.0025$

$[H^+] = 0.0025 / 75/1000 = 0.033$

$pH = 1.48$

3

$\text{mol } H^+ = 0.050 \times 100/1000 = 0.005$

$\text{mol } OH^- = 0.500 \times 50/1000 = 0.025$

$\text{XS } OH^- = 0.02$

$[OH^-] = 0.02 / 150/1000 = 0.133$

$[H^+] = 10^{-14} / 0.133 = 7.5 \times 10^{-14}$

$pH = 13.12$

4

$\text{mol } H^+ = 1.00 \times 13/1000 \times 2 = 0.02$

$\text{mol } OH^- = 1.00 \times 25/1000 = 0.025$

$\text{XS } OH^- = 0.005$

$[OH^-] = 0.005 / 35/1000 = 0.143$

$[H^+] = 10^{-14} / 0.143 = 7 \times 10^{-14}$

$pH = 13.15$

5 $\text{mol } H^+ = 0.250 \times \frac{50}{1000} = 0.0125$

$\text{mol } OH^- = 2 \times 0.100 \times \frac{50}{1000} = 0.0100$

$\text{excess mol } H^+ = 0.0125 - 0.010 = 0.0025$

$\text{excess } [H^+] = \frac{0.0025}{100/1000} = 0.025$

$pH = -\log [H^+] = -\log 0.025 = 1.60$

6 Initial flask: $pH = -\log 0.200 = 0.70$

after reaction: $\text{mol } H^+ = 0.200 \times \frac{100}{1000} = 0.0200$

$\text{mol } OH^- = 0.100 \times \frac{50}{1000} = 0.00500$

$\text{excess mol } H^+ = 0.0150$

$\text{excess } [H^+] = \frac{0.0150}{150/1000} = 0.100$

$pH = -\log 0.100 = 1.00$

∴ pH increases by 0.30

7 Initial flask: $[OH^-] = 0.150$

$[H^+] = K_w / [OH^-] = \frac{10^{-14}}{0.150} = 6.67 \times 10^{-14}$

$pH = -\log 6.67 \times 10^{-14} = 13.18$

After reaction: $\text{mol } OH^- = 0.150 \times \frac{50}{1000} = 0.0075$

$\text{mol } H^+ = 2 \times 0.100 \times \frac{50}{1000} = 0.0100$

$\text{excess } H^+ = 0.0025$

$\text{excess } [H^+] = \frac{0.0025}{100/1000} = 0.025$

$pH = -\log 0.025 = 1.60$

∴ pH decreases by 11.58

TASK 5 – A variety of pH calculations so far

1 a 13.48 b 0.70 c -0.48 d 12.70

2 a 6.63 b still neutral as $[H^+] = [OH^-]$

3 a 0.40 b 12.00

4 a 12.52 b 0.93

5 13.74

1 a) $[OH^-] = 0.300$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.300} = 3.33 \times 10^{-14}$

$pH = 13.48$

b) $[H^+] = 0.200$

$pH = 0.70$

c) $[H^+] = 3.00$

$pH = -0.48$

d) $[OH^-] = 0.0500$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.050} = 2 \times 10^{-13}$

$pH = 12.70$

2 a) $[H^+] = \sqrt{K_w} = \sqrt{5.48 \times 10^{-14}} = 2.34 \times 10^{-7}$

$pH = 6.63$

b) still neutral as $[H^+] = [OH^-]$

4 a) $\text{mol } H^+ = 0.2 \times \frac{100}{1000} = 0.020$

$\text{mol } OH^- = 0.5 \times \frac{50}{1000} = 0.025$

$\text{XS mol } OH^- = 0.005$

$\text{XS } [OH^-] = \frac{0.005}{150/1000} = 0.0333$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.0333} = 3 \times 10^{-13}$

$pH = 12.52$

b) mol $H^+ = 0.25 \times \frac{25}{1000} = 0.00625$

mol $OH^- = 0.10 \times \frac{15}{1000} = 0.00150$

XS $H^+ = 0.00475$

$[H^+] = \frac{0.00475}{60/1000} = 0.1188$

pH = 0.93

3 a) $[H^+] = 2 \times 1.0 \times \frac{20}{100} = 0.4$

pH = 0.40

b) $[OH^-] = 0.05 \times \frac{50}{250} = 0.01$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.01} = 1 \times 10^{-12}$

pH = 12.00

5 mol $OH^- = \frac{3.5}{40.0} \times 0.987 = 0.08636$

mol $H^+ = \frac{25}{1000} \times 2 \times 0.35 = 0.0175$

∴ XS $OH^- = 0.06886$

$[OH^-] = \frac{0.06886}{125/1000} = 0.551$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.551} = 1.81 \times 10^{-14}$

pH = 13.74

TASK 6 – The pH of weak acids

1 a 2.51 b 2.76 c 1.87

2 a 5.01×10^{-5} b 5.50×10^{-3} mol dm^{-3}

3 a ethanoic acid b propenoic acid

4 4.79×10^{-5} mol dm^{-3}

1 a) $[H^+] = \sqrt{K_a[HA]} = \sqrt{0.15 \times 10^{-4.20}} = 3.07 \times 10^{-3}$

pH = 2.51

b) $[H^+] = \sqrt{K_a[HA]} = \sqrt{0.20 \times 1.51 \times 10^{-5}} = 1.74 \times 10^{-3}$

pH = 2.76

c) $[H^+] = \sqrt{K_a[HA]} = \sqrt{1.00 \times 1.78 \times 10^{-4}} = 0.0133$

pH = 1.87

2 a) $[H^+] = 10^{-4.53} = 2.95 \times 10^{-5}$

$[HA] = \frac{[H^+]^2}{K_a} = 5.01 \times 10^{-5}$

b) $[H^+] = 10^{-3.56} = 2.75 \times 10^{-4}$

$[HA] = \frac{[H^+]^2}{K_a} = \frac{(2.75 \times 10^{-4})^2}{10^{-4.86}} = 5.50 \times 10^{-3}$

3 a) ethanoic acid stronger = lower pKa

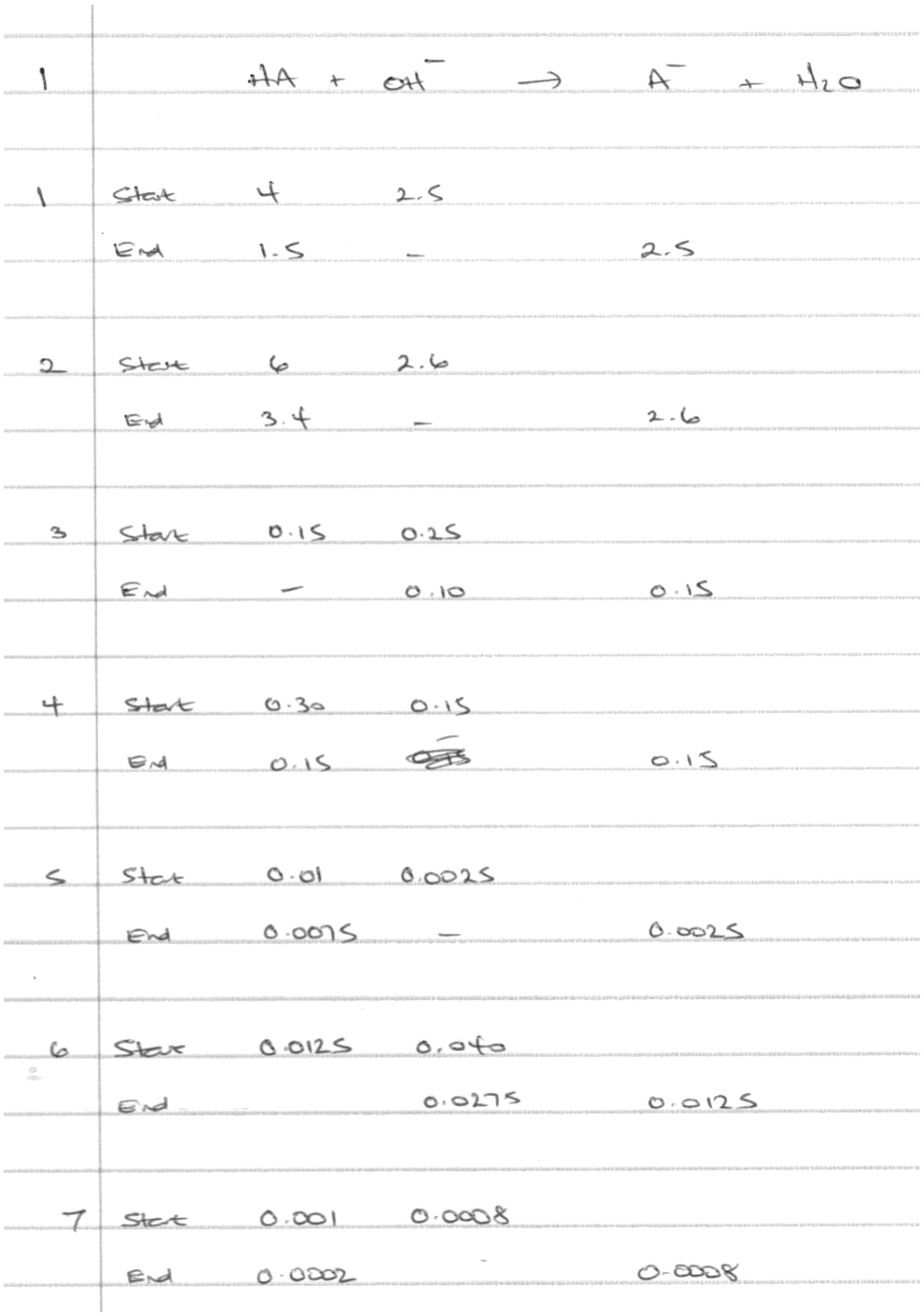
b) propenoic acid stronger = higher K_a

4 $[H^+] = 10^{-2.66} = 2.19 \times 10^{-3}$

$K_a = \frac{[H^+]^2}{[HA]} = 4.79 \times 10^{-5}$ mol dm^{-3}

TASK 7 – Reactions of weak acids

1) HA = 1.5, OH^- = 0, A^- = 2.5
2) HA = 3.4, OH^- = 0, A^- = 2.6
3) HA = 0, OH^- = 0.10, A^- = 0.15
4) HA = 0.15, OH^- = 0, A^- = 0.15
5) HA = 0.0075, OH^- = 0, A^- = 0.0025
6) HA = 0, OH^- = 0.0275, A^- = 0.0125
7) HA = 0.0002, OH^- = 0, A^- = 0.0008



TASK 8 – pH of mixtures of weak acids & strong bases

1	12.30	2	4.54	3	4.76	4	2.26	5	13.92	6	3.14

1

$HA + OH^- \rightarrow A^- + H_2O$

start $0.1 \times \frac{20}{1000}$ $0.08 \times \frac{40}{1000}$

0.002 0.0032

End – 0.0012 0.002

$[OH^-] = \frac{0.0012}{60/1000} = 0.02$

$[H^+] = \frac{K_w}{[H^+]} = \frac{10^{-14}}{0.02} = 5 \times 10^{-13}$ pH = 12.30

2

$HA + OH^- \rightarrow A^- + H_2O$

start $0.5 \times \frac{5}{1000}$ $0.08 \times \frac{100}{1000}$

0.025 0.008

end 0.017 – 0.008

$[H^+] = K_a \frac{[HA]}{[A^-]} = \frac{10^{-4.87} \times 0.017 \times 150/1000}{0.008 \times 150/1000} = 2.87 \times 10^{-5}$

pH = 4.54

3

$HA + OH^- \rightarrow A^- + H_2O$

Start $0.5 \times \frac{50}{1000}$ $0.25 \times \frac{50}{1000}$

0.025 0.0125

End 0.0125 – 0.0125

½ neutralisation pH = pK_a = 4.76

4

$HA + OH^- \rightarrow A^- + H_2O$

Start $0.5 \times \frac{50}{1000}$ $2 \times 0.1 \times \frac{25}{1000}$

0.025 0.005

End 0.020 — 0.005

$[H^+] = \frac{K_a [HA]}{[A^-]} = \frac{10^{-2.86} \times \frac{0.020}{75/1000}}{\frac{0.005}{75/1000}} = 5.52 \times 10^{-3}$

pH = 2.26

5

$HA + OH^- \rightarrow A^- + H_2O$

Start $1.5 \times \frac{50}{1000}$ $2.0 \times \frac{100}{1000}$

0.075 0.2

End — 0.125 0.075

$[OH^-] = \frac{0.125}{150/1000} = 0.833$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.833} = 1.20 \times 10^{-14}$

pH = 13.92

6

$HA + OH^- \rightarrow A^- + H_2O$

Start $1.0 \times \frac{25}{1000}$ $0.04 \times \frac{50}{1000}$

0.025 0.002

End 0.023 — 0.002

$[H^+] = \frac{K_a [HA]}{[A^-]} = \frac{10^{-4.20} \times \frac{0.023}{75/1000}}{\frac{0.002}{1000}} = 7.26 \times 10^{-4}$

pH = 3.14

TASK 9 – A variety of pH calculations so far

1	0.70	2	2.19	3	13.60	4	1.22	5	12.30
6	4.76	7	5.24	8	13.30				

1 $[H^+] = 0.2$

$pH = 0.70$

2 $[H^+] = \sqrt{K_a[HA]} = \sqrt{1.70\times10^{-4} \times 0.250} = 6.52 \times 10^{-3}$

$pH = 2.19$

3 $[OH^-] = 0.40$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.40} = 2.5 \times 10^{-14}$

$pH = 13.60$

4 $\text{mol } H^+ = 0.2 \times {}^{80}/_{1000} = 0.016$

$\text{mol } OH^- = 0.5 \times {}^{20}/_{1000} = 0.01$

$\text{xs } H^+ = 0.006$

$[H^+] = \frac{0.006}{{}^{100}/_{1000}} = 0.06$

$pH = 1.22$

5 $[OH^-] = 0.1 \times {}^{25}/_{125} = 0.02$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.02} = 5 \times 10^{-13}$

$pH = 12.30$

6 $\text{mol } HA = 0.2 \times {}^{25}/_{1000} = 0.005$

$\text{mol } OH^- = 0.1 \times {}^{25}/_{1000} = 0.0025$

½ neutralised

$\therefore pH = pK_a = 4.76$

7 mol HA = $0.100 \times \frac{100}{1000} = 0.01$

mol OH^- = $0.150 \times \frac{50}{1000} = 0.0075$

XS HA = 0.0025

∴ A^- = 0.0075

$HA + OH^- \rightarrow A^- + H_2O$

	HA	OH^-	A^-	H_2O
Start	0.01	0.0075		
End	0.0025	–	0.0075	

$[H^+] = K_a \frac{[HA]}{[A^-]} = \frac{10^{-4.76} \times \frac{0.0025}{150/1000}}{\frac{0.0075}{150/1000}}$

$= 5.79 \times 10^{-6}$

pH = 5.24

8 $HA + OH^- \rightarrow A^- + H_2O$

	HA	OH^-	A^-	H_2O
Start	$0.2 \times 50/1000$	$1.0 \times 25/1000$		
	0.01	0.025		
End	–	0.015	0.01	

$[OH^-] = \frac{0.015}{175/1000} = 0.2$

$[H^+] = \frac{10^{-14}}{0.2} = 5 \times 10^{-14}$

pH = 13.30

TASK 10a – Titration calculations

1	a	0.0752 mol dm^{-3}			b	3.01 g dm^{-3}							
2	a	0.050 mol dm^{-3}			b	7.10 g dm^{-3}							
3	1.13 g		4	87.8		5	2	6	K	7	87.7%	8	90.8%

1 a) mol H_2SO_4 = $0.05 \times \frac{18.8}{1000}$ = 0.00094

mol NaOH = 0.00188

conc NaOH = $\frac{0.00188}{25/1000}$ = 0.0752 mol dm^{-3}

b) conc NaOH = 40.0×0.0752 = 3.01 g dm^{-3}

2 a) mol NaOH = $0.10 \times \frac{37.5}{1000}$ = 0.00375

mol H_3AsO_4 = 0.00125

conc H_3AsO_4 = $\frac{0.00125}{25/1000}$ = 0.050 mol dm^{-3}

b) conc H_3AsO_4 = 141.9×0.050 = 7.10 g dm^{-3}

3 mol HCl = $0.100 \times \frac{28.2}{1000}$ = 0.00282

mol NaOH = 0.00282

mol NaOH in 250 cm^3 = 0.0282

mass NaOH = 40.0×0.0282 = 1.13 g

4 mol NaOH = $0.095 \times \frac{46.5}{1000}$ = 0.004418

mol HA = 0.004418

mol HA in 250 cm^3 = 0.04418

Mr = $\frac{3.88}{0.04418}$ = 87.8

5 mol NaOH = $0.160 \times \frac{15.6}{1000}$ = 0.002496

mol H_2A = 0.001248

mol H_2A in 250 cm^3 = 0.01248

Mr = $\frac{1.575}{0.01248}$ = 126

126 − 90 = 36 $n = \frac{36}{18}$ = 2

6 mol HCl = $0.10 \times \frac{27.0}{1000}$ = 0.0027

mol M_2CO_3 in 1000 cm^3 = $0.0027 \div 2 \times 40$ = 0.054

Mr = $\frac{7.46}{0.054}$ = 138.1

138.1 − 60 = 78.1 Ar M = $\frac{78.1}{2}$ = 39.1

M = K

7 mol NaOH = $0.1 \times 24.8/1000 = 0.00248$

∴ mol HCl in XS = 0.00248

mol HCl added to $CaCO_3$ = $0.2 \times 10/1000 = 0.02$

∴ mol HCl reacting with $CaCO_3$ = $0.02 - 0.00248 = 0.01752$

mol $CaCO_3$ = 0.00876

mass $CaCO_3$ = $100.1 \times 0.00876 = 0.877$ g

% = 87.7%

8 Conc NaOH : mol HCl = $0.20 \times 28.5/1000 = 0.0057$

mol NaOH = 0.0057

conc NaOH = $0.0057 / 25/1000 = 0.228$ mol dm^{-3}

XS HCl mol NaOH = $0.228 \times 10.9/1000 = 0.002485$

mol HCl in XS = 0.002485

Orgnd HCl mol HCl = $0.2 \times 100/1000 = 0.02$

∴ HCl reacting with $Ba(OH)_2$ = 0.01752

∴ mol $Ba(OH)_2$ = 0.00876

mass $Ba(OH)_2$ = $171.3 \times 0.00876 = 1.50$g

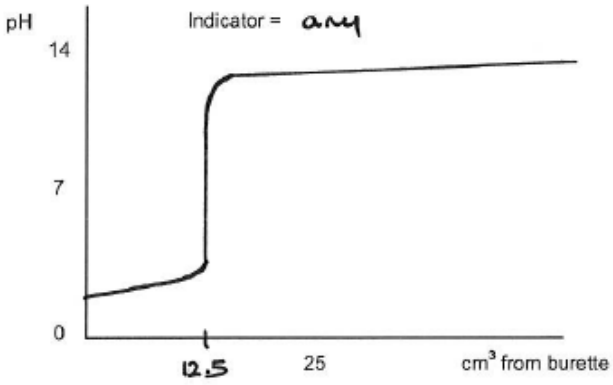
% purity = $\frac{1.50}{1.6524} \times 100 = 90.8\%$.

TASK 10b – Short-cut titration calculations

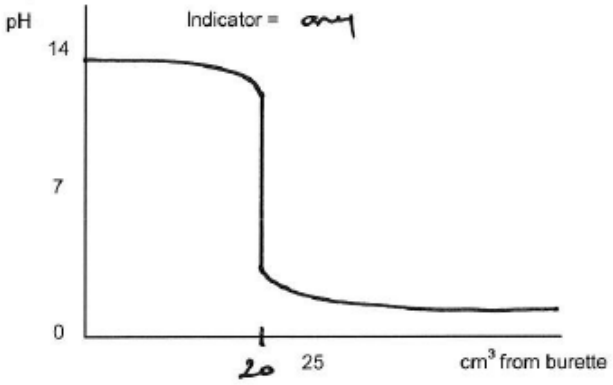
1	50 cm^3	2	80 cm^3	3	25 cm^3	4	12 cm^3	5	5 cm^3	6	10 cm^3
7	15 cm^3	8	25 cm^3	9	40 cm^3	10	100 cm^3				

TASK 11 – Sketching pH curves

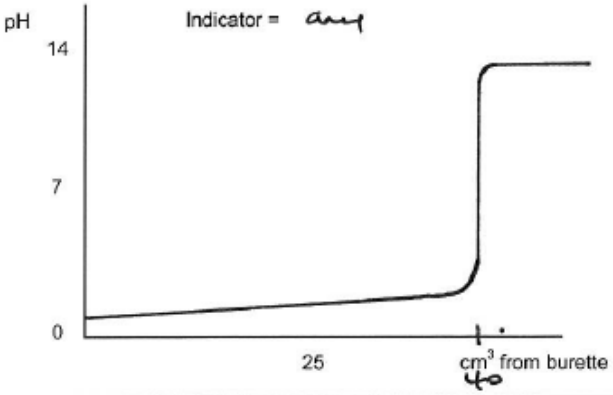
(1) Flask 25 cm^3 0.10 mol dm^{-3} HNO_3 pH 1
Burette 50 cm^3 0.20 mol dm^{-3} NaOH pH 13.3



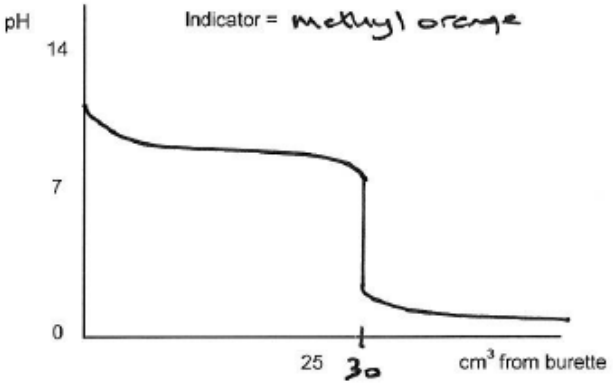
(2) Flask 20 cm^3 0.10 mol dm^{-3} NaOH 13 pH
Burette 50 cm^3 0.10 mol dm^{-3} HCl 1 pH



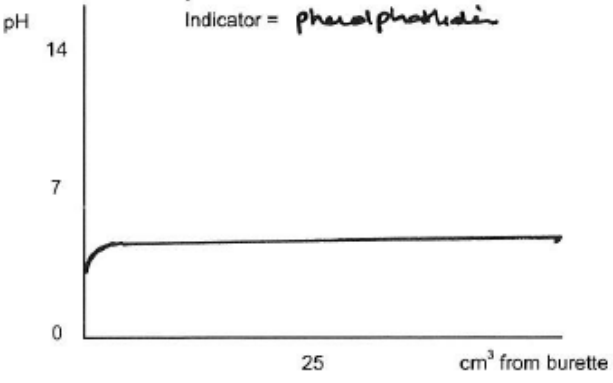
(3) Flask 10 cm^3 0.20 mol dm^{-3} HNO_3 0.7 pH
Burette 50 cm^3 0.05 mol dm^{-3} NaOH 12.7 pH



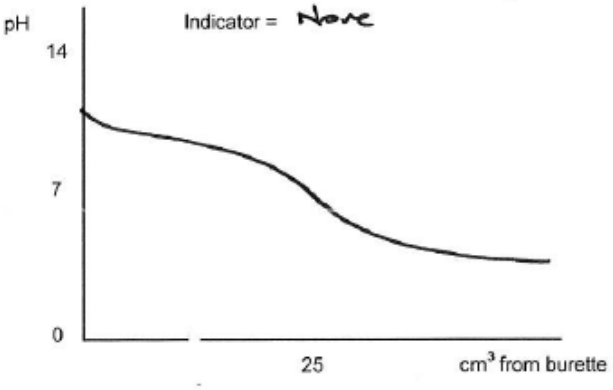
(4) Flask 30 cm^3 1.00 mol dm^{-3} NH_3 11-12ish pH
Burette 50 cm^3 1.00 mol dm^{-3} HCl 0 pH



(5) Flask 20 cm^3 0.20 mol dm^{-3} CH_3COOH pH 3ish
Burette 50 cm^3 0.05 mol dm^{-3} NaOH pH 12.7
(endpoint > 50 cm^3!)



(6) Flask 50 cm^3 0.500 mol dm^{-3} NH_3 pH 11-12ish
Burette 50 cm^3 1.00 mol dm^{-3} methanoic acid
pH 3ish



TASK 12 – Buffer solution calculations

1	a	3.35	b	4.72	c	4.77
2	a	4.35	b	3.57		
3	a	3.37	b	3.40		
4	a	4.43	b	4.40		
5	a	47.9 g	b	0.0113 g		
6	a	pH = 11.29, change = 4.29	b	pH = 4.60, change = 0.02		

1 a) $[HA] = 1.0 \times \frac{5}{10} = 0.714$

$[A^-] = 1.0 \times \frac{2}{10} = 0.286$

$[H^+] = K_a \frac{[HA]}{[A^-]} = 1.78 \times 10^{-4} \times \frac{0.714}{0.286} = 4.45 \times 10^{-4}$

$pH = 3.35$

b) $[HA] = 0.1 \times \frac{25}{45} = 0.0556$

$[A^-] = 0.1 \times \frac{20}{45} = 0.0444$

$[H^+] = K_a \frac{[HA]}{[A^-]} = 10^{-4.82} \times \frac{0.0556}{0.0444} = 1.89 \times 10^{-5}$

$pH = 4.72$

c) $\text{mol } KCH_3COO = \frac{1.00}{98.1} = 0.0102$

$[CH_3COO^-] = \frac{0.0102}{50/1000} = 0.204$

$[H^+] = K_a \frac{[HA]}{[A^-]} = 1.74 \times 10^{-5} \times \frac{0.200}{0.204} = 1.71 \times 10^{-5}$

$pH = 4.77$

2 a) $HA + OH^- \rightarrow A^- + H_2O$

Start	$0.5 \times 25/1000 = 0.0125$	$1.0 \times 19/1000 = 0.01$
End	0.0025	0.01

$$[H^+] = K_a \frac{[HA]}{[A^-]} = 1.78 \times 10^{-4} \times \frac{0.0025/V}{0.01/V}$$

$$= 4.45 \times 10^{-5}$$

$$pH = 4.35$$

b) $HA + OH^- \rightarrow A^- + H_2O$

Start	$1.0 \times 100/1000 = 0.100$	$0.8 \times 50/1000 = 0.04$
End	0.06	0.04

$$[H^+] = K_a \frac{[HA]}{[A^-]} = 1.78 \times 10^{-4} \times \frac{0.06/V}{0.04/V}$$

$$= 2.67 \times 10^{-4}$$

$$pH = 3.57$$

3 a) mol HA = $1.20 \times \frac{20.0}{1000} = 0.024$

mol A^- = $0.500 \times \frac{20.0}{1000} = 0.010$

$K_a = \frac{[H^+][A^-]}{[HA]}$ $\quad [H^+] = K_a \frac{[HA]}{[A^-]}$

$= \frac{1.78 \times 10^{-4} \times \frac{0.024}{40/1000}}{\frac{0.010}{40/1000}}$

$= 4.27 \times 10^{-4}$

pH = 3.37

b) mol OH^- added = $0.4 \times \frac{1.2}{1000} = 0.00048$

$HA + OH^- \rightarrow A^- + H_2O$

	HA	OH^-	A^-
Start	0.024	0.00048	0.010
End	0.02352	-	0.01048

$[H^+] = K_a \frac{[HA]}{[A^-]} = 1.78 \times 10^{-4} \times \frac{\frac{0.02352}{V}}{\frac{0.01048}{V}}$

$= 3.99 \times 10^{-4}$

$pH = -\log 3.99 \times 10^{-4}$

$= 3.40$

4 a) mol HA = $0.500 \times \frac{50.0}{1000} = 0.025$

mol OH^- = $0.800 \times \frac{10.0}{1000} = 0.008$

	HA + OH^-	→ A^- + H_2O
before	0.025 0.008	
after	0.017	0.008

$K_a = \frac{[H^+][A^-]}{[HA]}$ $[H^+] = K_a \frac{[HA]}{[A^-]}$

$= \frac{1.74 \times 10^{-5} \times \frac{0.017}{60/1000}}{\frac{0.008}{60/1000}}$

$= 3.70 \times 10^{-5}$

pH = 4.43

b) mol H^+ added = $0.20 \times \frac{2}{1000} = 0.0004$

	A^- + H^+	→ HA
before	0.008 0.0004	0.017
after	0.0076	0.0174

$[H^+] = \frac{1.74 \times 10^{-5} \times \frac{0.0174}{62/1000}}{\frac{0.0076}{62/1000}}$

$= 3.98 \times 10^{-5}$

pH = 4.40

5 a) $[A^-] = \frac{K_a[HA]}{[H^+]} = \frac{1.78 \times 10^{-4} \times 0.100}{10^{-5.2}} = 2.82$

mol A^- = conc × vol = $2.82 \times \frac{250}{1000} = 0.705$

mass $HCOONa$ = $0.705 \times 68.0 = 47.9$ g

b) $[A^-] = \frac{K_a[HA]}{[H^+]} = \frac{1.74 \times 10^{-5} \times 0.100}{10^{-3.5}} = 0.00550$

mol A^- = conc × vol = $0.00550 \times \frac{25}{1000} = 1.376 \times 10^{-4}$

mass CH_3COONa = $1.376 \times 10^{-4} \times 82.0 = 0.0113$ g

4 6 a) Initial pH = 7.00

Add NaOH: $[OH^-] = 0.1 \times \frac{2}{102} = 0.00196$

$[H^+] = \frac{10^{-14}}{0.00196} = 5.10 \times 10^{-12}$

pH = 11.29 pH change = 4.29.

b) Buffer:

$[H^+] = \frac{K_a[HA]}{[A^-]} = \frac{1.74 \times 10^{-5} \times 0.15}{0.10} = 2.61 \times 10^{-5}$

pH = 4.58

Add OH^-: $HA + OH^- \rightarrow A^- + H_2O$

	$0.15 \times \frac{100}{1000}$	$0.1 \times \frac{2}{100}$	$0.10 \times \frac{100}{1000}$
start	0.015	0.0002	= 0.01
end	0.0148		0.0102

$[H^+] = 1.74 \times 10^{-5} \times \frac{0.0148/V}{0.0102/V} = \frac{2.52 \times 10^{-5}}{0.0102}$ pH = 4.60

pH change = 0.02

c) $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

the OH^- added reacts with H^+

the equilibirum moves right to replace the H^+

as $[CH_3COOH]$ and $[CH_3COO^-]$ are much greater than $[H^+]$

so $[CH_3COOH]/ [CH_3COO^-]$ remains roughly constant

TASK 13 – One final lovely mixture of calculations just for fun

1	0.70	2	1.00	3	13.70	4	1.60	5	1.93	6	4.20
7	6.63, neutral as $[H^+] = [OH^-]$	8	4.45	9	0.88	10	12.52	11	4.06		
12	12.63										

1 H_2SO_4 = strong acid $[H^+] = 0.200$

$pH = -\log [H^+] = \underline{\underline{0.70}}$

2 HCl = strong acid

$[HCl) = 0.500 \times \frac{50}{250} = 0.100$

$[H^+] = 0.100$

$pH = \underline{\underline{1.00}}$

3 NaOH = strong base

$[OH^-] = 0.500$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.500} = 2 \times 10^{-14}$

$pH = \underline{\underline{13.70}}$

4 Solution at start: NaOH = strong base

$[OH^-] = 0.250$

$[H^+) = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.250} = 4 \times 10^{-14}$

$pH = 13.40$

Solution at end: diluted NaOH (strong base)

$[OH^-] = 0.25 \times \frac{25}{1000} = 0.00625$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.00625} = 1.6 \times 10^{-12}$

$pH = 11.80$

$\therefore$ pH change $= 13.40 - 11.80 = \underline{\underline{1.60}}$

5 Chloroethanoic acid = weak acid (low Ka value!) + carboxylic acid

$K_a = \frac{[H^+]^2}{[HA]}$

$[H^+] = \sqrt{K_a[HA]} = \sqrt{1.38 \times 10^{-3} \times 0.100} = 0.0117$

pH = <u>1.93</u>

6 Benzenecarboxylic acid = weak acid (low Ka value + carboxylic acid)

At ½ neutralisation $K_a = [H^+]$ or $pK_a = pH$

$[H^+] = 6.31 \times 10^{-5}$

pH = <u>4.20</u>

7 Pure water $K_w = [H^+]^2$

$[H^+] = \sqrt{K_w} = \sqrt{5.476 \times 10^{-14}} = 2.34 \times 10^{-7}$

pH = <u>6.63</u>

water is still neutral as $[H^+] = [OH^-]$

8 Buffer solution HA = ethanoic acid

A^- = ethanoate ion

mol $CH_3COONa = \frac{\text{mass}}{M_r} = \frac{1.00}{82} = 0.0122$

$[CH_3COO^-] = \frac{0.0122}{250/1000} = 0.0488$

Buffer: $[H^+] = K_a \frac{[HA]}{[A^-]} = 1.74 \times 10^{-5} \times \frac{0.100}{0.0488}$

$= 3.57 \times 10^{-5}$

pH = <u>4.45</u>

9 NaOH = strong base
HNO_3 = strong acid

mol H^+ = $0.25 \times {}^{50}/_{1000} = 0.0125$
mol OH^- = $0.10 \times {}^{25}/_{1000} = 0.0025$
XS mol H^+ = 0.0100
$[H^+]$ XS = $\dfrac{0.0100}{{}^{75}/_{1000}} = 0.133$

pH = <u>0.88</u>

10 NaOH = strong base
H_2SO_4 = strong acid

mol H^+ = $2 \times 0.15 \times {}^{20}/_{1000} = 0.006$
mol OH^- = $0.10 \times {}^{100}/_{1000} = 0.01$
∴ XS mol OH^- = 0.004
XS $[OH^-]$ = $\dfrac{0.004}{{}^{120}/_{1000}} = 0.0333$

$[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{10^{-14}}{0.0333} = 3.0 \times 10^{-13}$

pH = <u>12.52</u>

11 NaOH = strong base

CH_3COOH = weak acid (low Ka + carboxylic acid)

mol HA = $0.30 \times \frac{100}{1000} = 0.03$

mol OH^- = $0.10 \times \frac{50}{1000} = 0.005$

	HA	+ OH^-	→ A^-	+ H_2O
Start	0.03	0.005		
End	0.025	-	0.005	

$$[H^+] = K_a \frac{[HA]}{[A^-]} = 1.74 \times 10^{-5} \times \frac{\frac{0.025}{150/1000}}{\frac{0.005}{150/1000}}$$

$$= 8.7 \times 10^{-5}$$

pH = 4.06

12 $Ba(OH)_2$ = strong base

$HCOOH$ = weak acid (low Ka + carboxylic acid)

mol HA = $0.100 \times \frac{20}{1000} = 0.002$

mol OH^- = $2 \times 0.05 \times \frac{50}{1000} = 0.005$

∴ XS OH^- = 0.003

XS $[OH^-]$ = $\frac{0.003}{70/1000} = 0.0429$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{10^{-14}}{0.0429} = 2.33 \times 10^{-13}$

pH = 12.63