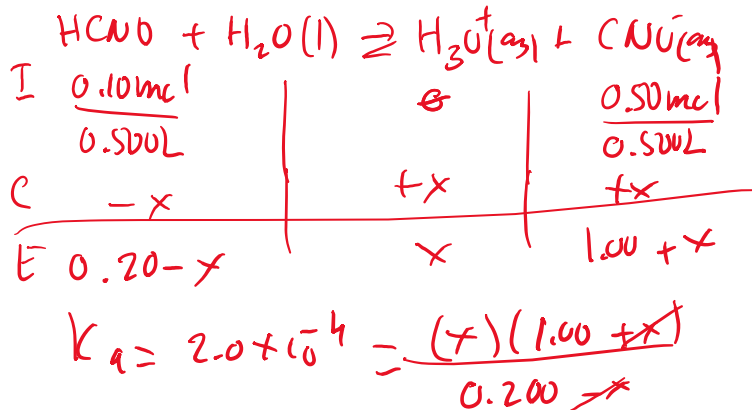


Due Thursday 4/30/20 at 6PM

*For each problem, show all necessary chemical reaction (hydrolysis in water and neutralization)

1. Calculate the pH of a buffer solution prepared by dissolving 0.10 moles of cyanic acid, HCNO, and 0.50 moles of sodium cyanate, NaCNO, in enough water to make 0.500 liter of solution. For HCNO, $K_a = 2.0 \times 10^{-4}$ at 25°C. Set up a chemical equation and an ICE table. Do not use Henderson-Hasselbalch.



$$x = 4.0 \times 10^{-5} \text{ (check } < 5\%)$$

$$\text{pH} = -\log(4.0 \times 10^{-5})$$

$$= \underline{4.40}$$

- a. Find the pH after 10.0 mL of 1.00 M KOH is added to the buffer from part (a). You can use Henderson-Hasselbalch after completing neutralization table.

mmol HCNO = 0.10 mol = 100 mmol
 mmol CNO⁻ = 0.50 mol = 500 mmol
 mmol OH⁻ = 10.0 mL × 1.00 M = 10.0 mmol

Neutralization



I	100 mmol	10.0 mmol	500 mmol
C	-10.0	-10.0	+10.0
E	90 mmol	\emptyset	510 mmol

Hydrolysis HCNO



I	$\frac{90 \text{ mmol}}{510 \text{ mL}}$	\emptyset	$\frac{510 \text{ mmol}}{510 \text{ mL}}$
C	$-x$	$+x$	$+x$
E	0.18 M	\emptyset	1.0 M + x

$$K_a = 2.0 \times 10^{-4} = \frac{(x)(1.0 + x)}{0.18}$$

$$x = 3.6 \times 10^{-5}$$

$$\text{pH} = -\log(3.6 \times 10^{-5})$$

$$= \underline{4.44} > 4.40$$

more basic

6. For the titration of 25.00 mL of 0.150 M HBr with 0.250 M NaOH:

a. Calculate the initial pH.

$$[H_3O^+] = 0.150 M$$

$$pH = -\log(0.150) = 0.824$$

b. How much NaOH is required to reach the equivalence point?

$$V_b = \frac{(0.150 M)(25.00 \text{ mL})}{0.250 M} = 15.0 \text{ mL}$$

c. What is the pH at the equivalent point?

$$pH = 7.00$$

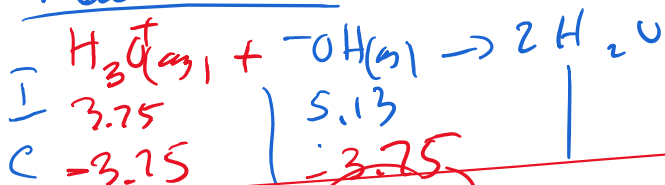
above equivalence

d. What is the pH after 20.50 mL of titrant has been added?

$$\text{mmol } H_3O^+ = 25.00 \text{ mL} \times 0.150 M = 3.75 \text{ mmol}$$

$$\text{mmol } -OH = 20.50 \text{ mL} \times 0.250 M = 5.13 \text{ mmol}$$

Neutralization



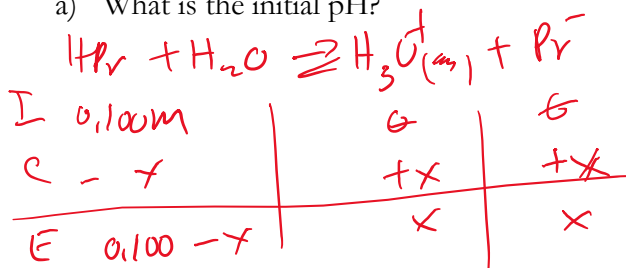
$$\text{new } [-OH] = \frac{1.38 \text{ mmol}}{45.5 \text{ mL}} = 0.0303$$

$$pH = -\log\left(\frac{1.0 \times 10^{-14}}{0.0303}\right) = 12.48$$

1. 40.0 mL of propionic acid (HPr) 0.100 M, $K_a = 1.3 \times 10^{-5}$, is titrated with 0.125 M NaOH.

Answer the following questions:

a) What is the initial pH?



$$pH = -\log(1.1 \times 10^{-3}) = 2.96$$

$$K_a = 1.3 \times 10^{-5} = \frac{(x)(x)}{0.100 - x}$$

$$x = 1.1 \times 10^{-3}$$

check < 5%

b) How many mL of NaOH are required to reach the equivalence point of this reaction?

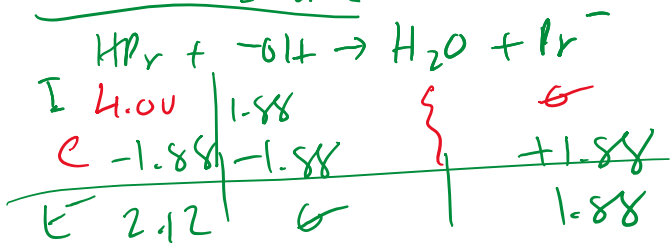
$$V_b = \frac{(40.0 \text{ mL})(0.100 M)}{0.125 M} = 32.0 \text{ mL}$$

c) What is the pH after adding 15.0 mL of NaOH?

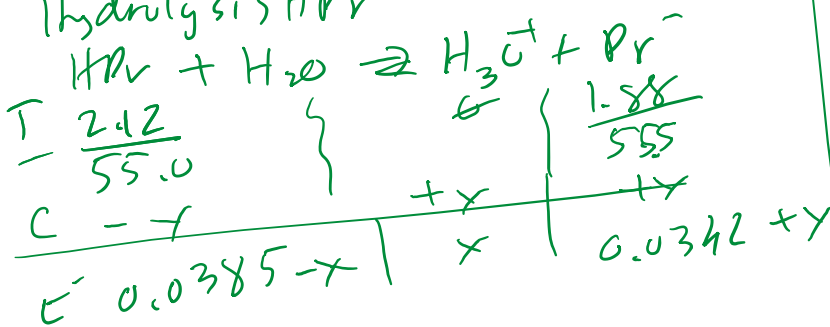
$$\text{mmol HPr} = 40.0 \text{ mL} \times 0.100 \text{ M} = 4.00 \text{ mmol}$$

$$\text{mmol OH}^- = 15.0 \text{ mL} \times 0.125 \text{ M} = 1.88 \text{ mmol}$$

Neutralization



Hydrolysis of HPr



$$K_a = 1.3 \times 10^{-5} = \frac{(x)(0.0342+x)}{0.0385-x}$$

$$x = 1.46 \times 10^{-5} \text{ (check)}$$

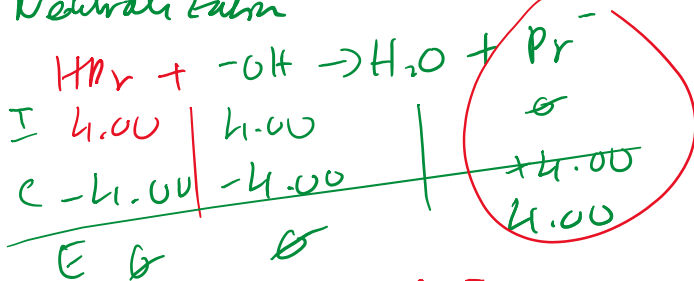
$$\text{pH} = -\log(1.46 \times 10^{-5})$$

$$= 4.83$$

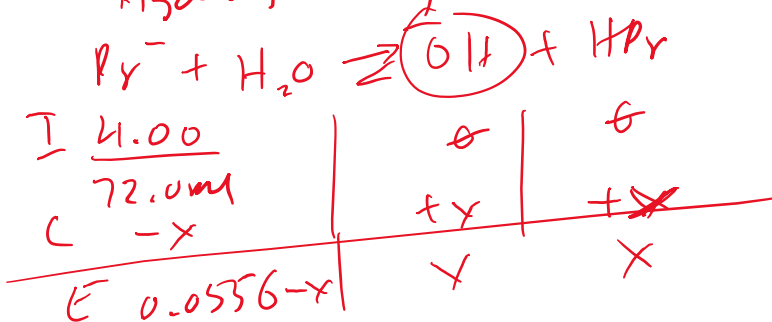
d) Calculate the pH at the equivalent point?

$$\text{mmol OH}^- = 32.0 \text{ mL} \times 0.125 \text{ M} = 4.00 \text{ mmol}$$

Neutralization



Hydrolysis of Pr⁻



$$K_b = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.6 \times 10^{-10}$$

$$K_b = 7.6 \times 10^{-10} = \frac{x^2}{0.0556-x}$$

$$x = 6.5 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = -\log(6.5 \times 10^{-6})$$

$$= 5.18$$

$$\text{pH} = 14.00 - 5.18$$

$$= 8.82$$

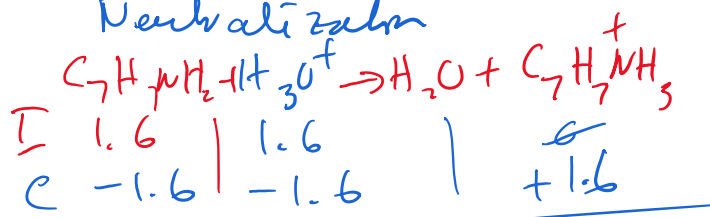
$$V_a = \frac{(25.0 \text{ mL})(0.065 \text{ M})}{0.050 \text{ M}} = 32.5 \text{ mL}$$

2. What is the pH of the solution obtained when 25.0 mL of 0.065 M benzylamine, $C_7H_7NH_2$, is titrated to the equivalence point with 0.050 M HCl. $K_b = 4.7 \times 10^{-10}$

$$\text{mmol } C_7H_7NH_2 = 25.0 \text{ mL} \times 0.065 \text{ M} = 1.6 \text{ mmol}$$

$$\text{mmol } H_3O^+ = 32.5 \text{ mL} \times 0.050 \text{ M} = 1.6 \text{ mmol}$$

Neutralization



$$K_a = \frac{1.0 \times 10^{-14}}{4.7 \times 10^{-10}} = 2.1 \times 10^{-5}$$

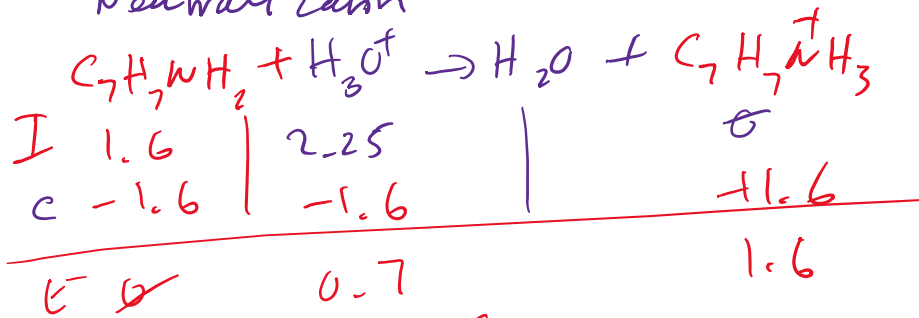
What is the pH at the midpoint?

$$\begin{aligned} \text{pH} &= \text{p}K_a \\ &= -\log(2.1 \times 10^{-5}) \\ &= 4.68 \end{aligned}$$

What is the pH after 45.0 mL of titrant was added?

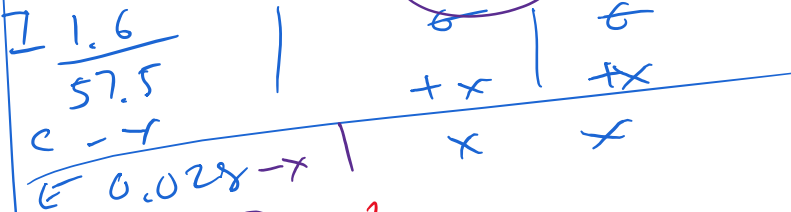
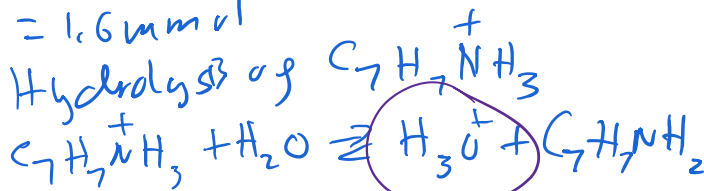
$$\text{mmol } H_3O^+ = 45.0 \text{ mL} \times 0.050 \text{ M} = 2.25 \text{ mmol}$$

Neutralization



stronger acid

$$\begin{aligned} [H_3O^+] &= \frac{0.7 \text{ mmol}}{85.0 \text{ mL}} \\ &= 0.008 \text{ M} \end{aligned}$$



$$2.1 \times 10^{-5} = \frac{x^2}{0.028-x}$$

$$\sqrt{5.9 \times 10^{-7}} = \sqrt{x^2}$$

$$x = 7.7 \times 10^{-4}$$

$$\text{pH} = -\log(7.7 \times 10^{-4}) = 3.11$$