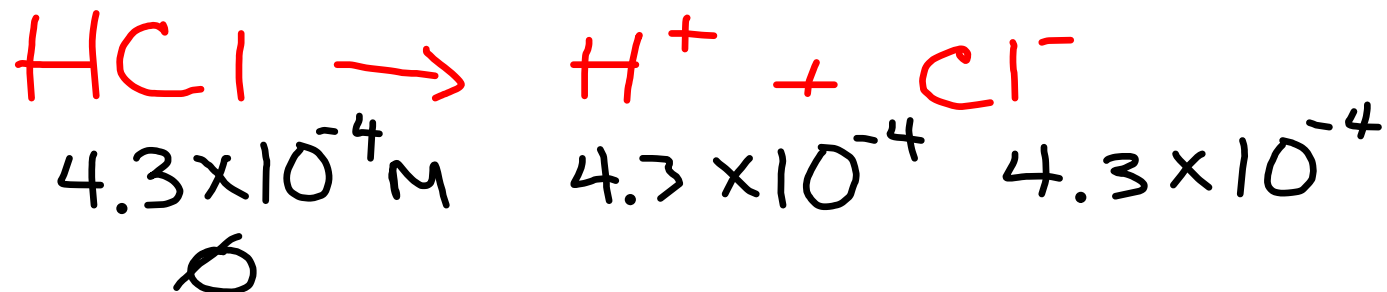


1.



$$[\text{H}^+] = 4.3 \times 10^{-4}$$

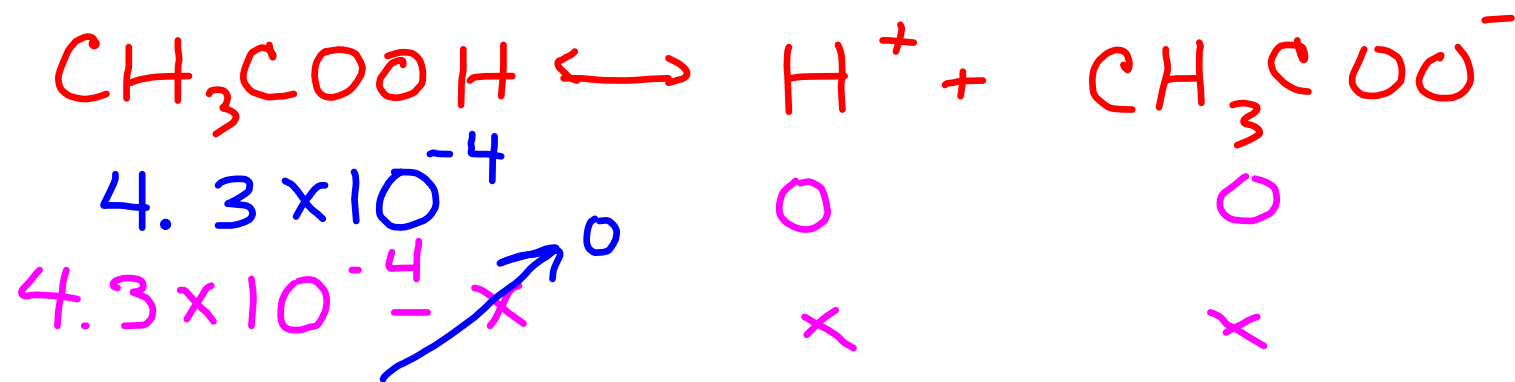
$$\text{pH} = 3.36$$

$$\text{pOH} = 10.64$$

$$[\text{OH}^-] = 2.29 \times 10^{-11}$$

2.3×10^{-11}

$$2. \quad K_a = 1.8 \times 10^{-5}$$



$$1.8 \times 10^{-5} = \frac{x^2}{4.3 \times 10^{-4}}$$

$$x = 8.79 \times 10^{-5}$$

$$[\text{H}^+] = 8.8 \times 10^{-5} \quad \text{pH} = 4.06$$

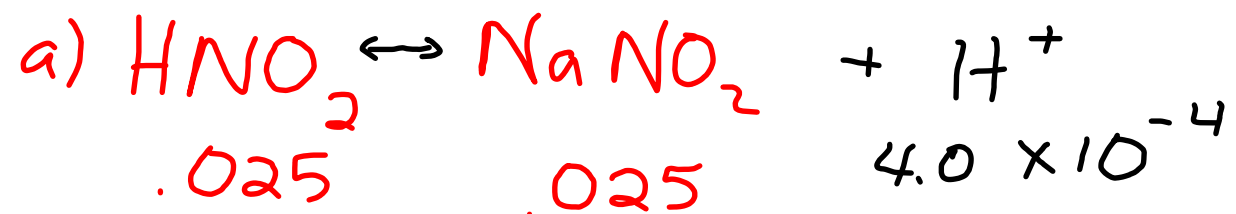
$$[\text{OH}^-] = 1.1 \times 10^{-10} \quad \text{pOH} = 9.94$$

$$\text{pH} = \text{pK}_a + \log \frac{[\text{B}]}{[\text{A}]}$$

↑
log of the K_a value.

$$\text{pH} = \text{pK}_a - \log \frac{[\text{A}]}{[\text{B}]}$$

1.



$$K_a \ 4.0 \times 10^{-4}$$

$$3.39 + \log\left(\frac{.025}{.025}\right)$$

We can lower or increase the pH by no more than 1.

log of equal amounts of
Acid to Base is 0.

$$\text{pH} = 3.39 \quad [\text{H}^+] = 4.0 \times 10^{-4}$$

$$\text{pOH} = 10.61 \quad [\text{OH}^-] = 2.5 \times 10^{-11}$$

1b

$$pH = pK_a + \log \frac{[B]}{[A]}$$

$$4.0 \times 10^{-4} = 3.39 + \log \frac{[.01]}{[.025]}$$

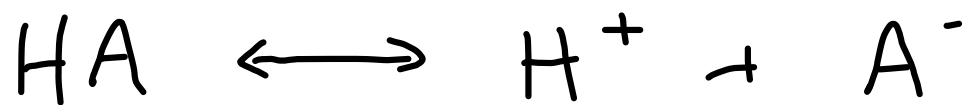
$$\text{3.39} - .39 = .3978$$

$$.40$$

$$pH = 3.00 \quad [H^+] = 1.0 \times 10^{-3}$$

$$pOH = 11.00 \quad [OH^-] = 1.0 \times 10^{-11}$$

23. K_a



.25

1.3489×10^{-2}

1.3489×10^{-2}

$pOH = 12.13$

$pH = 1.87$

$$= \frac{(1.3489 \times 10^{-2})^2}{.25}$$

7.278×10^{-4}

$K_a = 7.3 \times 10^{-4}$

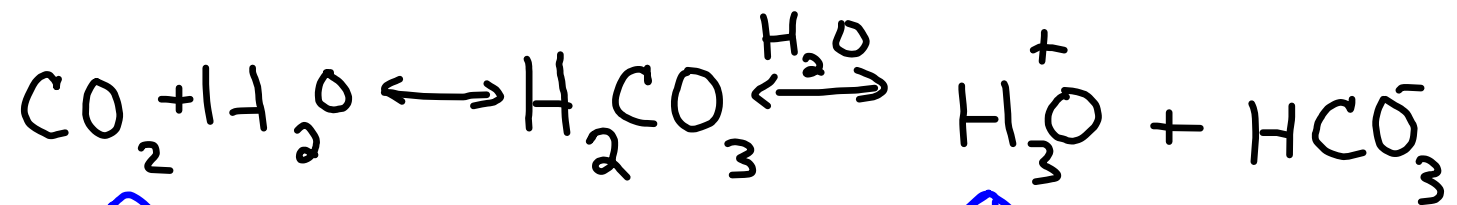
Example of sig figs.

$$\text{pH} = 4.\underline{73}8 \quad \text{pOH} = 9.\underline{26}2$$

$$[\text{H}^+] = 1.83 \times 10^{-5} \quad [\text{OH}^-] = 5.47 \times 10^{-10}$$

Blood Buffer

$$\text{pH} = 7.35 \text{ to } 7.45$$



↑
produced
as an end
product
of cellular
metabolism

↑
excess acid
enters blood
fluids and reacts
w/ HCO_3^-

3 c. buffer problem w/
weak acid and weak base

$$\text{pH} = 3.40 + \log\left(\frac{.057}{.132}\right)$$

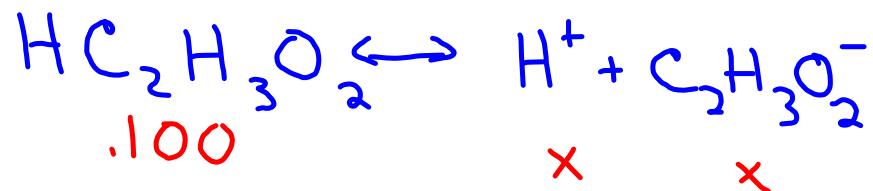
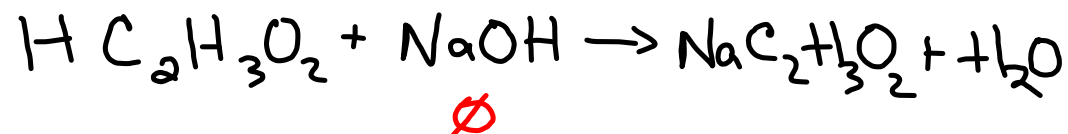
$$= 3.40 + (-.3646)$$

$$= 3.035$$

$$= 3.04$$

4 a.

0 mL



$$1.8 \times 10^{-5} = \frac{x^2}{.100}$$

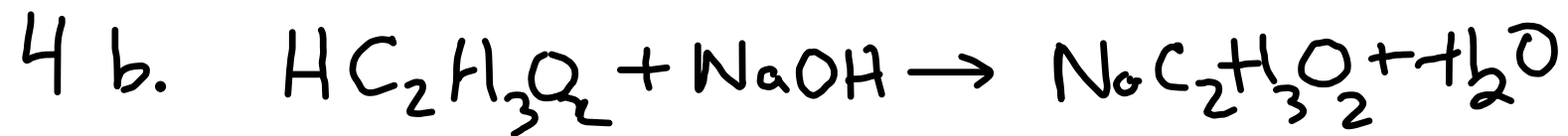
$$x = 1.342 \times 10^{-3}$$

$$[\text{H}^+] = 1.342 \times 10^{-3} \quad 1.3 \times 10^{-3}$$

$$\text{pH} = 2.87$$

$$\text{pOH} = 11.13$$

taken
from the
 K_a value



5 mL

$$.100 = \frac{x}{.03}$$

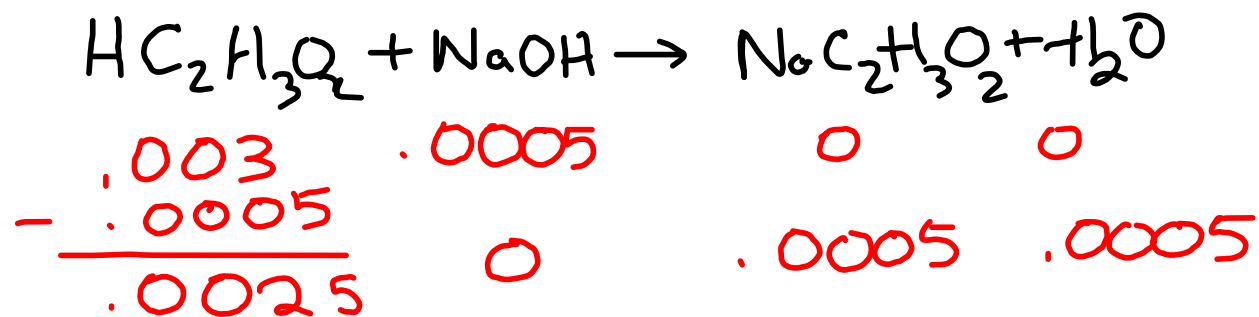
$$\frac{30 \text{ mL} \times 1 \text{ L}}{1000 \text{ mL}}$$

$$x = .003 \text{ mol}$$

$$\frac{x}{.005} = .100$$

$$\frac{5 \text{ mL} \times 1 \text{ L}}{1000 \text{ mL}}$$

$$x = .0005 \text{ mol}$$



$$\text{p}K_a = \log(1.8 \times 10^{-5})$$

$$\text{p}K_a = 4.745$$

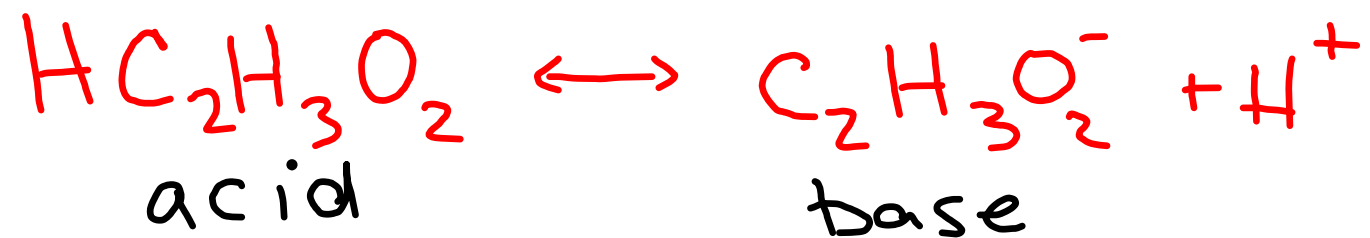
$$\text{pH} = \text{p}K_a + \log \frac{[\text{B}]}{[\text{A}]}$$

$$\text{pH} = 4.745 + \log \left(\frac{.0005}{.0025} \right)$$

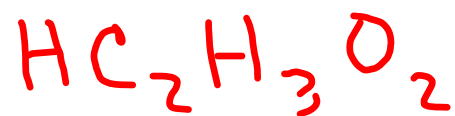
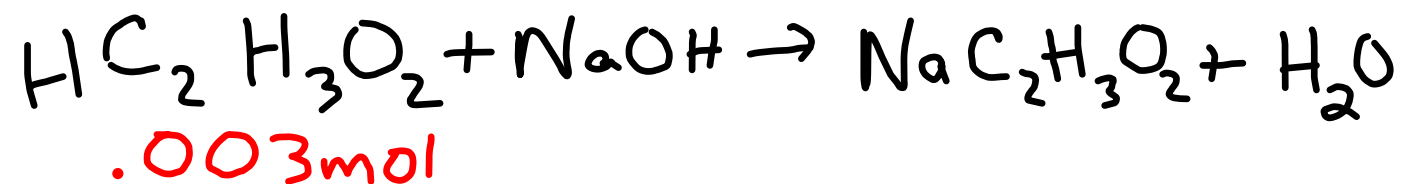
$$= 4.745 + (-.778)$$

$$\text{pH} = 4.046$$

$$\text{pH} = 4.05 \quad \text{pOH} = 9.95$$



4 c.



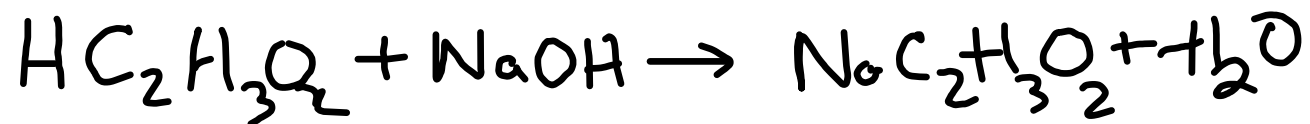
Stays as

.003

$$\frac{x}{.015} = .100$$

$$x = .0015 \text{ mol}$$

$$K_a = 1.8 \times 10^{-4}$$



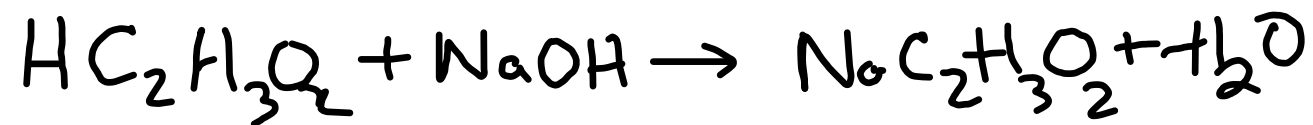
$$\begin{array}{r}
 .003 \\
 - .0015 \\
 \hline
 .0015
 \end{array}
 \quad
 \begin{array}{r}
 .0015 \\
 0
 \end{array}
 \quad
 \begin{array}{r}
 0 \\
 .0015
 \end{array}
 \quad
 \begin{array}{r}
 0 \\
 .0015
 \end{array}$$

$$\text{pH} = 4.745 + \log\left(\frac{.0015}{.0015}\right)$$

$$\text{pH} = \text{p}K_a = 4.75$$

$$\text{pOH} = 9.26$$

25 mL

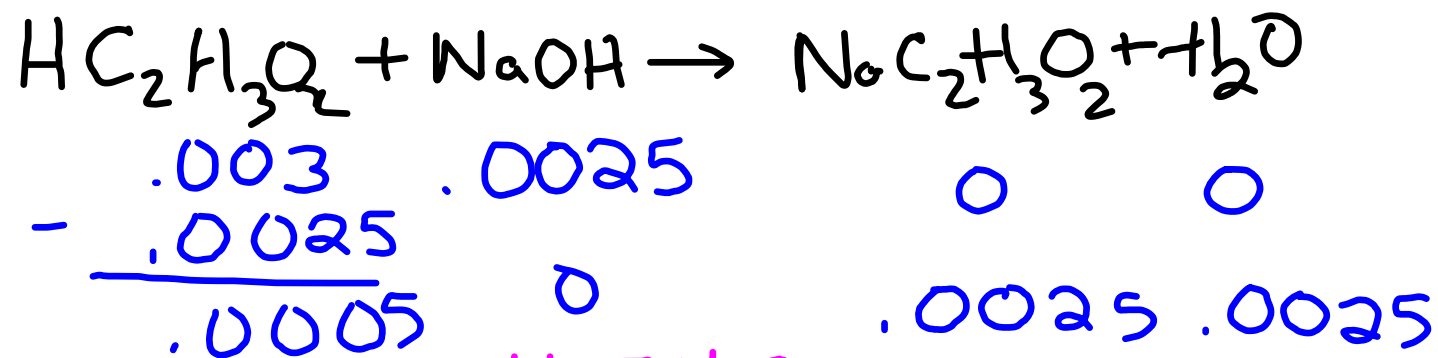


.003

$$\frac{x}{.025} = .100$$

$$x = .0025 \text{ mol}$$

25mL



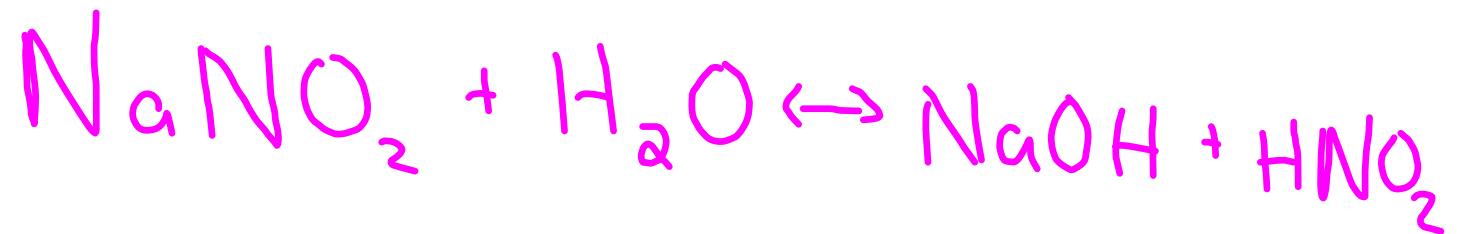
$$\text{pH} = \overset{4.745}{\text{p}K_a} + \log\left(\frac{.0025}{.0005}\right)$$

$$\text{pH} = 5.444$$

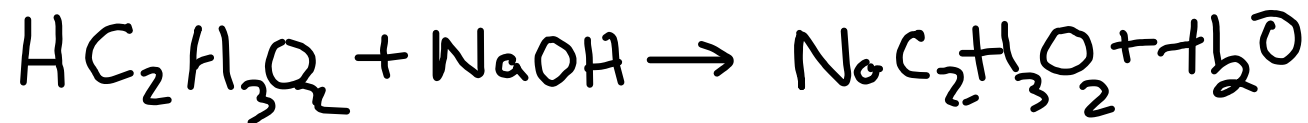
$$\text{pH} = 5.44 \quad \text{pOH} = 8.56$$

1) $\text{pH} = 4.814$

2) $\text{pH} = 8.75$



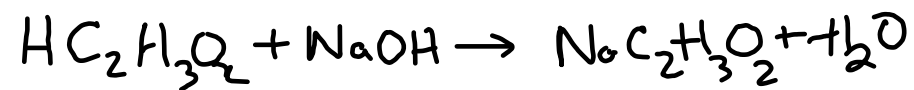
30 mL



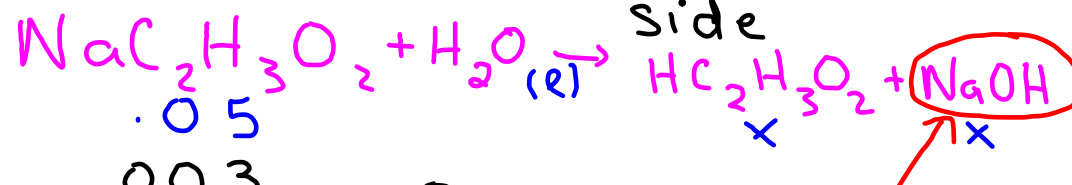
$$\begin{array}{r} .003 \\ .003 \\ \hline \times \\ \hline .03 \end{array} = .100$$
$$x = .003$$

$\begin{array}{r} .003 \\ \hline .060 \end{array}$
equivalence
pt.

The acid and
base are in
equal amounts



.003 has gone to the other side



$$\frac{.003}{.060} = .05 \text{ M}$$

$$K_b = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}}$$

$$K_b = 5.55\bar{5} \times 10^{-10}$$

$$\frac{x^2}{.05} = 5.556 \times 10^{-10}$$

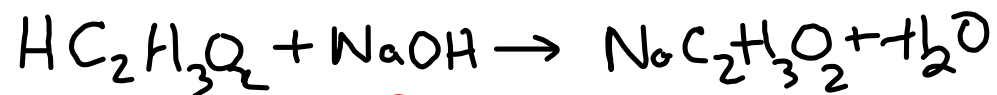
$$x = 5.27 \times 10^{-6}$$

$$\text{pOH} = 5.28 \quad \text{pH} = 8.72$$

strong base use K_b

31 mL

$$\frac{x}{.031} = .1 \quad x = .0031$$



.003 mol .0031 mol

.0031 mol .0001 mol

~~-.0031 mol~~

$$\frac{.0001 \text{ mol}}{.061 \text{ L}} = 1.6 \times 10^{-3} \text{ M}$$

30 mL + 31 mL =

pOH = 2.79 pH = 11.21