



$[\text{HCl}]_0 = [\text{H}^+] = 0.020 \text{ M}$

\rightarrow $[\text{H}^+]$ from HCl is negligible

$\text{pH} = -\log(0.020) = 1.70$



~~H^+~~

$\text{C}_3\text{H}_5\text{O}_2^-$

H^+

~~H^+~~



I 0.020 0.100 0

C -0.020 -0.020 +0.020

E 0 0.080 0.020

\rightarrow what's left is a buffer solution



I 0.020 0.080 0

C -x +x +x

E 0.020 - x 0.080 + x x

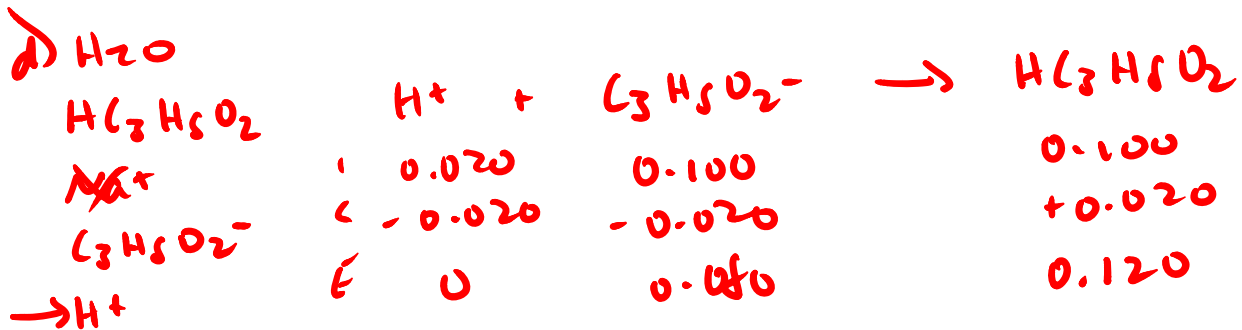
$K_a = \frac{[\text{C}_3\text{H}_5\text{O}_2^-][\text{H}^+]}{[\text{HC}_3\text{H}_5\text{O}_2]}$

$1.3 \times 10^{-5} = \frac{(0.080 + x)(x)}{0.020 - x}$

$x = 3.25 \times 10^{-6} \text{ M} = [\text{H}^+]$

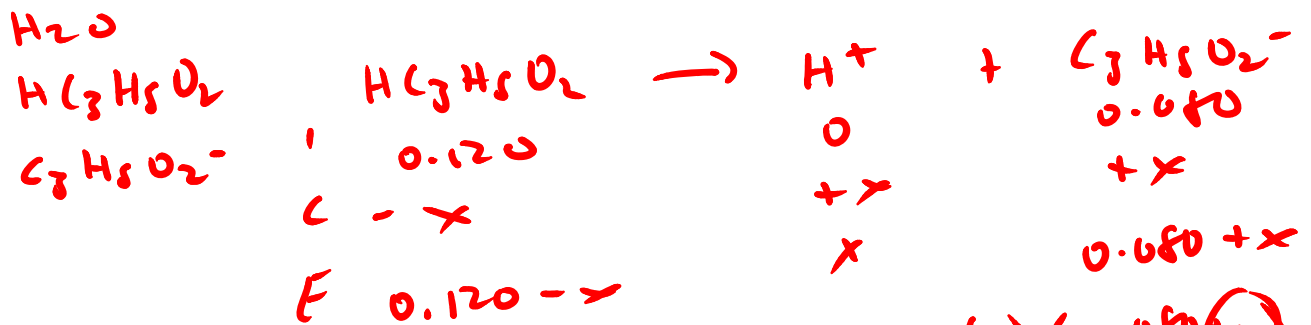
$\text{pH} = -\log(3.25 \times 10^{-6}) = 5.49$

$$c) [H^+] = 0.020 \text{ M} \quad pH = -\log(0.020) = 1.70$$



$\rightarrow H^+$
 Cl^-

the buffer consumed all the acid and excess buffer remains



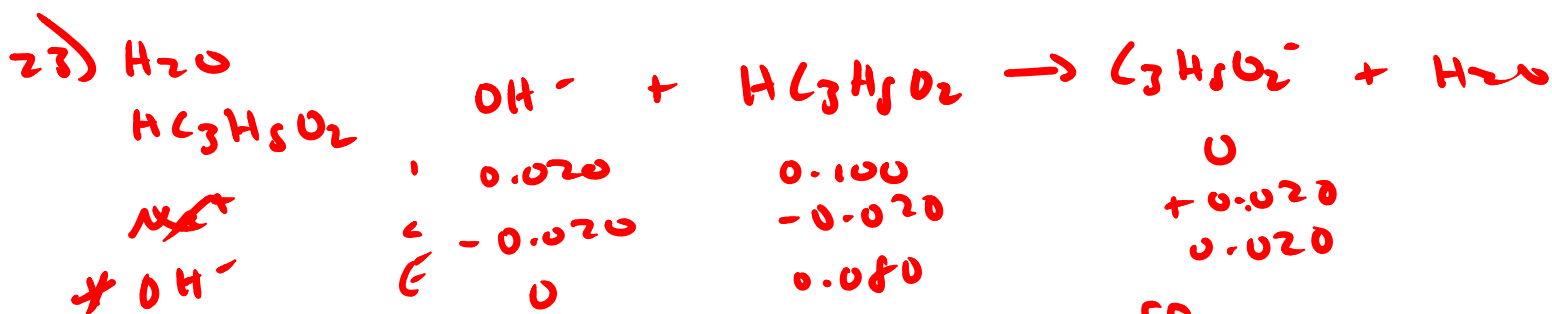
$$K_a = \frac{[H^+][C_3H_5O_2^-]}{[H(C_3H_5O_2)]} = 1.3 \times 10^{-5} = \frac{(x)(0.080+x)}{0.120-x}$$

$$x = 1.95 \times 10^{-5} \text{ M} = [H^+]$$

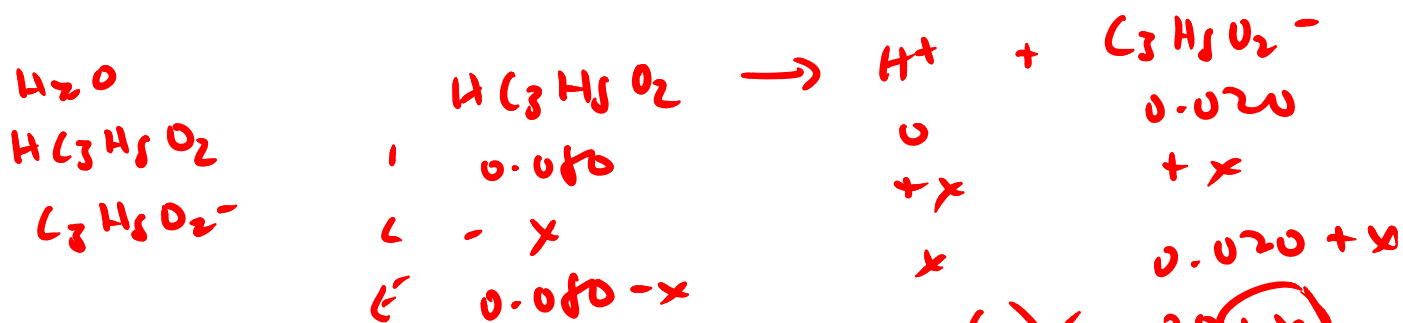
$$pH = -\log(1.95 \times 10^{-5}) = 4.70$$

$$\text{So } \frac{1.95 \times 10^{-5}}{0.120} \times 100 = 0.016\%$$

✓



what remains is a buffer

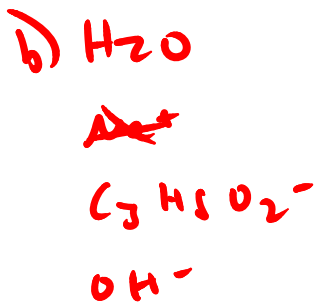


$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_2^-]}{[\text{HC}_3\text{H}_5\text{O}_2]} = 1.3 \times 10^{-5} = \frac{(x)(0.020 + x)}{0.080 - x}$$

$$x = 5.2 \times 10^{-5} \text{ M} = [\text{H}^+]$$

$$\text{pH} = -\log(5.2 \times 10^{-5}) = 4.28$$

$$\% = \frac{5.2 \times 10^{-5}}{0.080} \times 100 = 0.065\% \checkmark$$

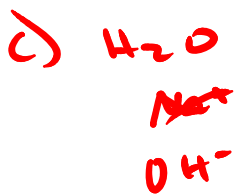


The amount of the OH^- from weak base is negligible relative to the OH^- from the strong base

$$[\text{NaOH}]_0 = [\text{OH}^-] = 0.020 \text{ M}$$

$$\text{pOH} = -\log(0.020) = 1.70$$

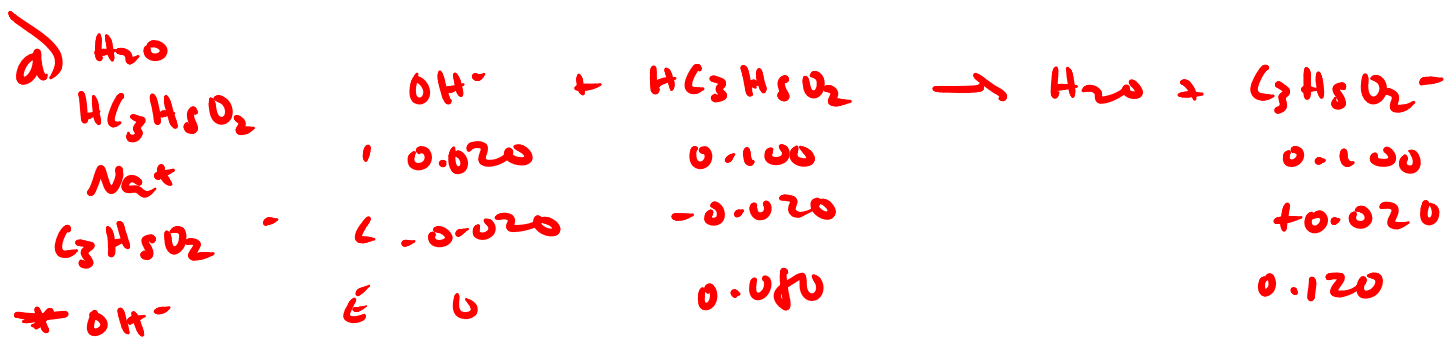
$$\text{pH} = 14 - 1.70 = 12.30$$



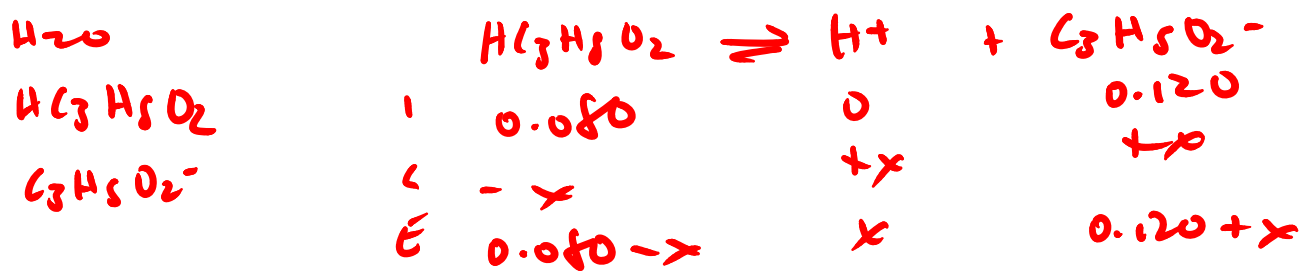
$$[\text{NaOH}]_0 = [\text{OH}^-] = 0.020 \text{ M}$$

$$\text{pOH} = -\log(0.020) = 1.70$$

$$\text{pH} = 14 - 1.70 = 12.30$$



what's left is a buffer solution



$$K_a = \frac{[H^+][C_3H_5O_2^-]}{[HCl_3H_5O_2]} = 1.3 \times 10^{-5} = \frac{(x)(0.120+x)}{(0.080-x)}$$

$$x: 8.67 \times 10^{-6} M = [H^+]$$

$$pH = -\log(8.67 \times 10^{-6}) = 5.06$$

$$5\% \quad \frac{8.67 \times 10^{-6}}{0.080} \times 100 = 0.011\% \quad \checkmark$$

25)

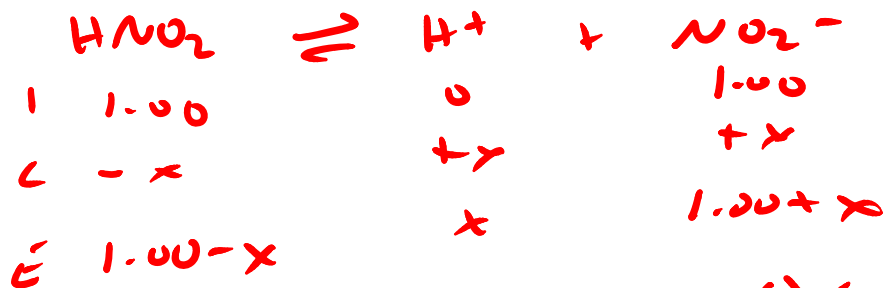
Solution	initial pH	after added H+	after added OH-
a	2.96	1.70	4.29
b	8.94	5.49	12.30
c	7.00	1.70	12.30
d	4.89	4.71	5.07

27) H₂O

HNO₂

NO₂⁺

NO₂⁻

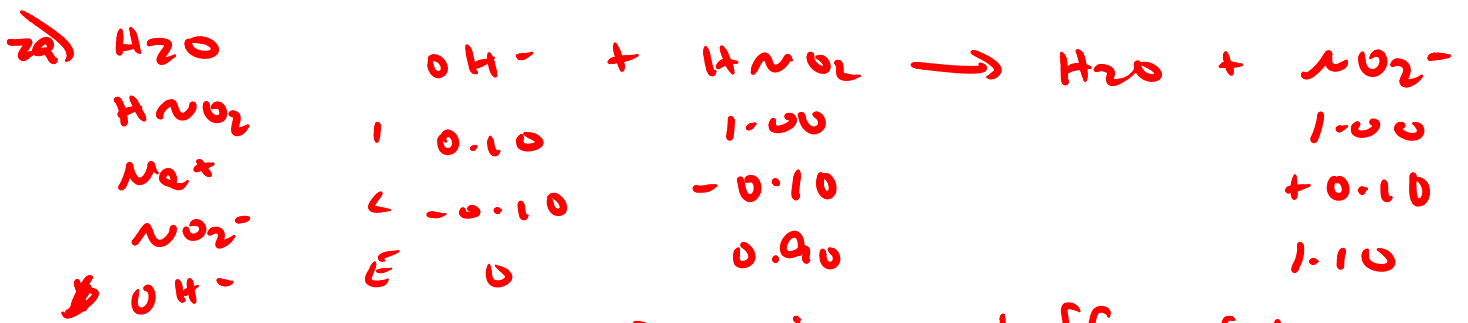


$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = 4.0 \times 10^{-4} = \frac{(x)(1.00+x)}{1.00-x}$$

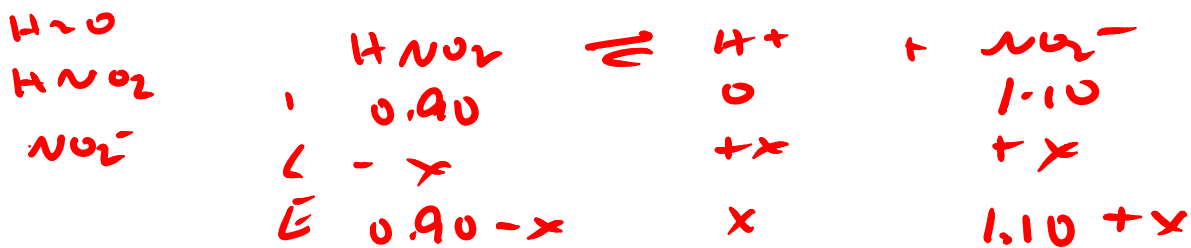
$$x = 4.0 \times 10^{-4} \text{ M} = [\text{H}^+]$$

$$\% \text{ ion} = \frac{4.0 \times 10^{-4}}{1.00} \times 100 = 4.0 \times 10^{-2} \% \checkmark$$

$$\text{pH} = -\log(4.0 \times 10^{-4}) = 3.40$$



what's left is a buffer solution

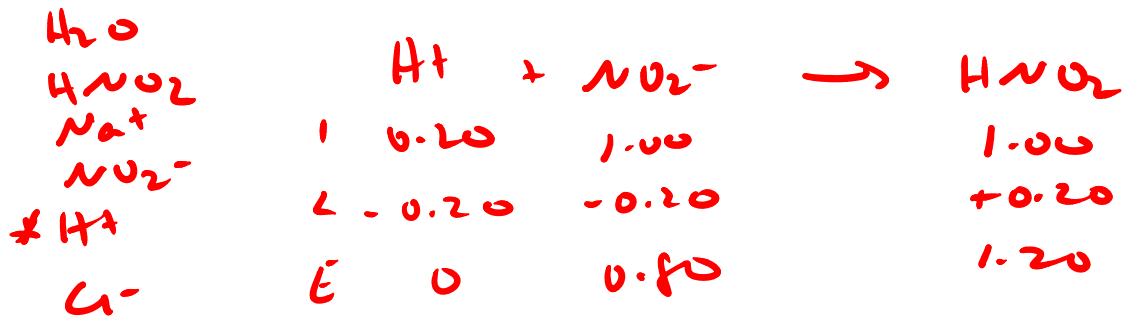


$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = 4.0 \times 10^{-4} = \frac{(x)(1.10+x)}{0.90-x}$$

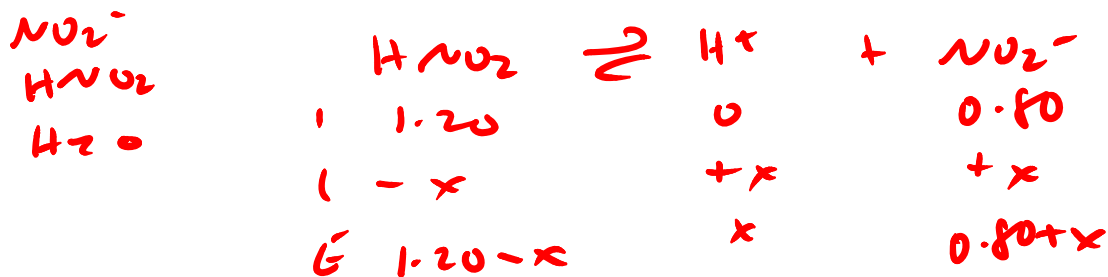
$$x = 3.3 \times 10^{-4} \text{ M} = [\text{H}^+]$$

$$\% = \frac{3.3 \times 10^{-4}}{0.90} \times 100 = 0.036\% \checkmark$$

$$\text{pH} = -\log(3.3 \times 10^{-4}) = 3.48$$



what's left is a buffer solution



$$K_a = \frac{[H^+][NO_2^-]}{[HNO_2]} = 4.0 \times 10^{-4} = \frac{(x)(0.80+x)}{1.20-x}$$

$$x = 6.0 \times 10^{-4} \text{ M} = [H^+]$$

$$\% = \frac{6.0 \times 10^{-4}}{1.20} \times 100 = 0.05\% \checkmark$$

$$pH = -\log(6.0 \times 10^{-4}) = 3.22$$