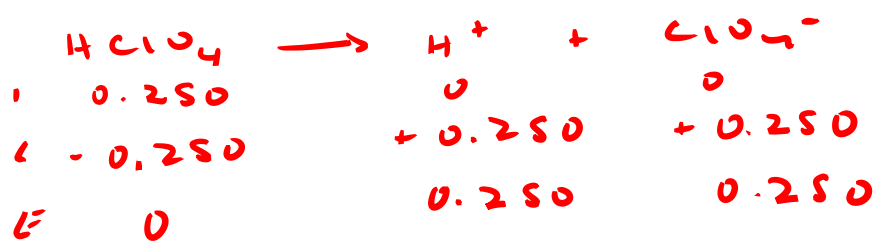
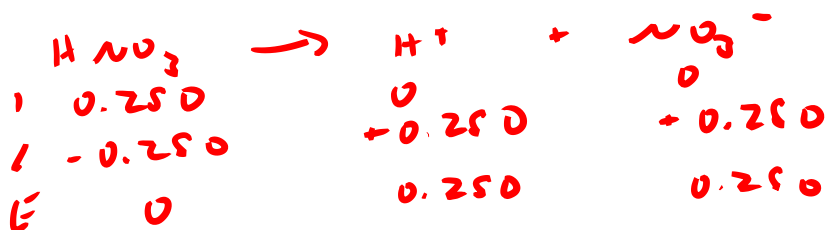


53) a) H^+ , ClO_4^- , H_2O



$$pH = -\log(0.250) = 0.602$$

b) H^+ , NO_3^- , H_2O



$$pH = -\log(0.250) = 0.602$$

55) a) $[HCl] = [H^+] = 0.10M$

$$pH = -\log(0.10) = 1.00$$

b) $[HClO_4] = [H^+] = 5.0M$

$$pH = -\log(5.0) = -0.70$$

c) $[HI] = [H^+] = 1.0 \times 10^{-11}$

$$pH = -\log(1.0 \times 10^{-11}) = 11.00 \text{ not possible}$$

* $[H^+]$ in H_2O is $1.00 \times 10^{-7}M$

$$pH = -\log(1.0 \times 10^{-7}) = 7$$

$$57) [H^+] = 10^{-pH} = 10^{-2.50} = 3.2 \times 10^{-3} M$$

$$[H^+] = [HI] = 3.2 \times 10^{-3} M$$

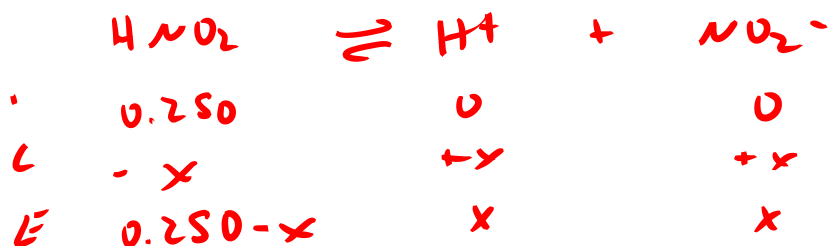
$$58) [H^+] = 10^{-1.50} = 3.16 \times 10^{-2} M$$

$$M_1 V_1 = M_2 V_2$$

$$(12 M)(x) = (3.16 \times 10^{-2} M)(1.6 L)$$

$$x = 4.2 \times 10^{-3} L$$

to 4.2 ml of 12 M HI, add enough water to make 1.6 L of solution



$$K_a = \frac{[H^+][NO_2^-]}{[HNO_2]} = 4.0 \times 10^{-4} = \frac{x^2}{(0.250-x)}$$

$$4.0 \times 10^{-4} = \frac{x^2}{0.250}$$

$$x = 0.010 M$$

$$pH = -\log(0.010)$$

$$pH = 2.00$$

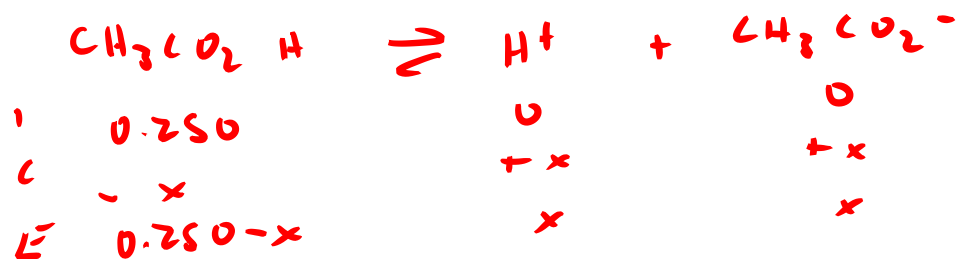
$$\% \text{ ionized} = \frac{0.010}{0.250} \times 100 = 4.0 \%$$

b)



$$1.8 \times 10^{-5}$$

$$1.0 \times 10^{-14}$$



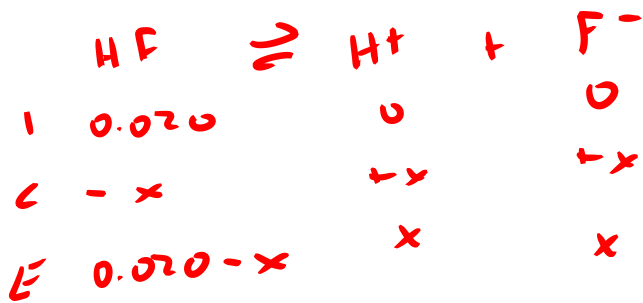
$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]} = 1.8 \times 10^{-5} = \frac{x^2}{0.250 - x}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.250} \quad x = 2.1 \times 10^{-3} \text{ M}$$

$$\% \text{ ion} = \frac{2.1 \times 10^{-3}}{0.250} \times 100 = 0.84\%$$

$$[\text{H}^+] = 2.1 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log(2.1 \times 10^{-3}) = 2.68$$



$$K_a = \frac{[H^+][F^-]}{[HF]} = 7.2 \times 10^{-4} = \frac{x^2}{0.020-x} \quad x = 3.79 \times 10^{-3}$$

$$\% = \frac{3.79 \times 10^{-3}}{0.020} \times 100 = 19\% \quad ;$$

quadratic

$$7.2 \times 10^{-4} = \frac{x^2}{0.020-x}$$

$$1.4 \times 10^{-5} - 7.2 \times 10^{-4}x = x^2$$

$$x^2 + 7.2 \times 10^{-4}x - 1.4 \times 10^{-5} = 0$$

$$\frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-7.2 \times 10^{-4} \pm \sqrt{(7.2 \times 10^{-4})^2 - 4(1)(-1.4 \times 10^{-5})}}{2(1)}$$

$$\frac{-7.2 \times 10^{-4} + \sqrt{5.65 \times 10^{-5}}}{2} = \frac{-7.2 \times 10^{-4} - \sqrt{5.65 \times 10^{-5}}}{2}$$

$$3.4 \times 10^{-3} \quad -4.7 \times 10^{-3}$$

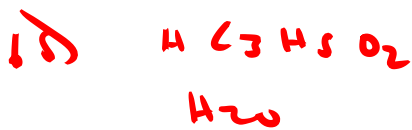
$$[H^+] = [F^-] = 3.4 \times 10^{-3} \text{ M}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{3.4 \times 10^{-3}} = 2.9 \times 10^{-12} \text{ M}$$

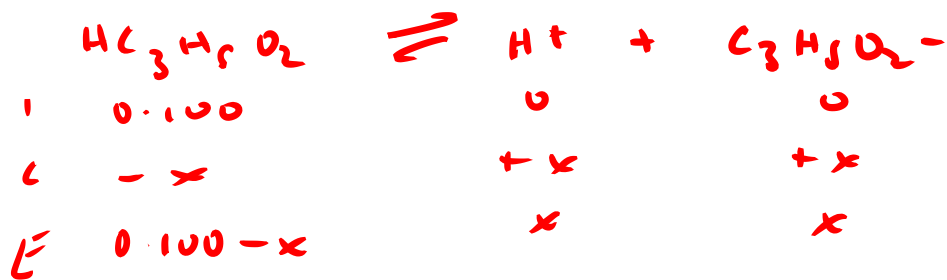
$$[HF] = 0.020 - 0.0034 = 0.017 \text{ M}$$

$$pH = -\log(3.4 \times 10^{-3})$$

$$= 2.46$$



$K_a = 1.3 \times 10^{-5}$ *
 $K_w = 1.0 \times 10^{-14}$



$$K_a = 1.3 \times 10^{-5} = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_2^-]}{[\text{HC}_3\text{H}_5\text{O}_2]} = \frac{x^2}{0.100 - x}$$

$$x = [\text{H}^+] = 1.1 \times 10^{-3} \text{ M}$$

$$\% \text{ dissociation} = \frac{1.1 \times 10^{-3}}{0.100} \times 100 = 1.1\% \checkmark \text{ \% dissociation}$$

$$[\text{H}^+] = [\text{C}_3\text{H}_5\text{O}_2^-] = 1.1 \times 10^{-3} \text{ M}$$

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

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{1.1 \times 10^{-3}} = 9.1 \times 10^{-12} \text{ M}$$

$$[\text{HC}_3\text{H}_5\text{O}_2] = 0.100 \text{ M}$$