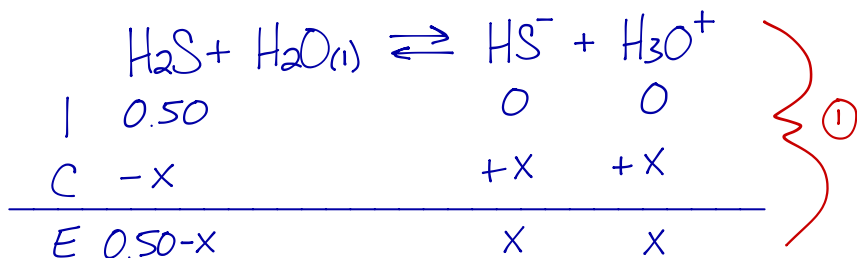


## CHEMISTRY 12 – CALCULATIONS INVOLVING $K_A$ & $K_B$ WORKSHEET

1) Calculate the pH of 0.50 M  $H_2S$ . (4 marks)



assume  
 $0.50 - x = 0.50$   
 (0.5)

$$K_a = \frac{[HS^-][H_3O^+]}{[H_2S]} \quad (0.5)$$

$$9.1 \times 10^{-8} = \frac{x^2}{0.50}$$

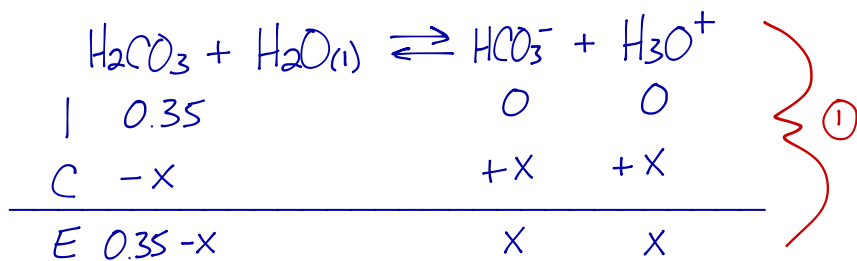
$$x = 2.1 \times 10^{-4} M \quad (1)$$

$$pH = -\log[H_3O^+]$$

$$= -\log(2.1 \times 10^{-4})$$

$$= 3.67 \quad (1)$$

2) Calculate the pOH of 0.35 M  $H_2CO_3$ . (5 marks)



assume  
 $0.35 - x = 0.35$   
 (0.5)

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]} \quad (0.5)$$

$$4.3 \times 10^{-7} = \frac{x^2}{0.35}$$

$$x = 3.9 \times 10^{-4} M \quad (1)$$

$$pH = -\log[H_3O^+]$$

$$= -\log(3.9 \times 10^{-4})$$

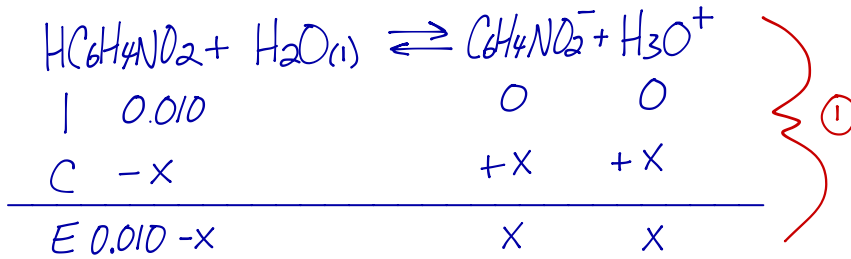
$$= 3.41 \quad (1)$$

$$pH + pOH = 14.00$$

$$3.41 + x = 14.00$$

$$x = 10.59 \quad (1)$$

- 3) Nicotinic acid,  $\text{HC}_6\text{H}_4\text{NO}_2$ , is a weak acid found in vitamin B. Calculate the pH of 0.010 M  $\text{HC}_6\text{H}_4\text{NO}_2$  ( $K_a = 1.4 \times 10^{-5}$ ). (4 marks)



assume  
 $0.010 - x = 0.010$  (0.5)

$$K_a = \frac{[\text{C}_6\text{H}_4\text{NO}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_6\text{H}_4\text{NO}_2]} \quad (0.5)$$

$$1.4 \times 10^{-5} = \frac{x^2}{0.010}$$

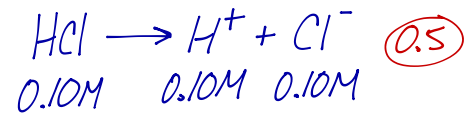
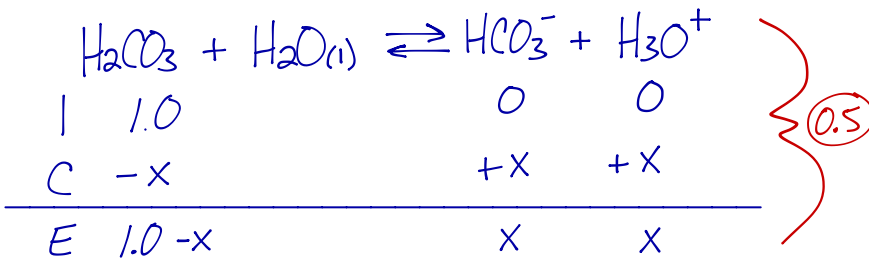
$$x = 3.7 \times 10^{-4} \text{ M} \quad (1)$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$= -\log(3.7 \times 10^{-4})$$

$$= 3.43 \quad (1)$$

- 4) Using calculations, show why the electrical conductivity of 1.0 M  $\text{H}_2\text{CO}_3$  will be less than that for 0.10 M HCl. (4 marks)



Total [ion] is:  
 $0.10 + 0.10 = 0.20 \text{ M} \quad (0.5)$

assume  
 $1.0 - x = 1.0$  (0.5)

$$K_a = \frac{[\text{HCO}_3^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{CO}_3]} \quad (0.5)$$

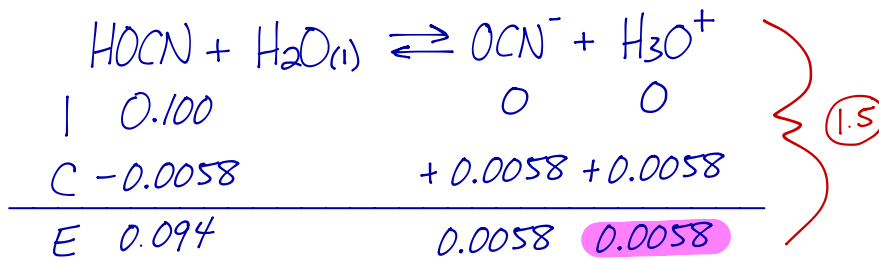
$$4.3 \times 10^{-7} = \frac{x^2}{1.0}$$

$$x = 6.6 \times 10^{-4} \text{ M} \quad (0.5)$$

Total [ion] is:  
 $6.6 \times 10^{-4} + 6.6 \times 10^{-4} = 1.3 \times 10^{-3} \text{ M} \quad (0.5)$

$\text{H}_2\text{CO}_3$  has a smaller [ion]  $\therefore$  will have a lower conductivity (0.5)

- 5) A solution of 0.100 M HO CN has a pH of 2.24. Calculate the  $K_a$  value for this acid. (4 marks)



$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$= 10^{-2.24}$$

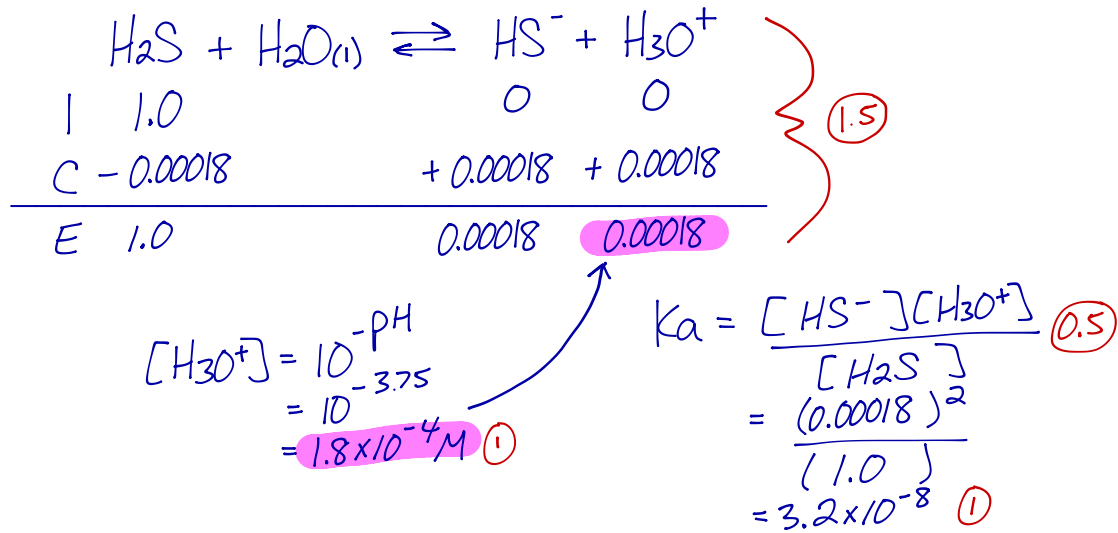
$$= 5.8 \times 10^{-3} \text{ M} \quad (1)$$

$$K_a = \frac{[\text{OCN}^-][\text{H}_3\text{O}^+]}{[\text{HO CN}]} \quad (0.5)$$

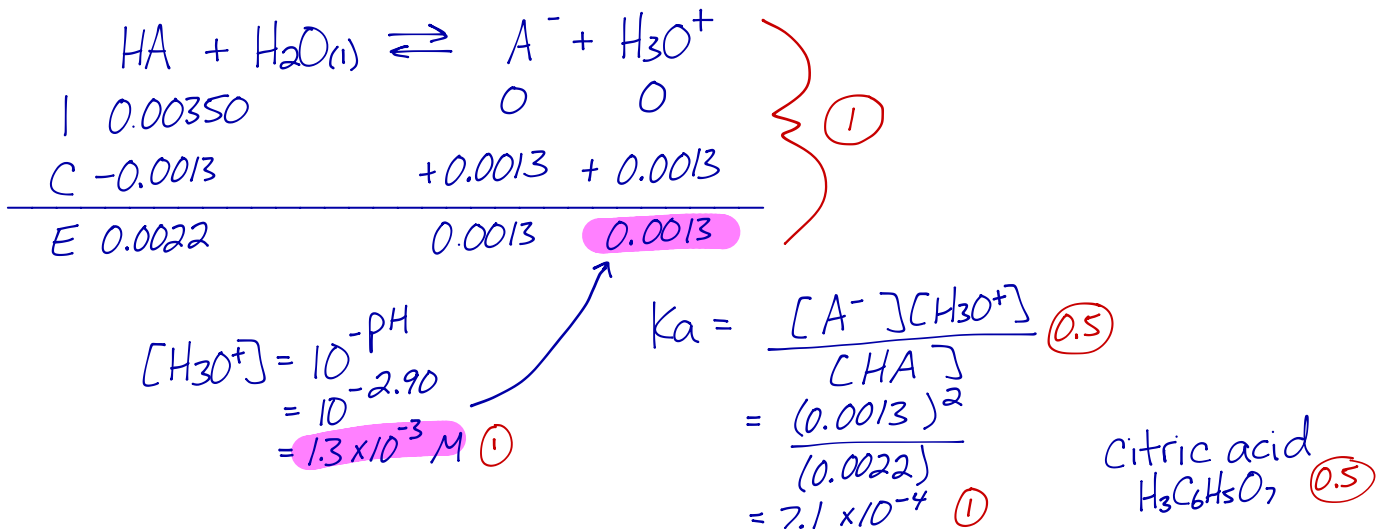
$$= \frac{(0.0058)^2}{(0.094)}$$

$$= 3.5 \times 10^{-4} \quad (1)$$

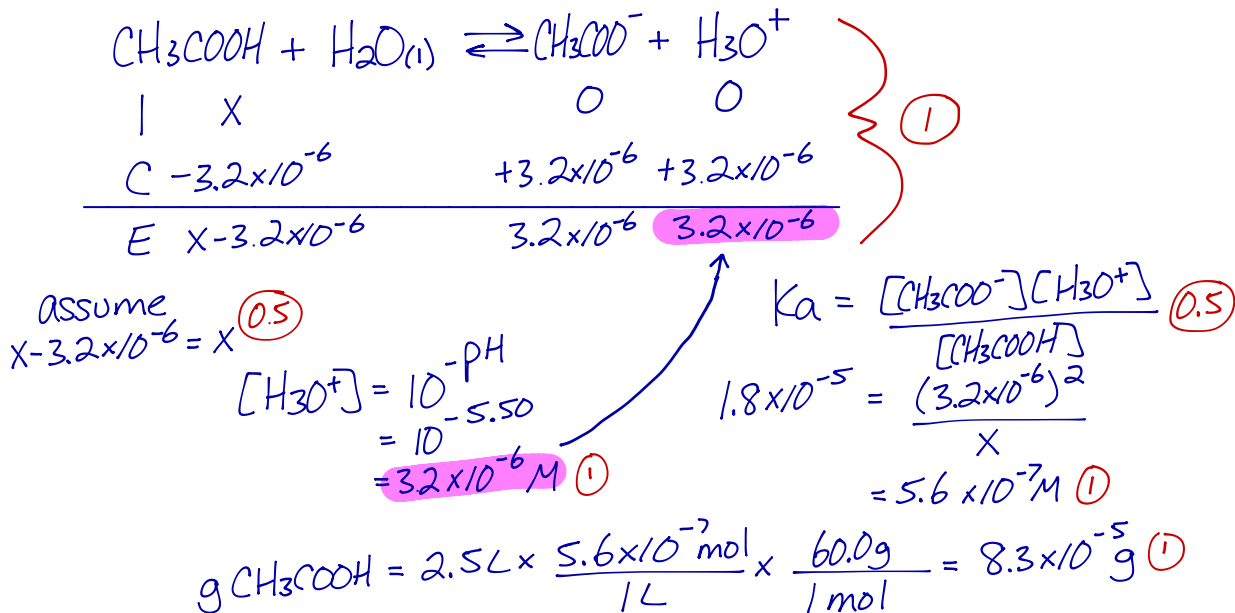
- 6) At a particular temperature a 1.0 M  $\text{H}_2\text{S}$  solution has a  $\text{pH} = 3.75$ . Calculate the value of  $K_a$  at this temperature. (4 marks)



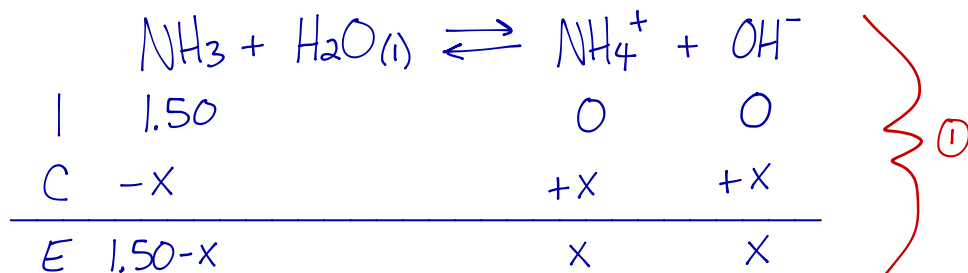
- 7) A  $3.50 \times 10^{-3}$  M sample of the unknown acid, HA, has a  $\text{pH}$  of 2.90. Calculate the value of  $K_a$  and identify this acid. (4 marks)



- 8) What mass of  $\text{CH}_3\text{COOH}$  will produce 2.5 L of a solution having a  $\text{pH}$  of 5.50? (5 marks)



9) Calculate the pH of 1.50 M  $\text{NH}_3$ . (5 marks)



assume ①  
 $1.50 - x = 1.50$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \quad \text{①}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.8 \times 10^{-5} \quad \text{①}$$

$$1.8 \times 10^{-5} = \frac{x^2}{1.50}$$

$$x = 5.2 \times 10^{-3} \text{ M} \quad \text{①}$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$= -\log(5.2 \times 10^{-3})$$

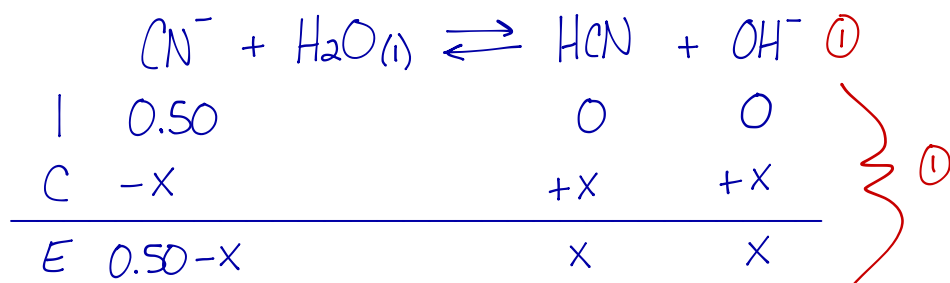
$$= 2.29$$

$$\text{pH} + \text{pOH} = 14.00$$

$$x + 2.29 = 14.00$$

$$x = 11.71$$

10) Calculate the  $[\text{OH}^-]$  in 0.50 M  $\text{CN}^-$ . (5 marks)



assume ①  
 $0.50 - x = 0.50$

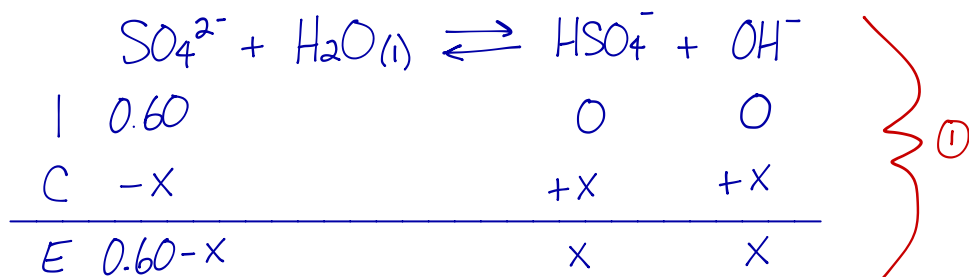
$$K_b = \frac{[\text{HCN}][\text{OH}^-]}{[\text{CN}^-]} \quad \text{①}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.0 \times 10^{-5} \quad \text{①}$$

$$2.0 \times 10^{-5} = \frac{x^2}{0.50}$$

$$x = 3.2 \times 10^{-3} \text{ M} \quad \text{①}$$

11) Calculate the  $[H_3O^+]$  in 0.60 M  $SO_4^{2-}$ . (5 marks)



assume (0.5)  
 $0.60 - x = 0.60$

$$K_b = \frac{[HSO_4^-][OH^-]}{[SO_4^{2-}]}$$

(0.5)

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.2 \times 10^{-2}} = 8.3 \times 10^{-13}$$

(1)

$$8.3 \times 10^{-13} = \frac{x^2}{0.60}$$

$$x = 7.1 \times 10^{-7} M$$

(1)

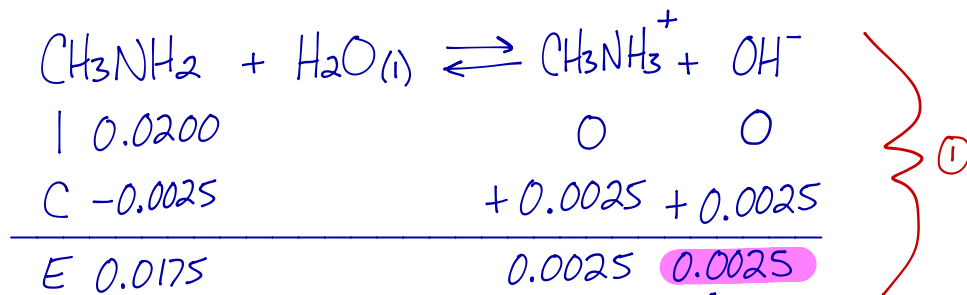
$$K_w = [H_3O^+][OH^-]$$

$$1.0 \times 10^{-14} = x(7.1 \times 10^{-7})$$

$$x = 1.4 \times 10^{-8} M$$

(1)

12) A 0.0200 M solution of methylamine,  $CH_3NH_2$ , has a pH = 11.40. Calculate the  $K_b$  for methylamine. (4 marks)



$$pH + pOH = 14.00$$

$$11.40 + x = 14.00$$

$$x = 2.60$$

(1)

$$[OH^-] = 10^{-pOH}$$

$$= 10^{-2.60}$$

$$= 2.5 \times 10^{-3} M$$

(1)

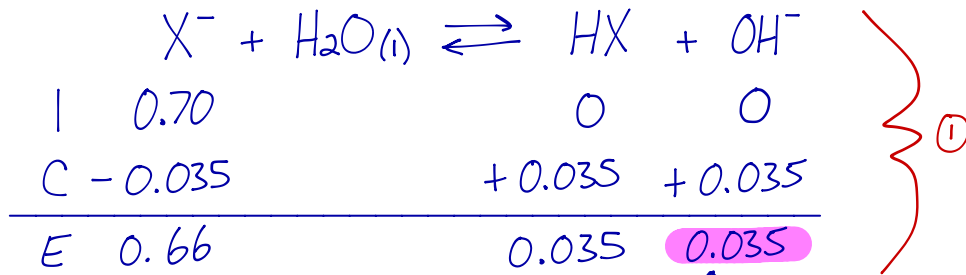
$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$$

$$= \frac{(0.0025)^2}{(0.0175)}$$

$$x = 3.6 \times 10^{-4}$$

(1)

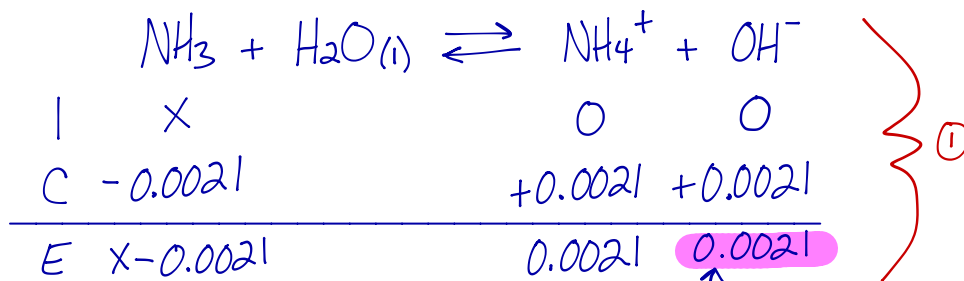
13) The pH of a 0.70 M solution of the weak base  $X^-$  is 12.55. What is the  $K_b$  for  $X^-$ ?



$$\begin{aligned}
 pH + pOH &= 14.00 \\
 12.55 + X &= 14.00 \\
 X &= 1.45 \quad (1) \\
 [OH^-] &= 10^{-pOH} \\
 &= 10^{-1.45} \\
 &= 0.035 \text{ M} \quad (1)
 \end{aligned}$$

$$\begin{aligned}
 K_b &= \frac{[HX][OH^-]}{[X^-]} \\
 &= \frac{(0.035)^2}{(0.66)} \\
 X &= 1.9 \times 10^{-3} \quad (1)
 \end{aligned}$$

14) Calculate the initial concentration of a solution of  $NH_3$  which has a pH = 11.33. (5 marks)



$$\begin{aligned}
 pH + pOH &= 14.00 \\
 11.33 + X &= 14.00 \\
 X &= 2.67 \quad (0.5) \\
 [OH^-] &= 10^{-pOH} \\
 &= 10^{-2.67} \\
 &= 0.0021 \text{ M} \quad (0.5)
 \end{aligned}$$

assume  
 $X - 0.0021 = X$   
 (0.5)

$$\begin{aligned}
 K_b &= \frac{[NH_4^+][OH^-]}{[NH_3]} \quad (0.5) \\
 1.8 \times 10^{-5} &= \frac{(0.0021)^2}{X} \\
 X &= 0.26 \text{ M} \quad (1)
 \end{aligned}$$

$$\begin{aligned}
 K_b &= \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} \\
 &= 1.8 \times 10^{-5} \quad (1)
 \end{aligned}$$