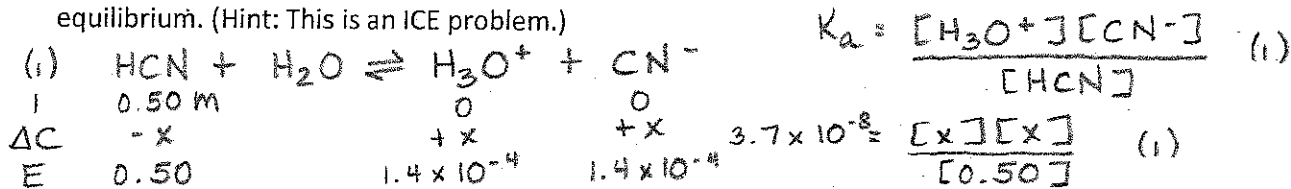


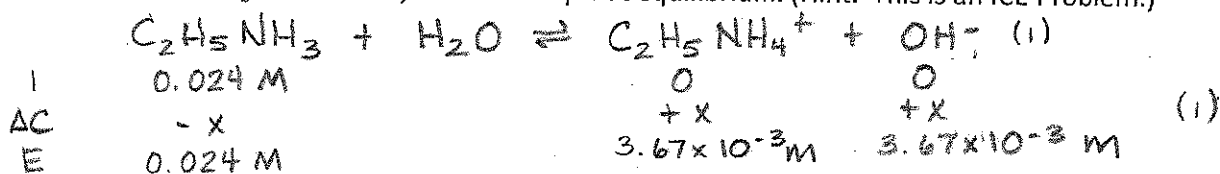
9. HCN has an initial molarity of 0.50 M, with a  $K_a$  value of  $3.7 \times 10^{-8}$ . Calculate its pH at equilibrium. (Hint: This is an ICE problem.)



(1)  $[H^+] = 1.36 \times 10^{-4} \text{ mol/L} \quad (1)$

$pH = 3.87 \quad (1)$

10. Ethylamine ( $C_2H_5NH_3$ ) is a weak Bronsted-Lowry base. If it has an initial molarity of 0.024 M and a  $K_b$  of  $5.6 \times 10^{-4}$ , calculate its pH at equilibrium. (Hint: This is an ICE Problem.)



$$K_b = \frac{[C_2H_5NH_4^+][OH^-]}{[C_2H_5NH_3]} \quad (1)$$

$$pOH = -\log(3.67 \times 10^{-3} M)$$

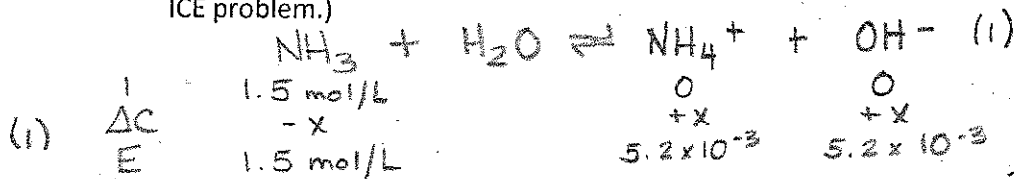
$$= 2.44$$

$$5.6 \times 10^{-4} = \frac{[x][x]}{[0.024]} = 3.67 \times 10^{-3} M = [OH^-] \quad (1)$$

$$pH = 14 - 2.44$$

$$= 11.56 \quad (1)$$

- (1) 11. A chemist adds 0.75 moles of  $NH_3$  to enough water to make 0.50 liters of solution.  $K_b$  of ammonia is  $1.8 \times 10^{-5}$ . Determine the pH of this solution at equilibrium. (Hint: This is an ICE problem.)



$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} \quad (1)$$

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

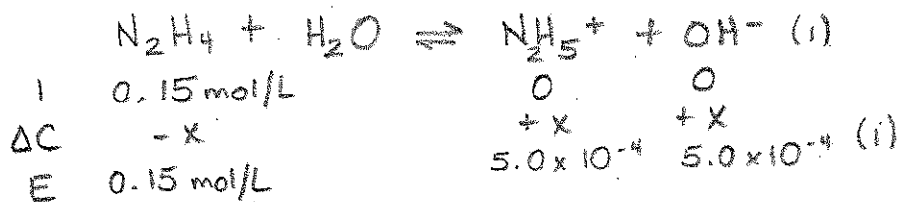
$$= 2.28$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{[1.5 \text{ mol/L}]} = 5.2 \times 10^{-3} \text{ mol/L} = [OH^-] \quad (1)$$

$$pH = 14 - 2.28$$

$$= 11.72 \quad (1)$$

12. Hydrazine,  $N_2H_4$ , has been used as a rocket fuel. Like ammonia, it is a Bronsted base. A 0.15 M solution has a pH of 10.70. What is the  $K_b$  for hydrazine?



$$pH = 10.70$$

$$pOH = 3.30$$

$$[OH^-] = 5.0 \times 10^{-4} \quad (1)$$

$$K_b = \frac{[5.0 \times 10^{-4}][5.0 \times 10^{-4}]}{[0.15]} \quad (1)$$

$$= 1.7 \times 10^{-6} \quad (1)$$