

# Acid-Base # 2

## (K calculations - Summary)

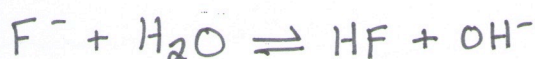
Write response in the space provided. Express your answer in correct units & sig figs where appropriate.

1.

(5 marks)

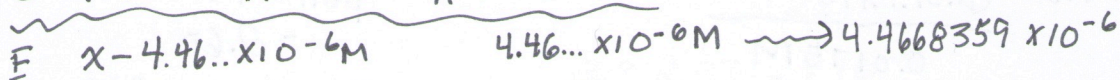
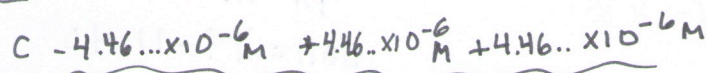
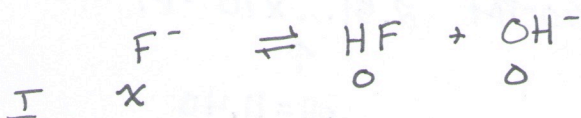
Calculate the initial concentration of a KF salt solution that has a pH = 8.65.

Begin by writing the equation for the predominant equilibrium reaction.



$$K_b = \frac{[HF][OH^-]}{[F^-]} = \frac{1.00 \times 10^{-14}}{3.5 \times 10^{-4}}$$

$$K_b = 2.857... \times 10^{-11}$$



$$\begin{aligned} \uparrow \\ pH &= 8.65 \\ pOH &= 14 - 8.65 \\ &= 5.35 \\ [OH^-] &= 10^{-5.35} \end{aligned}$$

$$K_b = \frac{(4.46... \times 10^{-6} M)(4.46... \times 10^{-6} M)}{(x - 4.46... \times 10^{-6} M)}$$

$$(4.46... \times 10^{-6} M)^2 = 2.857... \times 10^{-11} (x - 4.46... \times 10^{-6} M)$$

$$(4.46... \times 10^{-6} M)^2 + (4.46... \times 10^{-6} M)(2.857... \times 10^{-11}) = 2.857... \times 10^{-11} x$$

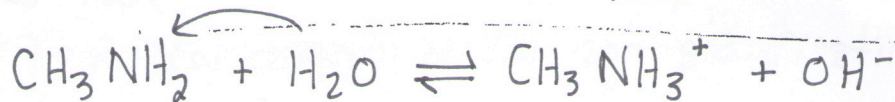
$$x = 0.6983...$$

$$x = [KF]_{\text{initial}} = \cancel{0.70} M$$

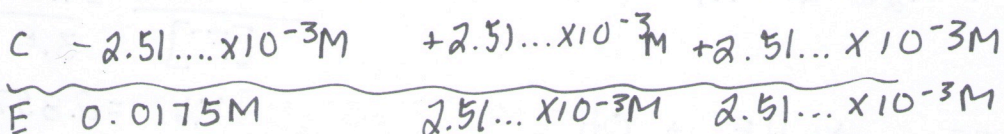
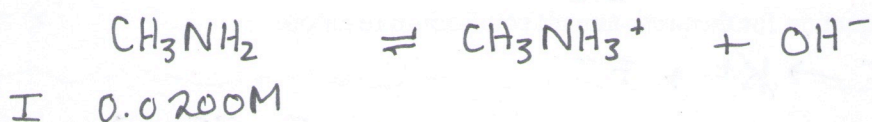
2. A 0.0200 M solution of methylamine,  $\text{CH}_3\text{NH}_2$ , has a  $\text{pH} = 11.40$ .  
 Calculate the  $K_b$  for methylamine.

base

(4 marks)



$$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = ?$$



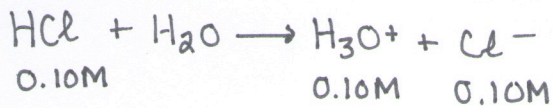
$$K_b = \frac{(2.51 \dots \times 10^{-3}\text{M})^2}{0.0175\text{M}}$$

$$\begin{array}{l} \uparrow \\ \text{pH} = 11.40 \\ \text{pOH} = 14 - 11.40 \\ \quad = 2.60 \\ [\text{OH}^-] = 10^{-2.60} \end{array}$$

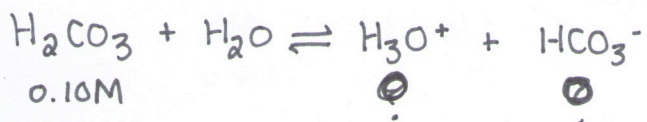
$$K_b = 3.6 \times 10^{-4} \quad (2 \text{ s.f.})$$

b/c  $\text{pH} = 11.40$   
is 2 s.f.)

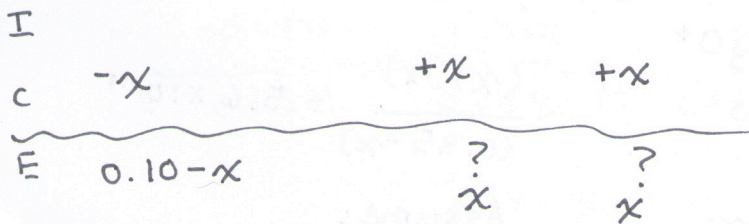
3. Using calculations, show why the electrical conductivity of 1.0M  $\text{H}_2\text{CO}_3$  will be less than that for 0.10M  $\text{HCl}$ . (4 marks)
- = higher concentration of ions



$$\begin{aligned} [\text{ions}] &= 0.10\text{M} + 0.10\text{M} \\ &= 0.20\text{M} \end{aligned}$$



$$K_a = \frac{[\text{HCO}_3^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{CO}_3]} = 4.3 \times 10^{-7}$$



$$= \frac{x^2}{0.10 - x} = 4.3 \times 10^{-7}$$

$$x^2 + 4.3 \times 10^{-7}x - 4.3 \times 10^{-8} = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = 2.07 \dots \times 10^{-4}$$

$$[\text{H}_3\text{O}^+] = x = 2.07 \dots \times 10^{-4}\text{M}$$

$$[\text{HCO}_3^-] = x = 2.07 \dots \times 10^{-4}\text{M}$$

$$[\text{ions}] = 4.14 \dots \times 10^{-4}\text{M}$$

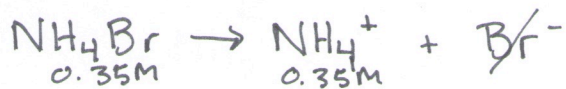
$$[\text{ions}] \text{ in HCl} = 0.20\text{M}$$

$$[\text{ions}] \text{ in H}_2\text{CO}_3 = 4.1 \times 10^{-4}\text{M}$$

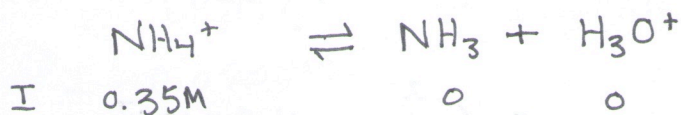
\* see back page  
in case you're  
wondering.....

4. (5 marks)

Calculate the pH of a 0.35 M solution of the salt ammonium bromide.  
Begin by writing the equation for the predominant equilibrium.



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.6 \times 10^{-10}$$



$$\frac{(x)(x)}{(0.35-x)} = 5.6 \times 10^{-10}$$

Assume:

$$0.35 - x \approx 0.35$$

$$\frac{x^2}{0.35} = 5.6 \times 10^{-10}$$

$$x = \sqrt{(5.6 \times 10^{-10})(0.35\text{M})}$$
$$= 1.4 \times 10^{-5}\text{M}$$

justification:

$$\% \text{ dissociation} = \frac{[\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} \times 100$$

$$= \frac{1.4 \times 10^{-5}\text{M}}{0.35\text{M}} \times 100$$

= 0.004%  
approximation  
is valid.

$$x = [\text{H}_3\text{O}^+]$$

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

$$= -\log (1.4 \times 10^{-5}\text{M})$$

$$\text{pH} = 4.85$$

A 0.20 M solution of a weak acid, HA, has a pH = 1.32.

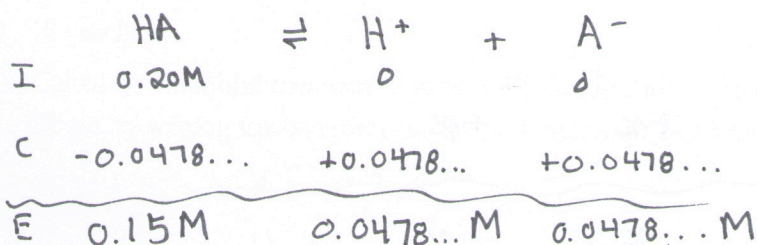
Use calculations and the table of "Relative Strengths of Brønsted-Lowry Acids and Bases" from the *Data Booklet* to determine the identity of the acid.

(5 marks)

you can identify an acid with its K<sub>a</sub>.



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = ?$$

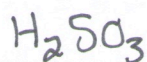


↑  
pH = 1.32  
[H<sup>+</sup>] = 10<sup>-1.32</sup>

$$K_a = \frac{(0.0478...M)^2}{0.15M}$$

$$K_a = 1.5 \times 10^{-3}$$

∴ the acid must be



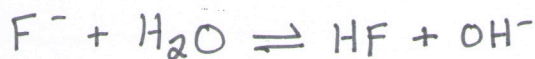
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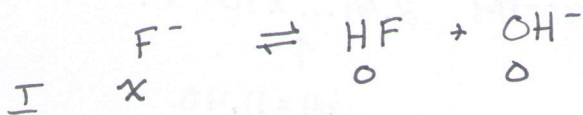
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$$C \quad \underline{4.46... \times 10^{-6} M} \quad + 4.46... \times 10^{-6} M \quad + 4.46... \times 10^{-6} M$$

$$E \quad x - 4.46... \times 10^{-6} M \quad \qquad \qquad 4.46... \times 10^{-6} M \quad \rightsquigarrow 4.4668359 \times 10^{-6}$$

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