

# Acids + Bases Problem Set

## Review:

1. For acid dissociation  $HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$ ,  
what is the  $K_a$  expression?

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

2. For base dissociation (association?)  $B(aq) + H_2O(l) \rightleftharpoons HB^+(aq) + OH^-(aq)$ ,  
what is the  $K_b$  expression?

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

3. What is the value of  $K_w$  at 298 K?

$$K_w = 1.0 \times 10^{-14} \text{ at } 298 \text{ K.}$$

4.  $K_w$  increases/decreases/stays constant with increasing temperature.  
choose one

5. What is the mathematical relationship between  $K_a$ ,  $K_b$ , and  $K_w$ ?

$$K_w = K_a K_b$$

## Problems:

1. Calculate the hydronium ion concentration in 50.0 mL of 0.10 M  $NaH_2AsO_4$ .  $[K_1 = 6.0 \times 10^{-3}, K_2 = 1.1 \times 10^{-7}, K_3 = 3.0 \times 10^{-12}]$

$H_2AsO_4^-$  can act as both an acid or base  $\rightarrow 2H_2AsO_4^- \rightleftharpoons H_3AsO_4 + HAsO_4^{2-}$

So,  $[H_2AsO_4^-] \approx [HAsO_4^{2-}]$

Since  $K_1 = \frac{[H^+][H_2AsO_4^-]}{[H_3AsO_4]}$ ,  $[H_3AsO_4] = \frac{[H^+][H_2AsO_4^-]}{K_1}$

Since  $K_2 = \frac{[H^+][HAsO_4^{2-}]}{[H_2AsO_4^-]}$ ,  $[HAsO_4^{2-}] = \frac{K_2[H_2AsO_4^-]}{[H^+]}$

So,  $\frac{[H^+][H_2AsO_4^-]}{K_1} = \frac{K_2[H_2AsO_4^-]}{[H^+]}$

$\rightarrow [H^+]^2 = K_1 K_2$  ← notice how  $[H^+]$  is independent of initial  $[H_2AsO_4^-]$ !

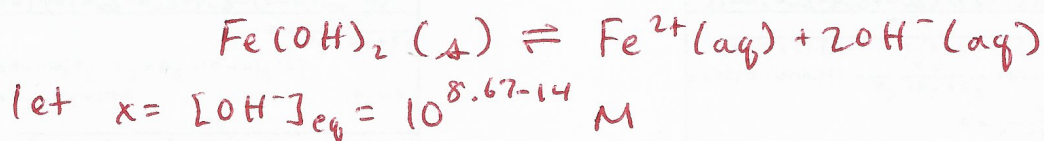
$$[H^+] = \sqrt{K_1 K_2} \approx \boxed{2.5 \times 10^{-5} \text{ M}}$$

2. 1.0 L of an aqueous solution of which  $[H_2CO_3] = [HCO_3^-] = 0.100 M$  has  $[H^+] = 4.2 \times 10^{-7} M$ . What is the  $[H^+]$  after 0.005 moles of NaOH have been added?

$$\begin{aligned} \text{HH equation: } pH &= pK_a + \frac{[HCO_3^-]}{[H_2CO_3]} \\ &= -\log [4.2 \times 10^{-7}] + \log \frac{0.105}{0.095} \\ &= 6.420 \end{aligned}$$

$$[H^+] = 10^{-6.420} = \boxed{3.80 \times 10^{-7}}$$

3. The pH of a saturated solution of  $Fe(OH)_2$  is 8.67. What is the  $K_{sp}$  of  $Fe(OH)_2$ ? (Hint:  $K_{sp} = [Fe^{2+}][OH^-]^2$ )



$$K_{sp} = \left(\frac{1}{2}x\right)(x)^2 = \boxed{5.1164 \times 10^{-17}}$$

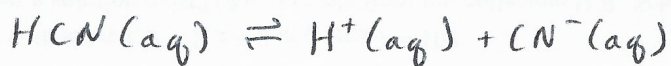
4. The Henderson-Hasselbalch equation describes pH of something called a buffer solution through altering concentrations of an acid and its conjugate base. Derive this equation from your answer to Review question 1.

$$\begin{aligned} \text{Henderson-Hasselbalch equation: } pH &= pK_a + \log \left( \frac{[A^-]}{[HA]} \right) \\ K_a &= \frac{[H_3O^+][A^-]}{[HA]} \quad \text{(really important!)} \\ -\log(K_a) &= -\log \left( \frac{[H_3O^+][A^-]}{[HA]} \right) = -\log [H_3O^+] + \left( -\log \frac{[A^-]}{[HA]} \right) \\ pK_a &= pH - \log \frac{[A^-]}{[HA]} \\ pH &= pK_a + \log \frac{[A^-]}{[HA]} \end{aligned}$$

## Food for Thought...

We've learned about the 5% rule and ice tables in order to solve for the pH of a given solution. Take a look at this problem:

Calculate the pH of a  $1.0 \times 10^{-5} \text{ M}$  HCN (hydrocyanic acid) solution.



(You should get  $\text{pH} \approx 7.11$  as your answer using ICE tables. But isn't HCN an acid? How is  $\text{pH} > 7$ ?)

$$K_a = 6.2 \times 10^{-10}$$

$$K_a = \frac{[\text{H}^+][\text{CN}^-]}{[\text{HCN}]} = \frac{(x)(x)}{10^{-5} - x} = 6.2 \times 10^{-10}; x > 0$$

$$x^2 = 6.2 \times 10^{-15} - 6.2 \times 10^{-10} x$$

$$x^2 + (6.2 \times 10^{-10}) x - (6.2 \times 10^{-15}) = 0$$

$$x = \frac{-(6.2 \times 10^{-10}) + \sqrt{(38.44 \times 10^{-20}) + (24.8 \times 10^{-15})}}{2}$$

$$= 7.84 \times 10^{-8}$$

↑ quadratic formula

$$\text{pH} = -\log x = \boxed{7.11}$$