

Chapter 16: Acid-Base Equilibria

Kahoot!

1. A Brønsted–Lowry acid is a/an ____ . **proton donor**, proton acceptor, electron-pair acceptor, electron-pair donor
2. A Brønsted–Lowry base is a/an ____ . electron-pair acceptor, hydroxide donor, **proton acceptor**, electron-pair donor
3. Which of the following is the conjugate acid of HPO_4^{2-} ? **H_3PO_4** , $H_2PO_4^-$, PO_4^{3-} , HPO_4^{2-}
4. Which of the following is the conjugate base of H_2SO_4 ? H_2SO_4 , SO_4^{2-} , $H_3SO_4^+$, **HSO_4^-**
5. The ____ the acid, the ____ its conjugate base. stronger;stronger, weaker;weaker, **stronger;weaker, weaker;stronger**
6. If the $[H_3O^+]$ is 10^{-5} what is the $[OH^-]$? 10^{-5} , 10^5 , **10^{-9}** , 10^9
7. A solution with pH of 6.00 would be considered ____ . **acidic**, basic, neutral, none of the above
8. What is the pH of a 0.0200 M aqueous solution of HBr? 1.00, **1.70**, 2.30, 12.30
9. What is the pH of a 0.0400 M aqueous solution of KOH? **12.60**, 10.30, 4.00, 1.40
10. A 0.30 M acid solution has a pH of 1.85. What is the %ionization? 21%, 16%, 6.2%, **4.7%**
11. Which of the following is NOT as weak base? NH_3 , NH_2OH , F^- , **OH^-**
12. For a conjugate acid-base pair, $K_w = K_a$ ____ K_b . +, -, **x**, /
13. A salt solution of RbF is considered ____ . **basic**, acidic, neutral, none of the above
14. Which of the following correctly lists the acids in order of decreasing strength? **$HClO_2 > HClO > HBrO > HIO$** , $HClO > HBrO > HIO > HClO_2$, $HIO > HBrO > HClO > HClO_2$, $HClO_2 > HIO > HBrO > HClO$
15. A Lewis acid is a/an ____ . proton donor, proton acceptor, electron-pair donor, **electron-pair acceptor**
16. A Lewis base is a/an ____ . proton donor, proton acceptor, **electron-pair donor**, electron-pair acceptor

Whiteboard Examples

Example: Determine the hydroxide concentration in a solution with

$$[H_3O^+] = 1.89 \times 10^{-4} \underline{M} .$$

$$K_w = [H_3O^+][OH^-] \rightarrow [OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{1.89 \times 10^{-4}} = 5.29 \times 10^{-11} \underline{M}$$

pKw derivation:

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

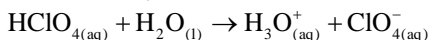
$$-\log K_w = -\log\{[OH^-][H_3O^+]\}$$

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

$$pK_w = pOH + pH = -\log(1.0 \times 10^{-14}) = 14$$

Strong Acid/Base Examples I: Write the balanced equation for each of the following and determine the pH.

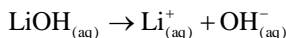
a.) 0.5000 M $HClO_4(aq)$



$$[H_3O^+] = [ClO_4^-] = 0.5000 \text{ M}$$

$$pH = -\log(0.5000 \text{ M}) = 0.30$$

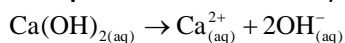
b.) 0.0256 M $LiOH(aq)$



$$[\text{OH}_{(aq)}^{-}] = [\text{LiOH}_{(aq)}] = 0.0256 \text{ M}$$

$$\text{pH} = 14 - \text{pOH} = 14 - (-\log(0.0256 \text{ M})) = 12.41$$

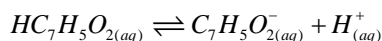
Strong Acid/Base Examples II: Determine the hydronium ion concentration for a 0.01500 M $\text{Ca}(\text{OH})_2$.



$$[\text{OH}_{(aq)}^{-}] = 2 \times [\text{Ca}(\text{OH})_{2(aq)}] = 2 \times 0.01500 \text{ M} = 0.03000 \text{ M}$$

$$[\text{H}_3\text{O}_{(aq)}^{+}] = \frac{K_w}{[\text{OH}_{(aq)}^{-}]} = \frac{1.0 \times 10^{-14}}{0.03000 \text{ M}} = 3.33 \times 10^{-13} \text{ M}$$

Calculating K_a from pH: What is the K_a of a 0.050 M solution of $\text{HC}_7\text{H}_5\text{O}_2$ if the pH of this solution is 2.75?



We begin with an ICE Table

	$\text{HC}_7\text{H}_5\text{O}_{2(aq)}$	$\text{C}_7\text{H}_5\text{O}_{2(aq)}^{-}$	$\text{H}_{(aq)}^{+}$
Initial	0.050	0.0	0.0
Change	-x	+x	+x
Eq	$0.050 - x$	+x	+x

Next, we write down symbolic representation of the equilibrium/acid dissociation expression:

$$K_a = \frac{[\text{C}_7\text{H}_5\text{O}_{2(aq)}^{-}][\text{H}_{(aq)}^{+}]}{[\text{HC}_7\text{H}_5\text{O}_{2(aq)}]}$$

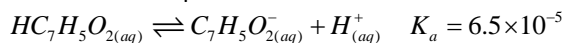
Now we remember we don't know K_a but need to find it which means we must know x:

$$[\text{H}_{(aq)}^{+}] = 10^{-\text{pH}} = 10^{-2.75} = 1.77 \times 10^{-3} \text{ M}$$

Finally we plug in x for our equilibrium expression and solve

$$K_a = \frac{[\text{C}_7\text{H}_5\text{O}_{2(aq)}^{-}][\text{H}_{(aq)}^{+}]}{[\text{HC}_7\text{H}_5\text{O}_{2(aq)}]} = \frac{x^2}{0.050 - x} = \frac{0.00177^2}{0.050 - 0.00177} = 6.5 \times 10^{-5}$$

Percent Dissociation Example: Determine the percent dissociation of 0.050M of benzoic acid.

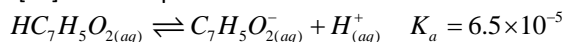


We already found $[\text{H}^{+}] = 1.77 \times 10^{-3} \text{ M}$ therefore

$$\frac{[\text{H}^{+}]}{[\text{HA}]} = \frac{1.77 \times 10^{-3} \text{ M}}{0.050 \text{ M}} \times 100\% = 3.54\%$$

It should be small since our K_a is so small

Weak Acid Example: Calculate $[\text{H}^{+}]$ and the pOH of 0.050M of benzoic acid.



	$\text{HC}_7\text{H}_5\text{O}_{2(aq)}$	$\text{C}_7\text{H}_5\text{O}_{2(aq)}^{-}$	$\text{H}_{(aq)}^{+}$
Initial	0.050	0.0	0.0
Change	-x	+x	+x
Eq	$0.050 - x$	+x	+x

$$K_a = \frac{[C_7H_5O_2^-][H^+]}{[HC_7H_5O_2]} = \frac{x^2}{0.050 - x} = 6.5 \times 10^{-5}$$

$$x^2 = 3.25 \times 10^{-6} - 6.5 \times 10^{-5}x \rightarrow x^2 + 6.5 \times 10^{-5}x - 3.25 \times 10^{-6} = 0$$

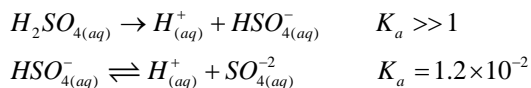
$$x = \frac{-6.5 \times 10^{-5} \pm \sqrt{(6.5 \times 10^{-5})^2 - 4 \times 1 \times (-3.25 \times 10^{-6})}}{2}$$

$$x = \frac{-6.5 \times 10^{-5} \pm 3.61 \times 10^{-3}}{2} \rightarrow x = 1.77 \times 10^{-3} M$$

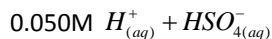
$$[H^+] = 1.77 \times 10^{-3} M$$

$$14 = pH + pOH \rightarrow pOH = 14 - pH = 14 + \log[H^+] = 11.25$$

Polyprotic Acid Example: Calculate the $[H^+]$ of 0.050M of sulfuric acid.



Initially all 0.050M of the H_2SO_4 dissociates completely into



	$HSO^-_{4(aq)}$	$H^+_{(aq)}$	$SO^{2-}_{4(aq)}$
Initial	0.050	0.050	0.0
Change	-x	+x	+x
Eq	0.050 - x	0.050 + x	+x

$$K_a = \frac{[H^+][SO^{2-}_{4(aq)}]}{[HSO^-_{4(aq)}]} = \frac{(0.050 + x)x}{(0.050 - x)} = 1.2 \times 10^{-2}$$

$$0.050x + x^2 = 6.0 \times 10^{-4} - 1.2 \times 10^{-2}x$$

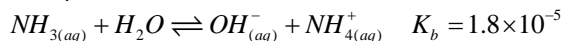
$$x^2 + 0.062x - 6.0 \times 10^{-4} = 0$$

$$x = \frac{-0.062 \pm \sqrt{(0.062)^2 - 4 \times 1 \times (-6.0 \times 10^{-4})}}{2}$$

$$x = \frac{-0.062 \pm 7.90 \times 10^{-2}}{2} \rightarrow x = 8.51 \times 10^{-3} M$$

$$[H^+] = 0.050 + 0.0085 = 0.059M$$

Weak Base Example: Calculate the pH of 0.050 M NH_3 .



	$NH_{3(aq)}$	$OH^-_{(aq)}$	$NH^+_{4(aq)}$
Initial	0.050	0.0	0.0
Change	-x	+x	+x
Eq	0.050 - x	+x	+x

$$K_b = \frac{[OH^-][NH^+_{4(aq)}]}{[NH_{3(aq)}]} = \frac{x^2}{0.050 - x} = 1.8 \times 10^{-5}$$

$$x^2 = 9 \times 10^{-7} - 1.8 \times 10^{-5}x \rightarrow x^2 + 1.8 \times 10^{-5}x - 9 \times 10^{-7} = 0$$

$$x = \frac{-1.8 \times 10^{-5} \pm \sqrt{(1.8 \times 10^{-5})^2 - 4 \times 1 \times (-9 \times 10^{-7})}}{2}$$

$$x = 9.4 \times 10^{-4} M$$

$$[OH^-] = 9.4 \times 10^{-4} M$$

$$pH = 14 + \log[OH^-] = 14 - 3.03 = 10.97$$

Conversion from Kb to Ka Example: Determine the K_b of HCN if $K_a = 4.9 \times 10^{-10}$.

$$K_w = K_a \times K_b \rightarrow K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.04 \times 10^{-5}$$

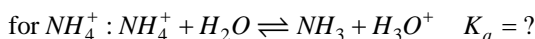
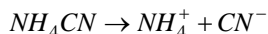
Salt Classification Example I: Classify the following solutions as basic, acidic, or neutral.

- a.) KBr b.) NaNO₂ c.) NH₄Cl

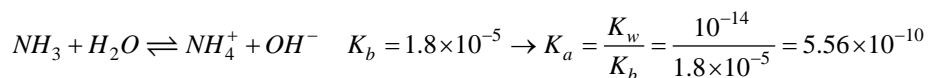
Answer:

- a.) neutral b.) basic c.) acidic

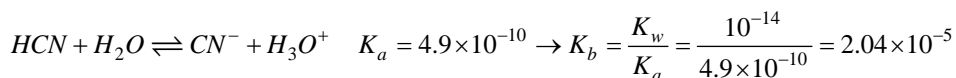
Salt Classification Example II: Calculate the K_a for the cation & the K_b for the anion in an aqueous solution containing NH₄CN. Is the solution acidic, basic, or neutral?



we will not find this K_a in a table BUT we can find the K_b of NH_3 to it:



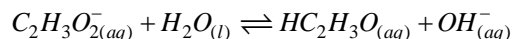
for CN^- we will have to use the K_a of HCN to get its K_b ,



$K_b(CN^-) > K_a(NH_4^+) \therefore$ the soln is basic

Salt Example: Calculate the pH of a 0.25M NaC₂H₃O₂, $K_a = 1.76 \times 10^{-5}$.

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.76 \times 10^{-5}} = 5.68 \times 10^{-10}$$



	$C_2H_3O_2^-(aq)$	$HC_2H_3O_2(aq)$	OH^-
Initial	0.250	0.0	0.0
Change	-x	+x	+x
Eq	0.250 - x	+x	+x

$$K_b = \frac{[HC_2H_3O_2(aq)][OH^-]}{[C_2H_3O_2^-(aq)]} = \frac{x^2}{0.250 - x} = 5.68 \times 10^{-10}$$

because we have a large concentration of acetate and a small K_b we will try and assume $0.250 \gg x$

$$\frac{x^2}{0.250 - x} \sim \frac{x^2}{0.250} = 5.68 \times 10^{-10} \rightarrow x = 1.19 \times 10^{-5} M$$

$$ck : \frac{1.19 \times 10^{-5}}{0.250} \times 100\% = 0.005\% < 5\%$$

therefore our assumption is valid and $[OH^-] = 4.77 \times 10^{-5} M$

$$pH = 14 - pOH$$

$$= 14 + \log[OH^-] = 9.08$$