8.6 - Finding the Hydronium Ion Concentration and pH of Strong and Weak Acid Solutions

If we wish to find the hydronium ion concentration ($[H_3O^+]$) and the pH of a solution, we need to know both the strength of the acid (or base) and the concentration of the acid (or base). We will find that we need to treat strong acids (and bases) differently than weak acids (and bases) based on the extent to which they react with water.

Strong Acids

By definition, strong acids are those acids with a $K_a \geq 1$. using this definition, we assume that strong acids will react completely with water, so that every molecule of acid reacts with a molecule of water to produce a hydronium ion and the conjugate base. (This assumption is, of course, not possible because there must always be at least a few particles of each reactant and each product at equilibrium, but our assumption is a valid one when K_a gets to be quite a bit larger than 1.):

$$HA_{(aq)} \; + \; H_2O_{(l)} \rightarrow H_3O^+_{(aq)} \; + \; A^-_{(aq)} \tag{1}$$

Based on our assumption, $[H_3O^+]_{eq} = [HA]_{initial}$, so we can carry out the following calculation:

$$pH = -log([H_3O^+]_{eq}) = -log([HA]_{initial}) \tag{2}$$

Example 1

What is the pH of a 0.175 M aqueous solution of HBr?

Solution

HBr is a strong acid ($K_a > 1$), so $[H_3O^+]_{eq} = [HBr]_{initial}$. Thus, pH = -log (0.175) = 0.757

Exercise 1

What is the pH of a 0.0043 M aqueous solution of HNO_3?

Answer

HNO_3 is a strong acid ($K_a > 1$), so $[H_3O^+]_{eq} = [HNO_3]_{initial}$. Thus, pH = -log (0.0043) = 2.37

Strong Bases

By definition, strong bases are those bases with a $K_b \geq 1$. using this definition, we assume that strong bases will react completely with water, so that every molecule of base reacts with a molecule of water to produce a hydroxide ion and the conjugate acid. (This assumption is, of course, not possible because there must always be at least a few particles of each reactant and each product at equilibrium, but our assumption is a valid one when K_b gets to be quite a bit larger than 1.):

$$B_{(aq)} \; + \; H_2O_{(l)} \rightarrow OH^-_{(aq)} \; + \; HB^+_{(aq)} \tag{3}$$

Based on our assumption, $[OH^-]_{eq} = [B]_{initial}$, so we can carry out the following calculation:

$$pOH = -log([OH^-]_{eq}) = -log([B]_{initial}) \tag{4}$$

To find the pH, we then subtract the pOH from 14.

Example 2

What is the pH of a 0.175 M aqueous solution of $NaNH_2$?

Solution

NH_2^- is a strong base ($K_b > 1$), so $[OH^-]_{eq} = [NH_2^-]_{initial}$. Thus, pOH = -log (0.175) = 0.757, and pH = 14.000 - 0.757 = 13.243

Exercise 2

What is the pH of a 0.0043 M aqueous solution of KOH?

Answer

KOH is a strong base ($K_b = 1$), so $[OH^-]_{eq} = [OH^-]_{initial}$. Thus, pOH = -log (0.0043) = 2.37, and pH = 14.00 - 2.37 = 11.63

Weak Acids

If the acid you are working with is weak ($K_a < 1$), you must use a logic chart (ICE diagram) to determine the $[H_3O^+]$ because you cannot tell how much H_3O^+ will form in solution simply by looking at the intial $[HA]$.

Example 3

What is the pH of a 0.025 M aqueous solution of HClO?

Solution

ClO^- is a weak base with $K_b = 3.5\ x\ 10^{-8}$. The law of mass action for the reaction of $HClO$ with water is $3.5\ x\ 10^{-8} = \frac{[ClO^-][H_3O^+]}{[HClO]}$

The ICE diagram will look like this:

HClO	H_2O	⇌	ClO^-	H_3O^+
0.025	-		0	0
-n	-		+n	+n
0.025-n	-		n	n

Thus,

$$3.5\ x\ 10^{-8} = \frac{[n][n]}{[0.025 - n]} \tag{5}$$

which rearranges to

$$0 = n^2 + 3.5\ x\ 10^{-8} n - 8.75\ x\ 10^{-10} \tag{6}$$

using the quadratic formula, the answer can be found to be 2.96 x 10^{-5} M H_3O+. Thus the pH is -log(2.96 x 10^{-5}) = 4.53

Exercise 3

What is the pH of a 0.111 M aqueous solution of acetic acid? $K_a = 1.8 \times 10^{-5}$

Answer

pH = 2.85

Weak Bases

If the base you are working with is weak ($K_b < 1$), you must use a logic chart (ICE diagram) to determine the $[OH^-]$ because you cannot tell how much OH^- will form in solution simply by looking at the intial $[B]$.

Example 4

What is the pH of a 0.025 M aqueous solution of NaClO?

Solution

ClO^- is a weak base with $K_b = 2.9\ x\ 10^{-7}$. The law of mass action for the reaction of ClO^- with water is $2.9\ x\ 10^{-7} = \frac{[HClO][OH^-]}{[ClO^-]}$

The ICE diagram will look like this:

ClO^-	H_2O	$\rightleftharpoons$	HClO	OH^-
0.025	-		0	0
-n	-		+n	+n
0.025-n	-		n	n

Thus,

$$2.9\ x\ 10^{-7} = \frac{[n][n]}{[0.025-n]} \tag{7}$$

which rearranges to

$$0 = n^2 + 2.9\ x\ 10^{-7} n - 7.25\ x\ 10^{-9} \tag{8}$$

using the quadratic formula, the answer can be found to be 2.68 x 10^{-5} M OH^-. Thus the pOH is -log(2.68 x 10^{-5}) = 4.57, and the pH is 14.00 -4.57 = 9.43.

Exercise 4

What is the pH of a 0.045 M aqueous solution of sodium phosphate. $K_b = 2.8 \times 10^{-2}$

Answer

12.38

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