

Chapter 10 - Monoprotic Acid-Base Equilibria

Homework: Due Monday, March 6

Problems 10-2, 10-3, 10-5, 10-8, 10-11, 10-19,
10-26, 10-27, 10-28, 10-29, 10-32,
10-36

ACID – BASE CHEMISTRY

Protic Acids and Bases (Bronsted-Lowry)

- Acid => increases $[\text{H}_3\text{O}^+]$ or $[\text{H}^+]$
- Base => increases $[\text{OH}^-]$
- Protic => transfer of H^+
- Bronsted acid => proton donor (like HCl)
- Bronsted base => proton acceptor (like NaOH)
- Salt => ionic solid made from reaction of acid and base
(like CaCl_2 – R_x of $2\text{HCl} + \text{Ca}(\text{OH})_2 = \text{CaCl}_2 + 2\text{HOH}$)
- Conjugate => acid-base pair
(like acetic acid $\text{H}_3\text{C}-\underset{\text{O}}{\underset{||}{\text{C}}}-\text{OH}$ and acetate ion $\text{H}_3\text{C}-\underset{\text{O}}{\underset{||}{\text{C}}}-\text{O}^-$)

Equilibrium of Water



- K_w allows you to calculate $[\text{H}^+]$ and $[\text{OH}^-]$
- $[\text{H}^+]$ and $[\text{OH}^-]$ describes the acid/base content of solution
- Since K_w is small, $[\text{H}^+] [\text{OH}^-]$ will also be small.
- pH ($-\log [\text{H}^+]$) or pOH ($-\log [\text{OH}^-]$) is a useful way to report $[\text{H}^+]$ or $[\text{OH}^-]$
- $K_w = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$ So....
 $\text{p}K_w = \text{pH} + \text{pOH} = 14$

Acid/Base Strength

Strong acids and bases **COMPLETELY** dissociates



(No reverse arrow as no HCl present in solution)

TABLE 6-2: Strong acids and Strong bases

Be able to recognize them (memorize)

Examples: Acids – HCl, HNO₃, H₂SO₄

Bases – NaOH, KOH, NH₄OH

Determining pH and pOH of strong acid/base

Example:

0.1 M HCl means $[H^+] = 0.1$ or 1×10^{-1}

$$\text{pH} = -\log(1 \times 10^{-1}) = 1$$

So...	<u>$[H^+]$</u>	<u>$[H^+]$</u>	<u>pH</u>
	0.01	1×10^{-2}	2
	0.001	1×10^{-3}	3
	0.0001	1×10^{-4}	4

0.035 M HCl $[H^+] = 3.5 \times 10^{-2}$ M $\text{pH} = -\log(3.5 \times 10^{-2}) = 1.46$

Given pH 5.67, what is the $[H^+]$?

$$[H^+] = \text{antilog}(-5.67) \Rightarrow 10^{-\text{pH}} = 10^{-5.67} = 2.14 \times 10^{-6} \text{ M}$$

Example: What is the pH of a 0.0045 M NaOH solution?

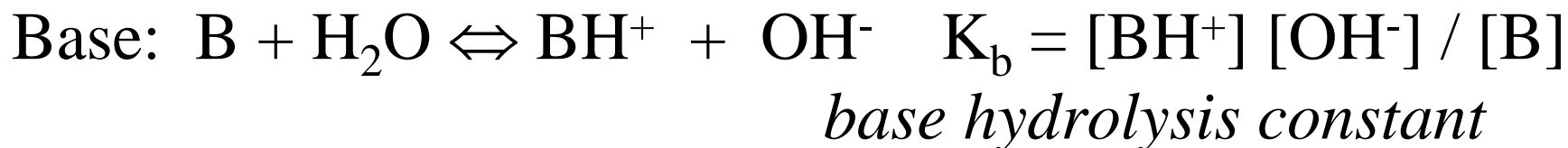
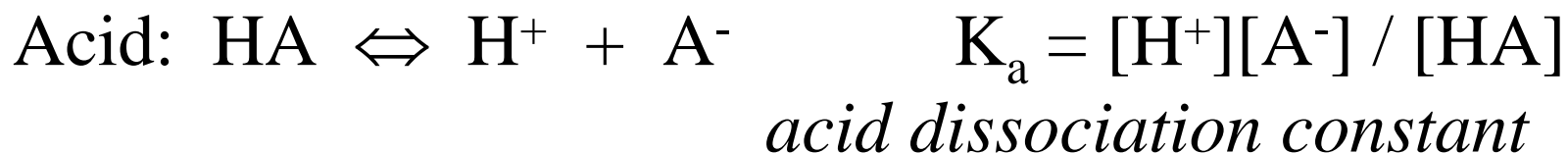
Example: What is the $[\text{OH}^-]$ of a solution at pH 4.3?

If concentration of strong acid or strong base is very small ($< 10^{-6}$ M), then you cannot simply use the concentration of the strong acid or strong base to determine pH or pOH. The reason is that at this low concentration, the dissociation of H_2O will also contribute to the $[\text{H}^+]$ and $[\text{OH}^-]$. ($\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$)

In the case of extremely small strong acid or strong base concentrations ($< 10^{-8}$), then the main contribution of $[\text{H}^+]$ or $[\text{OH}^-]$ is from the dissociation of H_2O so the solution has a pH 7 and pOH 7.

See pg. 180 and 181 of Harris

Weak acids and bases: PARTIAL dissociation



Relationship of K_a and K_b

The relationship between K_a and K_b for a conjugate pair is:

$$K_a \times K_b = K_w = 1 \times 10^{-14}$$

Also
$$pK_w = pK_a + pK_b$$

