Chem 103 lecture 2b

Strengths of acids and bases

K_a and K_b

pH and pH calculations

Strengths of acids and bases

Acids and Bases are categorized into two general types:

Strong and weak. Refers to the tendency to donate or accept H^+

Strong acids and bases dissociate 100%: HCl, HNO_3; NaOH, KOH

Weak acids and bases do not dissociate 100%: CH_3COOH, HF, NH_3

Among weak acids and bases, the rule of thumb is:

"the stronger the conjugate, the weaker the acid or base"

Strengths of acids & bases:

Conjugate acids		Conjugate bases	
HCl	↑ Increasing strength	Cl^-	↓ Increasing strength
HAc		Ac^-	
H_2O		OH^-	

Dissociation of strong acids in water

Strong acids dissociate 100% in water. (Classic example:HCl)

$HCl + H_2O \xrightarrow{100\%} H_3O^+ + Cl^-$

Other examples: $HNO_3 + H_2O \xrightarrow{100\%} H_3O^+ + NO_3^-$

HBr, HI, H_2SO_4 (remember all these strong acids)

Often we just write: $HCl \rightarrow H^+ + Cl^-$; $HNO_3 \rightarrow H^+ + NO_3^-$

Dissociation of weak acids in water

Weak acids do NOT ionize 100% (classic example: acetic acid; contains methyl -*CH_3*- attached to carboxyl group -*COOH*):

$CH_3\text{-}COOH + H_2O \Longleftrightarrow H_3O^+ + CH_3\text{-}COO^-$

Generic eq'n for weak acid: $HA + H_2O \Longleftrightarrow H_3O^+ + A^-$

Called " K_a equilibrium". K_a is "*acid* dissociation constant".

What is the expression for K_a ? $K_a = [H_3O^+][A^-] / [HA]$

The *larger* the K_a, the *stronger* the weak acid.

Usually K_a very small, ex : K_a for acetic acid = 1.8×10^{-5}

For NH_4^+ it is 5.6×10^{-10}. Which is the stronger weak acid?

Weak base equlibrium

For weak bases: "K_b equilibrium". K_b ="base ionization constant: Generic equation

$B + H_2O \Longleftrightarrow BH^+ + OH^-$

$K_b = [BH^+][OH^-] / [B]$

The greater K_b is, the stronger the base.

Usually K_b very small, example: NO_2^- has $K_b = 2.2 \times 10^{-11}$

$K_aK_b = K_w = 1.0x10^{-14}$

Note: $K_a = [H_3O^+][A^-]/[HA]$ for acid HA.

For it's conjugate, A^-, it is: $K_b = [HA][OH^-] / [A^-]$

So if multiply K_a of weak acid by K_b of its conjugate, then :

$K_aK_b = ([H_3O^+]\cancel{[A^-]}/\cancel{[HA]})\ (\cancel{[HA]}[OH^-] / \cancel{[A^-]})$

$\mathbf{K_aK_b = [H_3O^+][\ OH^-] = K_w = 1.0 \times 10^{-14}}$

pH

Because of wide range of $[H_3O^+]$, it's convenient to express concentration levels of $[H^+]$ by its exponent, using logarithmic scale. This is the "pH scale".

Definition of pH: $\mathbf{pH = -log[H^+] = -log[H_3O^+]}$

For example: what's the pH of pure water?

Note: pure water: $[H_3O^+] = 1.0x10^{-7}M$

$pH = -log\ \{[H_3O^+]\} = -log(1.0x10^{-7}M) = 7.00$

This is the *neutral* pH

pH of acidic solutions

If you have an acidic solution with $[H_3O^+] = 2.0 \times 10^{-4M}$. What's the pH?

$pH = -log(2.0 \times 10^{-4}M) = 3.70$

In general, if $pH < 7.00$ we say solution is acidic…

If $pH > 7.00$ the solution is basic.

What is $[H^+]$ if you are given the pH?

$\mathbf{[H^+] = 10^{-pH}}$

pOH

Definition: $pOH = -log[OH^-]$

Before: $[H_3O^+][OH^-] = 1.0x10^{-14}$

(so if you know $[H_3O^+]$, you can know $[OH^-]$)

take -log of both sides:

$-log\{[H_3O^+][OH^-]\} = -log\{1.0x10^{-14}\}$

$-log[H_3O^+] - log[OH^-] = 14.00$

Or, $pH + pOH = 14.00$

(so if you know pH, you can know pOH)

Summary

$[H_3O^+]\ [\ OH^-] = 1.0 \times 10^{-14}$

$pH = -log[H_3O^+]$ and $pOH = -log[OH^-]$

$pH + pOH = 14.00$

$K_a = [H^+][A^-]/[HA]$ $K_b = [HA][OH^-]/[A^-]$

$K_aK_b = 1.0 \times 10^{-14}$

$[H^+] = 10^{-pH}$ $[OH^-] = 10^{-pOH}$

Calculating pH of a solution

There are at least 5 scenarios we'll encounter involving pH calculations. You are expected to **MASTER** these calculations.

1) Strong acid solutions
2) Strong base solutions
3) Pure Weak acid solutions
4) Pure weak base solutions
5) Buffer solutions

Calculations of salt solutions

You are also expected to recognize that salts can have acid or base properties.

E.g.

NH_3 is a weak base. It is NOT a salt.

NH_4Cl is a salt. Is a solution of NH_4Cl an acidic or basic solution?

Well, NH_4^+ is the *conjugate acid* of NH_3 so NH_4^+ is a weal acid. Cl^- is conjugate base of HCl a *strong* acid. So, Cl^- is so weak as to have a negligible effect.

Net ionic equations and ICE

You are also expected to know how to write and recognize *net ionic equations*.

e.g. What is the K_a equilibrium for a NH_4Cl solution?

$NH_4^+ + H_2O \Longleftrightarrow NH_3 + H_3O^+$

Or even, $NH_4^+ \Longleftrightarrow NH_3 + H^+$

You are supposed to master ICE type calculations involving K_a and K_b equilibria.

I=initial, C=change, E= equilibrium

pH of a strong acid solution

As mentioned earlier, there are at least 5 "scenarios" involving pH calculations. Most simple : pH of a strong acid solution.

For example: What is the pH of a 0.020 M HCl solution?

Answer:

Since HCl is a strong acid, it dissociates completely:

$[HCl]_F = 0.020\ M = [H_3O^+]$

So, $pH = -\log [H_3O^+] = -\log [HCl]_F = -\log(.020) = 1.70$

pH of a strong base solution

2nd scenario:

What is the pH of a $5.2x10^{-4}$M NaOH solution?

Answer: Since NaOH is a strong base, we can write:

$[NaOH] = [OH^-] = 5.2\ x10^{-4}M$

It is easier to determine pOH first:

$pOH = -\log[OH^-] = -\log(5.2x10^{-4}) = 3.28$

pH = 14.00 - 3.28 = 10.72 (makes sense?)

Yes because it is a basic pH

pH of a weak acid solution(stopped here Wed)

What is the pH of a $2.5x10^{-3}$M acetic acid (HA, K_a=$1.8x10^{-5}$)

Answer:

Use the K_a equilibrium: $\underset{0.025-x}{HA} \Longleftrightarrow \underset{x}{H^+} + \underset{x}{A^-}$

$x^2/(.025-x) = 1.8x10^{-5} \Rightarrow x^2 + 1.8x10^{-5}x - 4.5x10^{-7} = 0$

Solve for x by the quadratic equation. $x = 6.6x10^{-4}$

$pH = -\log(6.6\ x10^{-4}) = 3.18$

pH of a weak acid solution

Redo same problem using approximation (5% rule):

What's pH of $2.5x10^{-2}$M acetic acid (HA, K_a=$1.8x10^{-5}$)

Answer:Use K_a equilibrium: $\underset{0.025-x}{HA} \Longleftrightarrow \underset{x}{H^+} + \underset{x}{A^-}$

$x^2/(.025-x) = 1.8x10^{-5} \Rightarrow x^2/(.025) \approx 1.8x10^{-5}$

Solve for x: $x = \{(0.025)(1.8x10^{-5})\}^{1/2} = 6.7\ x10^{-4}$

$pH = -\log(6.7x10^{-4}) = 3.17$ (compare with 3.18)

Note: $6.7x10^{-4}(100\%)/.025 = 3\% < 5\%$

pH of a weak base solution

What is the pH of 0.10 M ammonia? ($K_b = 1.8x10^{-5}M$)

Answer: Use the K_b equilibrium. $NH_3 + H_2O <==> NH_4^+ + OH^-$

0.10-x x x

$x^2/(0.10\text{-}x) = 1.8x10^{-5}$. Using 5% rule:

$x^2 \approx (1.8x10^{-5})(0.10) \Rightarrow x = 1.3x10^{-3}$ $=[OH^-]$

pOH = 2.87 => pH = 14.00-2.87 = 11.13

Polyprotic acids

*Poly*protic acids can donate more than 1 proton.

Examples: H_2CO_3 (carbonic acid); H_3PO_4 (phosphoric acid).

If you place H_2CO_3 in water you will have the following K_a equilibrium:

$H_2CO_3 + H_2O <==> H_3O^+ + HCO_3^-$ $K_{a1} = 4.2x10^{-7}$

If you place $NaHCO_3$ in water you have:

$HCO_3^- + H_2O <==> H_3O^+ + CO_3^{2-}$ $K_{a2} = 4.8x10^{-11}$

There are 2 K_a's