II) Weak Base Equilibrium & Kb

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II) Weak Base Equilibrium and K_b

Write a reaction for and find the pH for 0.10M NaOH. Why is it considered a 'strong' base?

$NaOH \xrightarrow{100\%} Na^+ + [OH^-]$ ← dissociates 100% (OH bases are ionic salts)
0.10M → 0.10M

$pOH = -\log(0.10M) = 1.00 \therefore pH = 14 - 1.00 = 13.00$

Write a reaction for 0.10M NH_3 solution . Why is it considered a 'weak' base? dissociate ≅ 5% (less)

How would you find the pH of a weak base solution?

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

weak Base (0.10M), acid (+H⁺); how much is dissoc.? [OH⁻] <<< 0.10M ~5% of 0.10M

Weak base problems such as the one previous can be solved using the K_b constant, a K_{eq} for weak bases (similar to K_a for weak acids).

The larger the K_b, the stronger the base (the more H^+ it will accept). → "greater tendency to accept protons"

Write a K_b expression: $NO_2^-{}_{(aq)} + H_2O_{(l)} \Leftrightarrow HNO_{2(aq)} + OH^-_{(aq)}$

[P]/[R]

$K_b = \frac{[OH^-][HNO_2]}{[NO_2^-]}$ but since $[OH^-] = [HNO_2]$… (equal; dissociate @ 1:1) $K_b = \frac{[OH^-]^2}{[NO_2^-]}$ ← any weak base.

Relationship of K_w, K_a, and K_b for a Conjugate Acid-Base Pair

Write the K_a for HF: $\frac{[H_3O^+][F^-]}{[HF]}$ Write the K_b for F^-: $\frac{[OH^-][HF]}{[F^-]}$

$HF + H_2O \rightleftharpoons F^- + H_3O^+$ (A, B, B, A)

multiply the K_a of HF by the K_b of F^-:

$$K_a \cdot K_b = \frac{[H_3O^+][F^-]}{[HF]} \cdot \frac{[OH^-][HF]}{[F^-]} = [H_3O^+][OH^-] = K_w$$

What results?

$$K_w = K_a \cdot K_b = 1.0 \times 10^{-14} \text{ (@ 25°C)}$$

Conclusion? $K_b = \frac{K_w}{K_a}$ ← use the K_a of the conjugate acid.

(look up the base on the BASE SIDE! (right)…. and use that K_a)

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The acid-base table only lists acid K_a values. Using what you learned above, how would you get the K_b for the corresponding conjugate base?

K_a (conjugate acid) $\cdot$ K_b (base) $= K_w = 1.0 \times 10^{-14}$

Example:
Determine K_b for the weak base SO_4^{2-}:

H^+ → base SO_4^{2-} — conj. acid HSO_4^- ⇐ $K_a = 1.2 \times 10^{-2}$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.2 \times 10^{-2}} = \boxed{8.3 \times 10^{-13}}$$

Example:
Determine K_b for HCO_3^-:
w. base

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = \boxed{2.3 \times 10^{-8}}$$

$H_2CO_3 = 4.3 \times 10^{-7}$

Homework

Practice Questions: Determine the K_b values for the following:

a) HPO_4^{2-} (accepting H^+)

conj. acid $= H_2PO_4$

$$K_b = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7}$$

b) $H_2PO_4^-$

conj. acid $= H_3PO_4$

$$K_b = \frac{1.0 \times 10^{-14}}{7.5 \times 10^{-3}} = 1.3 \times 10^{-12}$$

Assignment 2: Hebden p. 152 #77, 80, 82 & p.130 #35bce, 36

Weak base quantitative problems can be broken into three types.

Type 1 Problems: Finding the pH of a weak base solution

Example: Calculate the [OH$^-$] and pH for a 0.25M solution of the weak base HS^-. $[HS^-]_i$

R	$HS^- + H_2O \rightleftharpoons$	H_2S	$+ OH^-$
I	0.25M	Ø	Ø
C	$-x$		$+x$
E	$0.25 - x \cong 0.25M$		x

(1) (let $x = \Delta[OH^-]$)

assumption

(2) $Kb = \frac{[OH^-]^2_{eq}}{[HS^-]_{eq}} = \frac{x^2}{0.25} = Kb$

conj. acid H_2S → (3) $Kb = \frac{Kw}{Ka} = \frac{1.0\times10^{-14}}{9.1\times10^{-8}} = 1.1\times10^{-7}$

(4) $\therefore x = \sqrt{Kb \cdot 0.25} = \sqrt{(1.1\times10^{-7})(0.25M)}$

$x = 1.66\times10^{-4} M$

(5) $[OH^-] = 1.7\times10^{-4} M$

$pOH = -\log(1.66\times10^{-4}) = 3.78$

$\therefore pH = 14 - 3.78 = 10.22$

Sometimes when salts dissolve in water, one of the ions can act as a weak base in solution, such as the salt in the *next example.*

Example: Calculate the pH of a 0.100M solution of $K_2C_2O_4$.

$K_2C_2O_4 \longrightarrow 2K^+ + [C_2O_4^{-2}]$ act as a weak base.

(1) 0.100M   0.100M

(2)

K	$C_2O_4^{2-} + H_2O \rightleftharpoons$	$HC_2O_4^-$	$+ OH^-$
I	0.100M	Ø	Ø
C	$-x$	$+x$	$+x$
E	(4) $0.100M - x \cong 0.100M$		x

let $x = \Delta[OH^-]$ (3)

assume.

$Kb = \frac{[OH^-]^2}{[C_2O_4^{2-}]}$

(5) $Kb = \frac{Kw}{Ka} = \frac{1.0\times10^{-14}}{6.4\times10^{-5}}$

$Kb = 1.56\times10^{-10} = \frac{x^2}{0.100M}$

(6) $\therefore x = \sqrt{(1.56\times10^{-10})(0.100M)} = 3.9528\times10^{-6} = [OH^-]$

(7) $pOH = -\log(3.9528\times10^{-6}) = 5.4031$

$\therefore pH = 14 - 5.4031 = 8.60$

2 dp.

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Type 2 Problems: Calculating the Initial Concentration of a Weak Base

Example: A solution of NO_2^- has a pH of 8.900. Calculate the $[NO_2^-]$ that would have been required to make this solution.

Let $x = [NO_2^-]_i$

R	$NO_2^- + H_2O \rightleftharpoons$	OH^-	$+ HNO_2$
I	x	$\varnothing$	$\varnothing$
C	-7.94×10^{-6}	$+7.94\times10^{-6}$	
E	$(x - 7.94\times10^{-6}) \cong x$	7.94×10^{-6}	

(1) pH = 8.900
∴ pOH = 14 − 8.900 = 5.100
$[OH^-] = 10^{\wedge}(-5.100)$
$[OH^-] = 7.9433\times10^{-6}$ M (3 d.p. = 3 s.f.)

(4) $K_b = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{4.6\times10^{-4}} = 2.1739\times10^{-11}$ (K_a of HNO_2)

(5) $K_b = \frac{[OH^-]^2}{[NO_2^-]}$ ∴ $2.1739\times10^{-11} = \frac{(7.94\times10^{-6})^2}{(x - 7.94\times10^{-6})}$

$2.1739\times10^{-11}(x - 7.94\times10^{-6}) = 6.309\times10^{-11}$

$x = 2.9$ M $= [NO_2^-]$

Example: A solution of ammonia, NH_3, has a pH of 10.50. Calculate the $[NH_3]$ used to make the solution.

∴ pOH = 14 − 10.50 = 3.50
(1) $[OH^-] = 10^{\wedge}(-3.50) = 3.1623\times10^{-4}$

let $x = [NH_3]_i$

R	$NH_3 + H_2O \rightleftharpoons$	NH_4^+	$+ OH^-$
I	x		$\varnothing$
C	-3.16×10^{-4}	+	$+3.16\times10^{-4}$
E	$x - 3.16\times10^{-4}$ M		3.1623×10^{-4} M

$K_b = \frac{K_w}{K_a} = \frac{[OH^-]^2}{[NH_3]}$

(4) $\frac{1.0\times10^{-14}}{5.6\times10^{-10}} = \frac{(3.16\times10^{-4})^2}{(x - 3.16\times10^{-4})}$

$x = 5.9\times10^{-3} = [NH_3]_i$ (2 sf)

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Type 3 Problems: Finding the K_b of an Unknown Weak Base

Example: A 0.44M solution of the weak base B^- has a pH of 11.12. Calculate the K_b for this base, and the K_a for the conjugate acid, HB at 25°C.

R	$B^- + H_2O \rightleftharpoons$	HB	$+ OH^-$
I	0.44M	$\varnothing$	$\varnothing$
C	$-x$	$+x$	$+x$
E	$0.44 - x \cong 0.44$M		x

∴ $(0.44 - 1.318\times10^{-3})$

$x = 1.318\times10^{-3}$ M

(4) pH = 11.12
pOH = 14 − 11.12
pOH = 2.88
$[OH^-] = 10^{\wedge}(-2.88)$
$[OH^-] = 1.318\times10^{-3}$ M

(5) $K_b = \frac{[OH^-]^2}{[B^-]_{eq}} = \frac{(1.318\times10^{-3})^2}{(0.44 - 1.318\times10^{-3})} = 3.97\times10^{-6} \cong 4.0\times10^{-6} = K_b$

(6) $K_a = \frac{K_w}{K_b} = \frac{1.0\times10^{-14}}{3.97\times10^{-6}} = 2.5\times10^{-9} = K_a$

(not an acid/base pair listed on our table)

(not an acid / base
listed on our table)

Assignment 3: Hebden p.153 #84-89

III) Writing Formula (Molecular), Complete Ionic, and Net Ionic Equations for Acid/Base Reactions

1. **Strong Acid/Strong Base** (Neutralization):

F: $HCl_{(aq)}$ + $NaOH_{(aq)}$ → ________ + ________

C: →

N: →

[Ka Kb Kw pH pOH pKa pKb H+ OH- Calculations - Acids & Bases, Buffer Solutions , Chemistry Review]