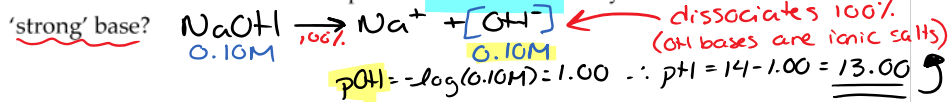


II) Weak Base Equilibrium & Kb

March 5, 2018 1:44 PM

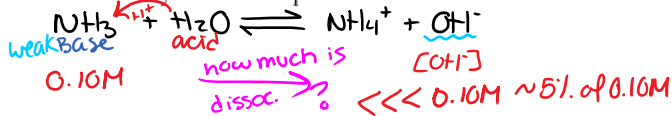
II) Weak Base Equilibrium and Kb

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



Write a reaction for 0.10M NH₃ solution. Why is it considered a 'weak' base?

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



dissociate ≈ 5% (less)

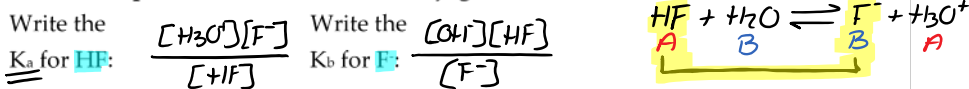
Weak base problems such as the one previous can be solved using the K_b constant, a K_b for weak bases (similar to K_a for weak acids).

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: $\text{NO}_2^-(\text{aq}) + \text{H}_2\text{O} \rightleftharpoons \text{HNO}_2(\text{aq}) + \text{OH}^-(\text{aq})$

$K_b = \frac{[\text{OH}^-][\text{HNO}_2]}{[\text{NO}_2^-]}$ but since *equal* $[\text{OH}^-] = [\text{HNO}_2] \dots$ $K_b = \frac{[\text{OH}^-]^2}{[\text{NO}_2^-]}$ *any weak base*
dissociate @ 1:1

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



multiply the K_a of HF by the K_b of F⁻:

$K_a \cdot K_b = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \cdot \frac{[\text{OH}^-][\text{HF}]}{[\text{F}^-]} = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$

What results?

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

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

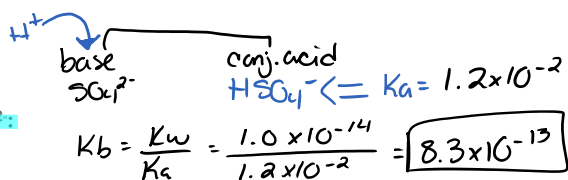
(look up the base on the BASE SIDE! (right).... and use that K_a)

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?

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

Example:

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



Example:

Determine K_b for

HCO_3^- :
w. base

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

$\text{H}_2\text{CO}_3 = 4.3 \times 10^{-7}$ ↑

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

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

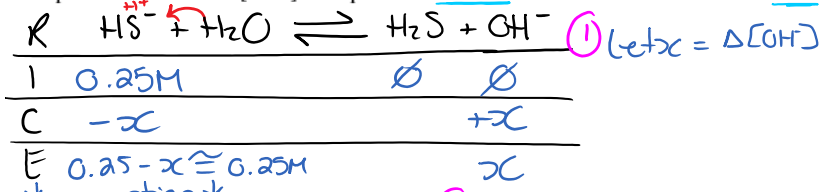
Homework

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⁻]:



assumption

$$K_b = \frac{[\text{OH}^-]^2}{[\text{HS}^-]_{\text{eq}}} = \frac{x^2}{0.25} = K_b \quad \therefore x = \sqrt{K_b \cdot 0.25} = \sqrt{(1.1 \times 10^{-7})(0.25M)}$$

$$x = 1.66 \times 10^{-4} \text{ M} \quad (5)$$

$$[\text{OH}^-] = 1.7 \times 10^{-4} \text{ M}$$

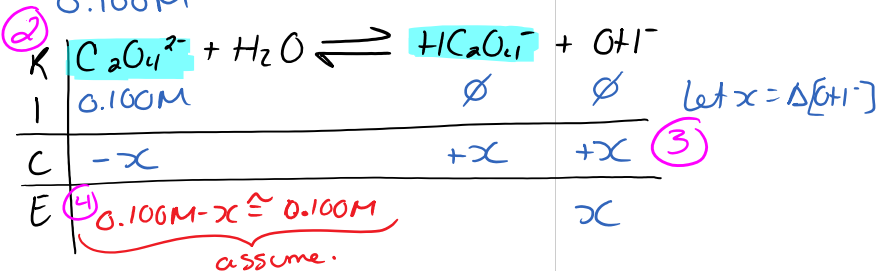
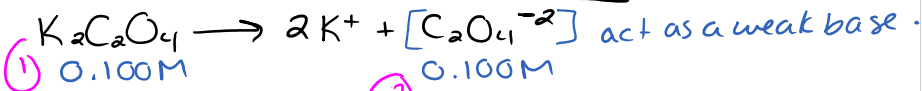
$$\text{pOH} = -\log(1.66 \times 10^{-4}) = 3.78$$

$$\therefore \text{pH} = 14 - 3.78 = 10.22$$

conj. acid H₂S $\rightarrow K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{9.1 \times 10^{-8}} = 1.1 \times 10^{-7}$ (3)

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 K₂C₂O₄.



$$K_b = \frac{[\text{OH}^-]^2}{[\text{C}_2\text{O}_4^{2-}]}$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.4 \times 10^{-5}}$$

$$K_b = 1.56 \times 10^{-10} = \frac{x^2}{0.100M} \quad \therefore x = \sqrt{(1.56 \times 10^{-10})(0.100M)} = 3.9528 \times 10^{-6} = [\text{OH}^-] \quad (6)$$

$$\text{pOH} = -\log(3.9528 \times 10^{-6}) = 5.4031 \quad (7)$$

$$\therefore \text{pH} = 14 - 5.4031 = 8.60$$

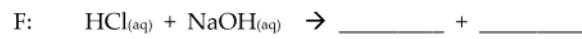
2 dp.

(not an acid, but will
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):



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

