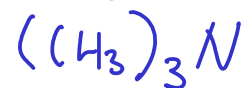


An aqueous solution of 0.25 M trimethylamine has a pH of 11.63. What's the value of  $K_b$ ?



$$K_b = \frac{[(\text{CH}_3)_3\text{NH}^+][\text{OH}^-]}{[(\text{CH}_3)_3\text{N}]} = ???$$

Species	[Initial]	$\Delta$	[Equilibrium]
$(\text{CH}_3)_3\text{NH}^+$	0	+X	X
$\text{OH}^-$	0	+X	X
$(\text{CH}_3)_3\text{N}$	0.25	-X	0.25 - X

$$K_b = \frac{(x)(x)}{(0.25 - x)}$$

$$K_b = \frac{x^2}{0.25 - x}$$

If we want to know what 'K<sub>b</sub>' is, we need to find the value of 'x', but NOT by solving this quadratic equation. (K<sub>b</sub> is also unknown)

$$K_b = \frac{x^2}{0.25 - x}$$

$$x = [\text{OH}^-]$$

... and hydroxide concentration is related to pH!

$$\text{pH} + \text{pOH} = 14.00$$

$$11.63 + \text{pOH} = 14.00$$

$$\text{pOH} = 2.37$$

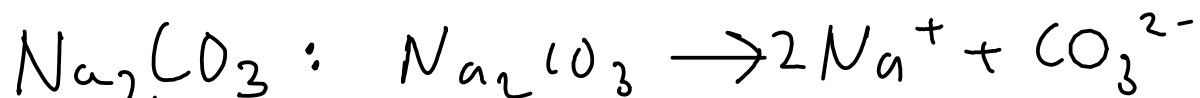
$$x = [\text{OH}^-] = 10^{-2.37}$$

$$x = 0.0042657952$$

$$K_b = \frac{x^2}{(0.25 - x)} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952} = 7.4 \times 10^{-5} = K_b$$

## SALTS

- Compounds that result from the reaction of an acid and a base.
- Salts are strong electrolytes (completely dissociate in water) IF SOLUBLE (not all salts dissolve appreciably).
- Most ionic compounds are considered salts (they can be made by some reaction between the appropriate acid and base)
- Salts have acidic and basic properties! The ions that form when salts are dissolved can be acidic, basic, or neutral.
  - Salts made from WEAK ACIDS tend to form BASIC solutions
  - Salts made from WEAK BASES tend to form ACIDIC solutions



Do any of these ions have acidic or basic properties?

$\text{Na}^+$ : neutral. Not a proton donor or a proton acceptor

$\text{CO}_3^{2-}$ : BASIC, since it can accept protons to form the weak acid CARBONIC ACID in solution.



ACID

BASE

## SALT OF A WEAK ACID

ex:  $\text{NaC}_2\text{H}_3\text{O}_2$ 

For this reaction to occur, HA MUST be stable in water. In other words, a weak acid.



The anion is a BASE. It can accept a proton from water to form the weak (therefore stable as a molecule!) acid HA

$$K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} \quad \left| \text{This is the base ionization constant for } \text{A}^- \right.$$

Since  $\text{A}^-$  and HA are a conjugate pair, the ionization constants are related!

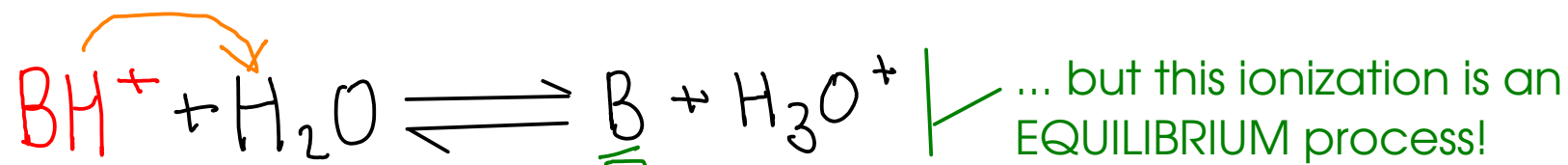
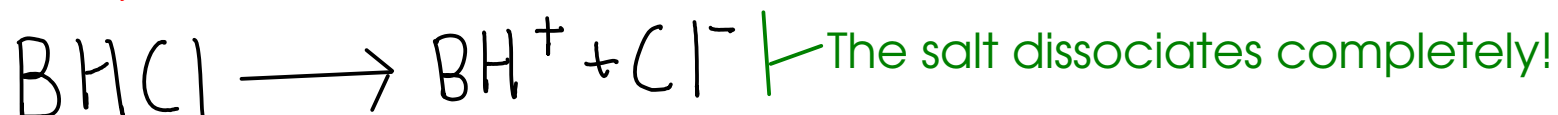
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

$1.0 \times 10^{-14}$

$$14 = \text{p}K_a + \text{p}K_b$$

You will generally not find both the  $K_a$  AND  $K_b$  for a conjugate pair in the literature, since one can be easily converted to the other!

## SALT OF A WEAK BASE

ex:  $\text{NH}_4\text{Cl}$ 

$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for BH}^+ \right.$$

$$K_w = (K_{a,\text{BH}^+})(K_{b,\text{B}})$$

$1.0 \times 10^{-14}$

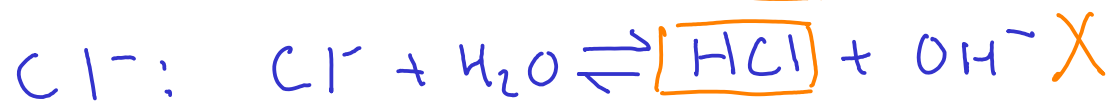
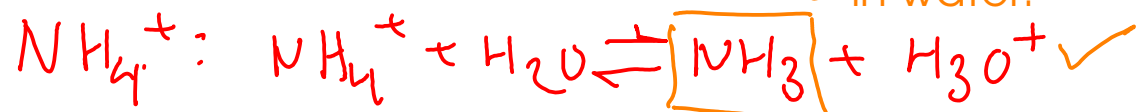
Find the pH for salt solutions just like you would find pH for any other weak acid or weak base solutions. Only trick is to find out whether the salt is actually acidic or basic!

0.100 M  $\text{NH}_4\text{Cl}$  ... Find the pH of the solution



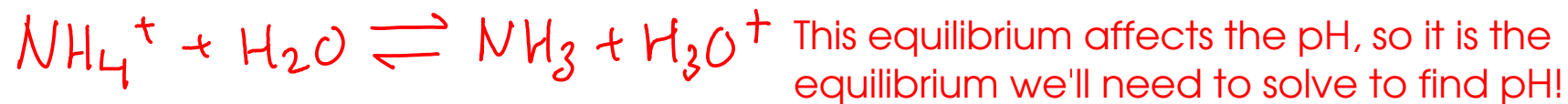
Acidic, basic, or neutral salt?

This is the WEAK BASE ammonia. Stable in water.



This is a STRONG ACID, which does not exist as a stable molecule in water.

The conjugate of a strong acid or base is NEUTRAL - does not affect pH!





$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

$$= 5.56 \times 10^{-10}$$

Where do we get this  $K_a$ ?

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$$

$$K_a \times K_b = 1.0 \times 10^{-14}$$



Species	[Initial]	$\Delta$	[Equilibrium]
$\text{NH}_3$	0	+X	X
$\text{H}_3\text{O}^+$	0	+X	X
$\text{NH}_4^+$	0.100	-X	0.100 - X

$$\frac{x^2}{0.100 - x} = 5.56 \times 10^{-10}$$

$$\downarrow \begin{array}{l} 0.100 - x \approx 0.100 \\ x \ll 0.100 \end{array}$$

$$\frac{x^2}{0.100} = 5.56 \times 10^{-10}$$

$$[\text{H}_3\text{O}^+] = 7.45 \times 10^{-6} \text{ M}$$

$$\text{pH} = -\log_{10}(7.45 \times 10^{-6})$$

$$= 5.13$$

Compare:

pH = 1.00 for 0.100 M strong acid

pH = 2.16 for 0.100 M nitrous acid

pH = 7.00 for pure distilled water

$$x = 7.45 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

0.100 M  $\text{NaC}_2\text{H}_3\text{O}_2$ , Find pH



Check these ions to see if they are acidic, basic, or neutral

$\text{Na}^+$ : Cannot be B-L acid (no protons), also not likely to be B-L base; positively charged.

$\text{C}_2\text{H}_3\text{O}_2^-$ : Has protons, but might not be as likely to donate one as to receive one due to the negative charge.



Acetic acid is a WEAK ACID and stable in water, so the acetate ion can function as a BASE.



$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = ?$$

$K_b$  for acetate ion isn't in our charts in the appendix, but we CAN find the  $K_a$  for acetic acid!

$$K_{a, \text{HC}_2\text{H}_3\text{O}_2} = 1.7 \times 10^{-5} \quad (\text{pA-13})$$

$$K_a + K_b = 1.00 \times 10^{-14}, \text{ so } K_b = 5.88 \times 10^{-10}$$

True for any conjugate pair  
(like acetic acid and acetate ion)





$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5.88 \times 10^{-10}$$

Species	[Initial]	$\Delta$	[Equilibrium]
$\text{HC}_2\text{H}_3\text{O}_2$	0	+x	x
$\text{OH}^-$	0	+x	x
$\text{C}_2\text{H}_3\text{O}_2^-$	0.100	-x	0.100 - x

We'll need to solve for the HYDROXIDE ION concentration, then use that to find the pH. (Hydroxide concentration is related to hydronium)!

$$\frac{x^2}{0.100 - x} = 5.88 \times 10^{-10}$$

$$\frac{x^2}{0.100} = 5.88 \times 10^{-10}$$

$$x = 7.67 \times 10^{-6}$$

$$[\text{OH}^-] = 7.67 \times 10^{-6}$$

$$[\text{OH}^-] = 7.67 \times 10^{-6}$$

$$\begin{aligned} \text{pOH} &= -\log_{10}(7.67 \times 10^{-6}) \\ &= 5.12 \end{aligned}$$

Convert pH to pOH using the water equilibrium relationship::

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} = 14.00 - 5.12$$

$$\text{pH} = 8.88$$

Compare:

pH = 7.00 for pure distilled water

pH = 13.00 for 0.100 M sodium hydroxide (strong base)