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

$(CH_3)_3N$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]} = ???$$

Species	[Initial]	Δ	[Equilibrium]
$(CH_3)_3NH^+$	0	+x	x
OH^-	0	+x	x
$(CH_3)_3N$	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 learn the value of Kb, we need to find some other way (other than the equation here) to find the value of 'x'!

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

We know that 'x' equals the HYDROXIDE ION CONCETRATION (see the chart on the previous page)

$pH = 11.63$; $pH + pOH = 14.00$

$pOH = 2.37$

$[OH^-] = 10^{-2.37} = 0.0042657952 = x$

Now, plug 'x' into the equilibtrium expression

$$K_b = \frac{x^2}{0.25 - x} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952}$$

$$\boxed{K_b = 7.4 \times 10^{-5}}$$

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

Na_2CO_3: $Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$

Do any of these ions have acidic or basic properties?

Na^+: neutral. Not a proton donor or a proton acceptor

CO_3^{2-}: BASIC, since it can accept protons to form the weak acid CARBONIC ACID in solution.

$$H_2CO_3 + 2H_2O \rightleftharpoons 2H_3O^+ + CO_3^{-2}$$

ACID BASE

SALT OF A WEAK ACID

ex: $NaC_2H_3O_2$

$$NaA \longrightarrow Na^+ + A^-$$

The salt dissolves completely!

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

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

... but the ionization of the salt's anion is an EQUILIBRIUM!

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

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

This is the base ionization constant for A^-

Since A^- and HA are a conjugate pair, the ionization constants are related!

$$\underset{1.0\times10^{-14}}{K_w} = (K_{a,HA})(K_{b,A^-})$$

$$14 = pK_a + pK_b$$

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

SALT OF A WEAK BASE

ex: NH_4Cl

$$BHCl \longrightarrow BH^+ + Cl^-$$ The salt dissociates completely!

$$BH^+ + H_2O \rightleftharpoons \underline{B} + H_3O^+$$ ... but this ionization is an EQUILIBRIUM process!

$$K_a = \frac{[B][H_3O^+]}{[BH^+]}$$ Acid ionization constant for BH^+

$$\underset{1.0\times10^{-14}}{K_w} = (K_{a,BH^+})(K_{b,B})$$

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 NH_4Cl ... Find the pH of the solution

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$

This is the WEAK BASE ammonia. Stable in water.

NH_4^+: $NH_4^+ + H_2O \rightleftharpoons \boxed{NH_3} + H_3O^+$

Cl^-: $Cl^- + H_2O \rightleftharpoons \boxed{HCl} + OH^-$

This is the STRONG ACID hydrochloric acid. It cdompletely ionizes in water, so this molecule won't form - it isn't stable in water.

The conjugate of a strong acid or base is neutral (it doesn't affect the pH)

We need to solve THIS equilibrium:

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

This equilibrium is the only one of the two above that can affect pH.

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

$$K_{a,NH_4^+} = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

We can get Ka by first getting the Kb for the conjugate (ammonia):

$$K_{b,NH_3} = 1.8 \times 10^{-5}$$

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

$$\text{So, } K_{a,NH_4^+} = 5.56 \times 10^{-10}$$

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	$+x$	x
H_3O^+	0	$+x$	x
NH_4^+	0.100	$-x$	$0.100 - x$

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

$x \ll 0.100$

$0.100 - x \approx 0.100$

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

$$x = 7.45 \times 10^{-6} = [H_3O^+]$$

$$\boxed{pH = 5.13}$$

Compare:

pH = 1.00 for 0.100 M strong acid

pH = 2.16 for 0.100 M nitrous acid (weak acid)

pH = 7.00 for distilled water