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

$(CH_3)_3N$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^- \; ; \; K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

Set up an equilibrium chart to reduce the number of variables.

Species	[Initial]	Δ	[Equilibrium]
$(CH_3)_3NH^+$	0	$+x$	x
OH^-	0	$+x$	x
$(CH_3)_3N$	0.25	$-x$	$0.25 - x$

Let "x" equal the change in trimethylammonium ion concentration.

It also equals the hydroxide concentration at equilibrium!

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

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

We need to find a way to calculate "x" WITHOUT using this equation! Then, we can plug the value of "x" in and find Kb.

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

Remember that "x" is the hydroxide concentration. We can use the pH (11.63) to help us find hydroxide concentration!

Find pOH using pH identities.

$$pH + pOH = 14.00$$

$$11.63 + pOH = 14.00$$

$$pOH = 2.37$$

Now, find (OH-) ...

$$[OH^-] = 10^{-pOH}$$

$$[OH^-] = 10^{-2.37} = 0.004265795\text{2 M } OH^- = x$$

Plug value for "x" into the Kb expression.

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

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

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

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.

$$\underbrace{H_2CO_3}_{\text{ACID}} + 2H_2O \rightleftharpoons 2H_3O^+ + \underbrace{CO_3^{-2}}_{\text{BASE}}$$

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!

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

1.0×10^{-14}

$$14 = pK_a + pK_b$$

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

Find the ions produced by the salt. Are they acids or bases? $NH_4Cl \rightarrow NH_4^+ + Cl^-$

$$NH_4^+ + H_2O \rightleftharpoons \underline{NH_3} + H_3O^+$$

This is ammonia, a WEAK BASE. It's water-stable, so this reaction is viable.

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

This is hydrochloric acid, a STRONG ACID. It's NOT water-stable, since it fully ionzes in water. This reaction won't go.

We need to solve the equilibrium of the ammonium ion.

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+ \; ; \; K_a = \frac{[NH_3][OH^-]}{[NH_4^+]} = \; ?$$

We can't find the Ka for ammonium ion in OpenStax Appendix H. We can, however, find the Kb of its conjugate (ammonia) in Appendix I:

$$K_{b,NH_3} = 1.8 \times 10^{-5} \qquad K_a \times K_b = 1.0 \times 10^{-14}$$

$$K_a (1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$$

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