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

$(CH_3)_3N$

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

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

We don't look up a value for Kb, since that's actually what we want to solve for ...

Let's set up a concentration table just like we have done for the previous acid-base problems...

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

$$K_b = \frac{[(CH_3^+)_3NH^+][OH^-]}{[(CH_3)_3N]} = \boxed{\frac{(X)(X)}{0.25-X} = K_b}$$

At the moment, we still can't solve for Kb because we don't know what the value of "x" is ...

If only there were some other way to solve for "x" ... Let's look at the pH and see if we can solve for "x" with it ...

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

We know the solution has a pH of 11.63. Using that, we can determine pOH ...

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

$$11.63 + pOH = 14.00 \text{ ; } pOH = 2.37$$

Now, find the concentration of hydroide ion. Based on how we defined "x" in the chart, hydroxide concentration also equals "x"...

$$[OH^-] = 10^{-pOH} = 10^{-2.37} = 0.004265795 2 = X$$

Plug into the equilibrium expression to find Kb ...

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

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

POLYPROTIC ACIDS

Find pH of 0,10 M H_3PO_4

... what's special about phosphoric acid?

① $H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$

② $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$

③ $HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$

Phosphoric acid has THREE acidic protons!

$K_{a1} = 6,9 \times 10^{-3}$

$K_{a2} = 6,2 \times 10^{-8}$

$K_{a3} = 4,8 \times 10^{-13}$

The first dissocation is dominant here, and for simple calculations of phosphoric acid in water, we will simply use the first ionization and ignore the other two.

Remember: This is a weak acid. It exists in water mostly as undissociated phosphoric acid molecules.

Solve the equilibrium of phosphoric acid's FIRST proton:

$H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$; $K_a = 6.9 \times 10^{-3}$

$$K_a = \frac{[H_2PO_4^-][H_3O^+]}{[H_3PO_4]} = 6.9 \times 10^{-3}$$

Species	[Initial]	Δ	[Equilibrium]
$H_2PO_4^-$	0	$+x$	x
H_3O^+	0	$+x$	x
H_3PO_4	0.10	$-x$	$0.10 - x$

$$\frac{x^2}{0.10 - x} = 6.9 \times 10^{-3}$$

assume $x \ll 0.10$

so $0.10 - x \simeq 0.10$

$$\frac{x^2}{0.10} = 6.9 \times 10^{-3}$$

$$x = 0.026267851 = [H_3O^+]$$

So, $pH = -\log(0.026267851)$

$$\boxed{pH = 1.58}$$