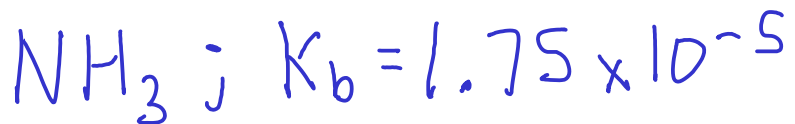


Consider an 0.100 M solution of the weak base ammonia:



What is the pH?



$$K_b = 1.75 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

Which term in this expression are we really interested in? Solve to get the HYDROXIDE concentration, since it can be easily related to hydronium (and pH).

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

Plug into the equilibrium expression

$$1.75 \times 10^{-5} = \frac{(x)(x)}{(0.100 - x)} = \frac{x^2}{0.100 - x}$$

Solve for 'x':

$$1.75 \times 10^{-5} = \frac{x^2}{0.100 - x}$$

This is a QUADRATIC EQUATION. But, we expect that 'x' will be small compared to 0.100. So we can simplify this equation

$$\downarrow x \ll 0.100, \text{ so } 0.100 - x \approx 0.100$$

$$1.75 \times 10^{-5} \approx \frac{x^2}{0.100}$$

$$0.0013228757 = x = [\text{OH}^-] \quad \text{HYDROXIDE ion concentration!}$$

$$-\log_{10}(0.0013228757) = 2.88 = \text{pOH}$$

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

$$\text{SO, pH} = 14.00 - 2.88 = \boxed{11.12}$$

If you had used the quadratic equation to solve this problem, you would have gotten a pH of 11.12 - no difference from this method, at least to two significant figures!

Compare pH to the pH of an 0.100 M solution of the strong base NaOH:

$$\text{pH}_{\text{NH}_3} \approx 11.12$$



$$[\text{OH}^-] = 0.100$$

$$\text{pOH} = 1.00$$

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

$$\boxed{\text{pH} = 13.00}$$

The higher the K_a or K_b value, the stronger the acid or base!

Find the pH and the degree of ionization for an 0.10 M solution of formic acid: HCHO_2



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = 1.7 \times 10^{-4}$$

Constant's value at 25 C
obtained from chart in
textbook, page A-13

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	+ x	x
CHO_2^-	0	+ x	x
HCHO_2	0.10	- x	0.10 - x

$$1.7 \times 10^{-4} = \frac{(x)(x)}{0.10 - x}$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10 - x}$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10 - x} \quad \text{Assume that } x \text{ is much smaller than } 0.10$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10}$$

This number is indeed much smaller than 0.10

$$x = 0.0041231056 = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log_{10}(0.0041231056) = 2.38 = \text{pH}$$

Degree of ionization? DEGREE OF IONIZATION is the fraction of a weak electrolyte (acid or base) that dissociates in water.

$$\frac{[\text{CHO}_2^-]}{[\text{HCHO}_2]} = \frac{[\text{H}_3\text{O}^+]}{[\text{HCHO}_2]} = \frac{0.0041231056}{0.10} = 0.041 = \text{D.O.I.}$$

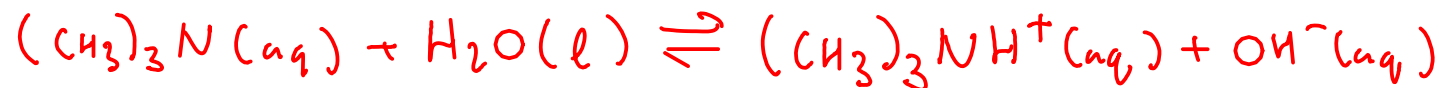
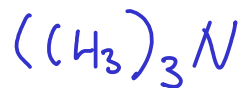
Sometimes, we express degree of ionization as a percent ... PERCENT IONIZATION

$$\% = \text{D.O.I.} \times 100\% = 0.041 \times 100\% = 4.1\% \text{ ionized}$$

... so about 96% of this acid exists in solution as undissociated formic acid molecules.

(WEAK acids exist in solution mostly as undissociated molecules!)

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]	Δ	[Equilibrium]
$(\text{CH}_3)_3\text{NH}^+$	0	+x	x
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_b is, we need to find the value of 'x', but NOT by solving this equation.

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

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

... but concentration of hydroxide is related to pH

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

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

$$\text{pOH} = 2.37$$

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

$$[\text{OH}^-] = 0.0042657952$$

$$\text{So, } x = 0.0042657952$$

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

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