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

$pH_{, NH_3} = 11.13$

$NaOH \rightarrow Na^+ + OH^-$

So, 0.100 M NaOH has $[OH^-] = 0.100$

$pOH = -\log_{10}(.100) = 1.00$

$pH = 14.00 - 1.00 = \boxed{13.00}$

The stronger the base:

- the higher the pH will be for a solution of given concentration
- the higher the HYDROXIDE concentration (compared to the nominal base concentration)

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

$$HCHO_2 + H_2O \rightleftharpoons CHO_2^- + H_3O^+ \; ; \; K_a = 1.7 \times 10^{-4}$$

$$K_a = \boxed{\frac{[CHO_2^-][H_3O^+]}{[HCHO_2]} = 1.7 \times 10^{-4}}$$

Constant's value was obtained at 25 C using the chart on page A-13 in Ebbing

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

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

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

$$\frac{x^2}{0.10-x} = 1.7\times10^{-4}$$

$$\Big\downarrow \; x \ll 0.10$$

$$\frac{x^2}{0.10} = 1.7\times10^{-4}$$

$$x = 0.0041231056 = [H_3O^+]$$

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

What about DEGREE OF IONIZATION? DEGREE OF IONIZATION is the fraction of a weak acid or base that ionizes in water. (Sometimes expressed as a percent - percent ionization)

$$\frac{[CHO_2^-]}{[HCHO_2]_{initial}} = \frac{[H_3O^+]}{[HCHO_2]_{initial}} = \frac{0.0041231056}{0.10} = \boxed{0.041 = DOI}$$

Sometimes, we express DOI as a percentage:

$$\% = DOI \times 100\% = \boxed{4.1\% \text{ ionized}}$$

Check in experiment 16A: As you dilute a solution of weak acid or weak base, the equilibrium will shift to the right (towards the production of more ions) - based on Le Chateleir's Principle. So, you should see an INCREASE in degree of/percent ionization as the solution is diluted - EVEN AS THE OVERALL pH GOES UP!

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

$(CH_3)_3N$

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

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

To get the number of unknowns down, let's express all the concentrations in terms of a single variable - like we'd do for a typical equilibrium problem.

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 find Kb, we need to find some way - OTHER THAN THIS EQUATION - to find the value of 'x'!

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

We know that 'x' equals HYDROXIDE concentration. Since hydroxide concentration is related to pH (which we DO know), we can use the pH to find 'x' ... and then find Kb.

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

$$pOH = 2.37$$

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

Now, plug 'x' back into the equilibrium 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!

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

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^+

$$\underset{1.0 \times 10^{-14}}{K_w} = (K_{a,BH^+})(K_{b,B})$$

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!