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: $HCHO_2$

$$HCHO_2 + H_2O \rightleftarrows H_3O^+ + CHO_2^-$$

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = 1.7\times10^{-4}$$

Value from page A-13 in Ebbing 10th edition...

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

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

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

Assume $x \ll 0.10$, so $0.10 - x \approx 0.10$

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

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

$$\boxed{pH = 2.38}$$

What is DEGREE OF IONIZATION? It's the fraction of a weak acid or base that ionizes in water!

$$\frac{[CHO_2^-]}{[HCHO_2]_{initial}} = \frac{[H_3O^+]}{[HCHO_2]_{initial}} = \frac{0.0041231056}{0.10} = \boxed{0.041 = 0.0\overline{0}}$$

Sometimes we express this as a percent to make it clearer ... this is called PERCENT IONIZATION:

$$\% = 0.0\overline{0} \times 100\% = \boxed{4.1\% \text{ ionized}}$$

If we DILUTED THE ACID, ...

$$HCHO_2 + H_2O \rightleftharpoons H_3O^+ + CHO_2^-$$

... amount of water goes up, equilibrium shifts to RIGHT.

When you do Experiment 16A. By Le Chateleir's Principle, adding water to the equilibrium should force it to the right - meaning that more acid will ionize - even as the pH goes up!. Therefore, the degree of (or percent) ionization should INCREASES as the concentration of the acid DECREASES. Check this with your experiment 16A data on acetic acid.

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

$(CH_3)_3N$

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

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

How do we find Kb? Let's set this up like a normal equilibrium problem...

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

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

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

If we want to find Kb, we must use other information to find 'x'. We know pH, which gives us a way to find hydronium and hydroxide concentrations...

$K_b = \frac{x^2}{0.25 - x}$ and $pH = 11.63$

$pOH = 14.00 - 11.63 = 2.37$

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

So, $x = 0.0042657952$

Now, just plug in 'x' to the equilibrium expression and we'll have Kb ...

$$K_b = \frac{(0.0042657952)^2}{0.25 - 0.0042657952} = \boxed{7.4 \times 10^{-5} = K_b}$$

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.

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

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.

$$\underline{A^-} + H_2O \rightleftharpoons HA + OH^-$$

... 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!

$$BH^+ + H_2O \rightleftharpoons B + H_3O^+$$

... 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!