Consider an 0.100 M solution of the weak base ammonia:

NH_3 ; $K_b = 1.8 \times 10^{-5}$ (p A-14, Ebbing 9th)

What is the pH?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$\boxed{K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}}$$

We need to solve this, but which one of these terms are we interested in?

We want to solve for the HYDROXIDE concentration, since it's closely related to - and can be easily converted to - 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:

$$1.8 \times 10^{-5} = \frac{(x)(x)}{(0.100 - x)} = \frac{x^2}{0.100 - x}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.100 - x}$$

This is a QUADRATIC equation. We could solve ir with the quadratic formula, BUT like the acid example, we expect that 'x' is small compared to 0.100 .

$0.100 - x \approx 0.100$

$$1.8 \times 10^{-5} = \frac{x^2}{0.100}$$

$$x = 0.001341641 = [OH^-]$$

Be careful! This is HYDROXIDE concentration, not HYDRONIUM concentration!

$$pOH = -\log_{10}(0.001341641) = 2.87$$

$$pH = 14.00 - pOH = 14.00 - 2.87$$

$$\boxed{pH = 11.13} \ *$$

* If you had solved this problem using the quadratic, you would have found the pH of the ammonia solution was 11.13 - no difference to two significant figures.

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 bas concentration)

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

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

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

Constant's value at 25C was obtained from the chart in Ebbing on page A-13 (9th edition)

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

$$\frac{(x)(x)}{(0.1-x)} = 1.7\times10^{-4}$$

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

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

$\downarrow$ $x << 0.10$

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

$x = 0.004123105\,6 = [H_3O^+]$

$pH = -\log_{10}(0.004123105\,6)$

$$\boxed{pH = 2.38}$$

DEGREE OF IONIZATION is the fraction of a weak acid or base that ionizes in water.

$$\frac{[CHO_2^-]}{[HCHO_2]_{initial}} = \frac{[H_3O^+]}{[HCHO_2]_{initial}} = \frac{0.004123105\,6}{0.10} = \boxed{0.041 = D.O.I.}$$

Sometimes, we express degree of ionization as a percentage - the PERCENT IONIZATION:

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

Check this in experiment 16A: A more dilute acid solution should have a HIGHER degree of ionization than a more concentrated one thanks to Le Chateleir's principle - even if the dilution causes the pH to increase overall.

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

$((CH_3)_3N$

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

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

Species	[Initial]	Δ	[Equlibrium]
$(CH_3)_3NH^+$	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 the value of Kb, then we need to come up with some OTHER way of finding the value of 'x'

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

We know that 'x equals the HYDROXIDE concentration. Since hydroxide concentration is related to pH, we can use the pH to find 'x'.

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

$pOH = 2.37$

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

$[OH^-] = x = 0.0042657952$

Now, plug 'x' 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.

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

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.

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

... 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!

$$\underset{1.0\times10^{-14}}{K_w} = (K_{a,HA})(K_{b,A^-})$$

$$14 = pK_a + pK_b$$

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

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$

Is this salt acid, basic, or neutral?

This is the WEAK BASE ammonia. Stable in water

NH_4^+ : $NH_4^+ + H_2O \rightleftharpoons \boxed{NH_3} + H_3O^+$

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

This is the STRONG ACID hydrochloric acid, which completely ionizes in water. This is not a stable molecule in water.

The conjugate of a STRONG acid or base is NEUTRAL - it does not affect the pH.

We will need to solve the following equilibrium:

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

This equilibrium is the only one of the two above that affects the pH!

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$K_{a,NH_4^+} = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

We can get Ka for the acid from Kb of the base (ammonia):

$K_{b,NH_3} = 1.8 \times 10^{-5}$

$K_a \times K_b = 1.0 \times 10^{-14}$

So, $K_a = 5.56 \times 10^{-10}$

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	$+x$	x
H_3O^+	0	$+x$	x
NH_4^+	0.100	$-x$	$0.100-x$

$$\frac{x^2}{0.100-x} = 5.56 \times 10^{-10}$$

$x \ll 0.100$

$0.100 - x \approx 0.100$

$$\frac{x^2}{0.100} = 5.56 \times 10^{-10}$$

$x = 7.46 \times 10^{-6}$

So, $[H_3O^+] = 7.46 \times 10^{-6}$

$$\boxed{pH = 5.13}$$

Compare:
pH = 1.00 for 0.100 M strong acid
pH = 2.16 for 0.100 M nitrous acid (weak acid)
pH = 7.00 for distilled water