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?

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

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

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

We want to solve for the HYDROXIDE concentration, since it's closely related to pH.

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

Now, 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 can solve it with the quadratic formula, but like the acid example, we expect that 'x' is small compared to 0.100 ...

$$0.100 - x \approx 0.100$$

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

$$x = 0.001341640 8 = [OH^-]$$

Be careful! This is the HYDROXIDE concentration, not the hydronium concentration!

$$pOH = -\log_{10}(0.0013416408) = 2.87$$

$$pH = 14.00 - pOH = \boxed{11.13} \ ^*$$

* If you had solved the problem with the quadratic equation, you would have calculated a pH of 11.13 - there was 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 base concentration)

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

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

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

Constant found on page A-13, Ebbing and Gammon 9th edition.

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

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

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

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

$x \ll 0.10$

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

$x = 0.0041231056 = [H_3O^+]$

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

$pH = 2.38$

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

$$\frac{[CHO_2^-]}{[HCHO_2]_{initial}} = \frac{[H_3O^+]}{[HCHO_2]_{initial}} = \frac{0.0041231056}{0.10} = 0.041 = D.O.I.$$

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

$\% = D.O.I. \times 100\% =$ 4.1% ionized

Check this in experiment 16A: A more dilute acid solution should have a HIGHER degree of ionization than a more concentrated one. Why?

(Le Chateleir's principle - adding a reactant (water) causes equilibrium shift to the right)

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

$(CH_3)_3N$

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

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

Species	[Init.]	Δ	[Equil.]
$(CH_3)_3NH^+$	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 know the value of Kb, we need to find some way of determining the value of 'x'.

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

We know 'x' equals the HYDROXIDE concentration. Since we know pH, we can calculate what the hydroxide concentration is.

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

$$pOH = 2.37$$

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

Now, just plug 'x' into the expression for Kb:

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

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

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!

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

... but this ionization is an EQUILIBRIUM process!

$$K_a = \frac{[B][H_3O^+]}{[BH^+]}$$

Acid ionization constant for BH^+

$$K_w = (K_{a,BH^+})(K_{b,B})$$

1.0×10^{-14}

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!