Compare:

- Weak acid HNO_2 : pH of 0.10 M solution = 2.17

Let's compare the pH of the weak nitrous acid with the pH of a strong acid like nitric acid:

0.10 M HNO_3, what is pH?

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$

$$[H_3O^+] = [HNO_3]_{nominal} = 0.10\ M$$

$$pH = 1.00$$

The stronger the acid:

- the lower the pH of a solution of given concentration will be
- the higher the concentration of hydronium ion (when compared to the nominal acid concentration)

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?

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

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

To get the pH, we'll first need to find out the HYDROXIDE concentration, since that's the only species in the equilibrium that can be related to pH.

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

Let "x" equal the change in hydroxide ion concentration.

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

Solve for "x" ...

$$\frac{(x)(x)}{(0.100-x)} = 1.8\times10^{-5}$$

$$\frac{x^2}{0.100-x} = 1.8\times10^{-5}$$

Assume x is small, so

$0.100 - x \approx 0.100$

$$\frac{x^2}{0.100} = 1.8\times10^{-5}$$

Be careful when solving for the pH of BASES, as the equilibrium calculation will give you HYDROXIDE concentration, and you'll need an extra step to get pH ...

$x = 0.0013416408\ M = [OH^-]$

$pOH = -\log_{10}(0.0013416408) = 2.87$ $\quad (pOH = -\log_{10}[OH^-])$

$pH = 14.00 - 2.87 = \boxed{11.13}$ $\quad (pH + pOH = 14.00)$

Solving this problem with the quadratic equation gives a pH of 11.13 ... same os the short solution!

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

$pH_{, NH_3} = 11.13$

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

$[OH^-] = [NaOH]_{nominal} = 0.100 M$

$pOH = 1.00$

$pH = 13.00$

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: $HCHO_2$

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

(value from p A-13 in textbook...)

Set up equilibrium to solve for hydronium ion concentration.

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

Let "x" equal the change in hydronium ion concentration

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

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

Assume $x \ll 0.10$

so $0.10 - x \approx 0.10$

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

$x = 0.0041231056 \text{ M} = [H_3O^+]$

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.

$$\frac{[CHO_2^-]}{[HCHO_2]_{nominal}} = \frac{[H_3O^+]}{[HCHO_2]_{nominal}} = \frac{0.004123056}{0.100} = \boxed{0.041 = DOI}$$

(or initial)

If we express DOI as a percentage, we call it the "percent ionization":

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

... so in this formic acid solution, about 96% of the formic acid molecules have NOT ionized and are still in molecular form!

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?

$(CH_3)_3N$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^- \quad ; \quad K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

Create an equilibrium chart to begin.

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

Let "x" equal the change in hydroxide ion concentration

Plug into the Kb expression ...

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

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

We still have two variables ... "x" and Kb. If we could find another way of determining "x", we could use this equation to solve for Kb.

The variable "x" we have in the chart as being equal to the hydroxide ion concentration. We know the pH is 11.63, so ...

$$11.63 + pOH = 14.00$$

$(pH + pOH = 14.00)$

$$pOH = 2.37$$

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

$([OH^-] = 10^{-pH})$

$$= 0.004265795\,2 \text{ M } OH^- = x$$

Now just plug in the value of "x" to find Kb ...

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

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

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

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.

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

... 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!