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:
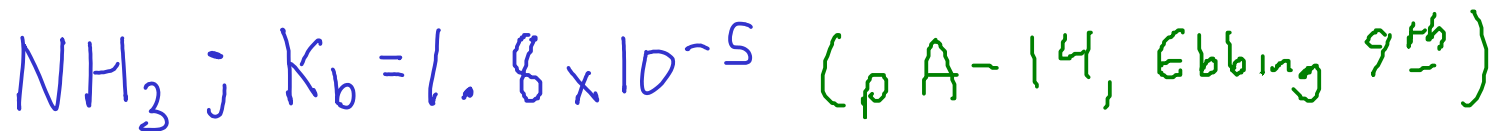
$$0.10 \text{ m } HNO_3, \text{ what is } pH?$$

$$HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3$$

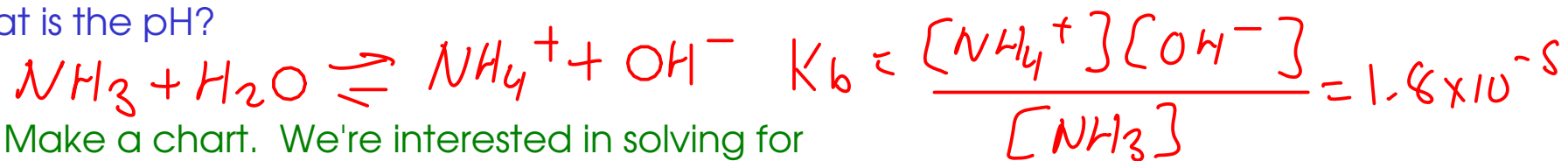$$[H_3O^+] = [HNO_3]_{nominal} = 0.10 \text{ M}$$

$$pH = 1.00$$

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:

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

What is the pH?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^- \quad K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$$

Make a chart. We're interested in solving for HYDROXIDE concentration since it relates to pH...

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

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

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

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

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

We can solve using the quadratic equation, but ...

$x \ll 0.100$, so

$0.100 - x \approx 0.100$

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

$x = 0.0013416408 \, M = [OH^-]$ (see chart on prev. page)

$pOH = 2.87$ $(pOH = -\log [OH^-])$

$pH + 2.87 = 14.00$ $(pH + pOH = 14.00)$

$\boxed{pH = 11.13}$

*If you'd solved this via the quadratic formula, the answer would've been pH = 11.13

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: $HCHO_2$

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

Let's set up an equlibrium chart so we can find hydroniumn ion concentration

(value from p A-14 in textbook...)

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

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

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

$X \ll 0.10$

$0.10 - x \approx 0.10$

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

$X = 0.00412310568 M = [H_3O^+]$

$$\boxed{pH = 2.38}$$

$(pH = -\log[H_3O^+])$

What is degree of ionization? It's the fraction of a weak acid or base that ionizes in solution.

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

$$\frac{0.0041231058 M}{0.10 m} = \boxed{0.041 = DOI}$$

Often we express DOI as a percentage, called "percent ionization".

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

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?

$$(CH_3)_3N$$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-; \quad K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$
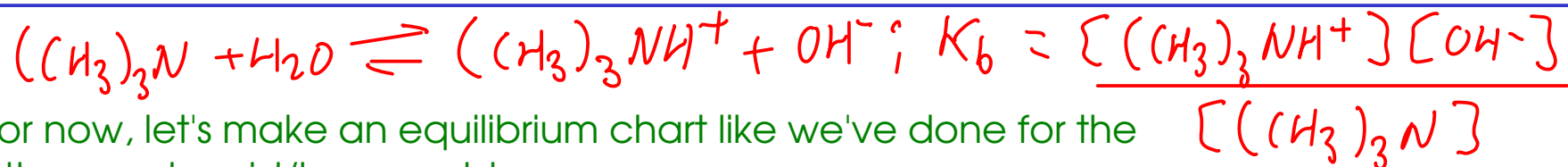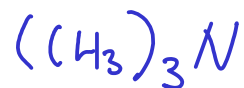
For now, let's make an equilibrium chart like we've done for the other weak acid/base problems.

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 concentration

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

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

At this point, we need to find out what "x" is ... but we can't find it from THIS equation.

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

SInce we know pH, we can use it to find "x". "x" is hydroxide concentration. First. let's use pH to find pOH.

$pH = 11.63$

$11.63 + pOH = 14.00 \quad (pH + pOH = 14.00)$

$pOH = 2.37$

$[OH^-] = x = 10^{-2.37} \quad ([OH^-] = 10^{-pOH})$

$x = 0.004265798752$

Plug "x" back into the expression for Kb, and we've got Kb.

$$\frac{(0.004265798752)^2}{0.25 - 0.004265798752} = \boxed{K_b = 7.4 \times 10^{-5}}$$

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 <u>WEAK ACIDS</u> tend to form <u>BASIC</u> solutions
- Salts made from <u>WEAK BASES</u> tend to form <u>ACIDIC</u> solutions

$$Na_2CO_3 : \quad Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$$
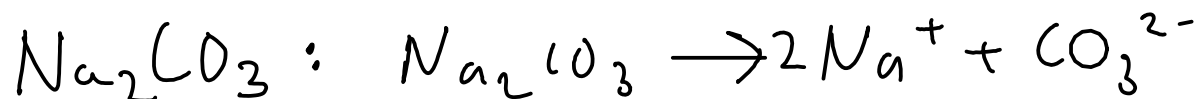
Do any of these ions have acidic or basic properties?

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^-

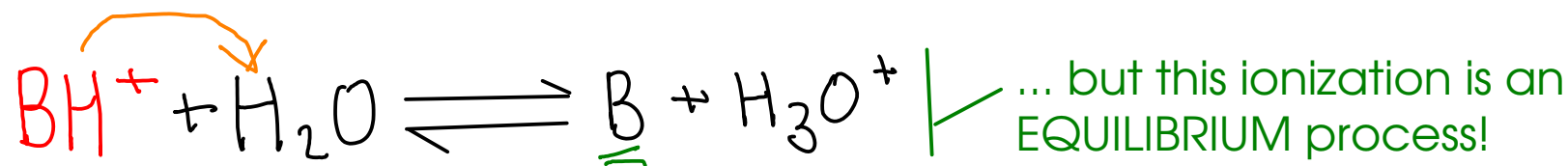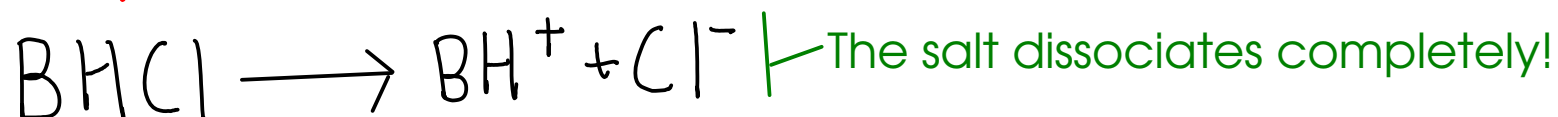Since A^- and HA are a conjugate pair, the ionization constants are related!

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

1.0×10^{-14}

$$14 \approx pK_a + pK_b$$

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^+

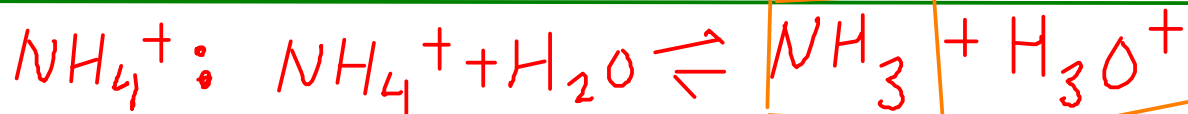$$K_w = \left(K_{a, BH^+}\right)\left(K_{b, B}\right)$$

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!

0.100 M NH_4Cl ...   Find the pH of the solution

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$
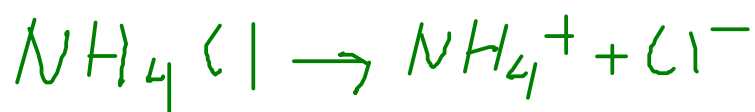
Is the salt acidic, basic, or neutral? Look at the IONS the salt releases when it dissolves!

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

Since ammonia is a WEAK base, that means it will exist in water as molecules ... meaning the reaction we wrote can occur and ammonium ion can act as an acid.

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

HCl is a strong acid. Since it completely ionizes, that means chloride ion isn't able to keep a proton attached, and isn't a base.

So the equilibrium we'll need to look at is the one with the acidic ammonium ion.

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

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = ?$$

$K_{a,NH_4^+} = ?$ We'll need to find Ka for ammonium by looking at Kb for its conjugate, ammonia.

$K_{b,NH_3} = 1.8 \times 10^{-5}$ $(pA-14)$

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

$K_{a,NH_4^+} = 5.56 \times 10^{-10}$

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+ \; ; \; K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$$

Set up and solve this equilibrium like any other weak acid problem!

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

Let "x" equal the change in ammonia concentration...

$$\frac{(X)(X)}{(0.100-X)} = 5.56 \times 10^{-10}$$

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

$X \ll 0.100$,

so $0.100 - X \approx 0.100$

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

$$X = 7.45 \times 10^{-6} M = [H_3O^+]$$

$$pH = -\log_{10}(7.45 \times 10^{-6})$$

$$\boxed{pH = 5.13}$$
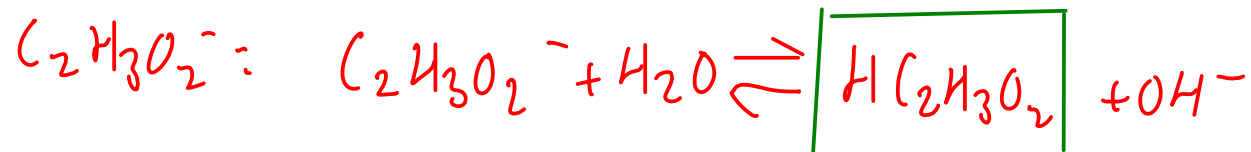
For comparison,
pH = 1.00 for a strong acid like HCl
pH = 2.17 for 0.10 M nitrous acid
pH = 7.00 for distilled water

$0.100 M \ NaC_2H_3O_2$, Find pH

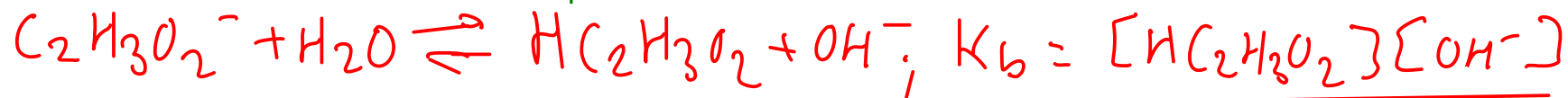$$NaC_2H_3O_2 \rightarrow Na^+ + C_2H_3O_2^-$$

Na^+: Neutral: Na+ has no H+ to donate, and its positive charge makes accepting a proton unlikely.

$C_2H_3O_2^-$: $C_2H_3O_2^- + H_2O \rightleftharpoons \boxed{HC_2H_3O_2} + OH^-$

↳ WEAK ACID, so we think acetate ion can hang on to an H+ and be a BASE.

We need to solve ACETATE ION's equilibrium:

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^- ; \quad K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$
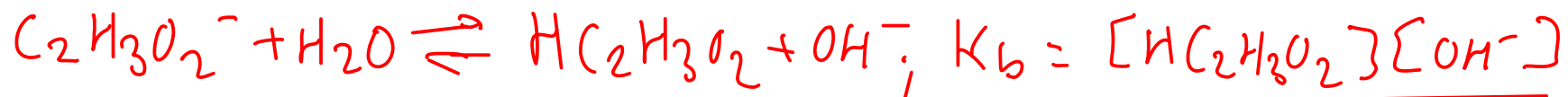
We can't find Kb foir acetate ion on page A-14, but we DO see the Ka for its conjugate (acetic acid) on page A-13. We can convert.

$$K_{a, HC_2H_3O_2} = 1.7 \times 10^{-5}$$

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

$$(1.7 \times 10^{-5}) K_b = 1.0 \times 10^{-14}$$

$$K_b = 5.89 \times 10^{-10}$$

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-; \quad K_b = \dfrac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

Set up an equilibrium chart!

Species	[Initial]	Δ	[Equilibrium]
OH^-	0	+X	X
$HC_2H_3O_2$	0	+X	X
$C_2H_3O_2^-$	0.100	−X	0.100−X

$$K_b = 5.89 \times 10^{-10}$$

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

$$\frac{(X)(X)}{0.100-X} = 5.89 \times 10^{-10}$$

$$\frac{X^2}{0.100-X} = 5.89 \times 10^{-10}$$

$\downarrow X << 0.100$, so $0.100 - X \approx 0.100$

$$\frac{X^2}{0.100} = 5.89 \times 10^{-10}$$

$$X = 7.67 \times 10^{-6} M = [OH^-]$$

$$pOH = 5.11$$

$$pH = 14.00 - 5.11$$

$$= \boxed{8.89}$$

For comparison:

    0.100 M sodium acetate, pH = 8.89

    0.100 M ammonia, pH = 11.13

    0.100 M NaOH (strong base), pH = 13.00

The acetate ion is basic, but it's a very weak base!