

What IS degree of ionization? The fraction of an acid or base that reacts (forms ions) in water.

$$\frac{[\text{CHO}_2^-]}{[\text{HCHO}_2]_{\text{nominal}}} = \frac{[\text{H}_3\text{O}^+]}{[\text{HCHO}_2]_{\text{nominal}}} = \frac{0.0041231056}{0.10} = \boxed{0.041231056}$$

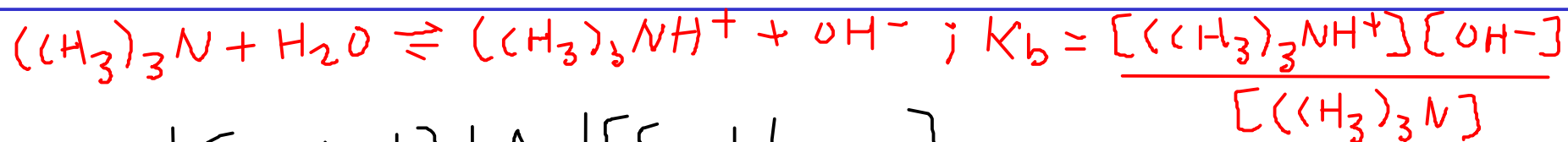
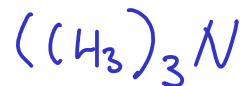
This can also be expressed as a percentage (the PERCENT IONIZATION). To get the percent ionization, just multiply the degree of ionization by 100.

$$0.041 \times 100 = \boxed{4.1\% \text{ ionized}}$$

(So, approximately 96 out of every hundred of the original formic acid molecules are still molecules. Only 4 out of every 100 forms hydronium and formate ions!)

When you do Experiment 16A. By Le Chateleur'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 INCREASE 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 K_b ?



Species	[Initial]	Δ	[Equilibrium]
$(\text{CH}_3)_3\text{NH}^+$	0	+x	x
OH^-	0	+x	x
$(\text{CH}_3)_3\text{N}$	0.25	-x	0.25 - x

Let "x" equal the increase in concentration of trimethylammonium ion...

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

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

We need to solve this expression, but it's got two variables. We need to find some other way to get the value of "x" ...

Since "x" equals hydroxide concentration (see the chart), we can use pH to find "x" ...

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

Find OH⁻ concentration from pH

$$11.63 + pOH = 14.00$$

$$(pH + pOH = 14.00)$$

$$pOH = 2.37$$

$$[OH^-] = 0.0042657952 = x$$

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

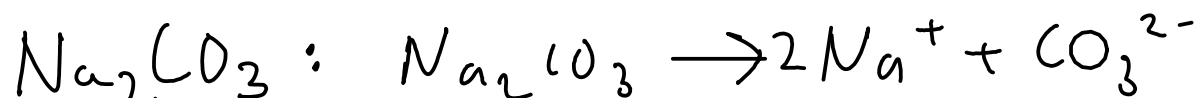
Plug "x" into the Kb expression to find Kb...

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

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



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.



ACID

BASE

SALT OF A WEAK ACID

ex: $\text{NaC}_2\text{H}_3\text{O}_2$ 

For this reaction to occur, HA MUST be stable in water. In other words, a weak acid.



The anion is a BASE. It can accept a proton from water to form the weak (therefore stable as a molecule!) acid HA

$$K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} \quad \text{— This is the base ionization constant for } \text{A}^-$$

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

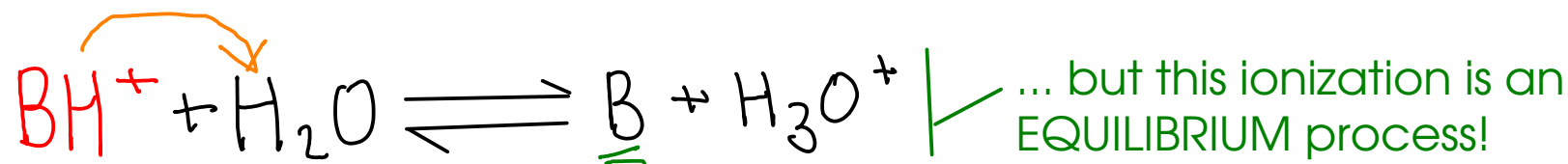
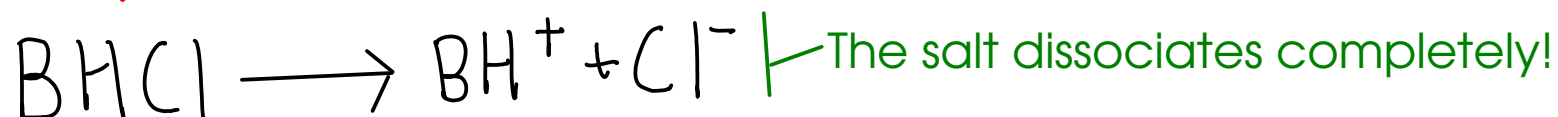
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

1.0×10^{-14}

$$14 \approx \text{p}K_a + \text{p}K_b$$

You will generally not find both the K_a AND K_b for a conjugate pair in the literature, since one can be easily converted to the other!

SALT OF A WEAK BASE

ex: NH_4Cl 

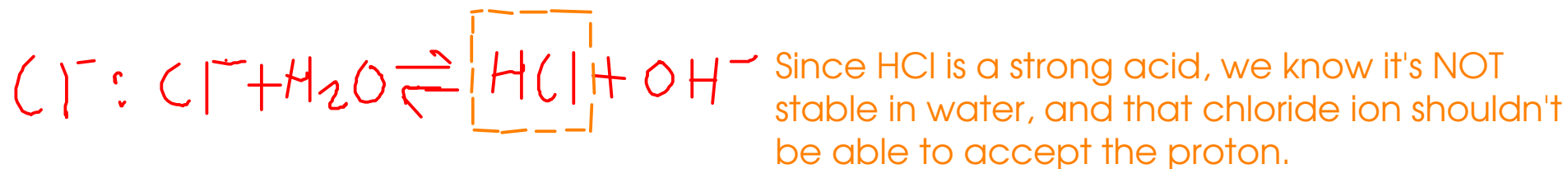
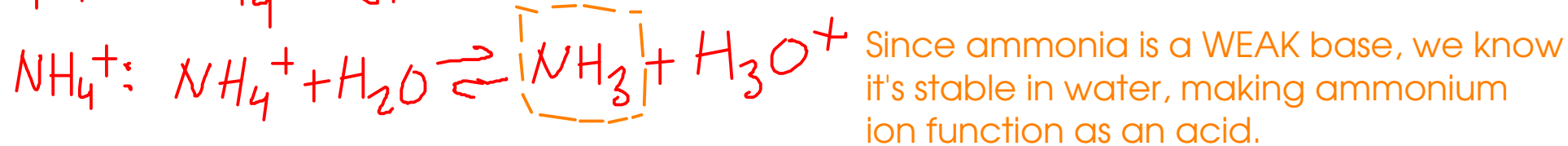
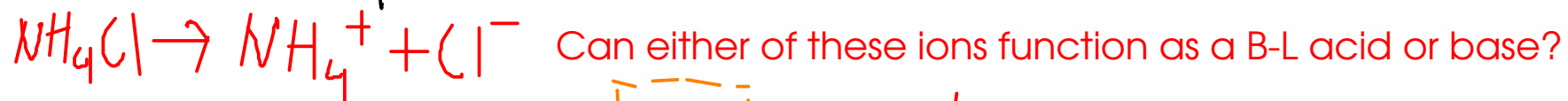
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for } \text{BH}^+ \right.$$

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

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!

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



So, to figure out the pH of the solution, we need to look at ammonium ion's equilibrium $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$; $K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$

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

$$\frac{(x)(x)}{(0.100 - x)} = K_a$$

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

$$\downarrow \begin{array}{l} x \ll 0.100, 50 \\ 0.100 - x \approx 0.100 \end{array}$$

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

$$x = 7.45 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\boxed{\text{pH} = 5.12}$$

We need K_a for ammonium ion. No value for this K_a on page A-13, but on page A-14 we find the K_b for ammonia, the conjugate of ammonium ion... $K_b, \text{NH}_3 = 1.8 \times 10^{-5}$

$$K_a \times K_b = K_w$$

$$K_a (1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$$

$$K_{a, \text{NH}_4^+} = 5.56 \times 10^{-10}$$