

Find the pH and the degree of ionization for an 0.10 M solution of formic acid:  $\text{HCHO}_2$



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = 1.7 \times 10^{-4}$$

The constant's value at 25C is on page A-13 of the textbook (10th ed)

Species	[Initial]	$\Delta$	[Equilibrium]
$\text{H}_3\text{O}^+$	0	+x	x
$\text{CHO}_2^-$	0	+x	x
$\text{HCHO}_2$	0.10	-x	0.10-x

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

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

If we assume 'x' is small compared to 0.10, then  $0.10-x \approx 0.10$

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

$$[\text{H}_3\text{O}^+] = x = 0.0041231056$$

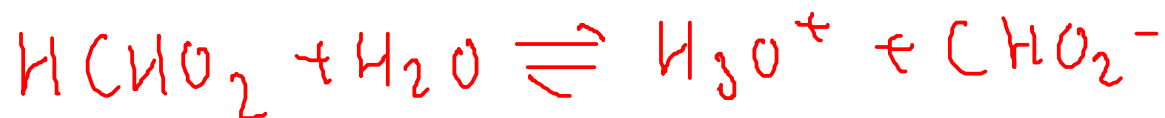
$$\text{pH} = 2.38$$

So what is DEGREE OF IONIZATION? DEGREE OF IONIZATION is the fraction of a weak acid or base that ionizes in water:

$$\frac{[\text{CHO}_2^-]}{[\text{HCHO}_2]_{\text{initial}}} \approx \frac{[\text{H}_3\text{O}^+]}{[\text{HCHO}_2]_{\text{initial}}} = \frac{0.0041231056}{0.10} = \boxed{0.041231056} \approx \boxed{0.041} = \text{DOI}$$

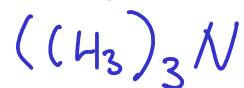
Sometimes, we express the degree of ionization as a percentage ... this is called PERCENT IONIZATION:

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



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 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  $K_b$ ?



$$K_b = \frac{[(\text{CH}_3)_3\text{NH}^+][\text{OH}^-]}{[(\text{CH}_3)_3\text{N}]} = \text{???}$$

Let's try to get rid of some variables by relating the concentrations to each other:

Species	[Initial]	$\Delta$	[Equil]
$(\text{CH}_3)_3\text{NH}^+$	0	+x	x
$\text{OH}^-$	0	+x	x
$(\text{CH}_3)_3\text{N}$	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  $K_b$ , 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 the pH, which gives us HYDRONIUM ION concentration. Since this is related to HYDROXIDE ION concentration, we can use pH to find 'x' ...

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

$$pOH = 2.37$$

$$[OH^-] = 10^{-2.37} = 0.0042657952$$

Since 'x' = the hydroxide concentration, ...

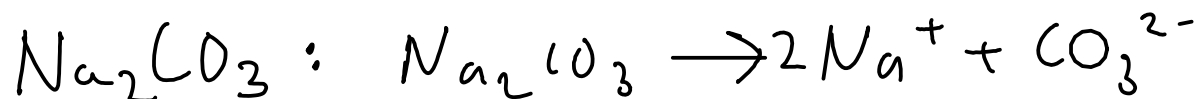
$$x = 0.0042657952$$

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

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

## 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



Do any of these ions have acidic or basic properties?

$\text{Na}^+$ : neutral. Not a proton donor or a proton acceptor

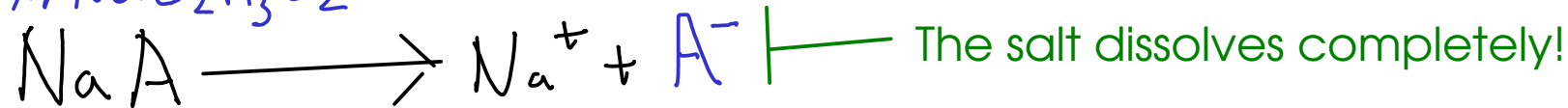
$\text{CO}_3^{2-}$ : BASIC, since it can accept protons to form the weak acid CARBONIC ACID in solution.



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

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

Since  $\text{A}^-$  and HA are a conjugate pair, the ionization constants are related!

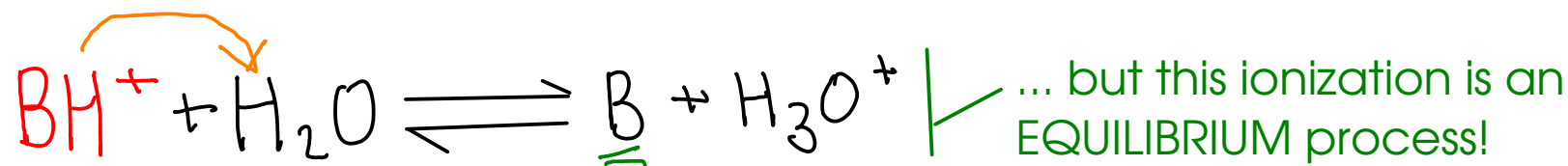
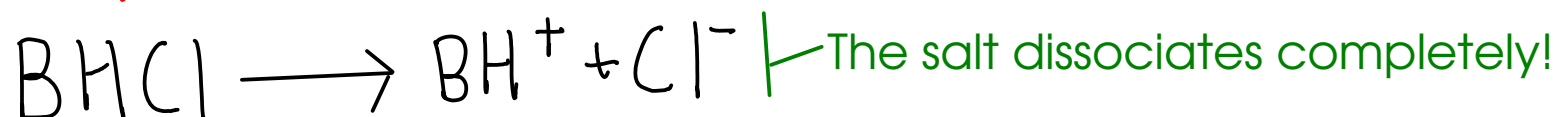
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

$1.0 \times 10^{-14}$

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

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:  $\text{NH}_4\text{Cl}$ 

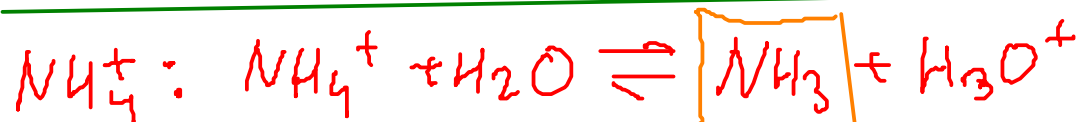
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for } \text{BH}^+ \right.$$

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

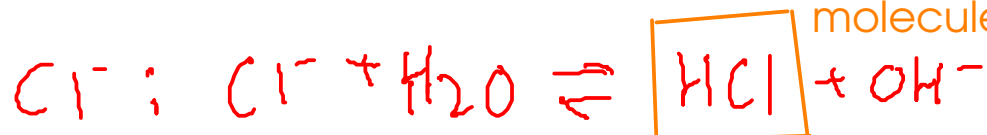
$1.0 \times 10^{-14}$

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  $\text{NH}_4\text{Cl}$  ... Find the pH of the solution



Is this molecule stable in water? YES ... this is the WEAK BASE ammonia. Since it's a weak base, it exists in water mostly as molecules!



HCl is a STRONG ACID, meaning that it completely ionizes in water. Since HCl is unstable in water, chloride ion is not going to function as a base!

So, we'll need to solve this equilibrium:



$$K_{a, \text{NH}_4^+} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

Since  $K_a$  for ammonium isn't in the chart on A-13, look up  $K_b$  for the conjugate base ammonia...

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$$

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

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





$$K_{a, \text{NH}_4^+} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.56 \times 10^{-10}$$

Species	[Initial]	$\Delta$	[Equil]
$\text{NH}_3$	0	+x	x
$\text{H}_3\text{O}^+$	0	+x	x
$\text{NH}_4^+$	0.100	-x	0.100 - x

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

$x \ll 0.100$   
 $\downarrow$   
 $0.100 - x \approx 0.100$

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

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

$$\text{So } \text{pH} = -\log(7.45 \times 10^{-6})$$

$$\text{pH} = 5.13$$

Compare:

pH = 1.00 for 0.100 M strong acid

pH = 2.17 for 0.100 M nitrous acid

pH = 7.00 for distilled water