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

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

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

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

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

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

$$[H_3O^+] = x = 0.0041231056$$

$$\boxed{pH = 2.38}$$

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

$$\frac{[CHO_2^-]}{[HCHO_2]_{initial}} \approx \frac{[H_3O^+]}{[HCHO_2]_{initial}} = \frac{0.00412310 56}{0.10} = \boxed{0.041 = DOI}$$

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

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

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

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?

$(CH_3)_3N$

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

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

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

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

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

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

If we want to find the value of Kb, 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 concnetration, we can use pH to find 'x' ...

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

$$pOH = 2.37$$

$$[OH^-] = 10^{-2.37} = 0.004265795$$

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

$$x = 0.004265795$$

$$K_b = \frac{x^2}{0.25 - x} = \frac{(0.004265795)^2}{0.25 - 0.004265795}$$

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

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 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.

$$\underbrace{H_2CO_3}_{\text{ACID}} + 2H_2O \rightleftharpoons 2H_3O^+ + \underbrace{CO_3^{-2}}_{\text{BASE}}$$

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.

$$\underline{A^-} + H_2O \rightleftharpoons HA + OH^-$$

... 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!

$$\underset{1.0\times10^{-14}}{K_w} = (K_{a,HA})(K_{b,A^-})$$

$$14 = pK_a + pK_b$$

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!

$$BH^+ + H_2O \rightleftharpoons \underline{B} + H_3O^+$$

... 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!

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

$NH_4Cl \rightarrow NH_4^+ + Cl^-$ Is the salt acidic, basic, or neutral?

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

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

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

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

So, we'll need to solve this equilibrium:

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

$$K_{a,NH_4^+} = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

Since Ka for ammonium isn;t in the chart on A-13, look up Kb for the conjugate base ammonia...

$K_{b,NH_3} = 1.8 \times 10^{-5}$

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

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

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

$$K_{a,NH_4^+} = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$$

Species	[Initial]	Δ	[Equil]
NH_3	0	$+x$	x
H_3O^+	0	$+x$	x
NH_4^+	0.100	$-x$	$.100-x$

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

$x \ll 0.100$

$0.100 - x \approx 0.100$

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

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

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

$$\boxed{pH = 5.13}$$

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