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^+

$$\underbrace{K_w}_{1.0\times10^{-14}} = (K_{a,BH^+})(K_{b,B})$$

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)

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

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

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+ ; \quad K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56\times10^{-10}$$

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...

$$\frac{(x)(x)}{(0.100-x)} = 5.56\times10^{-10}$$

$$\frac{x^2}{0.100-x} = 5.56\times10^{-10}$$

$x \ll 0.100$,
so $0.100 - x \approx 0.100$

$$\frac{x^2}{0.100} = 5.56\times10^{-10}$$

$x = 7.45\times10^{-6}\ M = [H_3O^+]$

$pH = -\log_{10}(7.45\times10^{-6})$

$$\boxed{pH = 5.13}$$

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^+: Cannot be B-L acid, as it's missing an H+ to donate. Has a positive charge, so unlikely to be a B-L base (would repel proton). Neutral.

$C_2H_3O_2^-$: Has hydrogens, but also has a negative charge. Negative charge is attractive to protons, so a base?

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

This is ACETIC ACID, a WEAK acid. Since acetic acid is weak (and stable in water), we expect the acetate ion to be BASIC.

So, we need to solve the equilbrium of acetate ion reacting as a base with water.

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} = ?$$

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

Since $K_a \times K_b = K_w$

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

$K_b = 5.88 \times 10^{-10}$

(data from pages A-13/A-14 in Ebbing)

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$

$(0.100\ M\ NaC_2H_3O_2)$

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} = 5.88 \times 10^{-10}$$

Set up an equilibrium chart and solve

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

Let "x" equal the change in acetic acid concentration.

$$\frac{(x)(x)}{(0.100 - \cancel{x})} = 5.88 \times 10^{-10}$$

Assume x << 0.100,
so 0.100-x = 0.100

$$\frac{x^2}{0.100} = 5.88 \times 10^{-10}$$

$$x = 7.6696 \times 10^{-6} = [OH^-]$$

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

$$pOH = -\log_{10}(7.6696 \times 10^{-6}) = 5.12$$

$$pH = 14 - 5.12 = \boxed{8.88}$$

$(pH + pOH = 14.00)$

For comparison:

0.100 M sodium acetate, pH = 8.88

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!

0.100 M NaCl, Find pH

$$NaCl \rightarrow Na^+ + Cl^-$$

Na^+: Neither a B-L acid nor base (no protons to donate, and the positive charge hinders acceptance of a proton). Neutral.

Cl^-: Base?

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

STRONG ACID, so chloride ion will not be a proton acceptor!

... so chloride ion is also neutral.

In this case, the pH of the solution is controlled by water's own ionization, so the solution would be pH = 7.00.

POLYPROTIC ACIDS

Find pH of 0,10 M H_3PO_4

... what's special about phosphoric acid?

① $H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$

② $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$

③ $HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$

Phosphoric acid has THREE acidic protons!

$K_{a1} = 6,9 \times 10^{-3}$

$K_{a2} = 6,2 \times 10^{-8}$

$K_{a3} = 4,8 \times 10^{-13}$

The first dissocation is dominant here, and for simple calculations of phosphoric acid in water, we will simply use the first ionization and ignore the other two.

Remember: This is a weak acid. It exists in water mostly as undissociated phosphoric acid molecules.

Solve the equilibrium of phosphoric acid's FIRST proton:

$$H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+ \; ; \; K_a = 6.9 \times 10^{-3}$$

$$K_a = \frac{[H_2PO_4^-][H_3O^+]}{[H_3PO_4]} = 6.9 \times 10^{-3}$$

Species	[Initial]	Δ	[Equilibrium]
$H_2PO_4^-$	0	$+x$	x
H_3O^+	0	$+x$	x
H_3PO_4	0.10	$-x$	$0.10 - x$

$$\frac{x^2}{0.10 - x} = 6.9 \times 10^{-3}$$

assume $x \ll 0.10$
so $0.10 - x \simeq 0.10$

$$\frac{x^2}{0.10} = 6.9 \times 10^{-3}$$

$$x = 0.026267851 = [H_3O^+]$$

So, $pH = -\log(0.026267851)$

$$\boxed{pH = 1.58}$$