$0.100\,M\ NH_4Cl$ ... Find the pH of the solution

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$

Acidic, basic, or neutral salt?

This is the WEAK BASE ammonia. Stable in water.

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

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

This is a STRONG ACID, which does not exist as a stable molecule in water.

The conjugate of a strong acid or base is NEUTRAL - does not affect pH!

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

This equilibrium affects the pH, so it is the equilibrium we'll need to solve to find pH!

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

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

Where do we get this Ka?

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

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

	initial	Δ	equilibrium
$[NH_3]$	0	$+x$	x
$[H_3O^+]$	0	$+x$	x
$[NH_4^+]$	0.100	$-x$	$0.100 - x$

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

$x << 0.100$

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

$$x = 7.46 \times 10^{-6}$$

We defined "x" as the concentration of hydronium ion made by the equilibrium!

$$x = 7.46 \times 10^{-6} = [H_3O^+]$$

$$pH = -\log_{10}(7.46 \times 10^{-6}) = \boxed{5.13 = pH}$$

Compare:

pH = 1.00 for 0.100 M strong acid

pH = 2.16 for 0.100 M nitrous acid

pH = 7.00 for pure distilled water

0.100 M $NaC_2H_3O_2$, Find pH

$$NaC_2H_3O_2 \longrightarrow \underline{Na^+} + \underline{C_2H_3O_2^-}$$

conjugate of acetic acid

$HC_2H_3O_2$

neutral

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

Since acetic acid is a weak acid and stable in water, we would expect acetate ion to be BASIC.

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

$K_{a, HC_2H_3O_2} = 1.7 \times 10^{-5}$; $K_a \times K_b = 1.0 \times 10^{-14}$ ... true for conjugate pairs

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

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

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

	initial	Δ	equilibrium
$[HC_2H_3O_2]$	0	$+x$	x
$[OH^-]$	0	$+x$	x
$[C_2H_3O_2^-]$	0.100	$-x$	$0.100-x$

Define "x" as the concentration of hydroxide produced by the equilibrium, since that's closely related to pH

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

$$\Big\downarrow \; x \ll 0.100$$

$$\frac{x^2}{0.100} = 5.88 \times 10^{-10} \; ; \; x = 7.67 \times 10^{-6}$$

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

$pOH = \log_{10}(7.67 \times 10^{-6})$

$pOH = 5.12$ ... this is pOH, but we need pH. Luckily for us, they are related very simply.

$pH + pOH = 14.00$

so, $pH = 14.00 - 5.12$

$pH = \boxed{8.88}$

Compare:
pH = 13.00 for 0.100 M strong base
pH = 7.00 for pure distilled water

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 exsits in water mostly as undissociated phosphoric acid molecules.

Solving the equilibrium of phosphoric acid's first proton:

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

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

	initial	Δ	equilibrium
$[H_2PO_4^-]$	0	$+x$	x
$[H_3O^+]$	0	$+x$	x
$[H_3PO_4]$	0.10	$-x$	$0.10-x$

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

For this problem, the assumption that $x \ll 0.10$ might not be safe (Ka is fairly large for a weak acid) , so we should solve the quadratic instead...

$$x = 0.023043 \text{ or } \cancel{-0.029943}$$

$$[H_3O^+] = \underline{0.023043\,M}$$

Less than 0.10, but not MUCH less!

$$\boxed{pH = 1.64}$$